NEHRU COLLEGE OF ENGINEERING AND RESEARCH CENTRE
(NAAC Accredited)
(Approved by AICTE, Affiliated to APJ Abdul Kalam Technological University, Kerala)

DEPARTMENT OF MECHATRONICS ENGINEERING
COURSE MATERIAL

MR 307 THERMODYNAMICS

VISION OF THE INSTITUTION
To mould true citizens who are millennium leaders and catalysts of change through excellence in education.
MISSION OF THE INSTITUTION
NCERC is committed to transform itself into a center of excellence in Learning and Research in Engineering and Frontier Technology and to impart quality education to mould technically competent citizens with moral integrity, social commitment and ethical values. We intend to facilitate our students to assimilate the latest technological know-how and to imbibe discipline, culture and spiritually, and to mould them in to technological giants, dedicated research scientists and intellectual leaders of the country who can spread the beams of light and happiness among the poor and the underprivileged.

ABOUT DEPARTMENT

- Established in: 2013
- Course offered: B.Tech Mechatronics Engineering
- Approved by AICTE New Delhi and Accredited by NAAC
- Affiliated to the University of Dr. A P J Abdul Kalam Technological University.

DEPARTMENT VISION

To develop professionally ethical and socially responsible Mechatronics engineers to serve the humanity through quality professional education.

DEPARTMENT MISSION

1) The department is committed to impart the right blend of knowledge and quality education to create professionally ethical and socially responsible graduates.

2) The department is committed to impart the awareness to meet the current challenges in technology.

3) Establish state-of-the-art laboratories to promote practical knowledge of mechatronics to meet the needs of the society

PROGRAMME EDUCATIONAL OBJECTIVES

I. Graduates shall have the ability to work in multidisciplinary environment with good professional and commitment.

II. Graduates shall have the ability to solve the complex engineering problems by applying electrical, mechanical, electronics and computer knowledge and engage in lifelong learning in their profession.

III. Graduates shall have the ability to lead and contribute in a team with entrepreneur skills, professional, social and ethical responsibilities.

IV. Graduates shall have ability to acquire scientific and engineering fundamentals necessary for higher studies and research.

PROGRAM OUTCOME (PO'S)

Engineering Graduates will be able to:

PO 1. Engineering knowledge: Apply the knowledge of mathematics, science, engineering fundamentals, and an engineering specialization to the solution of complex engineering problems.

PO 2. Problem analysis: Identify, formulate, review research literature, and analyze complex engineering problems reaching substantiated conclusions using first principles of mathematics, natural sciences, and engineering sciences.

PO 3. Design/development of solutions: Design solutions for complex engineering problems and design system components or processes that meet the specified needs with appropriate consideration for the public health and safety, and the cultural, societal, and environmental considerations.

PO 4. Conduct investigations of complex problems: Use research-based knowledge and research methods including design of experiments, analysis and interpretation of data, and synthesis of the information to provide valid conclusions.

PO 5. Modern tool usage: Create, select, and apply appropriate techniques, resources, and modern engineering and IT tools including prediction and modeling to complex engineering activities with an understanding of the limitations.

PO 6. The engineer and society: Apply reasoning informed by the contextual knowledge to assess societal, health, safety, legal and cultural issues and the consequent responsibilities relevant to the professional engineering practice.

PO 7. Environment and sustainability: Understand the impact of the professional engineering solutions in societal and environmental contexts, and demonstrate the knowledge of, and need for sustainable development.

PO 8. Ethics: Apply ethical principles and commit to professional ethics and responsibilities and norms of the engineering practice.

PO 9. Individual and team work: Function effectively as an individual, and as a member or leader in diverse teams, and in multidisciplinary settings.

PO 10. Communication: Communicate effectively on complex engineering activities with the engineering community and with society at large, such as, being able to comprehend and write effective reports and design documentation, make effective presentations, and give and receive clear instructions.

PO 11. Project management and finance: Demonstrate knowledge and understanding of the engineering and management principles and apply these to one’s own work, as a member and leader in a team, to manage projects and in multidisciplinary environments.

PO 12. Life-long learning: Recognize the need for, and have the preparation and ability to engage in independent and life-long learning in the broadest context of technological change.

PROGRAM SPECIFIC OUTCOME(PSO’S)

PSO 1: Design and develop Mechatronics systems to solve the complex engineering problem by integrating electronics, mechanical and control systems.

PSO 2: Apply the engineering knowledge to conduct investigations of complex engineering problem related to instrumentation, control, automation, robotics and provide solutions.

COURSE OUTCOME

After the completion of the course the student will be able to

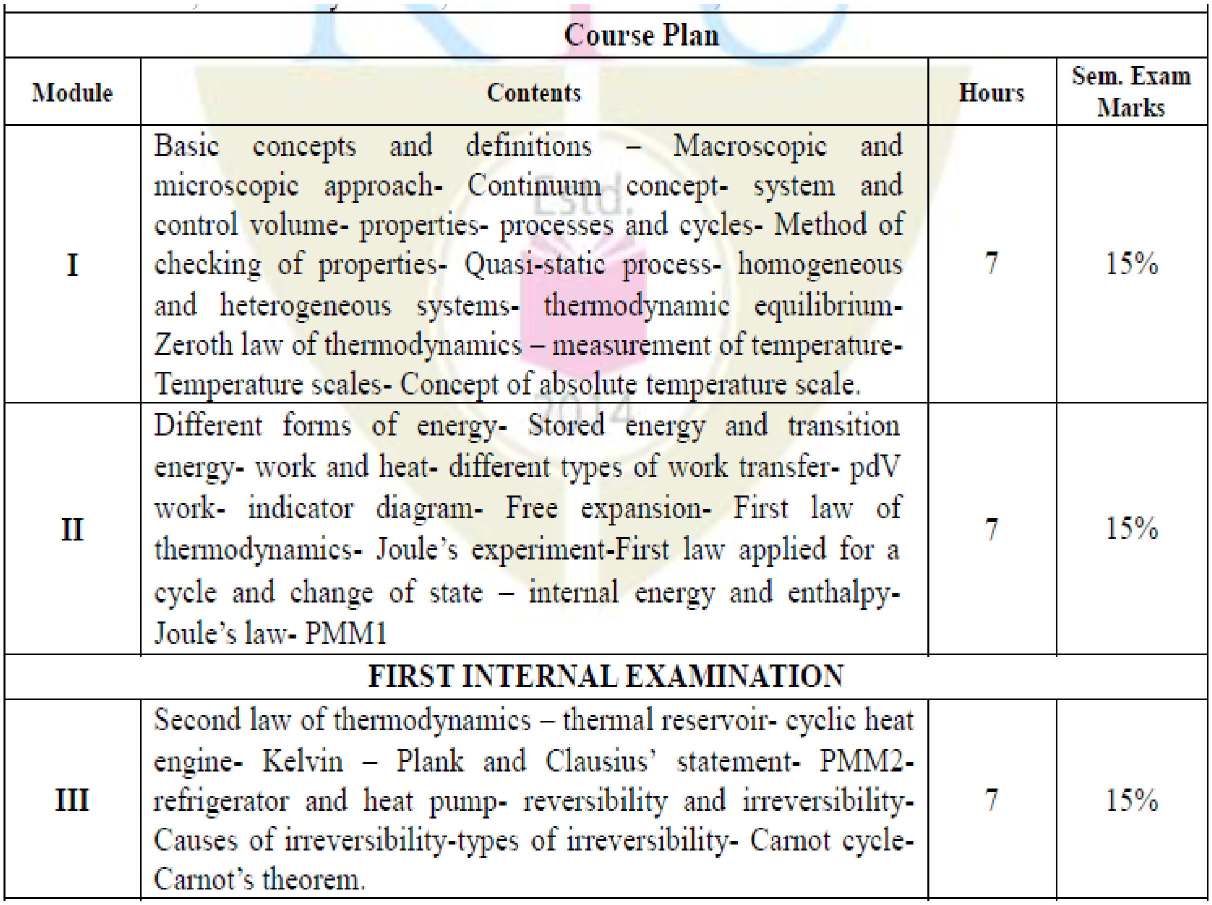
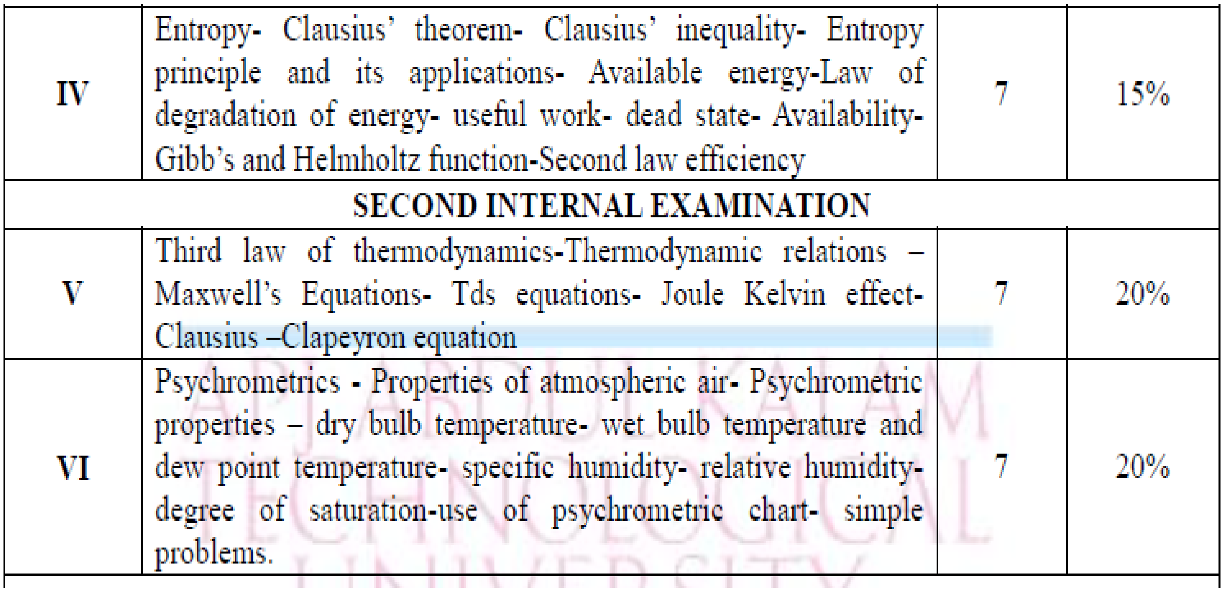
CO 1	Understand about the concept of thermodynamics
CO 2	Acquire knowledge about the concepts of energy and 1st law of thermodynamics
CO 3	Define the concepts of 2nd law of thermodynamics
CO 4	Understand the concepts of entropy and law of degradation of energy
CO 5	Identify the concepts of 3rd law of thermodynamics
CO 6	Analyze about the psychometric properties of atmospheric air

CO VS PO'S AND PSO'S MAPPING

CO	PO1	PO2	PO3	PO4	PO5	PO6	PO7	PO8	PO9	PO10	PO11	PO12	PS01	PSO2
C304.1	3	-	2	-	-	-	-	-	-	-	-	3	2	2
C304.2	3	2	-	-	-	-	-	-	-	-	-	3	2	-
C304.3	3	1	-	1	-	-	-	-	-	-	-	3	2	1
C304.4	3	-	1	-	-	-	-	-	-	-	-	3	2	-
C304.5	3	1	-	-	-	-	-	-	-	-	-	3	2	-
C304.6	3	2	-	-	-	-	-	-	-	-	-	3	2	-

Note: H-Highly correlated=3, M-Medium correlated=2, L-Less correlated=1

SYLLABUS

Course Plan

Module	Contents	Hours	Sem. Exam Marks
I	Basic concepts and definitions – Macroscopic and microscopic approach- Continuum concept- system and control volume- properties- processes and cycles- Method of checking of properties- Quasi-static process- homogeneous and heterogeneous systems- thermodynamic equilibrium- Zeroth law of thermodynamics – measurement of temperature- Temperature scales- Concept of absolute temperature scale.	7	15%
II	Different forms of energy- Stored energy and transition energy- work and heat- different types of work transfer- pdV work- indicator diagram- Free expansion- First law of thermodynamics- Joule's experiment-First law applied for a cycle and change of state – internal energy and enthalpy- Joule's law- PMM1	7	15%
FIRST INTERNAL EXAMINATION			
III	Second law of thermodynamics – thermal reservoir- cyclic heat engine- Kelvin – Plank and Clausius' statement- PMM2- refrigerator and heat pump- reversibility and irreversibility- Causes of irreversibility-types of irreversibility- Carnot cycle- Carnot's theorem.	7	15%
IV	Entropy- Clausius' theorem- Clausius' inequality- Entropy principle and its applications- Available energy-Law of degradation of energy- useful work- dead state- Availability- Gibb's and Helmholtz function-Second law efficiency	7	15%
SECOND INTERNAL EXAMINATION			
V	Third law of thermodynamics-Thermodynamic relations – Maxwell's Equations- Tds equations- Joule Kelvin effect- Clausius –Clapeyron equation	7	20%
VI	Psychrometrics - Properties of atmospheric air- Psychrometric properties – dry bulb temperature- wet bulb temperature and dew point temperature- specific humidity- relative humidity- degree of saturation-use of psychrometric chart- simple problems.	7	20%

QUESTION BANK

Q:NO:	QUESTIONS	CO	KL	PAGE NO:
	MODULE I			
1	What do you understand by macroscopic and microscopic viewpoints?	CO1	K2	10
2	What are thermodynamic systems?	CO1	K1	12
3	Explain the different thermodynamic system with example	CO1	K1	13
4	Define the terms : aChange of state b) Path c) Process	CO1	K1	13
5	Explain thermodynamics equilibrium	CO1	K1	14
6	What is quasistatic process	CO1	K1	15
7	Define Zeroth law of thermodynamics and explain its application	CO1	K3	18
8	Explain in detail about measurement of temperature	CO1	K1	19
9	What are the intensive and extensive properties	CO1	K1	25
10	Difference between extensive and intensive properties	CO1	K2	25
	MODULE II			
1	What is pdV work	CO2	K2	29
2	Discuss about path and point functions	CO2	K2	34
3	What is an indicator diagram? Explain	CO2	K2	37
4	What is free expansion	CO2	K1	45
5	Discuss about heat transfer	CO2	K2	46
6	What is a PMM1? Why is it impossible	CO2	K2	56
7	State the first law of thermodynamics for a closed system undergoing a cycle	CO2	K1	52
8	State the first law of thermodynamics for a closed system undergoing a change of state	CO2	K1	52

9	What is meant by specific heat	CO2	K1	47
10	Types of work transfer	CO2	K1	27
11	Problems	CO2	K5	57

MODULE III

1	Explain Kelvin Planck statement	CO3	K1	72
2	Explain Carnot cycle and elucidate the four processes that are associated with the cycle	CO3	K1	77
3	Elaborate second law of thermodynamics	CO3	K1	72
4	Explain cyclic heat engine.	CO3	K1	69
5	Explain carnot's theorem	CO3	K1	77
6	What are the causes of irreversibility	CO3	K2	76
7	Types of irreversibility	CO3	K1	76
8	Explain Clausius statement	CO3	K1	74
9	Discuss about refrigerator and heat pump	CO3	K1	71
10	Equivalence of KP and clausius statement	CO3	K1	74

MODULE IV

1	State and prove Clausius' theorem?	CO4	K1	80
2	Explain entropy principle and its applications	CO4	K3	92
3	Establish the inequality of Clausius.	CO4	K2	86
4	Discuss about available energy referred to a cycle	CO4	K1	93
5	What is meant by availability	CO4	K1	92
6	Explain available and unavailable energy	CO4	K1	92
7	Discuss about properties of entropy	CO4	K1	84
8	Explain entropy principle with suitable application	CO4	K3	92
9	Derive the inequality of clausius.	CO4	K3	86
10	Derive Helmholtz and Gibbs function	CO4	K3	94

MODULE V				
1	Derive Tds equations.	CO5	K3	102
2	What is third law of thermodynamics	CO5	K1	97
3	Derive Maxwell equations	CO5	K3	99
4	Explain Joule- Kelvin effect. What is inversion temperature	CO5	K2	104
5	Derive Clausius- Clapeyron equation	CO5	K3	103
6	What is Joule- Thomson coefficient? Why is it zero for an ideal gas?	CO5	K2	105
7	Explain about inversion temperature	CO5	K2	106
8	Explain about inversion line	CO5	K2	106
9	Explain about TdS equations	CO5	K3	102
10	Plot the T-P curve showing the important regions associated with JK effect	CO5	K3	106
MODULE VI				
1	What is relative humidity? How is it defined as the ratio of two mole fractions?	CO6	K1	107
2	What is degree of saturation? What are its limiting values	CO6	K1	108
3	Explain the phychrometric properties	CO6	K1	107
4	Explain specific humidity and relative humidity	CO6	K1	108
5	What do you understand by dry bulb temperature, wet bulb temperature and dew point temperature?	CO6	K2	107
6	What is degree of saturation? What are the limiting	CO6	K1	108

	values?			
7	What is relative humidity? How is it defined as the ratio of two mole fractions?	CO6	K1	108
8	Explain any three psychometric properties	CO6	K2	107
9	Define the terms in the psychometric processes (a) Sensible heating b) Sensible cooling c) By-pass factor	CO6	K1	109
10	What is meant by specific humidity, relative humidity and degree of saturation?	CO6	K1	108

APPENDIX 1

CONTENT BEYOND THE SYLLABUS

S:NO;	**TOPIC**	**PAGE NO:**
1	AIR CONDITIONING	113
2	HEATING, VENTILATION, AND AIR CONDITIONING	114

MODULE – 1

BASIC CONCEPTS & ZEROTH LAW

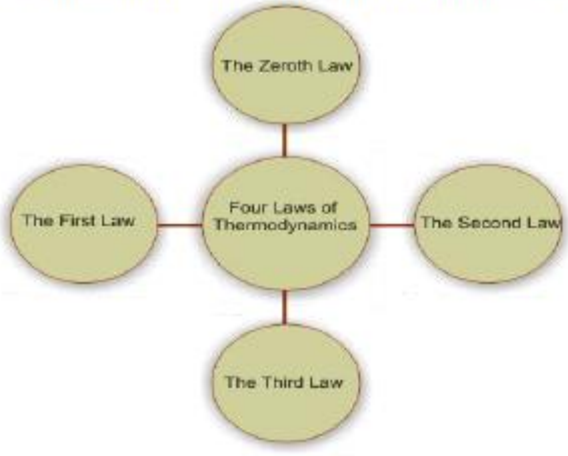
What is Thermodynamics?

- Thermodynamics is a science dealing with Energy and its transformation and its effect on the physical properties of substances.
- It deals with equilibrium and feasibility of a process.
- Deals with the relationship between heat and work and the properties of systems in equilibrium.

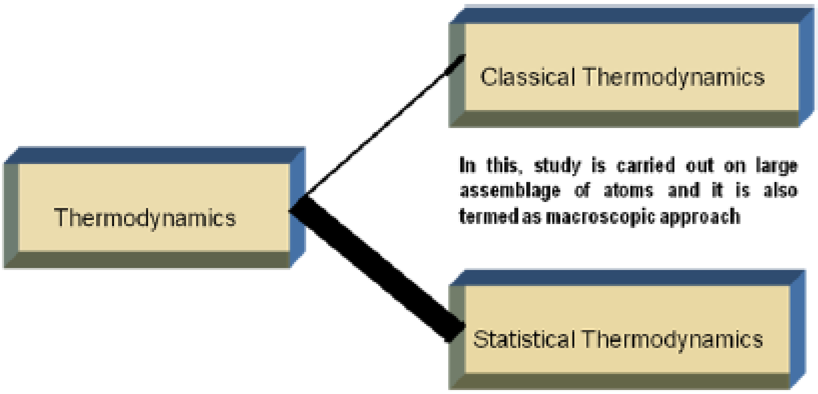
Scope of Thermodynamics:

- Steam power plant
- Separation and Liquification Plant
- Refrigeration
- Air-conditioning and Heating Devices.
- Internal combustion engine
- Chemical power plants
- Turbines
- Compressors, etc

The principles of thermodynamics are summarized in the form of four thermodynamic laws:

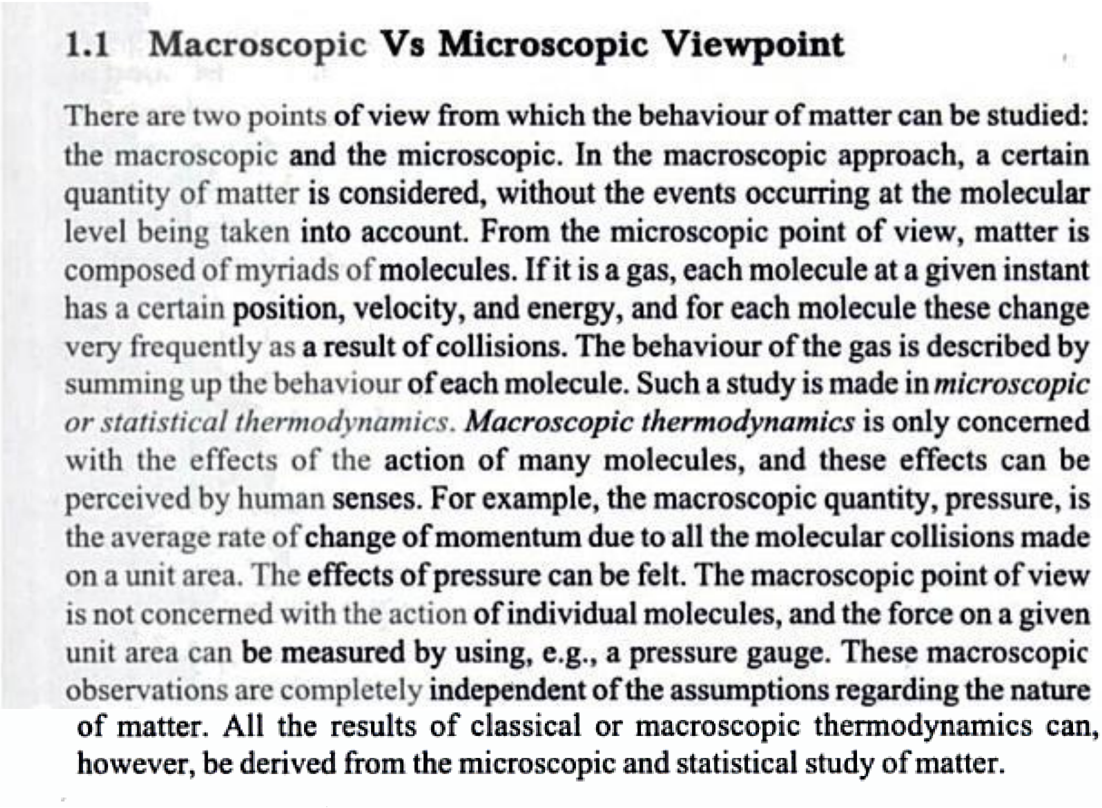


Different Approaches of Thermodynamics :



In this, study is carried out on large assemblage of atoms and it is also termed as macroscopic approach

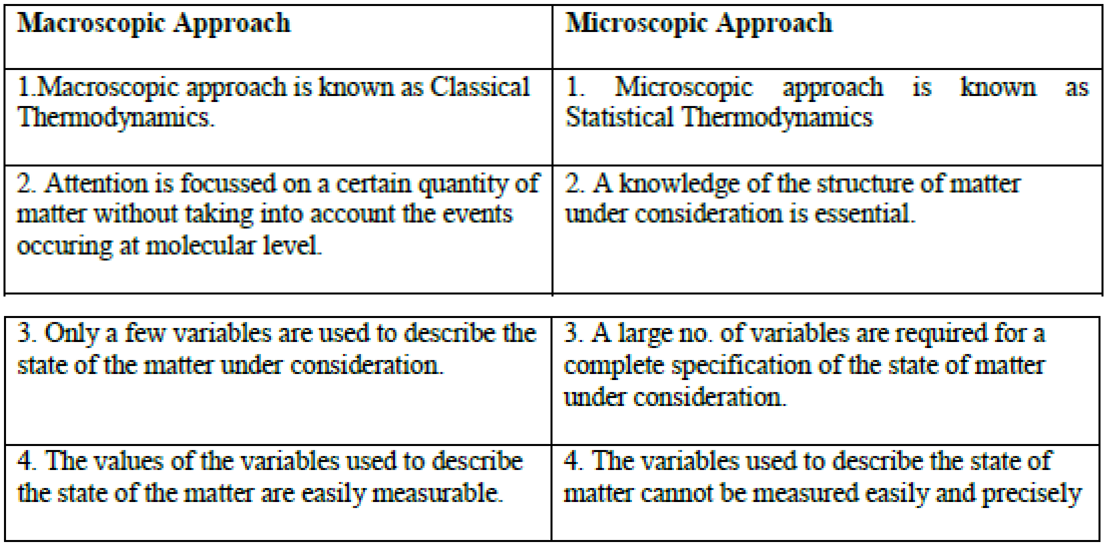
In this, study is carried out at the molecular level and it is also termed as microscopic approach



1.1 Macroscopic Vs Microscopic Viewpoint

There are two points of view from which the behaviour of matter can be studied: the macroscopic and the microscopic. In the macroscopic approach, a certain quantity of matter is considered, without the events occurring at the molecular level being taken into account. From the microscopic point of view, matter is composed of myriads of molecules. If it is a gas, each molecule at a given instant has a certain position, velocity, and energy, and for 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 a study is made in *microscopic or statistical thermodynamics*. *Macroscopic thermodynamics* is only concerned with the effects of the action of many molecules, and these effects can be perceived by human senses. For example, the macroscopic quantity, pressure, is the average rate of change of momentum due to all the molecular collisions made on a 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 microscopic and statistical study of matter.

Write the difference between Macroscopic and Microscopic approach of Thermodynamics:



Macroscopic Approach	**Microscopic Approach**
1.Macroscopic approach is known as Classical Thermodynamics.	1. Microscopic approach is known as Statistical Thermodynamics
2. Attention is focussed on a certain quantity of matter without taking into account the events occuring at molecular level.	2. A knowledge of the structure of matter under consideration is essential.
3. Only a few variables are used to describe the state of the matter under consideration.	3. A large no. of variables are required for a complete specification of the state of matter under consideration.
4. The values of the variables used to describe the state of the matter are easily measurable.	4. The variables used to describe the state of matter cannot be measured easily and precisely

1.2 Thermodynamic System and Control Volume

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 (Fig. 1.1). The boundary may be either *fixed* or *moving*. A system and its surroundings together comprise a *universe*.

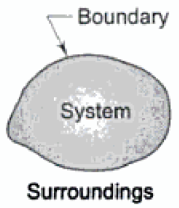


Fig. 1.1 *A thermodynamic system*

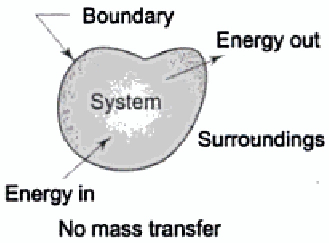


Fig. 1.2 *A closed system*

There are three classes of systems: (a) closed system, (b) open system and (c) isolated system. The *closed* system (Fig. 1.2) is a system of fixed mass. There is no mass transfer across the system boundary. There may be energy transfer into or out of the system. A certain quantity of fluid in a cylinder bounded by a piston constitutes a closed system. The *open system* (Fig. 1.3) is one in which matter crosses the boundary of the system. There may be energy transfer also. Most of the engineering devices are generally open systems, e.g., an air compressor in which air enters at low pressure and leaves at high pressure and there are energy transfers across the system boundary. The *isolated system* (Fig. 1.4) is one in which there is no interaction between the system and the surrounding. It is of fixed mass and energy, and there is no mass or energy transfer across the system boundary.

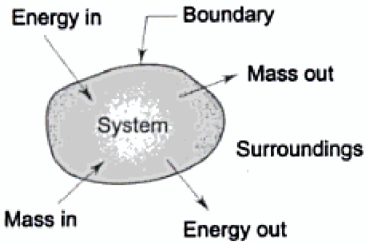


Fig. 1.3 *An open system*

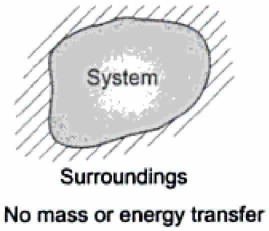


Fig. 1.4 *An isolated system*

If a system is defined as a certain quantity of matter, then the system contains the same matter and there can be no transfer of mass across its boundary.

However, if a system is defined as a region of space within a prescribed boundary, then matter can cross the system boundary. While the former is called a closed system, the latter is an open system.

For thermodynamic analysis of an open system, such as an air compressor (Fig. 1.5), attention is focussed on a certain volume in space surrounding the compressor, known as the *control volume*, bounded by a surface called the *control surface*. Matter as well as energy crosses the control surface.

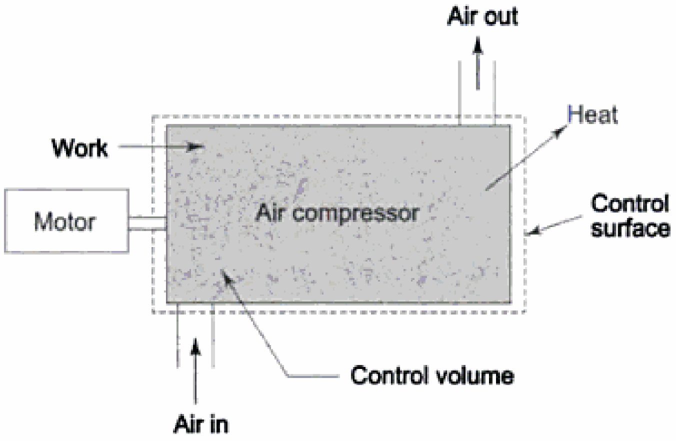


Fig. 1.5 *Control volume and control surface*

A closed system is a system closed to matter flow, though its volume can change against a flexible boundary. When there is matter flow, then the system is considered to be a volume of fixed identity, the control volume. There is thus no difference between an open system and a control volume.

1.3 Thermodynamic Properties, Processes 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 the system. These are all *macroscopic* in nature. When all the 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*, e.g., a constant pressure process. A thermodynamic *cycle* is defined as a series of state changes such that the final state is identical with the initial state (Fig. 1.6)

Properties may be of two types. *Intensive properties* are independent of the mass in the system, e.g., pressure, temperature, etc. *Extensive properties* are related to mass, e.g., volume, energy, etc. If mass is increased, the values of the extensive properties also increase. Specific extensive properties, i.e., extensive properties per unit mass, are intensive properties, e.g., specific volume, specific energy, density, etc.

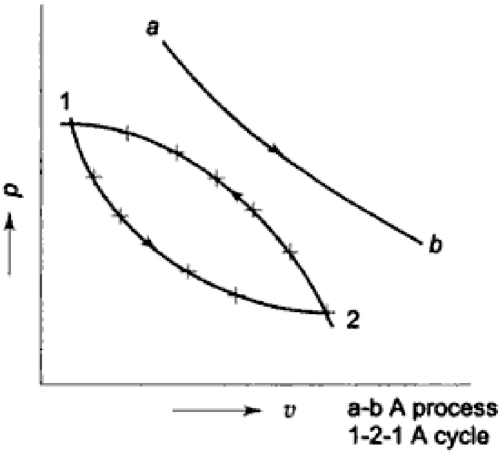


Fig. 1.6 *A process and a cycle*

1.4 Homogeneous and Heterogeneous Systems

A quantity of matter homogeneous throughout in chemical composition and physical structure is called a *phase*. Every substance can exist in any one of the three phases, viz., solid, liquid and gas. A system consisting of a single phase is called a *homogeneous system*, while a system consisting of more than one phase is known as a *heterogeneous system*.

1.5 Thermodynamic Equilibrium

A system is said to exist in a state of *thermodynamic equilibrium* when no change in any macroscopic property is registered, if the system is isolated from its surroundings.

An isolated system always reaches in course of time a state of thermodynamic equilibrium and *can never depart from it spontaneously*.

Therefore, there can be no *spontaneous change in any macroscopic property* if the system exists in an equilibrium state. Thermodynamics studies mainly the properties of physical systems that are found in equilibrium states.

A system will be in a state of thermodynamic equilibrium, if the conditions for the following three types of equilibrium are satisfied:

(a) Mechanical equilibrium
(b) Chemical equilibrium
(c) Thermal equilibrium

In the absence of any unbalanced force within the system itself and also between the system and the surroundings, the system is said to be in a state of *mechanical equilibrium*. If an unbalanced force exists, either the system alone or both the system and the surroundings will undergo a change of state till mechanical equilibrium is attained.

If there is no chemical reaction or transfer of matter from one part of the system to another, such as diffusion or solution, the system is said to exist in a state of *chemical equilibrium*.

When a system existing in mechanical and chemical equilibrium is separated from its surroundings by a diathermic wall (diathermic means 'which allows heat to flow') and if there is no spontaneous change in any property of the system, the system is said to exist in a state of *thermal equilibrium*. When this is not satisfied, the system will undergo a change of state till thermal equilibrium is restored.

When the conditions for any one of the three types of equilibrium are not satisfied, a system is said to be in a *nonequilibrium state*. If the nonequilibrium of the state is due to an unbalanced force in the interior of a system or between the system and the surrounding, the pressure varies from one part of the system to another. There is no single pressure that refers to the system as a whole. Similarly, if the nonequilibrium is because of the temperature of the system being different from that of its surroundings, there is a nonuniform temperature distribution set up within the system and there is no single temperature that stands for the system as a whole. It can thus be inferred that when the conditions for thermodynamic equilibrium are not satisfied, the states passed through by a system cannot be described by thermodynamic properties which represent the system as a whole.

Thermodynamic properties are the macroscopic coordinates defined for, and significant to, only thermodynamic equilibrium states. Both classical and statistical thermodynamics study mainly the equilibrium states of a system.

1.6 Quasi-Static Process

Let us consider a system of gas contained in a cylinder (Fig. 1.7). The system initially is in equilibrium state, represented by the properties p_1, v_1, t_1. The weight on the piston just balances the upward force exerted by the gas. If the weight is removed, there will be an unbalanced force between the system and the surroundings, and under gas pressure, the piston will move up till it hits the stops.

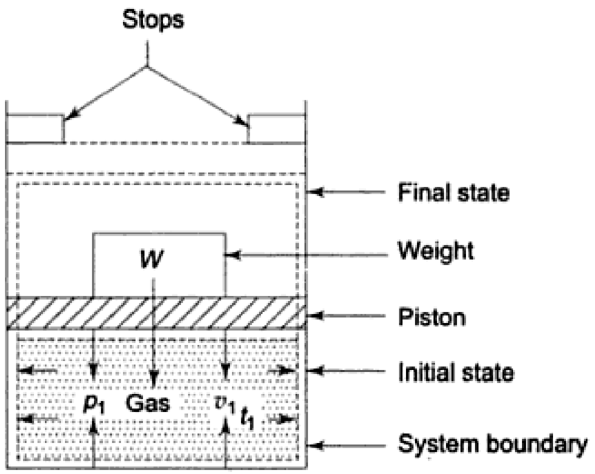


Figure 1.14 shows points 1 and 2 as the initial and final equilibrium states joined by a dotted line, which has got no meaning otherwise. Now if the single weight on the piston is made up of many very small pieces of weights (Fig. 1.15), and these weights are removed one by one very slowly from the top of the piston, at any instant of the upward travel of the piston, if the gas system is isolated, *the departure of the state of the system from the thermodynamic equilibrium state will be infinitesimally small*. So every state passed through by the system will be an equilibrium state. Such a process, which is but a locus of all the equilibrium points passed through by the system, is known as a *quasi-static process* (Fig. 1.16), 'quasi' meaning 'almost'. *Infinite slowness is the characteristic feature of a quasi-static process.* A quasi-static process 1–2 (Fig. 1.16) is thus a succession of equilibrium states.

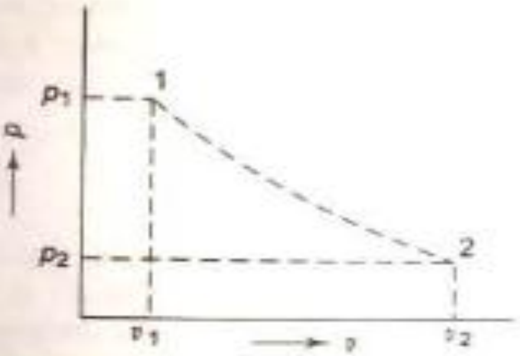


Fig. 1.14 *Plot representing the transition between two equilibrium states*

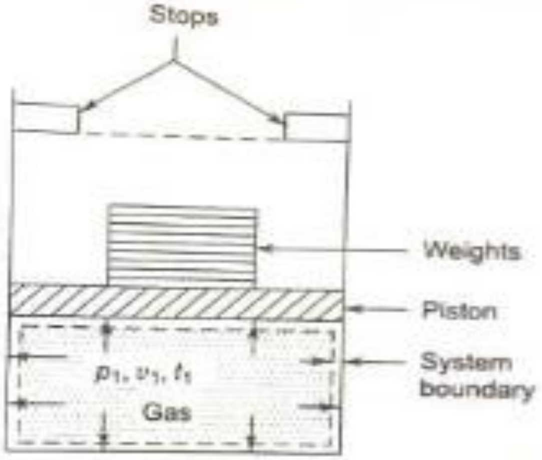


Fig. 1.15 *Infinitely slow transition of a system by infinitesimal force*

If the same slices or very small pieces of weights are now placed slowly one by one on the top of the piston, the piston will move down slowly from state 2 to state 1 by following the same quasi-static path of the succession of equilibrium states (Figs. 1.15 and 1.16). So a quasi-static process is also a *reversible process*. A reversible process is performed in both ways from 1 to 2 and from 2 to 1 involving *infinite time* in executing each of the processes.

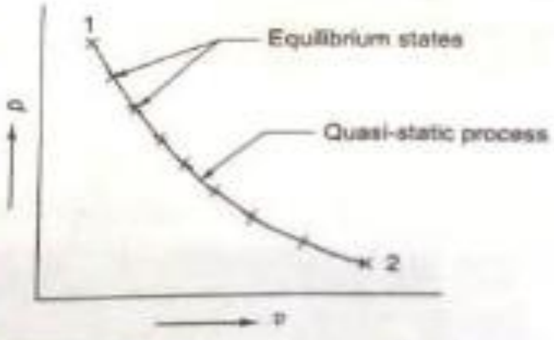
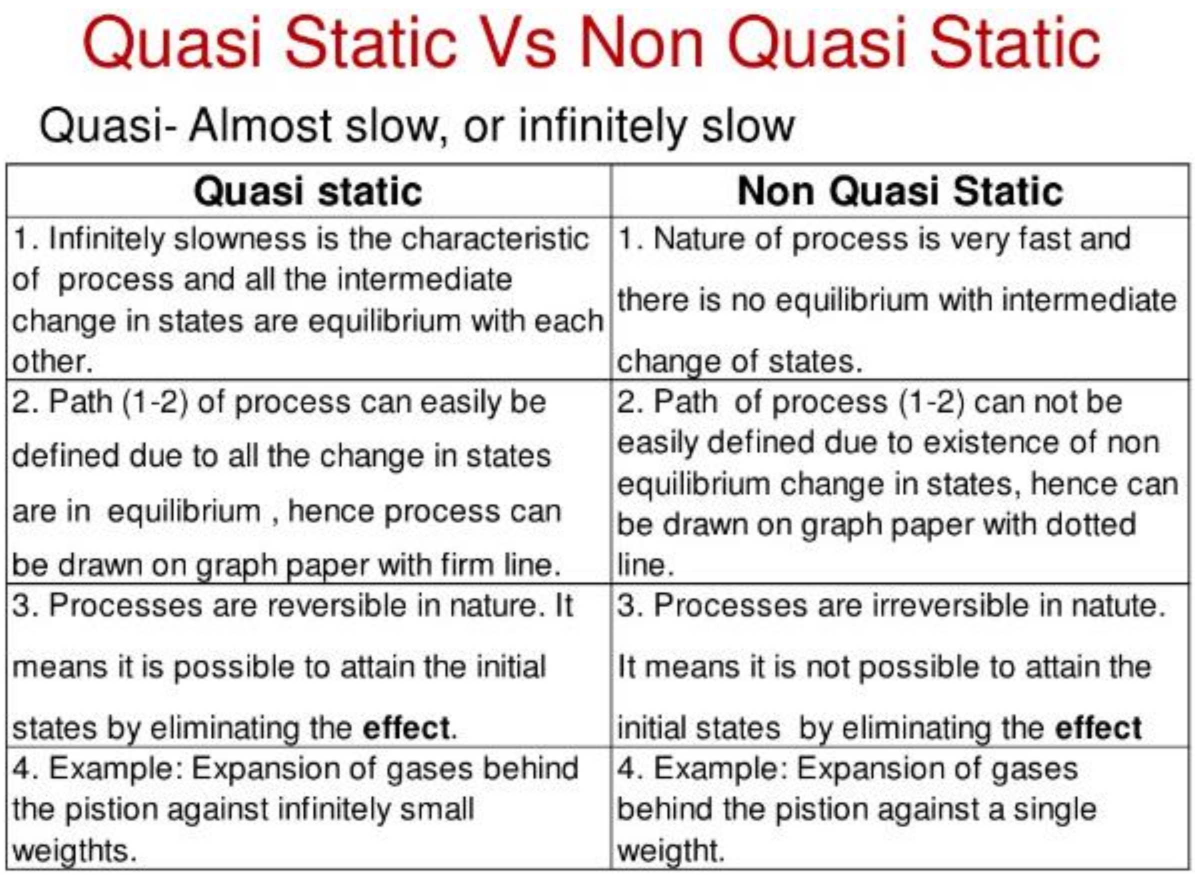


Fig. 1.16 *A quasi-static process*



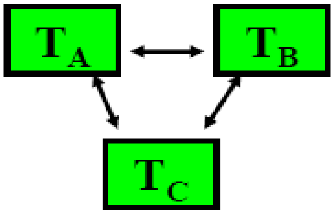
Quasi Static Vs Non Quasi Static

Quasi- Almost slow, or infinitely slow

Quasi static	Non Quasi Static
1. Infinitely slowness is the characteristic of process and all the intermediate change in states are equilibrium with each other.	1. Nature of process is very fast and there is no equilibrium with intermediate change of states.
2. Path (1-2) of process can easily be defined due to all the change in states are in equilibrium , hence process can be drawn on graph paper with firm line.	2. Path of process (1-2) can not be easily defined due to existence of non equilibrium change in states, hence can be drawn on graph paper with dotted line.
3. Processes are reversible in nature. It means it is possible to attain the initial states by eliminating the **effect**.	3. Processes are irreversible in natute. It means it is not possible to attain the initial states by eliminating the **effect**
4. Example: Expansion of gases behind the pistion against infinitely small weigths.	4. Example: Expansion of gases behind the pistion against a single weigtht.

Zeroth Law of Thermodynamics

- If two systems (say A and B) are in thermal equilibrium with a third system (say C) separately (that is A and C are in thermal equilibrium; B and C are in thermal equilibrium) then they are in thermal equilibrium themselves (that is A and B will be in thermal equilibrium

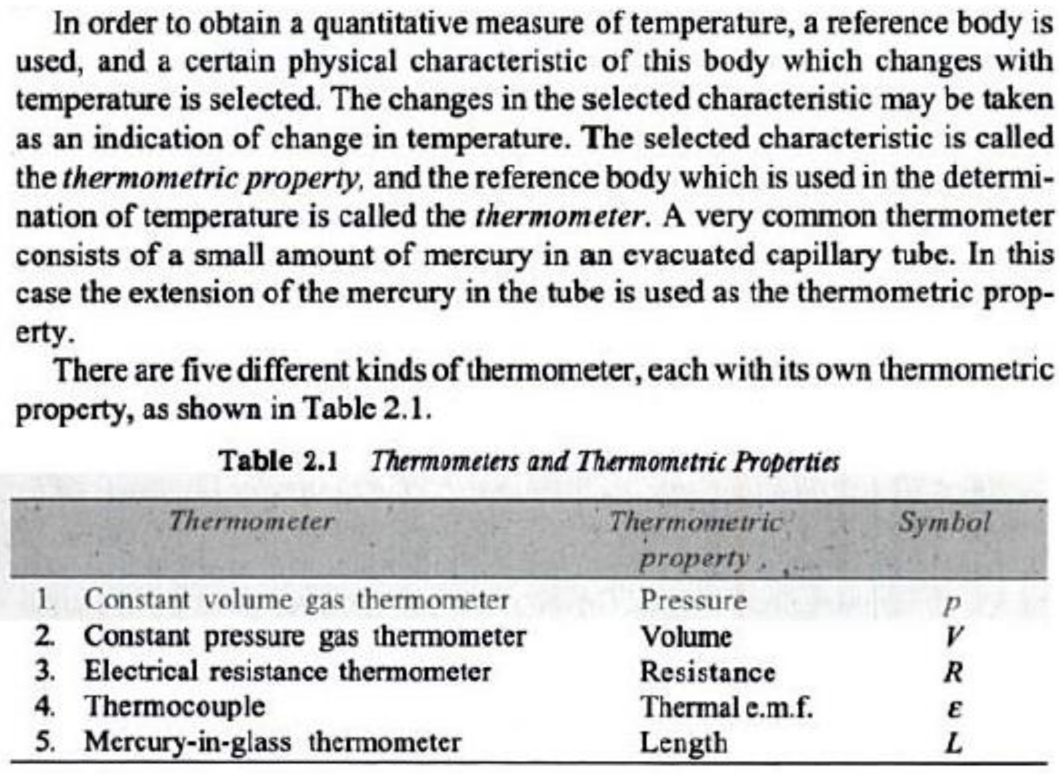


- Let us say T_A,T_B and T_C are the temperatures of A,B and C respectively.
- A and c are in thermal equilibrium. $T_a = t_c$
- B and C are in thermal equilibrium. $T_b = t_c$

Consequence of of ‘0’th law

- A and B will also be in thermal equilibrium $T_A = T_B$
- Looks very logical
- All temperature measurements are based on this LAW.

Measurement of Temperature (Instruments)



In order to obtain a quantitative measure of temperature, a reference body is used, and a certain physical characteristic of this body which changes with temperature is selected. The changes in the selected characteristic may be taken as an indication of change in temperature. The selected characteristic is called the *thermometric property*, and the reference body which is used in the determination of temperature is called the *thermometer*. A very common thermometer consists of a small amount of mercury in an evacuated capillary tube. In this case the extension of the mercury in the tube is used as the thermometric property.

There are five different kinds of thermometer, each with its own thermometric property, as shown in Table 2.1.

Table 2.1 *Thermometers and Thermometric Properties*

Thermometer	*Thermometric property*	*Symbol*
1. Constant volume gas thermometer	Pressure	p
2. Constant pressure gas thermometer	Volume	V
3. Electrical resistance thermometer	Resistance	R
4. Thermocouple	Thermal e.m.f.	ε
5. Mercury-in-glass thermometer	Length	L

2.5 Gas Thermometers

A schematic diagram of a constant volume gas thermometer is given in Fig. 2.1. A small amount of gas is enclosed in bulb B which is in communication via the capillary tube C with one limb of the mercury manometer M. The other limb of the mercury manometer is open to the atmosphere and can be moved vertically to adjust the mercury levels so that the mercury just touches lip L of the capillary. The pressure in the bulb is used as a thermometric property and is given by

$$p = p_0 + \rho_M Zg$$

where ρ_0 is the atmospheric pressure and ρ_M is the density of mercury.

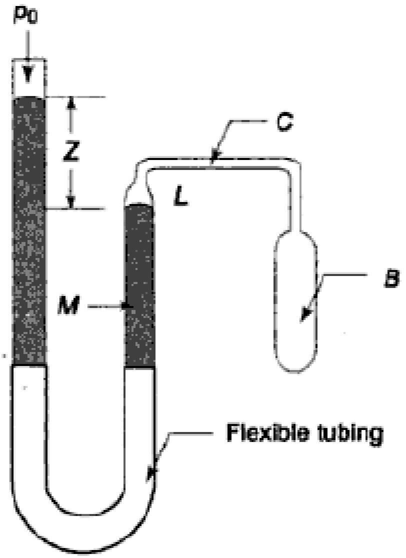


Fig. 2.1 *Constant volume gas thermometer*

When the bulb is brought in contact with the system whose temperature is to be measured, the bulb, in course of time, comes in thermal equilibrium with the system. The gas in the bulb expands, on being heated, pushing the mercury downward. The flexible limb of the manometer is then adjusted so that the mercury again touches the lip L. The difference in mercury level Z is recorded and the pressure p of the gas in the bulb is estimated. Since the volume of the trapped gas is constant, from the ideal gas equation,

$$\Delta T = \frac{V}{R} \Delta p \tag{2.10}$$

i.e. the temperature increase is proportional to the pressure increase.

In a constant pressure gas thermometer, the mercury levels have to be adjusted to keep Z constant, and the volume of gas V, which would vary with the temperature of the system, becomes the thermometric property.

$$\Delta T = \frac{p}{R} \Delta V \tag{2.11}$$

i.e. the temperature increase is proportional to the observed volume increase. The constant volume gas thermometer is, however, mostly in use, since it is simpler in construction and easier to operate.

2.8 Electrical Resistance Thermometer

In the resistance thermometer (Fig. 2.3) the change in resistance of a metal wire due to its change in temperature is the thermometric property. The wire, frequently platinum, may be incorporated in a Wheatstone bridge circuit. The platinum resistance thermometer measures temperature to a high degree of accuracy and sensitivity, which makes it suitable as a standard for the calibration of other thermometers.

In a restricted range, the following quadratic equation is often used

$$R = R_0(1 + At + Bt^2)$$

where R_0 is the resistance of the platinum wire when it is surrounded by melting ice and A and B are constants.

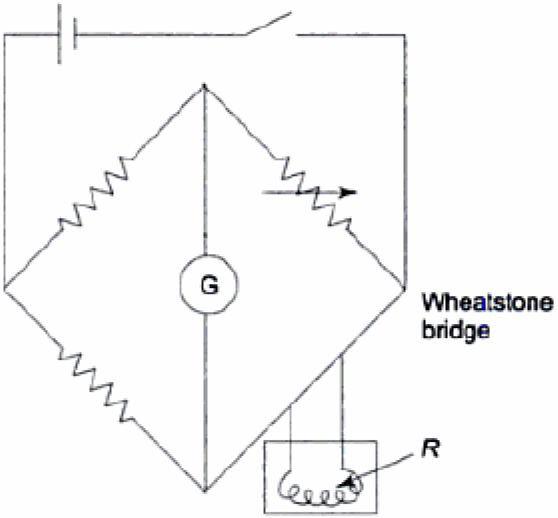


Fig. 2.3 *Resistance thermometer*

2.9 Thermocouple

A thermocouple circuit made up from joining two wires *A* and *B* made of dissimilar metals is shown in Fig. 2.4. Due to the Seeback effect, a net e.m.f. is generated in the circuit which depends on the difference in temperature between the hot and cold junctions and is, therefore, a thermometric property of the circuit. This e.m.f. can be measured by a microvoltmeter to a high degree of accuracy. The choice of metals depends largely on the temperature range to be investigated, and copper-constantan, chromel-alumel and platinum-platinum-rhodium are typical combinations in use.

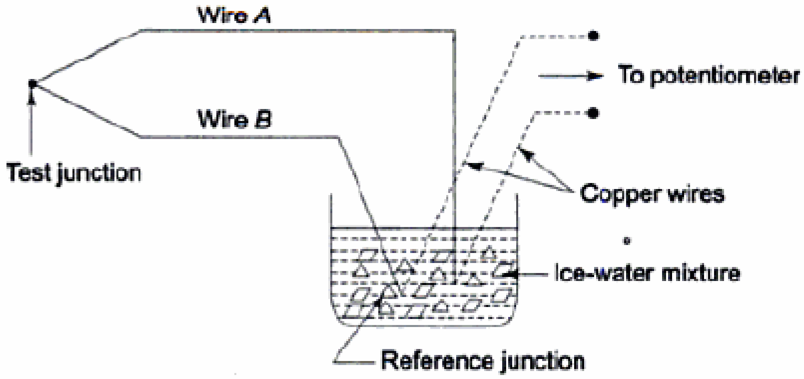


Fig. 2.4 *Thermocouple*

A thermocouple is calibrated by measuring the thermal e.m.f. at various known temperatures, the reference junction being kept at 0°C. The results of such measurements on most thermocouples can usually be represented by a cubic equation of the form

$$\varepsilon = a + bt + ct^2 + dt^3$$

where ε is the thermal e.m.f. and the constants a, b, c and d are different for each thermocouple.

The advantage of a thermocouple is that it comes to thermal equilibrium with the system, whose temperature is to be measured, quite rapidly, because its mass is small.

2.7 Celsius Temperature Scale

The Celsius temperature scale employs a degree of the same magnitude as that of the ideal gas scale, but its zero point is shifted, so that the Celsius temperature of the triple point of water is 0.01 degree Celsius or 0.01°C. If t denotes the Celsius temperature, then

$$t = T - 273.15^\circ$$

Thus the Celsius temperature t_s at which steam condenses at 1 atm. pressure

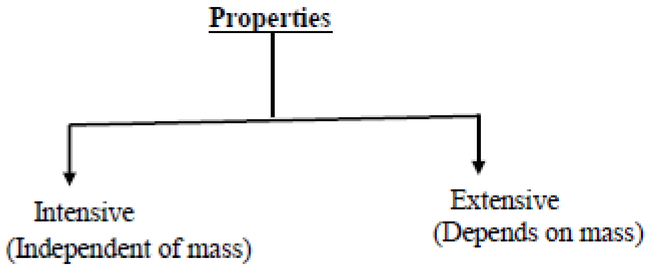
$$\begin{aligned} t_s &= T_s - 273.15^\circ \\ &= 373.15 - 273.15 \\ &= 100.00^\circ\text{C} \end{aligned}$$

Similar measurements for ice points show this temperature on the Celsius scale to be 0.00°C. The only Celsius temperature which is fixed by definition is that of the triple point.

SIMPLIFIED NOTES FOR REFERENCE

What do you mean by Property?

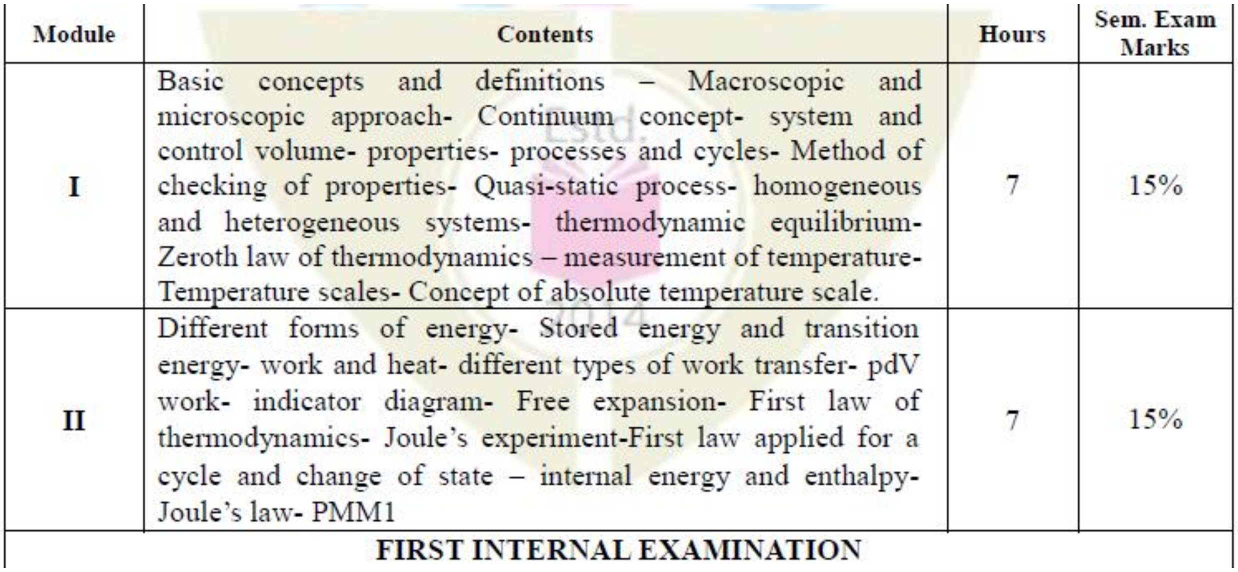
Any observable characteristics required to describe the conditions or state of a system is known as Thermodynamic property of a system.



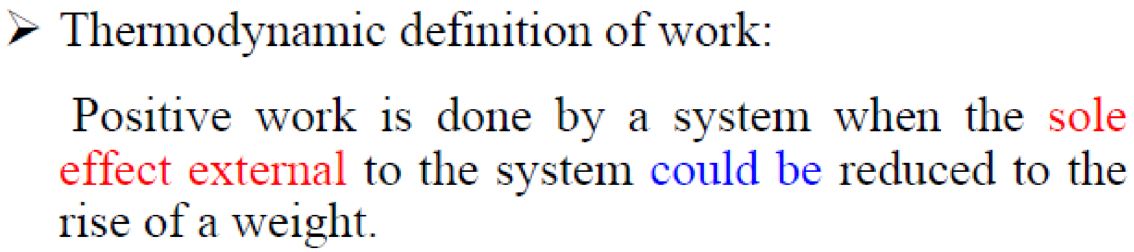
Differentiate Intensive and Extensive Property?

Extensive Property	**Intensive Property**
1. Extensive properties are dependent on the mass of a system.	1. Intensive properties are independent of the mass of a system.
2.Extensive properties are additive.	2. Intensive properties are not additive.

MODULE 2
WORK,HEAT & 1st LAW OF THERMODYNAMICS

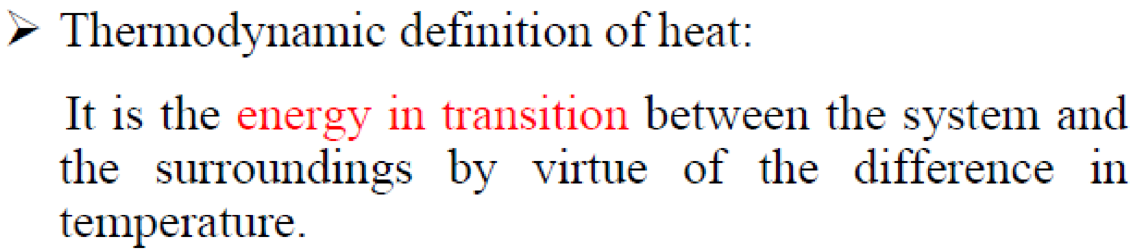


Module	Contents	Hours	Sem. Exam Marks
I	Basic concepts and definitions – Macroscopic and microscopic approach- Continuum concept- system and control volume- properties- processes and cycles- Method of checking of properties- Quasi-static process- homogeneous and heterogeneous systems- thermodynamic equilibrium- Zeroth law of thermodynamics – measurement of temperature- Temperature scales- Concept of absolute temperature scale.	7	15%
II	Different forms of energy- Stored energy and transition energy- work and heat- different types of work transfer- pdV work- indicator diagram- Free expansion- First law of thermodynamics- Joule's experiment-First law applied for a cycle and change of state – internal energy and enthalpy- Joule's law- PMM1	7	15%
FIRST INTERNAL EXAMINATION			



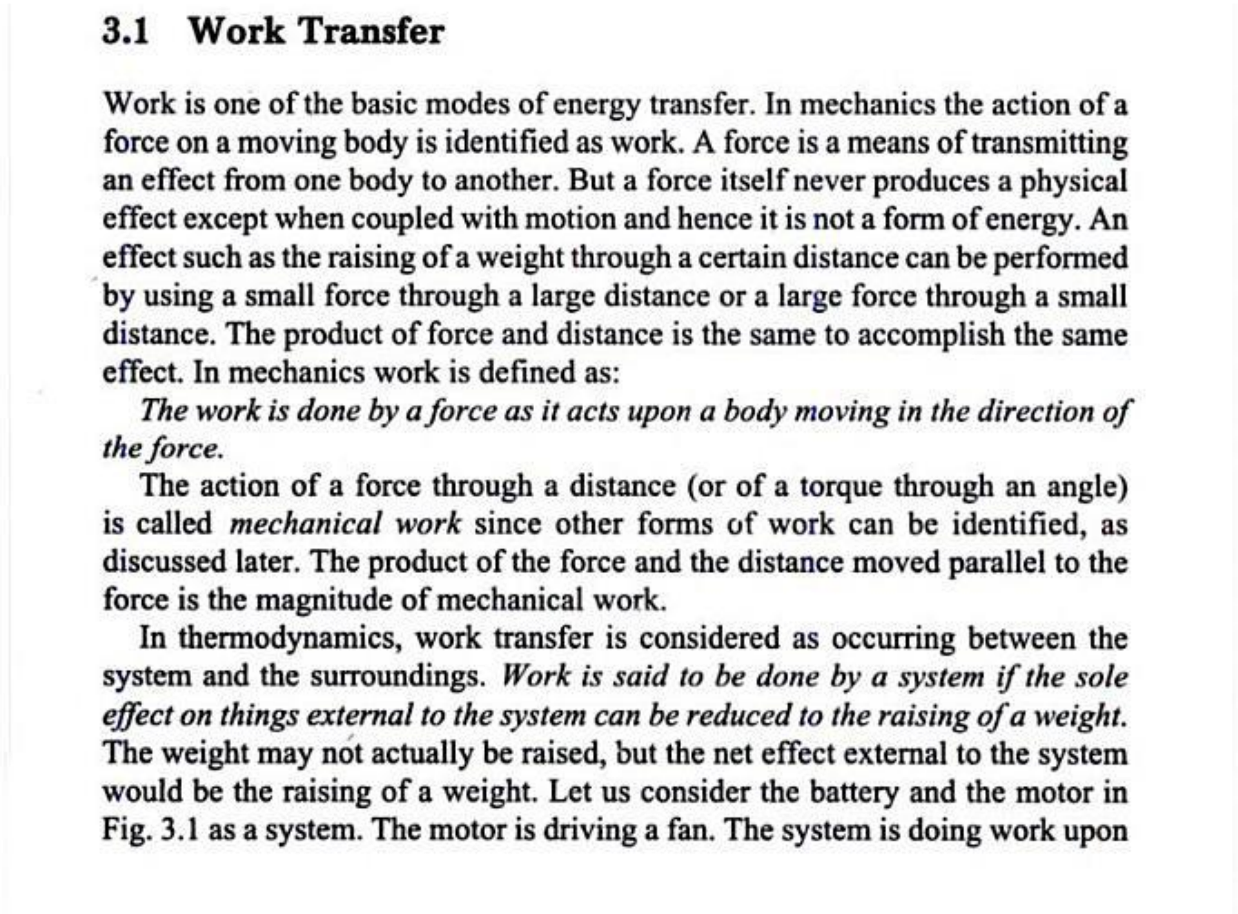
- Thermodynamic definition of work:

Positive work is done by a system when the sole effect external to the system could be reduced to the rise of a weight.



- Thermodynamic definition of heat:

It is the energy in transition between the system and the surroundings by virtue of the difference in temperature.



3.1 Work Transfer

Work is one of the basic modes of energy transfer. In mechanics the action of a force on a moving body is identified as work. A force is a means of transmitting an effect from one body to another. But a force itself never produces a physical effect except when coupled with motion and hence it is not a form of energy. An effect such as the raising of a weight through a certain distance can be performed by using a small force through a large distance or a large force through a small distance. The product of force and distance is the same to accomplish the same effect. In mechanics work is defined as:

The work is done by a force as it acts upon a body moving in the direction of the force.

The action of a force through a distance (or of a torque through an angle) is called *mechanical work* since other forms of work can be identified, as discussed later. The product of the force and the distance moved parallel to the force is the magnitude of mechanical work.

In thermodynamics, work transfer is considered as occurring between the system and the surroundings. *Work is said to be done by a system if the sole effect on things external to the system can be reduced to the raising of a weight.* The weight may not actually be raised, but the net effect external to the system would be the raising of a weight. Let us consider the battery and the motor in Fig. 3.1 as a system. The motor is driving a fan. The system is doing work upon

the surroundings. When the fan is replaced by a pulley and a weight, as shown in Fig. 3.2, the weight may be raised with the pulley driven by the motor. The sole effect on things external to the system is then the raising of a weight.

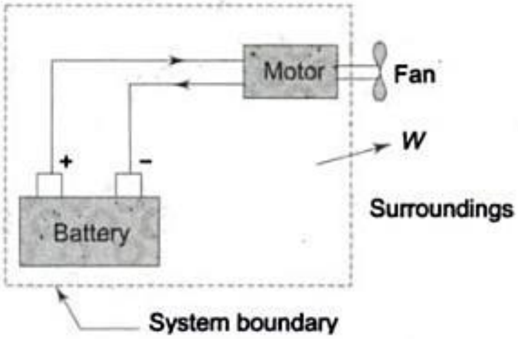


Fig. 3.1 ***Battery-motor system driving a fan***

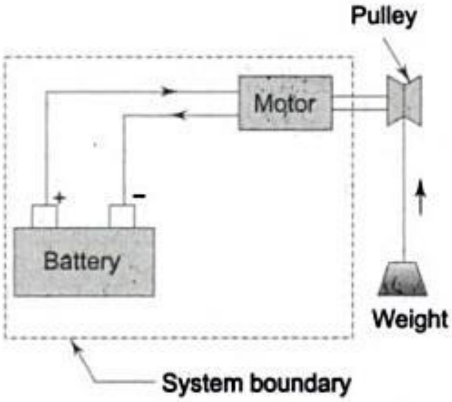
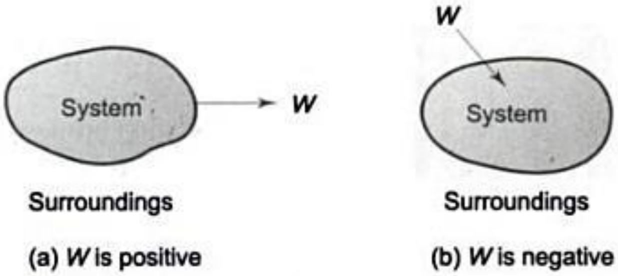


Fig. 3.2 ***Work transfer from a system***

When work is done by a system, it is arbitrarily taken to be positive, and when work is done on a system, it is taken to be negative (Fig. 3.3.). **The symbol *W* is used for work transfer.**



(a) *W* is positive

(b) *W* is negative

Fig. 3.3 ***Work interaction between a system and the surroundings***

The unit of work is N.m or Joule [1 Nm = 1 Joule]. The rate at which work is done by, or upon, the system is known as *power*. The unit of power is J/s or watt.

Work is one of the forms in which a system and its surroundings can interact with each other. There are various types of work transfer which can get involved between them.

3.2 $p\mathrm{d}V$–Work or Displacement Work

Let the gas in the cylinder (Fig. 3.4) be a system having initially the pressure p_1 and volume V_1. The system is in thermodynamic equilibrium, the state of which is described by the coordinates p_1, V_1. The piston is the only boundary which moves due to gas pressure. Let the piston move out to a new final position 2, which is also a thermodynamic equilibrium state specified by pressure p_2 and volume V_2. At any intermediate point in the travel of the piston, let the pressure be p and the volume V. This must also be an equilibrium state, since macroscopic properties p and V are significant only for equilibrium states. When the piston moves an infinitestimal distance dl, and if 'a' be the area of the piston, the force F acting on the piston $F = p.a.$ and the infinitesimal amount of work done by the gas on the piston

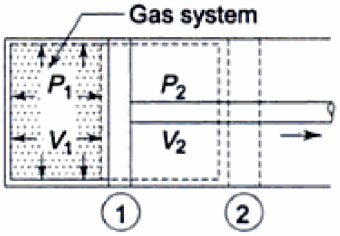


Fig. 3.4 *pdV work*

$$\text{đ}W = F \cdot \mathrm{d}l = pa\mathrm{d}l = p\mathrm{d}V \tag{3.1}$$

where $\mathrm{d}V = a\mathrm{d}l$ = infinitesimal displacement volume. The differential sign in dW with the line drawn at the top of it will be explained later.

When the piston moves out from position 1 to position 2 with the volume changing from V_1 to V_2, the amount of work W done by the system will be

$$W_{1-2} = \int_{V_1}^{V_2} p\mathrm{d}V$$

The magnitude of the work done is given by the area under the path 1–2, as shown in Fig. 3.5. Since p is at all times a thermodynamic coordinate, all the states passed through by the system as the volume changes from V_1 to V_2 must be equilibrium states, and the path 1–2 must be *quasi-static*. The piston moves infinitely slowly so that every state passed through is an equilibrium state. *The integration* $\int p\mathrm{d}V$ *can be performed only on a quasi-static path.*

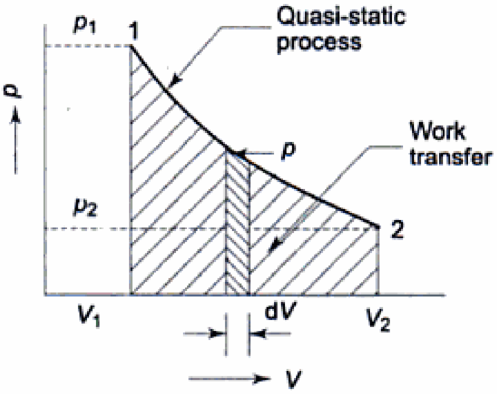
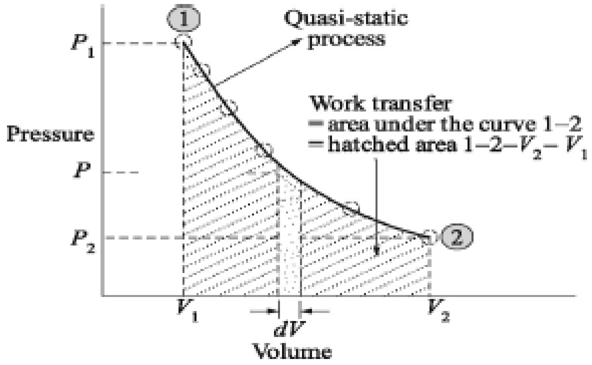


Fig. 3.5 *Quasi-static pdV work*

SIMPLIFIED NOTES ON pdV WORK IN VARIOUS QUASISTATIC PROCESSES



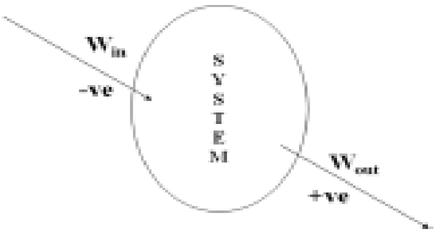
Expression for pdV Work:

$$\begin{aligned} W &= \int dW \\ &= \int (p * \text{piston area} * displacement) \\ &= \int p * A * dL \\ &= \int pdV \end{aligned}$$

$$W_{12} = \int_1^2 pdV$$

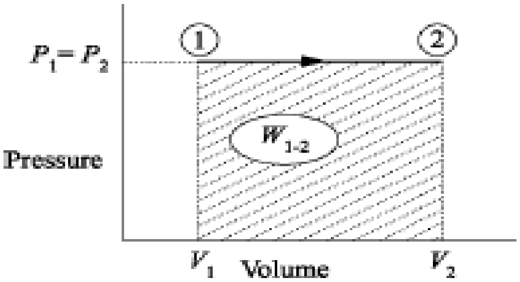
Sign Convention for work transfer:

Work done by the system is **positive and** Work done on the system is **negative.**



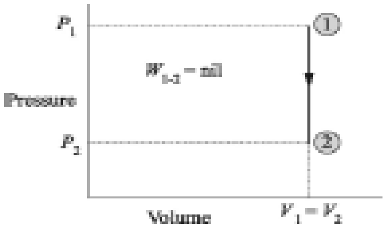
Work done during various Quasi-static Processes:

Constant Pressure Proces:



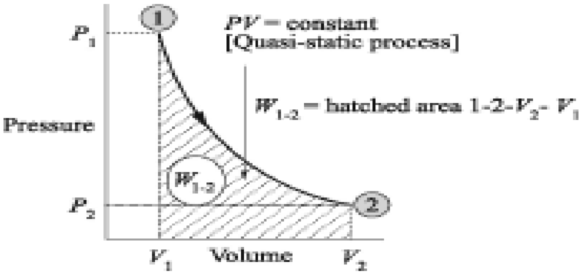
$$W_{1-2} = \int_{V_1}^{V_2} pdV = p(V_2 - V_1)$$

Constant Volume Process:



$$W_{1-2} = \int_{V_1}^{V_2} pdV = 0$$

Constant Temperature Process:



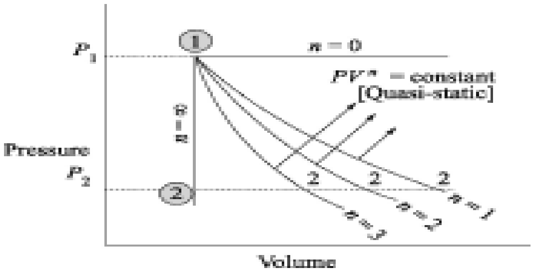
$$W_{1-2} = \int_{V_1}^{V_2} pdV$$

$$pV = p_1V_1 = C$$

$$p = \frac{p_1V_1}{V}$$

$$W_{1-2} = p_1V_1 \int_{V_1}^{V_2} \frac{dV}{V} = p_1V_1 \ln \frac{V_2}{V_1}$$

Polytropic Process $pV^n = C$, **where n is a constant**



$$pV^n = p_1V_1^n = p_2V_2^n = C$$

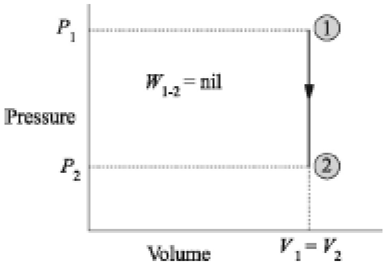
$$p = \frac{(p_1V_1^n)}{V^n}$$

$$W_{1-2} = \int_{V_1}^{V_2} pdV$$

$$= \int_{V_1}^{V_2} \frac{p_1V_1^n}{V^n} dV$$

$$= \frac{p_1V_1 - p_2V_2}{n-1}$$

Constant Volume Process:



3.2.1 *Path Function and Point Function*

With reference to Fig. 3.6, it is possible to take a system from state 1 to state 2 along many quasi-static paths, such as A, B or C. Since the area under each curve represents the work for each process, the amount of work involved in each case is not a function of the end states of the process, and it depends on the path the system follows in going from state 1 to state 2. For this reason, work is called a *path function, and* $đW$ *is an inexact or imperfect differential.*

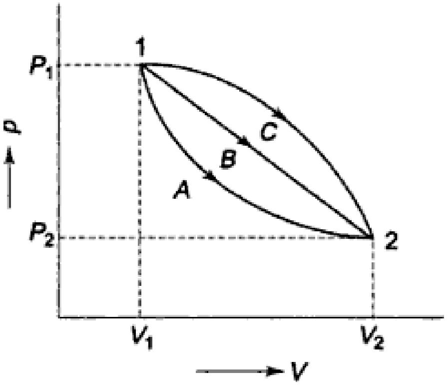


Fig. 3.6 *Work—a path function*

Thermodynamic properties are *point functions,* since for a given state, there is a definite value for each property. The change in a thermodynamic property of a system in a change of state is independent of the path the system follows during the change of state, and depends only on the initial and final states of the system. The differentials of point functions are *exact or perfect differentials,* and the integration is simply

$$\int_{V_1}^{V_2} dV = V_2 - V_1$$

The change in volume thus depends only on the end states of the system irrespective of the path the system follows.

On the other hand, work done in a quasi-static process between two given states depends on the path followed.

$$\int_1^2 đW \neq W_2 - W_1$$

Rather,

$$\int_1^2 đW = W_{1-2} \text{ or } {}_1W_2$$

To distinguish an inexact differential $đW$ from an exact differential dV or dp the differential sign is being cut by a line at its top.

From Eq. (3.1),

$$dV = \frac{1}{p} đW \tag{3.2}$$

Here, $1/p$ is called the *integrating factor*. Therefore, an inexact differential $đW$ when multiplied by an integrating factor $1/p$ becomes an exact differential dV.

For a cyclic process, the initial and final states of the system are the same, and hence, the change in any property is zero, i.e.

$$\oint dV = 0, \oint dp = 0, \oint dT = 0 \tag{3.3}$$

where the symbol $\oint$ denotes the cyclic integral for the closed path. Therefore, *the cyclic integral of a property is always zero.*

3.2.2 *pdV-Work in Various Quasi-Static Processes*

(a) Constant pressure process (Fig. 3.7) (isobaric or isopiestic process)

$$W_{1-2} = \int_{V_1}^{V_2} p dV = p(V_2 - V_1) \tag{3.4}$$

(b) Constant volume process (Fig. 3.8) (isochoric process)

$$W_{1-2} = \int p dV = 0 \tag{3.5}$$

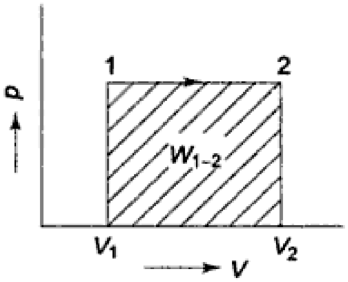


Fig. 3.7 *Constant pressure process*

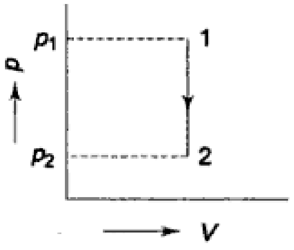


Fig. 3.8 *Constant volume process*

(c) Process in which $pV = C$ (Fig. 3.9)

$$\therefore \quad W_{1-2} = \int_{V_1}^{V_2} p dV, \quad pV = p_1 V_1 = C$$

$$p = \frac{(p_1 V_1)}{V}$$

$$W_{1-2} = p_1 V_1 \int_{V_1}^{V_2} \frac{dV}{V} = p_1 V_1 \ln \frac{V_2}{V_1}$$

$$= p_1 V_1 \ln \frac{p_1}{p_2} \tag{3.6}$$

(d) Process in which $pV^n = C$, where n is a constant (Fig. 3.10).

$$pV^n = p_1 V_1^n = p_2 V_2^n = C$$

$$\therefore \quad p = \frac{(p_1 V_1^n)}{V^n}$$

$$\therefore \quad W_{1-2} = \int_{V_1}^{V_2} p\,dV$$

$$= \int_{V_1}^{V_2} \frac{p_1 V_1^n}{V^n} \cdot dV$$

$$= (p_1 V_1^n) \left[\frac{V^{-n+1}}{-n+1} \right]_{V_1}^{V_2}$$

$$= \frac{p_1 V_1^n}{1-n} (V_2^{1-n} - V_1^{1-n})$$

$$= \frac{p_2 V_2^n \times V_2^{1-n} - p_1 V_1^n \times V_1^{1-n}}{1-n}$$

$$= \frac{p_1 V_1 - p_2 V_2}{n-1} = \frac{p_1 V_1}{n-1} \left[1 - \left(\frac{p_2}{p_1} \right)^{n-1/n} \right] \qquad (3.7)$$

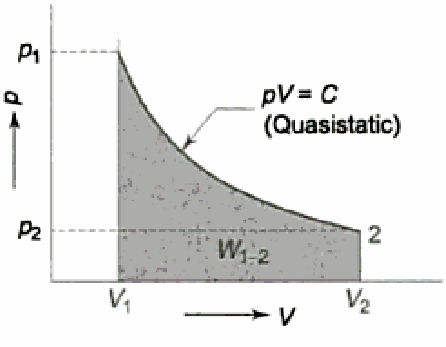


Fig. 3.9 *Process in which $pV = constant$*

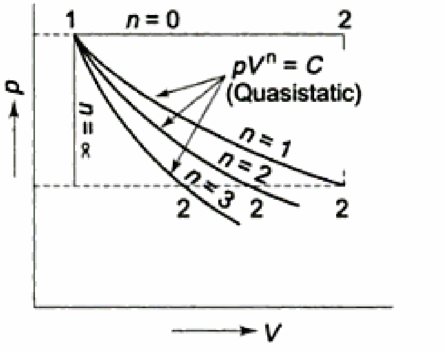
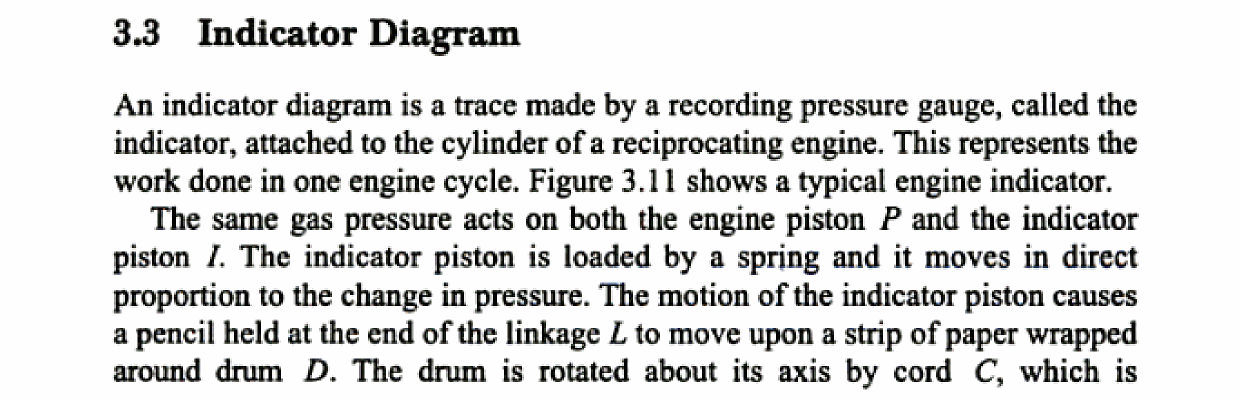


Fig. 3.10 *Process in which $pV^n = constant$*



3.3 Indicator Diagram

An indicator diagram is a trace made by a recording pressure gauge, called the indicator, attached to the cylinder of a reciprocating engine. This represents the work done in one engine cycle. Figure 3.11 shows a typical engine indicator.

The same gas pressure acts on both the engine piston *P* and the indicator piston *I*. The indicator piston is loaded by a spring and it moves in direct proportion to the change in pressure. The motion of the indicator piston causes a pencil held at the end of the linkage *L* to move upon a strip of paper wrapped around drum *D*. The drum is rotated about its axis by cord *C*, which is

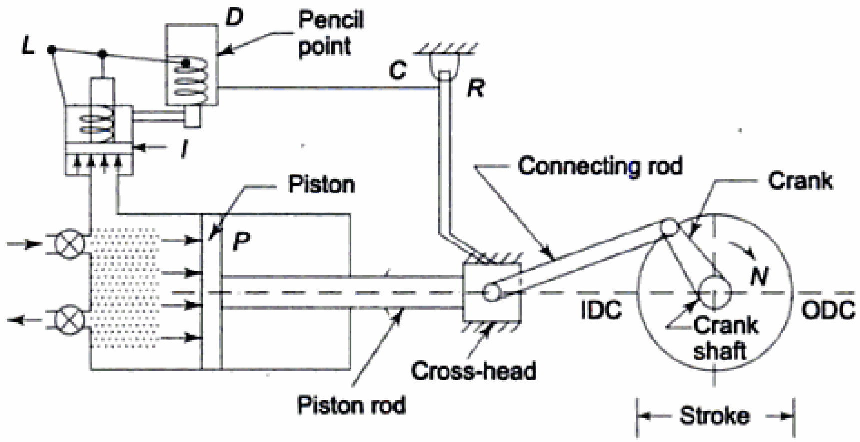


Fig. 3.11 *Engine indicator*

connected through a reducing motion R to the piston P of the engine. The surface of drum D moves horizontally under the pencil while the pencil moves vertically over the surface and a plot of pressure upon the piston vs. piston travel is obtained.

Before tracing the final indicator diagram, a pressure reference line is recorded by subjecting the indicator to the atmosphere and tracing a line at a constant pressure of one atmosphere.

The area of the indicator diagram represents the magnitude of the net work done by the system in one engine cycle. The area under the path 1–2 represents work done by the system and the area under the path 2–1 represents work done upon the system (Fig. 3.12). The area of the diagram, a_d, is measured by means of a planimeter, and the length of the diagram, l_d, is also measured. The *mean effective pressure* (m.e.p.) p_m is defined in the following way

$$p_m = \frac{a_d}{l_d} \times K$$

where K is the indicator spring constant (N/cm^2 × cm travel). Work done in one engine cycle

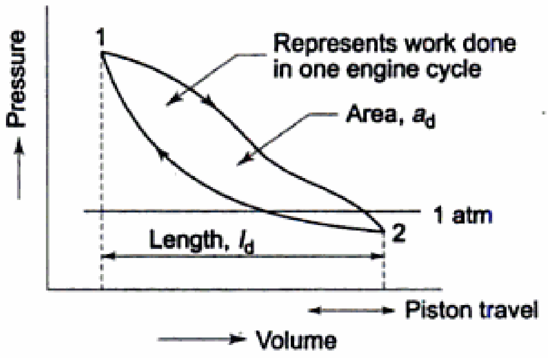


Fig. 3.12 *Indicator diagram*

$$= (p_m \cdot A)\, L$$

where A = cross-sectional area of the cylinder

$= \frac{\pi}{4} D^2$, where D is the cylinder diameter

and L = stroke of piston, or length of cylinder.

Let N be the revolutions per minute (r.p.m.) of the crankshaft. In a two stroke cycle, the engine cycle is completed in two strokes of the piston or in one revolution of the crankshaft. In a four-stroke cycle, the engine cycle is completed in four strokes of the piston or two revolutions of the crankshaft.

For a two-stroke engine, work done in one minute = $p_m ALN$, and for a four-stroke engine, work done in one minute = $p_m ALN/2$.

The power developed inside the cylinder of the engine is called *indicated power* (IP),

$$\therefore \qquad \text{IP} = \frac{p_m AL\left(N \text{ or } \frac{N}{2}\right)n}{60} \text{ kW} \tag{3.8}$$

where p_m is in kPa and n is the number of cylinders in the engine.

The power available at the crankshaft is always less than this value (IP) due to friction, etc. and is called the *brake power* (BP) or *shaft power* (SP). If ω is the angular velocity of the crankshaft in radian/sec, then

$$\text{BP} = T\,\omega \tag{3.9}$$

where T is the torque transmitted to the crankshaft in mN.

$$\therefore \qquad \text{BP} = \frac{2\pi TN}{60} \tag{3.10}$$

where N is the number of revolutions per minute (rpm).

The mechanical efficiency of the engine, η_{mech}, is defined as

$$\eta_{\text{mech}} = \frac{\text{BP}}{\text{IP}} \tag{3.11}$$

An engine is said to be *double-acting*, if the working fluid is made to work on both sides of the piston. Such an engine theoretically develops twice the amount of work developed in a single-acting engine. Most reciprocating steam engines are double-acting, and so are many marine diesel engines. Internal combustion engines for road transport are always single-acting.

3.4 Other Types of Work Transfer

There are forms of work other than $p\mathrm{d}V$ or displacement work. The following are the additional types of work transfer which may get involved in system-surroundings interactions.

(a) Electrical Work When a current flows through a resistor (Fig. 3.13), taken as a system, there is work transfer into the system. This is because the current can drive a motor, the motor can drive a pulley and the pulley can raise a weight.

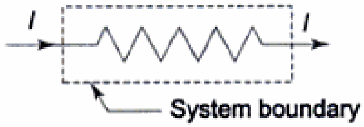


Fig. 3.13 *Electrical work*

The current flow, I, in amperes, is given by

$$I = \frac{\mathrm{d}C}{\mathrm{d}\tau}$$

where C is the charge in coulombs and τ is time in seconds. Thus $\mathrm{d}C$ is the charge crossing a boundary during time $\mathrm{d}\tau$. If E is the voltage potential, the work is

$$\begin{aligned} \text{đ}W &= E \cdot \mathrm{d}C \\ &= EI\,\mathrm{d}\tau \end{aligned}$$

$$\therefore \qquad W = \int_1^2 EI\mathrm{d}\tau \tag{3.12}$$

The electrical power will be

$$\dot{W} = \lim_{\mathrm{d}\tau \to 0} \frac{\mathrm{d}W}{\mathrm{d}\tau} = EI \tag{3.13}$$

This is the rate at which work is transferred.

(b) Shaft Work When a shaft, taken as the system (Fig. 3.14), is rotated by a motor, there is work transfer into the system. This is because the shaft can rotate a pulley which can raise a weight. If T is the torque applied to the shaft and $\mathrm{d}\theta$ is the angular displacement of the shaft, the shaft work is

$$W = \int_1^2 T\,\mathrm{d}\theta \tag{3.14}$$

and the shaft power is

$$\dot{W} = \int_1^2 T\frac{\mathrm{d}\theta}{\mathrm{d}\tau} = T\,\omega \tag{3.15}$$

where ω is the angular velocity and T is considered a constant in this case.

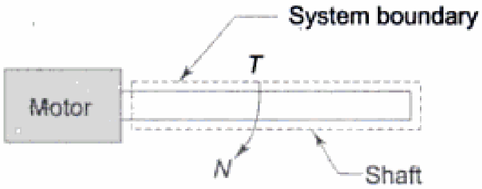


Fig. 3.14 *Shaft work*

(c) Paddle-Wheel Work or Stirring Work As the weight is lowered, and the paddle wheel turns (Fig. 3.15), there is work transfer into the fluid system which gets stirred. Since the volume of the system remains constant, $\int p\mathrm{d}V = 0$. If m is the mass of the weight lowered through a distance $\mathrm{d}z$ and T is the torque transmitted by the shaft in rotating through an angle $\mathrm{d}\theta$, the differential work transfer to the fluid is given by

$$\mathrm{d}W = mg\mathrm{d}z = T\mathrm{d}\theta$$

and the total work transfer is

$$W = \int_1^2 mg\mathrm{d}z = \int_1^2 W'\mathrm{d}z = \int_1^2 T\mathrm{d}\theta \qquad (3.15)$$

where W' is the weight lowered.

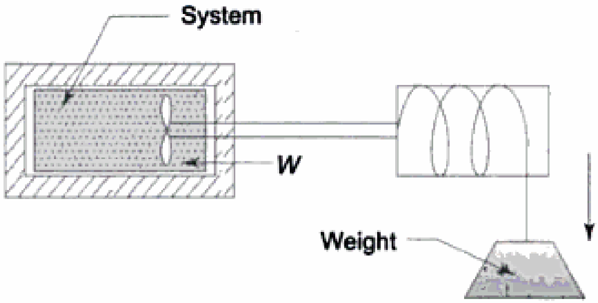


Fig. 3.15 *Paddle-wheel work*

(d) Flow Work The flow work, significant only in a flow process or an open system, represents the energy transferred across the system boundary as a result of the energy imparted to the fluid by a pump, blower or compressor to make the fluid flow across the control volume. Flow work is analogous to displacement work. Let p be the fluid pressure in the plane of the imaginary piston, which acts in a direction normal to it (Fig. 3.16). The work done on this imaginary piston by the external pressure as the piston moves forward is given by

$$đ W_{\text{flow}} = p\mathrm{d}V, \qquad (3.16)$$

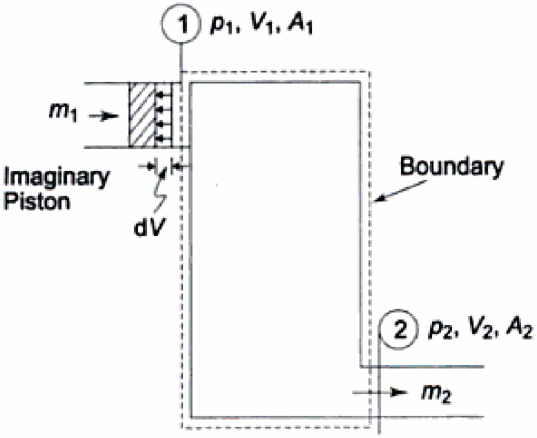


Fig. 3.16 *Flow work*

where dV is the volume of fluid element about to enter the system.

$$\therefore \qquad đW_{\text{flow}} = pv\,dm \tag{3.17}$$

where $dV = v\,dm$

Therefore, flow work at inlet (Fig. 3.16),

$$(dW_{\text{flow}})_{\text{in}} = p_1 v_1 dm_1 \tag{3.18}$$

Equation (3.18) can also be derived in a slightly different manner. If the normal pressure p_1 is exerted against the area A_1, giving a total force $(p_1\,A_1)$ against the piston, in time $d\tau$, this force moves a distance $\mathbf{V}_1 d\tau$, where $\mathbf{V}_1$ is the velocity of flow (piston). The work in time $d\tau$ is $p_1\,A_1\,\mathbf{V}_1 d\tau$, or the work per unit time is $p_1\,A_1\,\mathbf{V}_1$. Since the flow rate

$$w_1 = \frac{A_1 \mathbf{V}_1}{v_1} = \frac{dm_1}{d\tau}$$

the work done in time $d\tau$ becomes

$$(đW_{\text{flow}})_{\text{in}} = p_1\,v_1\,dm_1$$

Similarly, flow work of the fluid element leaving the system is

$$(đW_{\text{flow}})_{\text{out}} = p_2\,v_2\,dm_2 \tag{3.19}$$

The flow work per unit mass is thus

$$W_{\text{flow}} = pv \tag{3.20}$$

It is the displacement work done at the moving system boundary.

(e) Work Done in Stretching a Wire Let us consider a wire as the system. If the length of the wire in which there is a tension $\mathcal{F}$ is changed from L to $L + dL$, the infinitesimal amount of work that is done is equal to

$$đW = -\mathcal{F}\,dL$$

The minus sign is used because a positive value of dL means an expansion of the wire, for which work must be done on the wire, i.e., negative work. For a finite change of length,

$$W = -\int_1^2 \mathcal{F}\,dL \tag{3.21}$$

If we limit the problem to within the elastic limit, where E is the modulus of elasticity, s is the stress, ε is the strain, and A is the cross-sectional area, then

$$\mathcal{F} = sA = E\varepsilon A, \text{ since } \frac{s}{\varepsilon} = E$$

$$d\varepsilon = \frac{dL}{L}$$

$$đW = -\mathcal{F}\,dL = -E\varepsilon\,AL\,d\varepsilon$$

$$\therefore \qquad W = -A\varepsilon L \int_1^2 \varepsilon \, d\varepsilon = -\frac{AEL}{2}\left(\varepsilon_2^2 - \varepsilon_1^2\right) \tag{3.22}$$

(f) Work Done in Changing the Area of a Surface Film A film on the surface of a liquid has a surface tension, which is a property of the liquid and the surroundings. The surface tension acts to make the surface area of the liquid a minimum. It has the unit of force per unit length. The work done on a homogeneous liquid film in changing its surface area by an infinitesimal amount dA is

$$đW = -\sigma dA$$

where σ is the surface tension (N/m).

$$\therefore \qquad W = -\int_1^2 \sigma dA \tag{3.23}$$

(g) Magnetization of a Paramagnetic Solid The work done per unit volume on a magnetic material through which the magnetic and magnetization fields are uniform is

$$đW = -H dI$$

$$\text{and} \qquad W_{1-2} = -\int_{I_1}^{I_2} H dI \tag{3.24}$$

where H is the field strength, and I is the component of the magnetization field in the direction of the field. The minus sign provides that an increase in magnetization (positive dI) involves negative work.

The following equations summarize the different forms of work transfer:

Displacement work

$$\text{(compressible fluid)} \qquad W = \int_1^2 p dV$$

$$\text{Electrical work} \qquad W = \int_1^2 E dC = \int_1^2 EI \, d\tau$$

$$\text{Shaft work} \qquad W = \int_1^2 T d\theta$$

$$\text{Surface film} \qquad W = -\int_1^2 \sigma dA \tag{3.25}$$

$$\text{Stretched wire} \qquad W = -\int_1^2 \mathcal{F} dL$$

Magnetised solid $W = -\int_1^2 H\,dI$

It may be noted in the above expressions that the work is equal to the integral of the product of an intensive property and the change in its related extensive property. These expressions are valid only for infinitesimally slow quasi-static processes.

There are some other forms of work which can be identified in processes that are not quasi-static, for example, the work done by shearing forces in a process involving friction in a viscous fluid.

3.5 Free Expansion with Zero Work Transfer

Work transfer is identified only at the boundaries of a system. It is a boundary phenomenon, and a form of energy in transit crossing the boundary. Let us consider a gas separated from the vacuum by a partition (Fig. 3.17). Let the partition be removed. The gas rushes to fill the entire volume. The expansion of a gas against vacuum is called *free expansion*. If we neglect the work associated with the removal of partition, and consider the gas and vacuum together as our system (Fig. 3.17a), there is no work transfer involved here, since no work crosses the system boundary, and hence

$$\int_1^2 đW = 0, \text{ although } \int_1^2 p\,dV \neq 0$$

If only the gas is taken as the system (Fig. 3.17), when the partition is removed there is a change in the volume of the gas, and one is tempted to calculate the work from the expression $\int_1^2 p\,dV$. However, this is not a quasistatic process, although the initial and final end states are in equilibrium. Therefore, the work cannot be calculated from this relation. The two end states can be located on the p–V diagram and these are joined by a dotted line (Fig. 3.17c) to indicate that the process had occurred. However, if the vacuum space is divided into a large number of small volumes by partitions and the partitions are removed one by one slowly (Fig. 3.17d), then every state passed through by the system is an equilibrium state and the work done can then be estimated from the relation $\int_1^2 p\,dV$ (Fig. 3.17e), Yet, in free expansion of a gas, there is no resistance to the fluid at the system boundary as the volume of the gas increases to fill up the vacuum space. Work is done by a system to overcome some resistance. Since vacuum does not offer any resistance, there is no work transfer involved in free expansion.

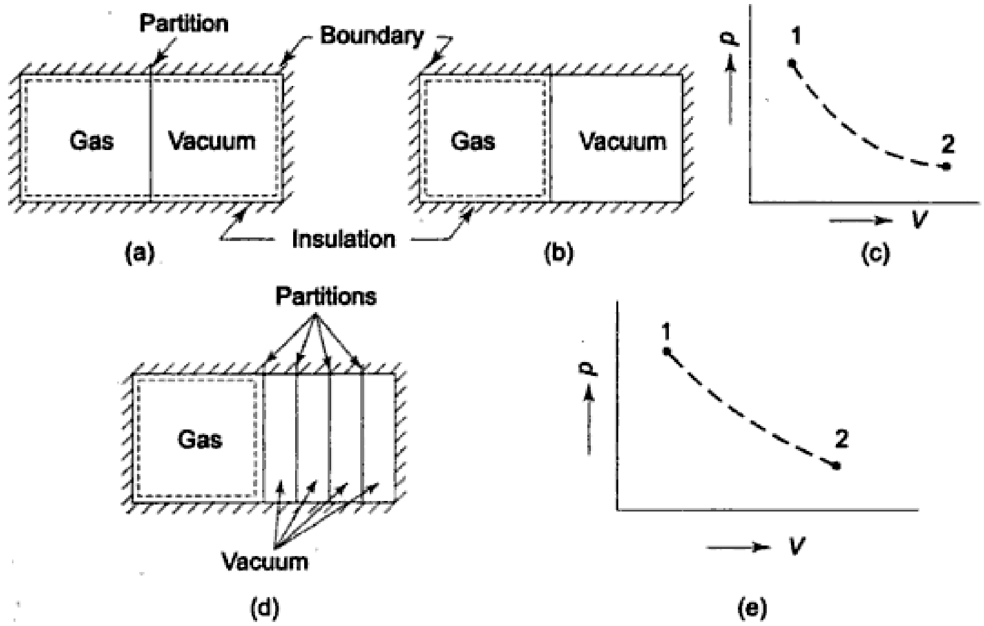


Fig. 3.17 *Free expansion*

3.6 Net Work Done by a System

Often different forms of a work transfer occur simultaneously during a process executed by a system. When all these work interactions have been evaluated, the total or net work done by the system would be equal to the algebraic sum of these as given below

$$W_{\text{total}} = W_{\text{displacement}} + W_{\text{shear}} + W_{\text{electrical}} + W_{\text{stirring}} + \ldots$$

3.7 Heat Transfer

Heat is defined as the form of energy that is transferred across a boundary *by virtue of a temperature difference*. The temperature difference is the 'potential' or 'force' and heat transfer is the 'flux'.

The transfer of heat between two bodies in direct contact is called *conduction*. Heat may be transferred between two bodies separated by empty space or gases by the mechanism of *radiation* through electromagnetic waves. A third method of heat transfer is *convection* which refers to the transfer of heat between a wall and a fluid system in motion.

The direction of heat transfer is taken from the high temperature system to the low temperature system. *Heat flow into a system is taken to be positive, and heat flow out of a system is taken as negative* (Fig. 3.18). The symbol Q is used for heat transfer, i.e., the quantity of heat transferred within a certain time.

Heat is a form of energy in transit (like work transfer). It is a boundary phenomenon, since it occurs only at the boundary of a system. Energy transfer by virtue of temperature difference only is called heat transfer. All other energy interactions may be termed as work transfer.

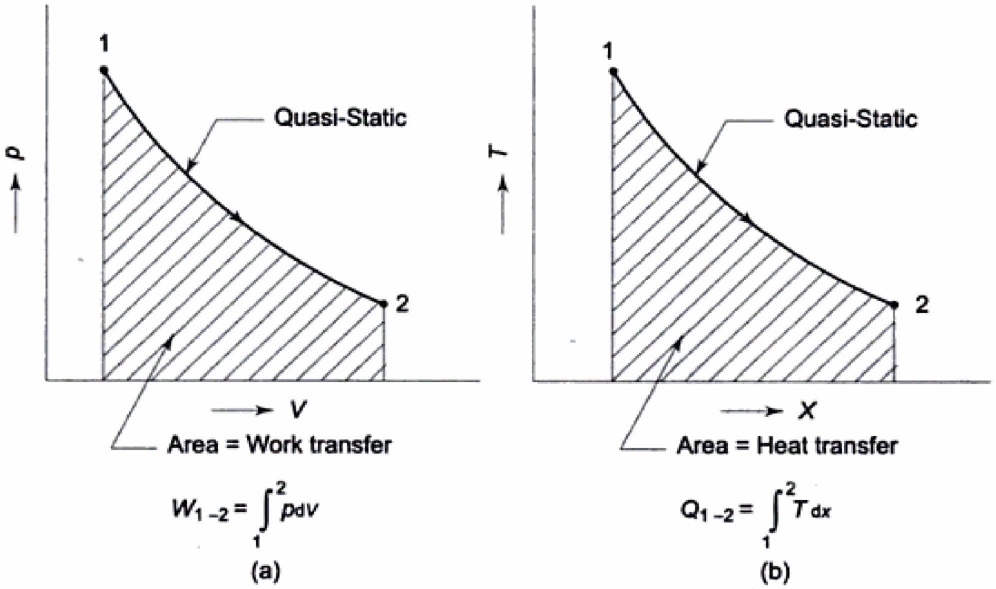


Fig. 3.19 *Representation of work transfer and heat transfer in quasi-static processes on p-v and T-x coordinates*

Just like displacement work, the heat transfer can also be written as the integral of the product of the intensive property T and the differential change of an extensive property, say X (Fig.3.19b).

$$Q_{1-2} = \int_1^2 đQ = \int_1^2 T\,dX \tag{3.26}$$

It must also be valid for a quasi-static process only, and the heat transfer involved is represented by the area under the path 1–2 in T–X plot (Fig. 3.19b). Heat transfer is, therefore, a path function, i.e., the amount of heat transferred when a system changes from a state 1 to a state 2 depends on the path the system follows (Fig. 3.19b). Therefore, $đQ$ is an inexact differential. Now,

$$đQ = T\,dX$$

where X is an extensive property and dX is an exact differential.

$$\therefore \qquad dX = \frac{1}{T}\,đQ \tag{3.27}$$

To make $đQ$ integrable, i.e., an exact differential, it must be multiplied by an integrating factor which is, in this case, $1/T$. The extensive property X is yet to be defined. It has been introduced in Chapter 7 and it is called ‘entropy’.

3.9 Specific Heat and Latent Heat

The *specific heat* of a substance is defined as the amount of heat required to raise a unit mass of the substance through a unit rise in temperature. The symbol c will be used for specific heat.

$$\therefore \qquad c = \frac{Q}{m \cdot \Delta t} \text{ J/kg K}$$

where Q is the amount of heat transfer (J), m, the mass of the substance (kg), and Δt, the rise in temperature (K).

Since heat is not a property, as explained later, so the specific heat is qualified with the process through which exchange of heat is made. For gases, if the process is at constant pressure, it is c_p, and if the process is at constant volume, it is c_v. For solids and liquids, however, the specific heat does not depend on the process. An elegant manner of defining specific heats, c_v and c_p, in terms of properties is given in Secs 4.5 and 4.6.

The product of mass and specific heat (mc) is called the *heat capacity* of the substance. The capital letter C, C_p or C_v, is used for heat capacity.

The latent heat is the amount of heat transfer required to cause a phase change in unit mass of a substance *at a constant pressure and temperature*. There are three phases in which matter can exist: solid, liquid, and vapour or gas. *The latent heat of fusion* (l_{fu}) is the amount of heat transferred to melt unit mass of solid into liquid, or to freeze unit mass of liquid to solid. The *latent heat of vaporization* (l_{vap}) is the quantity of heat required to vaporize unit mass of liquid into vapour, or condense unit mass of vapour into liquid. *The latent heat of sublimation* (l_{sub}) is the amount of heat transferred to convert unit mass of solid to vapour or vice versa. l_{fu} is not much affected by pressure, whereas l_{vap} is highly sensitive to pressure.

3.10 Points to Remember Regarding Heat Transfer and Work Transfer

(a) Heat transfer and work transfer are the *energy interactions*. A closed system and its surroundings can interact in two ways: by heat transfer and by work transfer. Thermodynamics studies how these interactions bring about property changes in a system.

(b) The same effect in a closed system can be brought about either by heat transfer or by work transfer. Whether heat transfer or work transfer has taken place depends on what constitutes the system.

(c) Both heat transfer and work transfer are boundary phenomena. Both are observed at the boundaries of the system, and both represent energy crossing the boundaries of the system.

(d) *It is wrong to say 'total heat' or 'heat content' of a closed system*, because heat or work is not a property of the system. Heat, like work, cannot be stored by the system. Both heat and work are the energy in transit.

(e) Heat transfer is the energy interaction due to temperature difference only. All other energy interactions may be termed as work transfer.

(f) Both heat and work are path functions and inexact differentials. The magnitude of heat transfer or work transfer depends upon the path the system follows during the change of state.

Example 3.1 Gas from a bottle of compressed helium is used to inflate an inelastic flexible balloon, originally folded completely flat to a volume of 0.5 m^3. If the barometer reads 760 mm Hg, what is the amount of work done upon the atmosphere by the balloon? Sketch the system before and after the process.

Solution The firm line P_1 (Fig. Ex. 3.1) shows the boundary of the system before the process, and the dotted line P_2 shows the boundary after the process. The displacement work

$$W_d = \int_{\text{Balloon}} p dV + \int_{\text{Bottle}} p dV = p\Delta V + 0$$

$$= 101.325 \frac{\text{kN}}{\text{m}^2} \times 0.5 \text{ m}^3$$

$$= 50.66 \text{ kJ}$$

This is positive because work is done by the system. Work done by the atmosphere is –50.66 kJ. Since the wall of the bottle is rigid, there is no $p\,dV$–work involved in it.

It is assumed that the pressure in the balloon is atmospheric at all times, since the balloon fabric is light, inelastic and unstressed. If the balloon were elastic and stressed during the filling process, the work done by the gas would be greater than 50.66 kJ by an amount equal to the work done in stretching the balloon, although the displacement work done by the atmosphere is still – 50.66 kJ. However, if the system includes both the gas and the balloon, the displacement work would be 50.66 kJ, as estimated above.

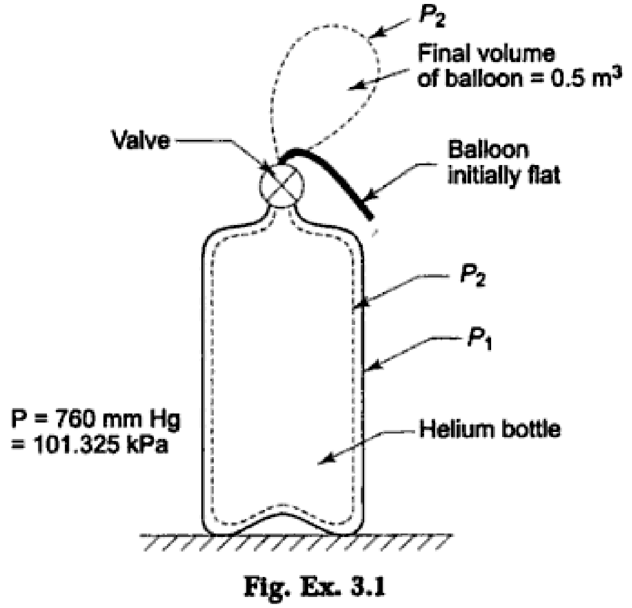


Fig. Ex. 3.1

Example 3.2 When the value of the evacuated bottle (Fig. Ex. 3.2) is opened, atmospheric air rushes into it. If the atmospheric pressure is 101.325 kPa, and 0.6 m^3 of air (measured at atmospheric conditions) enters into the bottle, calculate the work done by air.

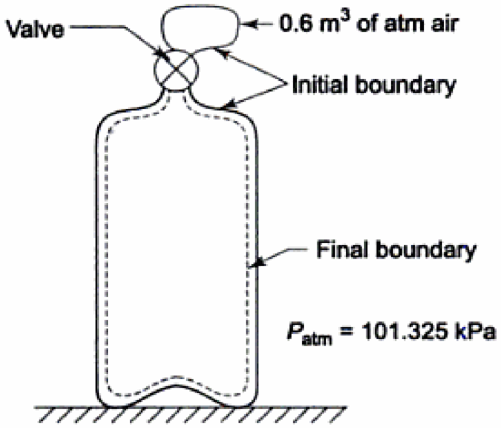


Fig. Ex. 3.2

Solution The displacement work done by air

$$W_d = \int_{\text{Bottle}} p\,dV + \int_{\substack{\text{Free-air}\\ \text{boundary}}} p\,dV$$

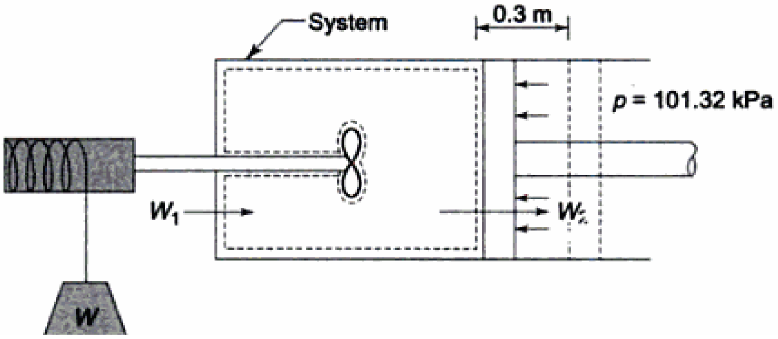
$$= 0 + p\,\Delta V$$

$$= 101.325\ \text{kN/m}^2 \times 0.6\ \text{m}^3$$

$$= 60.8\ \text{kJ}$$

Since the free-air boundary is contracting, the work done by the system is negative (ΔV being negative), and the surroundings do positive work upon the system.

Example 3.3 A piston and cylinder machine containing a fluid system has a stirring device in the cylinder (Fig. Ex. 3.3). The piston is frictionless, and it is held down against the fluid due to the atmospheric pressure of 101.325 kPa. The stirring device is turned 10,000 revolutions with an average torque against the fluid of 1.275 mN. Meanwhile the piston of 0.6 m diameter moves out 0.8 m. Find the net work transfer for the system.



Solution Work done by the stirring device upon the system (Fig. Ex. 3.3).

$$W_1 = 2\pi TN$$
$$= 2\pi \times 1.275 \times 10{,}000 \text{ Nm}$$
$$= 80 \text{ kJ}$$

This is negative work for the system.

Work done by the system upon the surroundings

$$W_2 = (pA) \cdot L$$
$$= 101.325 \frac{\text{kN}}{\text{m}^2} \times \frac{\pi}{4} (0.6)^2 \text{ m}^2 \times 0.80 \text{ m}$$
$$= 22.9 \text{ kJ}$$

This is positive work for the system. Hence, the net work transfer for the system

$$W = W_1 + W_2 = -80 + 22.9 = -57.1 \text{ kJ}$$

Example 3.4 The following data refer to a 12-cylinder, single-acting, two-stroke marine diesel engine:

Speed—150 rpm

Cylinder diameter—0.8 m

Stroke of piston—1.2 m

Area of indicator diagram—5.5×10^{-4} m^2

Length of diagram—0.06 m

Spring value—147 MPa per m

Find the net rate of work transfer from the gas to the pistons in kW.

Solution Mean effective pressure, p_m, is given by

$$p_m = \frac{a_d}{l_d} \times \text{spring constant}$$
$$= \frac{5.5 \times 10^{-4} \text{ m}^2}{0.06} \times 147 \frac{\text{MPa}}{\text{m}}$$
$$= 1.35 \text{ MPa}$$

One engine cycle is completed in two strokes of the piston or one revolution of the crank-shaft.

∴ Work done in one minute

$$= p_m LAN$$
$$= 1.35 \times \frac{\pi}{4}(0.8)^2 \times 1.2 \times 150 = 122 \text{ MJ}$$

Since the engine is single-acting, and it has 12 cylinders, each contributing an equal power, the rate of work transfer from the gas to the piston is given by

$$\mathbf{W} = 122 \times 12 \text{ MJ/min}$$
$$= 24.4 \text{ MJ/s}$$
$$= 24.4 \text{ MW} = 24{,}400 \text{ kW} \qquad \textit{Ans.}$$

Energy can be in two forms: (a) energy in *transit*, like heat and work transfer observed at the boundaries of a system, and (b) energy in *storage*, where energy is stored either *macroscopically* by virtue of motion, position or configuration of the system, or *microscopically* in the molecules or atoms constituting the system.

4.1 First Law for a Closed System Undergoing a Cycle

The transfer of heat and the performance of work may both cause the same effect in a system. Heat and work are different forms of the same entity, called energy, which is conserved. Energy which enters a system as heat may leave the system as work, or energy which enters the system as work may leave as heat.

Let us consider a closed system which consists of a known mass of water contained in an adiabatic vessel having a thermometer and a paddle wheel, as shown in Fig. 4.1. Let a certain amount of work W_{1-2} be done upon the system by the paddle wheel. The quantity of work can be measured by the fall of weight which drives the paddle wheel through a pulley. The system was initially at temperature t_1, the same as that of atmosphere, and after work transfer let the temperature rise to t_2. The pressure is always 1 atm. The process 1–2 undergone by the system is shown in Fig. 4.2 in generalized thermodynamic coordinates X, Y. Let the insulation now be removed. The system and the surroundings interact by heat transfer till the system returns to the original temperature t_1, attaining the condition of thermal equilibrium with the atmosphere. The amount of heat transfer Q_{2-1} from the system during this process, 2–1, shown in Fig. 4.2, can be estimated. The system thus executes 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 the heat Q_{2-1}, and the constant of proportionality is called the Joule's equivalent

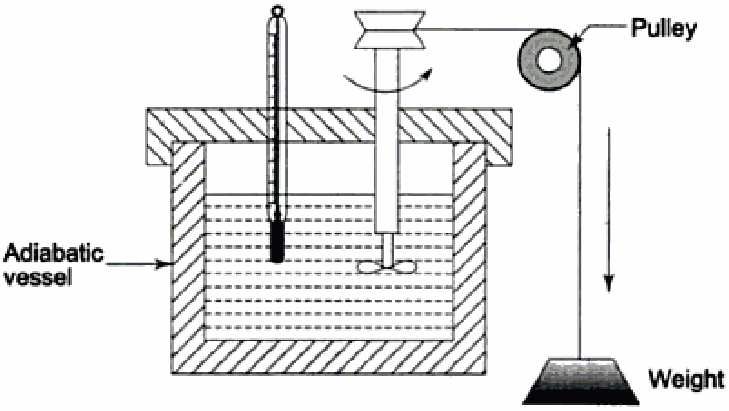


Fig. 4.1 *Adiabatic work*

or the *mechanical equivalent of heat*. In the simple example given here, there are only two energy transfer quantities as the system performs a thermodynamic cycle. If the cycle involves many more heat and work quantities, the same result will be found. Expressed algebraically.

$$(\Sigma W)_{\text{cycle}} = J\,(\Sigma Q)_{\text{cycle}} \tag{4.1}$$

where J is the Joule's equivalent. This is also expressed in the form

$$\oint \mathrm{d}W = J \oint \mathrm{d}Q$$

where the symbol $\oint$ denotes the cyclic integral for the closed path. This is the *first law for a closed system undergoing a cycle*. It is accepted as a *general law of nature*, since no violation of it has ever been demonstrated.

In the S.I. system of units, both heat and work are measured in the derived unit of energy, the Joule. The constant of proportionality, J, is therefore unity ($J = 1$ Nm/J).

The first law of thermodynamics owes much to J.P. Joule who, during the period 1840–1849, carried out a series of experiments to investigate the equivalence of work and heat. In one of these experiments, Joule used an apparatus similar to the one shown in Fig. 4.1. Work was transferred to the

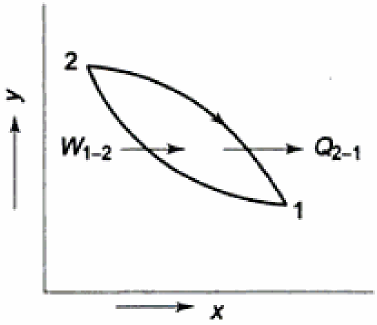


Fig. 4.2 *Cycle completed by a system with two energy interactions: adiabatic work transfer W_{1-2} followed by heat transfer Q_{2-1}*

measured mass of water by means of a paddle wheel driven by the falling weight. The rise in the temperature of water was recorded. Joule also used mercury as the fluid system, and later a solid system of metal blocks which absorbed work by friction when rubbed against each other. Other experiments involved the supplying of work in an electric current. In every case, he found the same ratio (J) between the amount of work and the quantity of heat that would produce identical effects in the system.

Prior to Joule, heat was considered to be an invisible fluid flowing from a body of higher calorie to a body of lower calorie, and this was known as the *caloric theory of heat*. It was Joule who first established that heat is a form of energy, and thus laid the foundation of the first law of thermodynamics.

4.2 First Law for a Closed System Undergoing a Change of State

The expression $(\Sigma W)_{\text{cycle}} = (\Sigma Q)_{\text{cycle}}$ applies only to systems undergoing cycles, and the algebraic summation of all energy transfer across system boundaries is zero. But if a system undergoes a change of state during which both heat transfer and work transfer are involved, the *net* energy transfer will be stored or accumulated within the system. If Q is the amount of heat transferred to the system and W is the amount of work transferred from the system during the process (Fig. 4.3), the net energy transfer $(Q - W)$ will be stored in the system. Energy in storage is neither heat nor work, and is given the name *internal energy* or simply, the *energy* of the system.

Therefore $$Q - W = \Delta E$$

where ΔE is the increase in the energy of the system

or $$Q = \Delta E + W \tag{4.2}$$

Here Q, W, and ΔE are all expressed in the same units (in joules). Energy may be stored by a system in different modes, as explained in Article 4.4.

If there are more energy transfer quantities involved in the process, as shown in Fig. 4.4, the first law gives

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

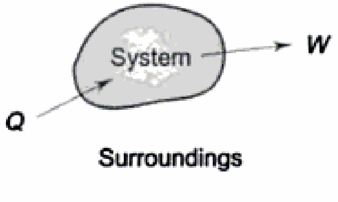


Fig. 4.3 *Heat and work interactions of a system with its surroundings in a process*

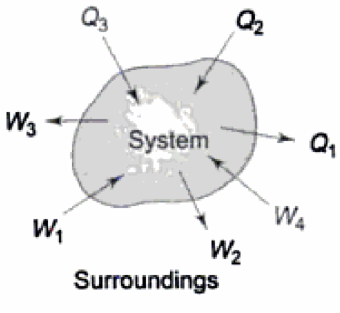


Fig. 4.4 *System-surroundings interaction in a process involving many energy fluxes*

Energy is thus conserved in the operation. The first law is a particular formulation of the principle of the conservation of energy. Equation (4.2) may also be considered as the definition of energy. This definition does not give an absolute value of energy E, but only the change of energy ΔE for the process. It can, however, be shown that the energy has a definite value at every state of a system and is, therefore, a property of the system.

4.3 Energy–A Property of the System

Consider a system which changes its state from state 1 to state 2 by following the path A, and returns from state 2 to state 1 by following the path B (Fig. 4.5). So the system undergoes a cycle. Writing the first law for path A

$$Q_A = \Delta E_A + W_A \tag{4.3}$$

and for path B

$$Q_B = \Delta E_B + W_B \tag{4.4}$$

The processes A and B together constitute a cycle, for which

$$(\Sigma W)_{\text{cycle}} = (\Sigma Q)_{\text{cycle}}$$

or

$$W_A + W_B = Q_A + Q_B$$

or

$$Q_A - W_A = W_B - Q_B \tag{4.5}$$

From equations (4.3), (4.4), and (4.5), it yields

$$\Delta E_A = -\Delta E_B \tag{4.6}$$

Similarly, had the system returned from state 2 to state 1 by following the path C instead of path B

$$\Delta E_A = -\Delta E_C \tag{4.7}$$

From equations (4.6) and (4.7)

$$\Delta E_B = \Delta E_C \tag{4.8}$$

Therefore, it is seen that the change in energy between two states of a system is the same, whatever path the system may follow in undergoing that change of state. If some arbitrary value of energy is assigned to state 2, the value of energy at state 1 is fixed independent of the path the system follows. Therefore, energy has a definite value for every state of the system. Hence, it is a *point function and a property of the system.*

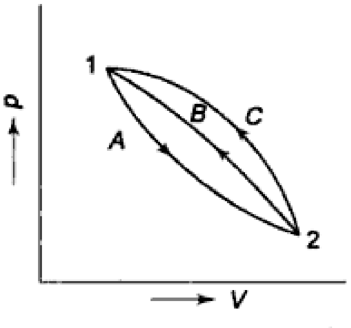


Fig. 4.5 *Energy–a property of a system*

The energy E is an extensive property. The *specific energy*, $e = E/m$ (J/kg), is an intensive property.

The cyclic integral of any property is zero, because the final state is identical with the initial state. $\oint dE = 0$, $\oint dV = 0$, etc. So for a cycle, the equation (4.2) reduces to equation (4.1).

4.9 Perpetual Motion Machine of the First Kind–PMM1

The first law states the general principle of the conservation of energy. *Energy is neither created nor destroyed, but only gets transformed from one form to another*. There can be no machine which would continuously supply mechanical work without some other form of energy disappearing simultaneously (Fig. 4.8). Such a *fictitious machine* is called a *perpetual motion machine of the first kind*, or in brief, PMM1. *A PMM1 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 (Fig. 4.9).

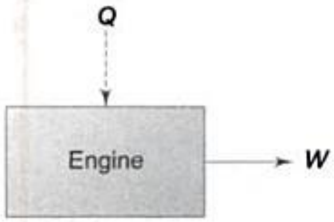


Fig. 4.8 *A PMMI*

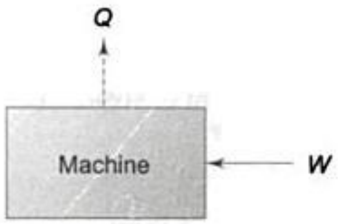


Fig. 4.9 *The converse of PMMI*

4.10 Limitations of the First Law

The first law deals with the amounts of energy of various forms transferred between the system 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 second law which assigns a quality to different forms of energy, and also indicates the direction of any spontaneous process.

SOLVED EXAMPLES

Example 4.1 A stationary mass of gas is compressed without friction from an initial state of 0.3 m^3 and 0.105 MPa to a final state of 0.15 m^3 and 0.105 MPa, the pressure remaining constant during the process. There is a transfer of 37.6 kJ of heat from the gas during the process. How much does the internal energy of the gas change?

Solution First law for a stationary system in a process gives

$$Q = \Delta U + W$$

or

$$Q_{1-2} = U_2 - U_1 + W_{1-2} \qquad (1)$$

Here

$$W_{1-2} = \int_{V_1}^{V_2} p\,dV = p(V_2 - V_1)$$

$$= 0.105\ (0.15 - 0.30)\ \text{MJ}$$

$$= -\ 15.75\ \text{kJ}$$

$$Q_{1-2} = -37.6\ \text{kJ}$$

∴ Substituting in equation (1)

$$-\ 37.6\ \text{kJ} = U_2 - U_1 - 15.75\ \text{kJ}$$

$$\therefore \qquad U_2 - U_1 = -\ 21.85\ \text{kJ} \qquad \textit{Ans.}$$

The internal energy of the gas decreases by 21.85 kJ in the process.

Example 4.2 When a system is taken from state *a* to state *b*, in Fig. Ex. 4.2, along path *acb*, 84 kJ of heat flow into the system, and the system does 32 kJ of work. (a) How much will the heat that flows into the system along path *adb* be, if the work done is 10.5 kJ? (b) When the system is returned from *b* to *a* along the curved path, the work done on the system is 21 kJ. Does the system absorb or liberate heat, and how much of the heat is absorbed or liberated? (c) If $U_a = 0$ and $U_d = 42$ kJ, find the heat absorbed in the processes *ad* and *db*.

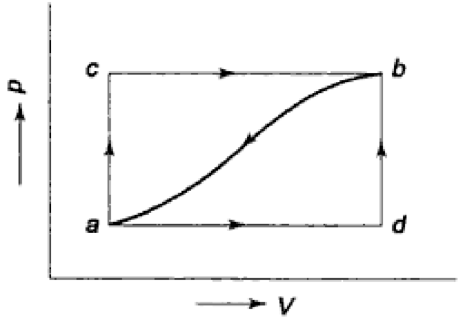


Fig. Ex. 4.2

Solution

$$Q_{acb} = 84 \text{ kJ}$$
$$W_{acb} = 32 \text{ kJ}$$

We have

$$Q_{acb} = U_b - U_a + W_{acb}$$

∴ $$U_b - U_a = 84 - 32 = 52 \text{ kJ}$$ *Ans.*

(a) $$Q_{adb} = U_b - U_a + W_{adb} = 52 + 10.5 = 62.5 \text{ kJ}$$ *Ans.*

(b) $$Q_{b-a} = U_a - U_b + W_{b-a} = -52 - 21 = -73 \text{ kJ}$$ *Ans.*

The system liberates 73 kJ of heat.

(c) $$W_{adb} = W_{ad} + W_{db} = W_{ad} = 10.5 \text{ kJ}$$

∴ $$Q_{ad} = U_d - U_a + W_{ad} = 42 - 0 + 10.5 = 52.5 \text{ kJ}$$

Now $$Q_{adb} = 62.5 \text{ kJ} = Q_{ad} + Q_{db}$$

∴ $$Q_{db} = 62.5 - 52.5 = 10 \text{ kJ}$$ *Ans.*

Example 4.3 A piston and cylinder machine contains a fluid system which passes through a complete cycle of four processes. During a cycle, the sum of all heat transfers is –170 kJ. The system completes 100 cycles per min. Complete the

following table showing the method for each item, and compute the net rate of work output in kW.

Process	*Q (kJ/min)*	*W (kJ/min)*	*ΔE (kJ/min)*
a–b	0	2,170	—
b–c	21,000	0	—
c–d	–2,100	—	–36,600
d–a	—	—	—

Solution Process *a–b*:

$$Q = \Delta E + W$$

$$0 = \Delta E + 2170$$

$$\therefore \quad \Delta E = -2170 \text{ kJ/min}$$

Process *b–c*:

$$Q = \Delta E + W$$

$$21{,}000 = \Delta E + 0$$

$$\therefore \quad \Delta E = 21{,}000 \text{ kJ/min}$$

Process *c–d*:

$$Q = \Delta E + W$$

$$-2100 = -36{,}600 + W$$

$$\therefore \quad W = 34{,}500 \text{ kJ/min}$$

Process *d–a*:

$$\sum_{\text{cycle}} Q = -170 \text{ kJ}$$

The system completes 100 cycles/min.

$$\because \quad Q_{ab} + Q_{bc} + Q_{cd} + Q_{da} = -17{,}000 \text{ kJ/min}$$

$$0 + 21{,}000 - 2{,}100 + Q_{da} = -17{,}000$$

$$\therefore \quad Q_{da} = -35{,}900 \text{ kJ/min}$$

Now $\oint dE = 0$, since cyclic integral of any property is zero.

$$\therefore \quad \Delta E_{a-b} + \Delta E_{b-c} + \Delta E_{c-d} + \Delta E_{d-a} = 0$$

$$-2{,}170 + 21{,}000 - 36{,}600 + \Delta E_{d-a} = 0$$

$$\therefore \quad \Delta E_{d-a} = 17{,}770 \text{ kJ/min}$$

$$\therefore \quad W_{d-a} = Q_{d-a} - \Delta E_{d-a}$$

$$= -35{,}900 - 17{,}770$$

$$= -53{,}670 \text{ kJ/min}$$

The table becomes

Process	*Q* (kJ/min)	*W* (kJ/min)	*ΔE* (kJ/min)
a–b	0	2,170	– 2,170
b–c	21,000	0	21,000

c–d	– 2,100	34,500	– 36,600
d–a	– 35,900	– 53,670	17,770

Since

Rate of work output $$\sum_{\text{cycle}} Q = \sum_{\text{cycle}} W$$

$$= -\,17{,}000 \text{ kJ/min}$$

$$= -\,283.3 \text{ kW}$$ *Ans.*

Example 4.4 The internal energy of a certain substance is given by the following equation

$$u = 3.56\, pv + 84$$

where u is given in kJ/kg, p is in kPa, and v is in m^3/kg.

A system composed of 3 kg of this substance expands from an initial pressure of 500 kPa and a volume of 0.22 m^3 to a final pressure 100 kPa in a process in which pressure and volume are related by $pv^{1.2}$ = constant.

(a) If the expansion is quasi-static, find Q, ΔU, and W for the process.
(b) In another process the same system expands according to the same pressure-volume relationship as in part (a), and from the same initial state to the same final state as in part (a), but the heat transfer in this case is 30 kJ. Find the work transfer for this process.
(c) Explain the difference in work transfer in parts (a) and (b).

Solution

(a) $$u = 3.56\, pv + 84$$

$$\Delta u = u_2 - u_1 = 3.56\,(p_2 v_2 - p_1 v_1)$$

$\therefore$ $$\Delta U = 3.56\,(p_2 V_2 - p_1 V_1)$$

Now $$p_1 V_1^{\,1.2} = p_2 V_2^{\,1.2}$$

$\therefore$ $$V_2 = V_1\left(\frac{p_1}{p_2}\right)^{1/1.2} = 0.22\left(\frac{5}{1}\right)^{1/1.2}$$

$$= 0.22 \times 3.83 = 0.845 \text{ m}^3$$

$\therefore$ $$\Delta U = 356\,(1 \times 0.845 - 5 \times 0.22) \text{ kJ}$$

$$= -356 \times 0.255 = -91 \text{ kJ}$$ *Ans.* (a)

For a quasi-static process

$$W = \int p\mathrm{d}V = \frac{p_2 V_2 - p_1 V_1}{1 - n}$$

$$= \frac{(1 \times 0.845 - 5 \times 0.22)100}{1 - 1.2} = 127.5 \text{ kJ}$$

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

$$= -\,91 + 127.5 = 36.5 \text{ kJ}$$ *Ans.* (a)

(b) Here $Q = 30$ kJ

Since the end states are the same, ΔU would remain the same as in (a).

$$\therefore \quad W = Q - \Delta U$$
$$= 30 - (-91)$$
$$= 121 \text{ kJ} \qquad \textit{Ans.} \text{ (b)}$$

(c) The work in (b) is not equal to $\int p\mathrm{d}V$ since the process is not quasi-static.

Example 4.5 A fluid is confined in a cylinder by a spring-loaded, frictionless piston so that the pressure in the fluid is a linear function of the volume ($p = a + bV$). The internal energy of the fluid is given by the following equation

$$U = 34 + 3.15\, pV$$

where U is in kJ, p in kPa, and V in cubic metre. If the fluid changes from an initial state of 170 kPa, 0.03 m^3 to a final state of 400 kPa, 0.06 m^3, with no work other than that done on the piston, find the direction and magnitude of the work and heat transfer.

Solution The change in the internal energy of the fluid during the process.

$$U_2 - U_1 = 3.15\,(p_2V_2 - p_1V_1)$$
$$= 315\,(4 \times 0.06 - 1.7 \times 0.03)$$
$$= 315 \times 0.189 = 59.5 \text{ kJ}$$

Now

$$p = a + bV$$
$$170 = a + b \times 0.03$$
$$400 = a + b \times 0.06$$

From these two equations

$$a = -60 \text{ kN/m}^2$$
$$b = 7667 \text{ kN/m}^5$$

Work transfer involved during the process

$$W_{1-2} = \int_{V_1}^{V_2} p\,\mathrm{d}V = \int_{V_1}^{V_2} (a + bV)\,\mathrm{d}V$$
$$= a(V_2 - V_1) + b\,\frac{V_2^2 - V_1^2}{2}$$
$$= (V_2 - V_1)\left[a + \frac{b}{2}(V_1 + V_2)\right]$$
$$= 0.03 \text{ m}^3\left[-60 \text{ kN/m}^2 + \frac{7667}{2}\,\frac{\text{kN}}{\text{m}^5} \times 0.09 \text{ m}^3\right]$$
$$= 8.55 \text{ kJ}$$

Work is done by the system, the magnitude being 8.55 kJ.

∴ Heat transfer involved is given by

$$\begin{aligned} Q_{1-2} &= U_2 - U_1 + W_{1-2} \\ &= 59.5 + 8.55 \\ &= 68.05 \text{ kJ} \end{aligned}$$

68.05 kJ of heat flow into the system during the process.

4.6 Enthalpy

The enthalpy of a substance, h, is defined as

$$h = u + pv \tag{4.24}$$

It is an intensive property of a system (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 (4.22), it is possible to derive an expression for the heat transfer in a constant pressure process involving no work other than pdV work. In such a process in a closed stationary system of unit mass of a pure substance

$$đQ = \mathrm{d}u + p\mathrm{d}v$$

At constant pressure

$$p\mathrm{d}v = \mathrm{d}(pv)$$

∴ $$(đQ)_\mathrm{p} = \mathrm{d}u + \mathrm{d}(pv)$$

or $$(đQ)_\mathrm{p} = \mathrm{d}(u + pv)$$

or $$(đQ)_\mathrm{p} = \mathrm{d}h \tag{4.25}$$

where $h = u + pv$ is the *specific enthalpy*, a property of the system.

Heat transferred at constant pressure increases the enthalpy of a system.

For an ideal gas, the enthalpy becomes

$$h = u + RT \tag{4.26}$$

Since the internal energy of an ideal gas depends only on the temperature (Eq. 4.11), the enthalpy of an ideal gas also depends on the temperature only, i.e.

$$h = f(T) \text{ only} \tag{4.27}$$

Total enthalpy $H = mh$

Also $$H = U + pV$$

and $$h = H/m \text{ (J/kg)}$$

SIMPLIFIED NOTES ON 1st LAW OF TD

FIRST LAW OF THERMODYNAMICS

- This is based on Law of Conservation of Energy.
- This is also called as First Principle.

For a closed system, undergoing a cycle

Sum of all Work transfers = Sum of all Heat Transfers

$$(W_1+W_2+W_3+\ldots..) = \Sigma(Q_1+Q_2+Q_3+\ldots\ldots\ldots)$$

$$\Sigma(W) = \Sigma(Q)$$

$$\oint dW = \oint dQ$$

For a closed system,undergoing a Process

Whenever heat is absorbed by a system it increases its internal energy and does some work.

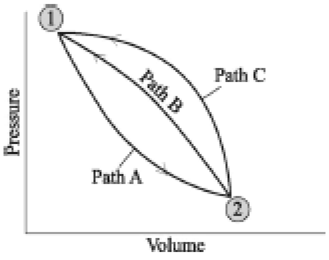
$$Q = \Delta E + W$$

Where Q – heat absorbed by the system

W – Work output from the system

ΔE – Change in Stored Energy of the system

Show that Energy is a property of the system



For path A,

$$Q_A = W_A + \Delta E_A \qquad (1)$$

For path B,

$$Q_B = W_B + \Delta E_B \qquad (2)$$

For path B,

$$Q_C = W_C + \Delta E_C \qquad (3)$$

For Cycle 1-A-2-B-1,

$$W_A + W_B = Q_A + Q_B \qquad (4)$$

$$Q_A - W_A = -(Q_B - W_B)$$

$$\Delta E_A = -\Delta E_B \qquad (A)$$

For Cycle 1-A-2-C-1,

$$W_A + W_C = Q_A + Q_C$$

$$Q_A - W_A = -(Q_C - W_C)$$

$$\Delta E_A = -\Delta E_C \qquad (C)$$

Comparing A and C

$$\Delta E_B = \Delta E_C$$

Enthalpy:

- It is the energy content of the flowing fluid.
- It is defined by the summation of internal energy and flow work.

H = U + PV

<u>Note:</u> For an ideal gas h = u + Pv.

= u + RT

So, h = f(T)

Define C_v with the help internal energy and Temperature:

The amount of heat required to raise the temperature of unit mass of a substance by 1° C in a reversible constant volume process.

$$C_v = \left(\frac{\partial u}{\partial T}\right)_v$$

C_v is also defined as the change of internal energy of the substance per unit change in temperature at constant volume.

Define C_p with the help enthalpy and Temperature:

The amount of heat required to raise the temperature of unit mass of a substance by 1° C in a reversible constant pressure process.

$$C_p = \left(\frac{\partial h}{\partial T}\right)_p$$

C_p is also defined as the change of internal energy of the substance per unit change in temperature at constant pressure.

Application of First law to different Thermodynamic process:

Process	**Index=n**	Q	W= $\int PdV$	P-V-T Relation
Rev. Const.Vol.	∞	$Q = \Delta U$ $= mC_v(T_2 - T_1)$	W=0	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$
Rev.Const.pressure	n=0	$Q = \Delta H$ $= mC_p(T_2 - T_1)$	$W = P(V_2 - V_1)$ $= mR(T_2 - T_1)$	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Rev. Isothermal	n=1	$Q = W = P_1V_1 \ln\left(\frac{V_2}{V_1}\right)$	$W = P_1V_1 \ln\left(\frac{V_2}{V_1}\right)$	$P_1V_1 = P_2V_2$
Rev.Adiabatic	n=γ	Q=0	$W = \frac{P_1V_1 - P_2V_2}{\gamma - 1}$	$P_1V_1^\gamma = P_2V_2^\gamma$
Rev.Polytropic	n	$Q = \Delta U + W$	$W = \frac{P_1V_1 - P_2V_2}{n - 1}$	$P_1V_1^n = P_2V_2^n$

3. Its value for an overall system is the sum of its values for the parts into which the system is divided.	3. Its value remains the same whether one considers the whole system or only a part of it.
4.Example:mass(m),volume(V),Energy(E),Enthalpy(H) etc.	4.Example:Pressure(P),Temperature(T),Density etc.
5. Uppercase letters are used for extensive properties except mass.	5. Lowercase letters are used for intensive properties except pressure(P) and temp.(T)

Specific property= Extensive property/mass.

Example: Specific volume (**v**) = Volume(**V**)/mass(**m**)

Specific enthalpy (**h**) = Enthalpy(**H**)/mass(**m**)

Specific entropy (s) = Entropy(S)/mass(**m**)

Path and Process:

- The series of state through a system passes during a change of state is ***Path of the system***.
- If the path followed by the system during change of state is specified or defined completely, then it is called a process.

We can allow one of the properties to remain a constant during a process.

Isothermal ⟹ Constant Temperature (**T**)

Isobaric ⟹ Constant Pressure (**P**)

Isochoric ⟹ Constant Volume (**V**)

Isentropic ⟹ Constant Entropy (s)

Isenthalpic ⟹ Constant Enthalpy (**h**)

Concept of continuum

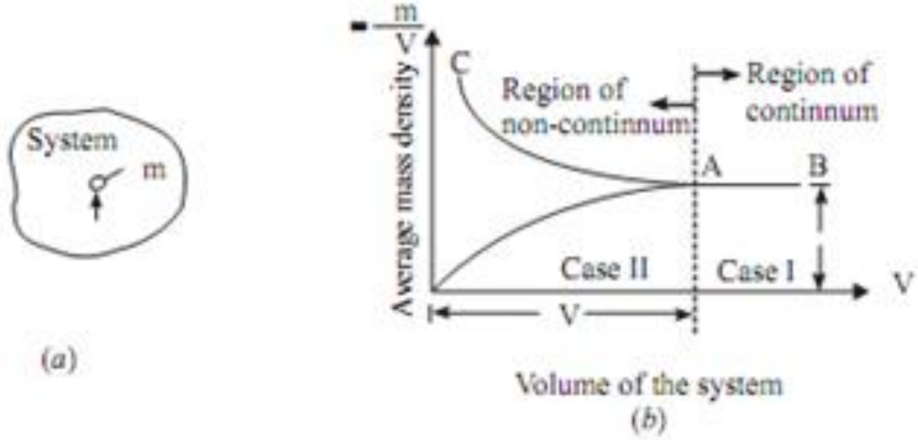
Even the simplification of matter into molecules, atoms, electrons, etc. is too complex a picture for many problems of thermodynamics. Thermodynamics doesn't make any hypotheses about the structure of the matter of system. The volumes of system considered are quite large as compared to molecular dimensions. The system can be regarded as continuum. The system is supposed to contain continuous distribution of matter. There are no voids and cavities present. The pressure, temperature, density and other properties are average values of action of several molecules and atoms. This kind of idealization is a must for solving most of the problems. The laws and concepts of thermodynamics are not dependent of structure of matter.

In accordance to this concept there is minimum limit of volume up to which the property remains continuum. Below this volume, there is sudden change in the value of the property. This type of region is called as region of discrete particles and the region for which the property are maintain is called as region of continuum. The volume up to which continuum properties are maintained is called as continuum limit.

For Example: If we measure the density of a substance for a large volume (ν_1), the value of density is(ρ_1). If we go on reducing the volume by v', below which ratio äm/äv deviates from its actual value and value of äm/äv is large or small.

Thus according to this concept design could be defined as

$$\rho = \lim \delta v - \delta v' \, [\delta m / \delta v]$$



(*a*) (*b*)

Thermodynamic Equilibrium

A system is said to exist in a state of Thermodynamic Equilibrium when no changes in macroscopic property is observed if the system is isolated from its surrounding.

At the state of equilibrium, the properties of the system are uniform and only one value can be assigned to it.

A system will be in a state of thermodynamic equilibrium, if the condition for following three types of equilibrium are satisfied.

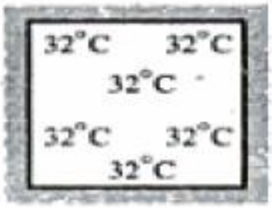
Thermal Equilibrium

Mechanical Equilibrium

Chemical Equilibrium

Thermal Equilibrium (Equality of Temperature):

A state of thermal equilibrium can be described as one in which the temperature of the system is uniform.



Mechanical Equilibrium(Equality of Pressure):

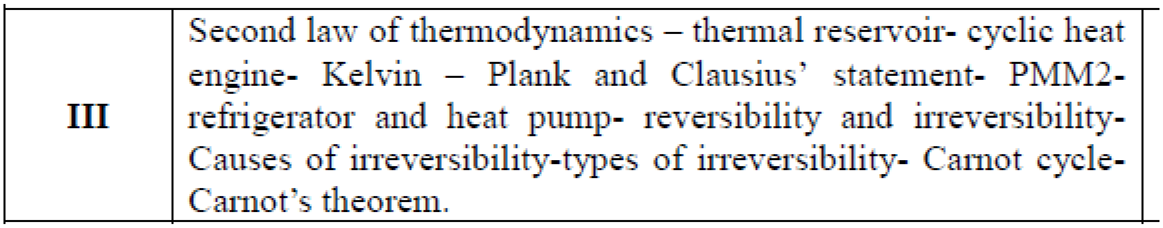
In the absence of any unbalanced force within the system itself and also between the system and the surroundings, the system is said to be in a state of mechanical equilibrium.

- Mechanical equilibrium is related to pressure.
- A system is in mechanical equilibrium if there is no change in pressure at any point of the system .

Chemical Equilibrium(Equality of chemical potential):

- A system is in chemical equilibrium when its chemical composition does not change with time, that is no chemical reaction occurs .
- It is related to chemical potential.

MODULE – 3
2ND LAW OF THERMODYNAMICS



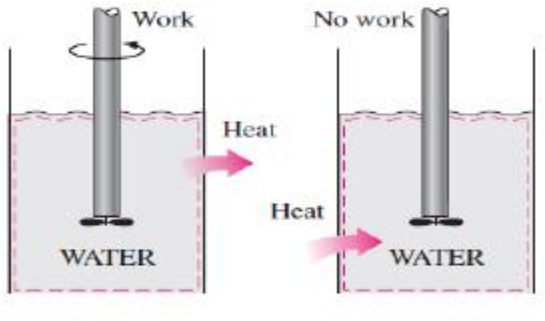
III	Second law of thermodynamics – thermal reservoir- cyclic heat engine- Kelvin – Plank and Clausius' statement- PMM2- refrigerator and heat pump- reversibility and irreversibility- Causes of irreversibility-types of irreversibility- Carnot cycle- Carnot's theorem.

THERMAL RESERVOIR

- Thermal Reservoir is a hypothetical body with a relatively large thermal energy capacity that can supply or absorb finite amounts of heat without undergoing any change in temperature. Such a body is called a thermal energy reservoir, or just a **reservoir**.
- A reservoir that supplies energy in the form of heat is called a source, and one that absorbs energy in the form of heat is called a **sink**
- Thermal energy reservoirs are often referred to as heat reservoirs since they supply or absorb energy in the form of heat.

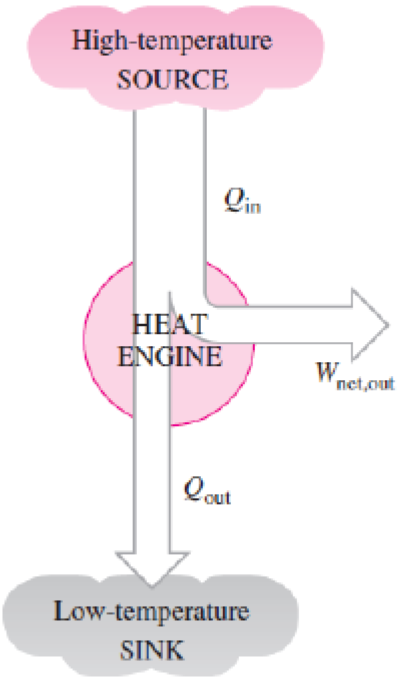
HEAT ENGINE

- Work can easily be converted to other forms of energy, but converting other forms of energy to work is not that easy.
- The mechanical work done by the shaft for example, is first converted to the internal energy of the water.
- This energy may then leave the water as heat.
- From experience that any attempt to reverse this process will fail.
- That is, transferring heat to the water does not cause the shaft to rotate.



Work can always be converted to heat directly and completely, but the reverse is not true.

From this and other observations, we conclude that work can be converted to heat directly and completely, but converting heat to work requires the use of some special devices.
These devices are called **heat engines**



$$W_{net,out} = W_{out} - W_{in}$$

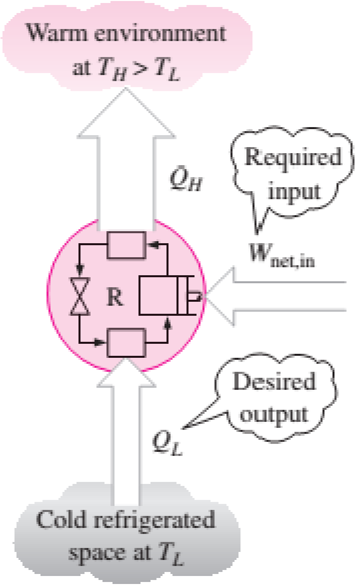
$$\text{Thermal efficiency} = \frac{\text{Net work output}}{\text{Total heat input}}$$

$$\eta_{th} = \frac{W_{net,out}}{Q_{in}}$$

$$\eta_{th} = 1 - \frac{Q_{out}}{Q_{in}}$$

REFRIGERATORS AND HEAT PUMPS

- The transfer of heat from a low-temperature medium to a high-temperature one requires special devices called refrigerators



Coefficient of Performance

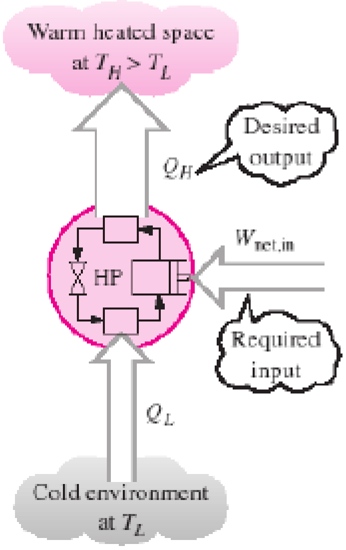
- The efficiency of a refrigerator is expressed in terms of the coefficient of performance (COP), denoted by COP_R.

$$COP_R = \frac{\text{Desired output}}{\text{Required input}} = \frac{Q_L}{W_{net,in}}$$

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

HEAT PUMP

- Another device that transfers heat from a lowtemperature medium to a high-temperature one is the heat pump.
- Refrigerators and heat pumps operate on the same cycle but differ in their objectives.
- The objective of a refrigerator is to maintain the refrigerated space at a low temperature by removing heat from it.
- The objective of a heat pump is to maintain a heated space at a high temperature.



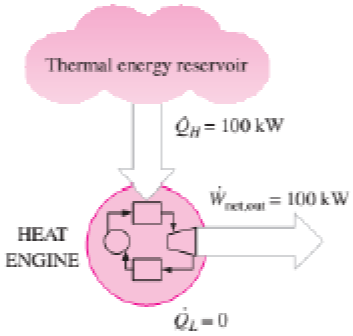
The objective of a heat pump is to supply heat Q_H into the warmer space.

$$\mathrm{COP_{HP}} = \frac{\text{Desired output}}{\text{Required input}} = \frac{Q_H}{W_{\text{net,in}}}$$

$$\mathrm{COP_{HP}} = \mathrm{COP_R} + 1$$

The Second Law of Thermodynamics:
Kelvin–Planck Statement

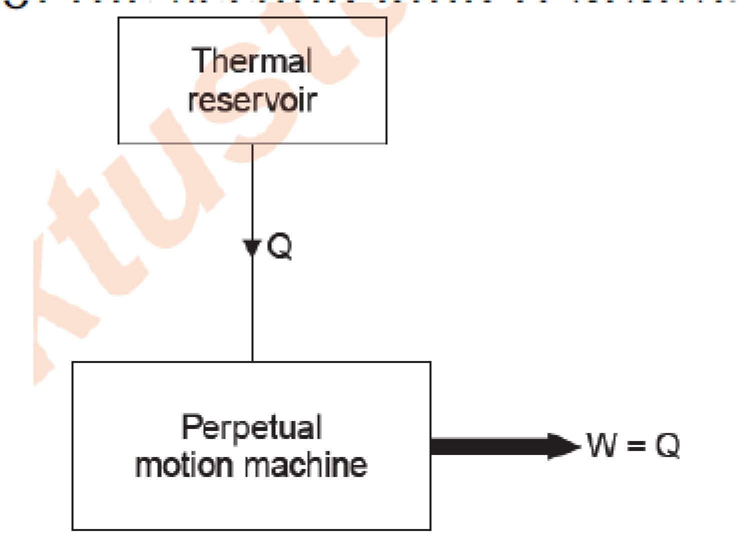
- It is impossible for any device that operates on a cycle to receive heat from a single reservoir and produce a net amount of work
- No heat engine can have a thermal efficiency of 100 percent.



PMM2-Violates 2nd Law

PMM2

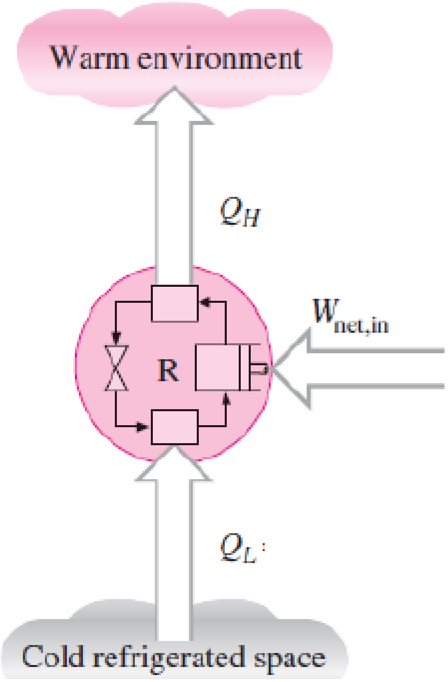
- A device that violates the second law of thermodynamics is called a perpetual-motion machine of the second kind (PMM2).



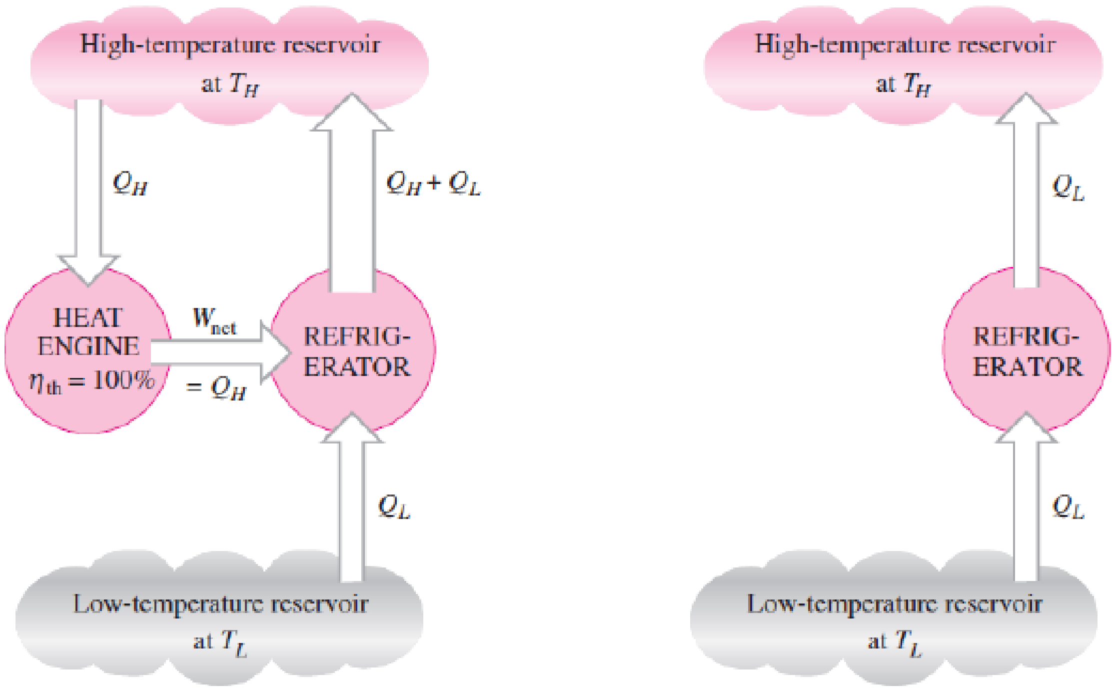
Perpetual motion machine of second kind (PMM2).

THE SECOND LAW OF THERMODYNAMICS : CLAUSIUS STATEMENT

- It is impossible to construct a device that operates in a cycle and produces no effect other than the transfer of heat from a lower-temperature body to a higher-temperature body.



EQUIVALENCE OF TWO STATEMENTS



(*a*) A refrigerator that is powered by a 100 percent efficient heat engine

(*b*) The equivalent refrigerator

Consider the heat-engine-refrigerator combination operating between the same two reservoirs.

- The heat engine is assumed to have, in violation of the Kelvin–Planck statement, a thermal efficiency of 100 percent, and therefore it converts all the heat Q_H it receives to work W.
- This work is now supplied to a refrigerator that removes heat in the amount of QL from the lowtemperature reservoir and rejects heat in the amount of Q_L+Q_H to the high-temperature reservoir.
- During this process, the high temperature reservoir receives a net amount of heat Q_L (the difference between Q_L+Q_H and Q_H).

Thus, the combination of these two devices can be viewed as a refrigerator, fig (b), that transfers heat in an amount of QL from a cooler body to a warmer one without requiring any input from outside.

- This is clearly a violation of the Clausius statement.
- Therefore, a violation of the Kelvin–Planck statement results in the violation of the Clausius statement.
- It can also be shown in a similar manner that a violation of the Clausius statement leads to the violation of the Kelvin–Planck statement.

REVERSIBLE AND IRREVERSIBLE PROCESSES

- A reversible process is defined as a process that can be reversed without leaving any trace on the surroundings
- That is, both the system and the surroundings are returned to their initial states at the end of the reverse process.
- This is possible only if the net heat and net work exchange between the system and the surroundings is zero for the combined (original and reverse) process.
- Processes that are not reversible are called irreversible processes.
- For irreversible processes, the surroundings usually do some work on the system

Reversible processes actually do not occur in nature.
They are merely *idealizations* of actual processes.

- Reversible processes can be approximated by actual devices, but they can never be achieved. That is, all the processes occurring in nature are irreversible
- Reversible processes are easy to analyze, since a system passes through a series of equilibrium states during a reversible process; second, they serve as idealized models to which actual processes can be compared.

CAUSES OF IRREVERSIBILITY

- The factors that cause a process to be irreversible are called irreversibilities.
- They include friction, unrestrained expansion, mixing of two fluids, heat transfer across a finite temperature difference, electric resistance, inelastic deformation of solids, and chemical reactions.
- The presence of any of these effects renders a process irreversible.
- A reversible process involves none of these.
- Some of the frequently encountered irreversibilities are discussed briefly below

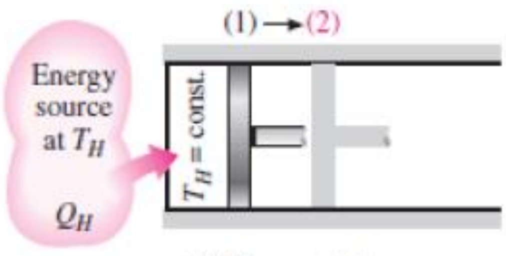
TYPES OF IRREVERSIBIITY

- Friction
- unrestrained expansion
- mixing of two fluids
- heat transfer across a finite temperature difference
- electric resistance
- inelastic deformation of solids
- chemical reactions

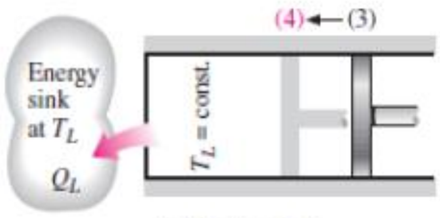
THE CARNOT CYCLE

Probably the best known reversible cycle is the Carnot cycle, first proposed in 1824 by French engineer Sadi Carnot.

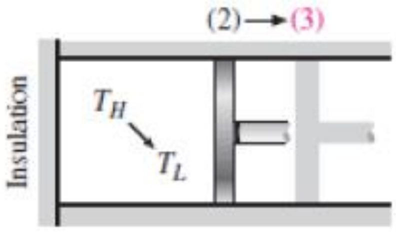
- The theoretical heat engine that operates on the Carnot cycle is called the Carnot heat engine.
- The Carnot cycle is composed of four reversible processes—two isothermal and two adiabatic—and it can be executed either in a closed or a steadyflow system.
- Consider a closed system that consists of a gas contained in an adiabatic piston–cylinder device,
- The insulation of the cylinder head is such that it may be removed to bring the cylinder into contact with reservoirs to provide heat transfer.
- The four reversible processes that make up the Carnot cycle are as follows:



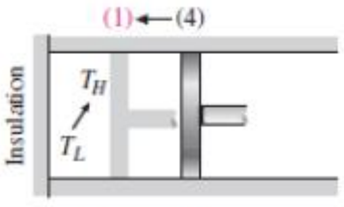
(*a*) Process 1-2



(*c*) Process 3-4



(*b*) Process 2-3



(*d*) Process 4-1

Reversible Isothermal Expansion (process 1-2, T_H constant).

- Initially(state 1), the temperature of the gas is T_H and the cylinder head is in close contact with a source at temperature T_H.
- The gas is allowed to expand slowly, doing work on the surroundings.
- As the gas expands, the temperature of the gas tends to decrease.

- But as soon as the temperature drops by an infinitesimal amount dT, some heat is transferred from the reservoir into the gas, raising the gas temperature to T_H.
- Thus, the gas temperature is kept constant at T_H.
- Since the temperature difference between the gas and the reservoir never exceeds a differential amount dT, this is a reversible heat transfer process.

Reversible Adiabatic Expansion (process 2-3, temperature drops from T_H to T_L).

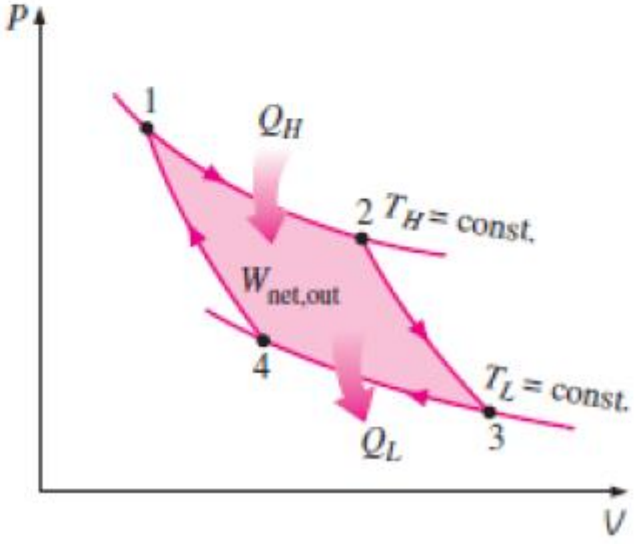
- At state 2, the reservoir that was in contact with the cylinder head is removed and replaced by insulation so that the system becomes adiabatic.
- The gas continues to expand slowly, doing work on the surroundings until its temperature drops from T_H to T_L (state 3).
- The piston is assumed to be frictionless and the process to be quasiequilibrium, so the process is reversible as well as adiabatic

Reversible Isothermal Compression (process 3-4, T_L constant).

- At state 3, the insulation at the cylinder head is removed, and the cylinder is brought into contact with a sink at temperature T_L.
- Now the piston is pushed inward by an external force, doing work on the gas.
- As the gas is compressed, its temperature tends to rise.
- But as soon as it rises by an infinitesimal amount dT, heat is transferred from the gas to the sink, causing the gas temperature to drop to T_L.
- Thus, the gas temperature remains constant at T_L.
- Since the temperature difference between the gas and the sink never exceeds a differential amount dT, this is a reversible heat transfer process.
- It continues until the piston reaches state 4

Reversible Adiabatic Compression (process 4-1, temperature rises from T_L to T_H).

- State 4 is such that when the low-temperature reservoir is removed, the insulation is put back on the cylinder head, and the gas is compressed in a reversible manner, the gas returns to its initial state (state 1).
- The temperature rises from TL to TH during this reversible adiabatic compression process, which completes the cycle.



P-*V* diagram of the Carnot cycle.

MODULE – 4

ENTROPY

IV	Entropy- Clausius' theorem- Clausius' inequality- Entropy principle and its applications- Available energy-Law of degradation of energy- useful work- dead state- Availability-Gibb's and Helmholtz function-Second law efficiency	

CLAUSIUS THEOREM

Let a system be taken from an equilibrium state i to another equilibrium state f by following the reversible path i-f (Fig. 7.2). Let a reversible adiabatic i-a be drawn through i and another reversible adiabatic b-f be drawn through f. Then a reversible isotherm a-b is drawn in such a way that the area under i-a-b-f is equal to the area under i-f. Applying the first law for

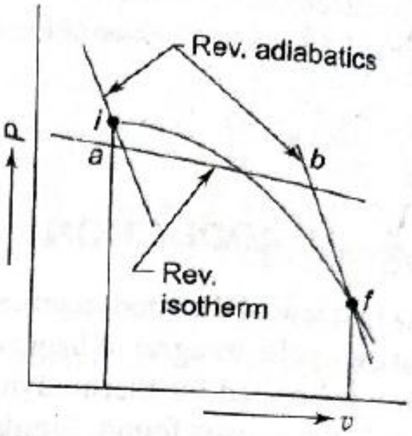


Fig. 7.2 *Reversible Path Substituted by Two Reversible Adiabatics and a Reversible Isotherm*

Process i-f

$$Q_{i\text{-}f} = U_f - U_i + W_{if} \qquad (7.1)$$

Process i-a-b-f

$$Q_{iabf} = U_f - U_i + W_{iabf} \qquad (7.2)$$

Since

$$W_{if} = W_{iabf}$$

$\therefore$ From Eqs. (7.1) and (7.2)

$$Q_{if} = Q_{iabf}$$

$$= Q_{ia} + Q_{ab} + Q_{bf}$$

Since $Q_{ia} = 0$ and $Q_{bf} = 0$

$$Q_{if} = Q_{ab}$$

Heat transferred in the process *i-f* is equal to the heat transferred in the isothermal process *a-b*.

Thus any reversible path may be substituted by a reversible zigzag path, between the same end states, consisting of a reversible adiabatic followed by a reversible isotherm and then by a reversible adiabatic, such that the heat transferred during the isothermal process is the same as that transferred during the original process.

Let a smooth closed curve representing a reversible cycle (Fig. 7.3) be considered. Let the closed cycle be divided into a large number of strips by means of reversible adiabatics. Each strip may be closed at the top and bottom by reversible isotherms. The original closed cycle is thus replaced by a zigzag closed path consisting of alternate adiabatic and isothermal processes, such that the heat transferred during all the isothermal processes is equal to the heat transferred in the original cycle. Thus the original cycle is replaced by a large number of Carnot cycles. If the adiabatics are close to one another and the number of Carnot cycles is large, the saw-toothed zig-zag line will coincide with the original cycle.

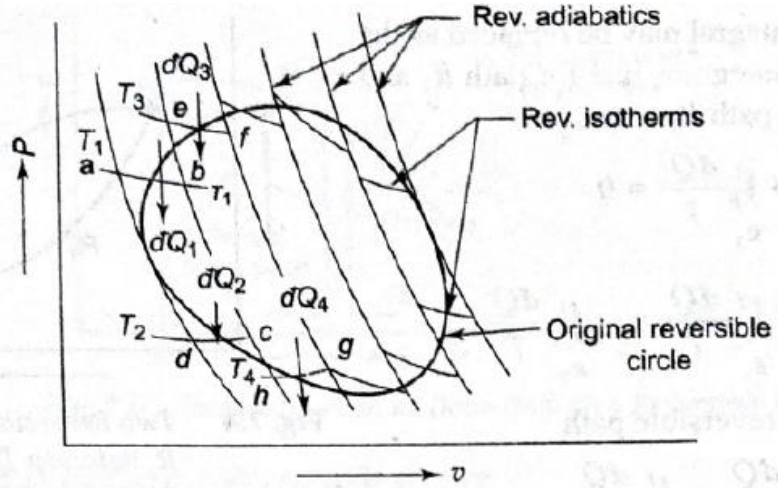


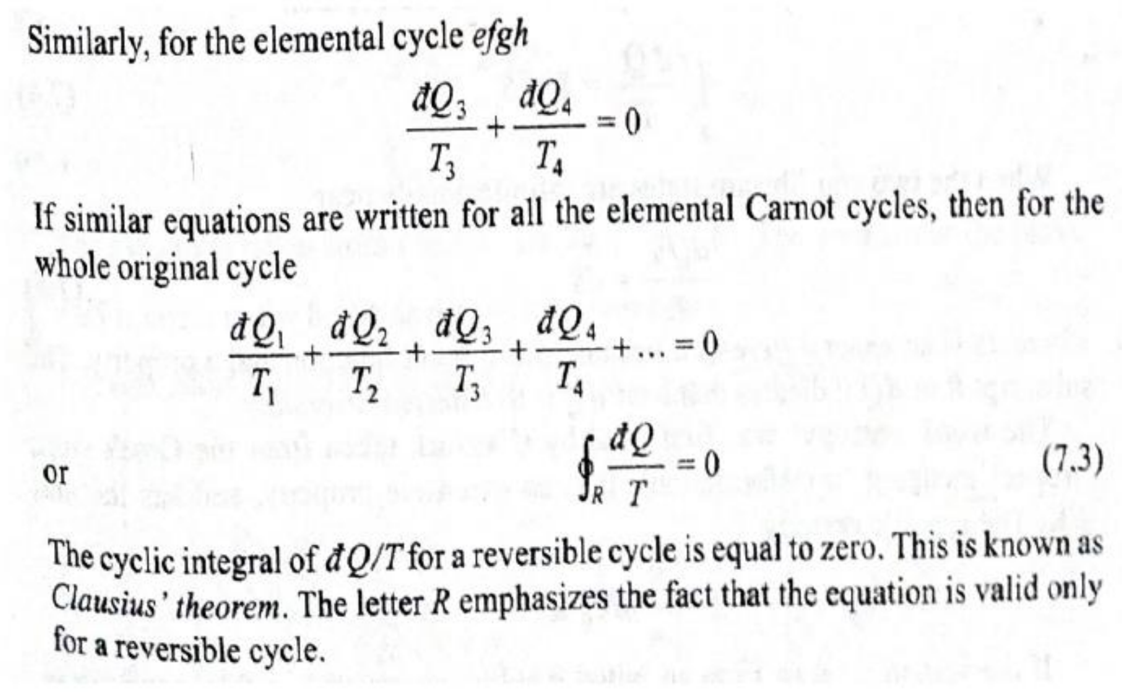
Fig. 7.3 *A Reversible Cycle Split into a Large Number of Carnot Cycles*

For the elemental cycle *abcd* $đQ_1$ heat is absorbed reversibly at T_1, and $đQ_2$ heat is rejected reversibly at T_2

$$\frac{đQ_1}{T_1} = \frac{đQ_2}{T_2}$$

If heat supplied is taken as positive and heat rejected as negative

$$\frac{đQ_1}{T_1} + \frac{đQ_2}{T_2} = 0$$



Similarly, for the elemental cycle *efgh*

$$\frac{\text{đ}Q_3}{T_3} + \frac{\text{đ}Q_4}{T_4} = 0$$

If similar equations are written for all the elemental Carnot cycles, then for the whole original cycle

$$\frac{\text{đ}Q_1}{T_1} + \frac{\text{đ}Q_2}{T_2} + \frac{\text{đ}Q_3}{T_3} + \frac{\text{đ}Q_4}{T_4} + \ldots = 0$$

or

$$\oint_R \frac{\text{đ}Q}{T} = 0 \qquad (7.3)$$

The cyclic integral of $\text{đ}Q/T$ for a reversible cycle is equal to zero. This is known as *Clausius' theorem*. The letter R emphasizes the fact that the equation is valid only for a reversible cycle.

THE PROPERTY OF ENTROPY

Let a system be taken from an initial equilibrium state i to a final equilibrium state f by following the reversible path R_1 (Fig. 7.4). The system is brought back from f to i by following another reversible path R_2. Then the two paths R_1 and R_2 together constitute a reversible cycle. From Clausius' theorem

$$\oint_{R_1 R_2} \frac{đQ}{T} = 0$$

The above integral may be replaced as the sum of two integrals, one for path R_1 and the other for path R_2

$$\underset{R_1}{\int_i^f} \frac{đQ}{T} + \underset{R_2}{\int_f^i} \frac{đQ}{T} = 0$$

or

$$\underset{R_1}{\int_i^f} \frac{đQ}{T} = -\underset{R_2}{\int_f^i} \frac{đQ}{T}$$

Since R_2 is a reversible path

$$\underset{R_1}{\int_i^f} \frac{đQ}{T} = \underset{R_2}{\int_i^f} \frac{đQ}{T}$$

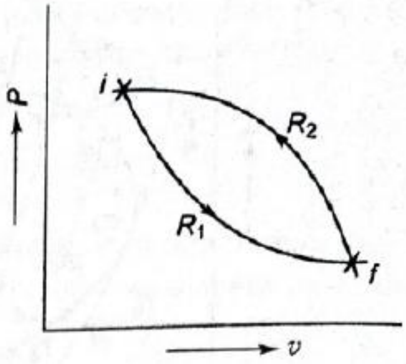


Fig. 7.4 *Two Reversible Paths R_1 and R_2 Between Two Equilibrium States i and f*

Since R_1 and R_2 represent any two reversible paths, $\int_{i\,R}^{f} \frac{đQ}{T}$ is independent of the reversible path connecting i and f. Therefore, there exists a property of a system whose value at the final state f minus its value at the initial state i is equal to $\int_{i\,R}^{f} \frac{đQ}{T}$. This property is called entropy, and is denoted by S. If S_i is the entropy at the initial state i, and S_f is the entropy at the final state f, then

$$\int_{i\,R}^{f} \frac{đQ}{T} = S_f - S_i \tag{7.4}$$

When the two equilibrium states are infinitesimally near

$$\frac{đQ_R}{T} = dS \tag{7.5}$$

where dS is an exact *differential* because S is a point function and a property. The subscript R in $đQ$ indicates that heat $đQ$ is transferred *reversibly*.

The word 'entropy' was first used by Clausius, taken from the Greek word 'tropee' meaning 'transformation'. It is an extensive property, and has the unit J/K. The specific entropy

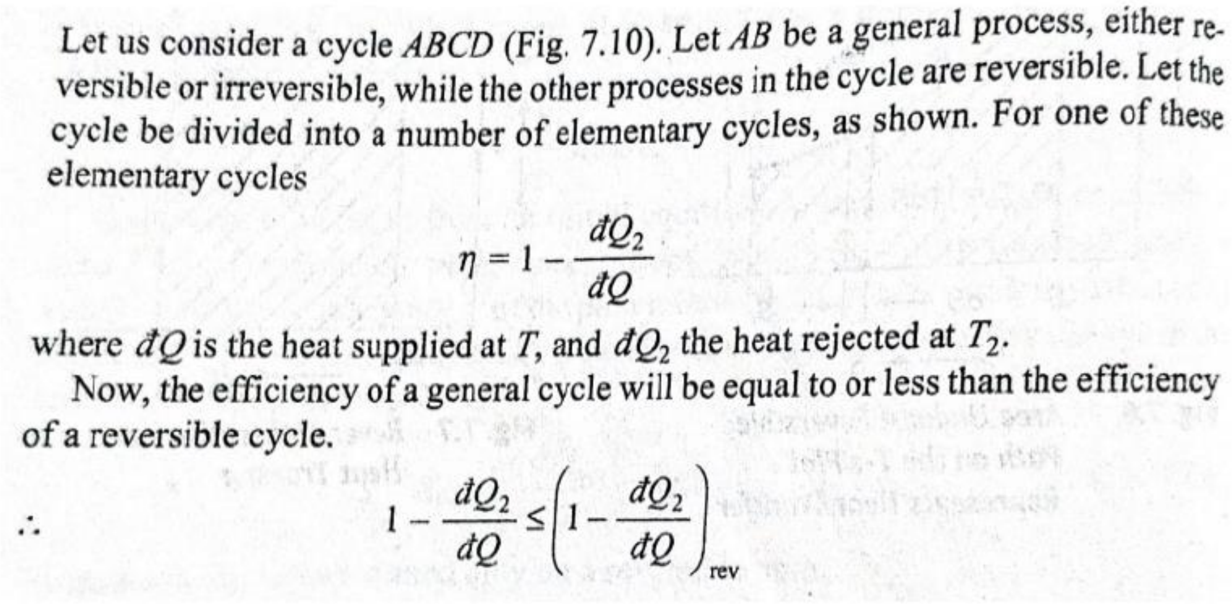
$$s = \frac{S}{m} \text{ J/kg K}$$

If the system is taken from an initial equilibrium state i to a final equilibrium state f by an *irreversible path*, since entropy is a point or state function, and the entropy change is independent of the path followed, the non-reversible path is to be replaced by a reversible path to integrate for the evaluation of entropy change in the irreversible process (Fig. 7.5).

$$S_f - S_i = \int_i^f \frac{đQ_{\text{rev}}}{T} = (\Delta S)_{\text{irrev path}} \tag{7.6}$$

Integration can be performed only on a reversible path.

CLAUSIUS INEQUALITY



Let us consider a cycle *ABCD* (Fig. 7.10). Let *AB* be a general process, either reversible or irreversible, while the other processes in the cycle are reversible. Let the cycle be divided into a number of elementary cycles, as shown. For one of these elementary cycles

$$\eta = 1 - \frac{\text{đ}Q_2}{\text{đ}Q}$$

where $\text{đ}Q$ is the heat supplied at T, and $\text{đ}Q_2$ the heat rejected at T_2.

Now, the efficiency of a general cycle will be equal to or less than the efficiency of a reversible cycle.

$$\therefore \qquad 1 - \frac{\text{đ}Q_2}{\text{đ}Q} \leq \left(1 - \frac{\text{đ}Q_2}{\text{đ}Q}\right)_{\text{rev}}$$

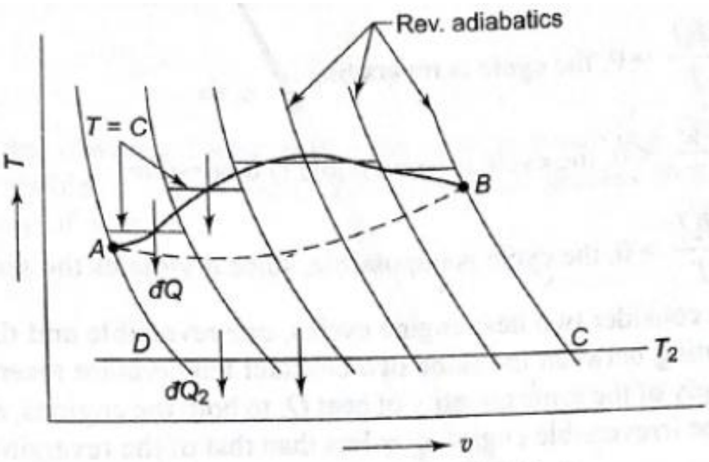


Fig. 7.10 *Inequality of Clausius*

or $$\frac{đQ_2}{đQ} \geq \left(\frac{đQ_2}{đQ}\right)_{\text{rev}}$$

or $$\frac{đQ}{đQ_2} \leq \left(\frac{đQ}{đQ_2}\right)_{\text{rev}}$$

Since $$\left(\frac{đQ}{đQ_2}\right)_{\text{rev}} = \frac{T}{T_2}$$

∴ $$\frac{đQ}{đQ_2} \leq \frac{T}{T_2}$$

or $\frac{đQ}{T} \leq \frac{đQ_2}{T_2}$, for any process AB, reversible or irreversible.

For a reversible process

$$ds = \frac{đQ_{\text{rev}}}{T} = \frac{đQ_2}{T_2} \tag{7.7}$$

Hence, for any process AB

$$\frac{đQ}{T} \leq ds \tag{7.8}$$

Then for any cycle

$$\oint \frac{đQ}{T} \leq \oint ds$$

Since entropy is a property and the cyclic integral of any property is zero

$$\oint \frac{đQ}{T} \leq 0 \tag{7.9}$$

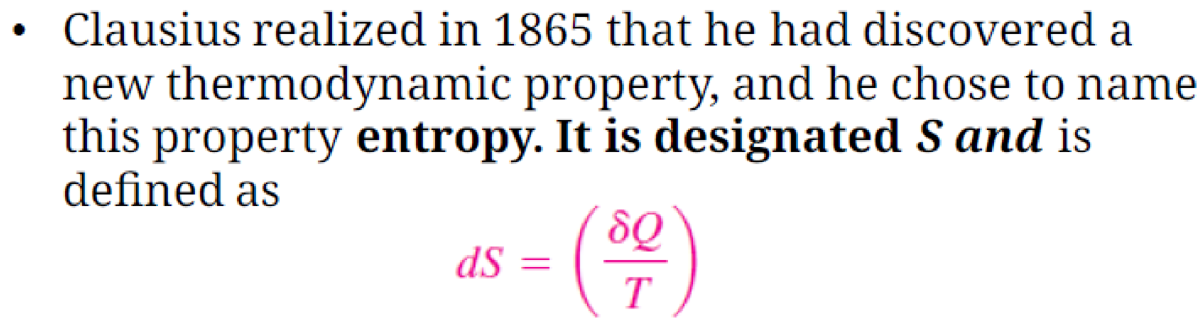
This equation is known as the *inequality of Clausius*. It *provides the criterion of the reversibility of a cycle*.

If $\oint \frac{đQ}{T} = 0$, the cycle is reversible,

$\oint \frac{đQ}{T} < 0$, the cycle is irreversible and possible

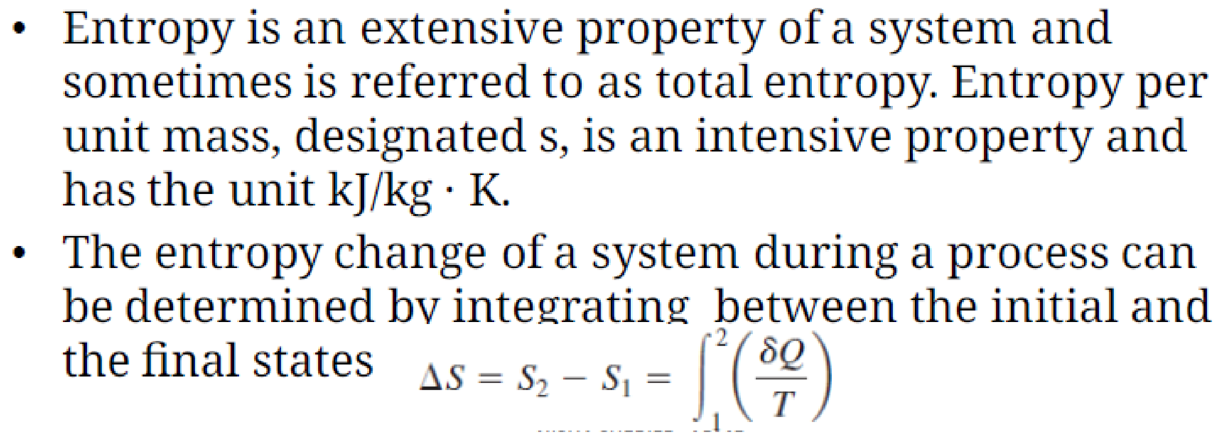
$\oint \frac{đQ}{T} > 0$, the cycle is impossible, since it violates the second law.

ENTROPY



- Clausius realized in 1865 that he had discovered a new thermodynamic property, and he chose to name this property **entropy. It is designated *S and*** is defined as

$$dS = \left(\frac{\delta Q}{T}\right)$$



- Entropy is an extensive property of a system and sometimes is referred to as total entropy. Entropy per unit mass, designated s, is an intensive property and has the unit kJ/kg · K.
- The entropy change of a system during a process can be determined by integrating between the initial and the final states $\Delta S = S_2 - S_1 = \int_1^2 \left(\frac{\delta Q}{T}\right)$

- Note that entropy is a property, and like all other properties, it has fixed values at fixed states.
- Therefore, the entropy change S between two specified states is the same no matter what path, reversible or irreversible, is followed during a process
- Also note that the integral of dQ/T gives us the value of entropy change only if the integration is carried out along an internally reversible path between the two states.
- The integral of dQ/T along an irreversible path is not a property, and in general, different values will be obtained when the integration is carried out along different irreversible paths.
- Therefore, even for irreversible processes, the entropy change should be determined by carrying out this integration along some convenient imaginary internally reversible path between the specified states.

CHARACTERISTICS OF ENTROPY

1. It increases when heat is supplied irrespective of the fact whether temperature changes or not.
2. It decrease when heat is removed whether temperature changes or not.
3. It remains unchanged in all adiabatic frictionless processes.
4. It increases if temperature of heat is lowered without work being done as in a throttling process.

ENTROPY OF THE UNIVERSE

$$(dS)_{universe} \geqq 0$$

$$\text{where } (dS)_{universe} = (dS)_{system} + (dS)_{surroundings}.$$

ENTROPY PRINCIPLE

$$\Delta S_{\text{isolated}} \geq 0$$

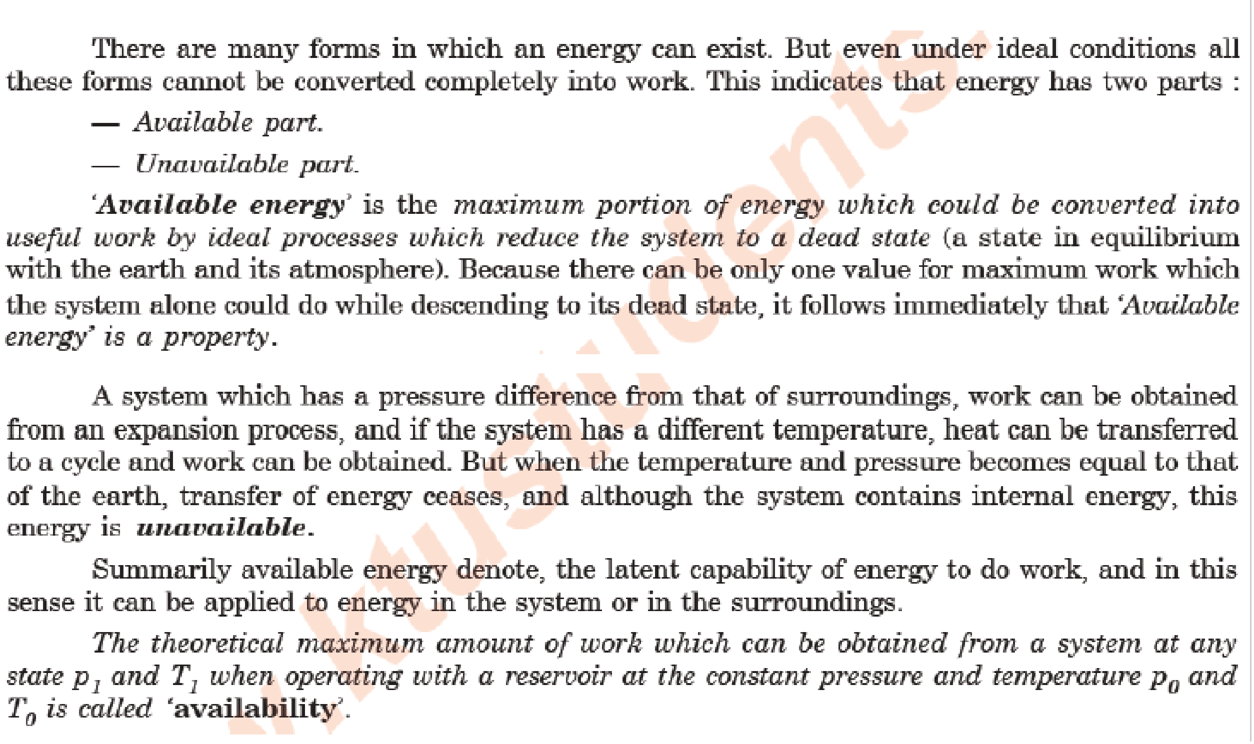
This equation can be expressed as the entropy of an isolated system during a process always increases or, in the limiting case of a reversible process, remains constant.

- In other words, it never decreases.
- This is known as the **increase of entropy principle.**
- **Note that in the absence of any heat transfer,** entropy change is due to irreversibilities only.

APPLICATIONS OF ENTOPY PRINCIPLE

1. Transfer of heat through a finite temperature difference
2. Mixing of two fluids
3. Maximum work obtainable from two finite bodies at temperatures T_1 and T_2
4. Maximum work obtainable from a finite body and thermal energy reservoir
5. Processes exhibiting external mechanical irreversibility

AVAILABLE AND UNAVAILABLE ENERGY



There are many forms in which an energy can exist. But even under ideal conditions all these forms cannot be converted completely into work. This indicates that energy has two parts :

— *Available part.*

— *Unavailable part.*

'***Available energy***' is the *maximum portion of energy which could be converted into useful work by ideal processes which reduce the system to a dead state* (a state in equilibrium with the earth and its atmosphere). Because there can be only one value for maximum work which the system alone could do while descending to its dead state, it follows immediately that '*Available energy*' *is a property*.

A system which has a pressure difference from that of surroundings, work can be obtained from an expansion process, and if the system has a different temperature, heat can be transferred to a cycle and work can be obtained. But when the temperature and pressure becomes equal to that of the earth, transfer of energy ceases, and although the system contains internal energy, this energy is ***unavailable***.

Summarily available energy denote, the latent capability of energy to do work, and in this sense it can be applied to energy in the system or in the surroundings.

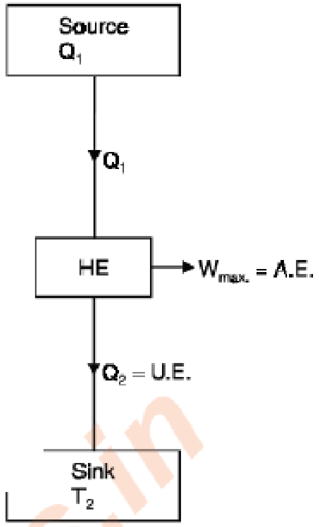
The theoretical maximum amount of work which can be obtained from a system at any state p_1 and T_1 when operating with a reservoir at the constant pressure and temperature p_0 and T_0 is called '**availability**'.

AVAILABLE ENERGY REFFERED TO A CYCLE

The ***available energy (A.E.)*** or the available part of the energy supplied is the maximum work output obtainable from a certain heat input in a cyclic heat engine (Fig. 6.1). The minimum energy that has to be rejected to the sink by the second law is called the ***unavailable energy (U.E.),*** or the unavailable part of the energy supplied.

$$Q_1 = \text{A.E.} + \text{U.E.}$$

$$W_{max} = \text{A.E.} = Q_1 - \text{U.E.}$$



Available and unavailable energy in a cycle.

For the given values of the source temperature T_1 and sink temperature T_2, the reversible efficiency,

$$\eta_{rev.} = 1 - \frac{T_2}{T_1}$$

For a given T_1, $\eta_{rev.}$ will increase with the decrease of T_2. *The lowest practicable temperature of heat rejection is the temperature of the surroundings,* T_0.

$$\therefore \qquad \eta_{max} = 1 - \frac{T_0}{T_1}$$

and

$$W_{max} = \left(1 - \frac{T_0}{T_1}\right) Q_1$$

Consider a finite process l-m, in which heat is supplied reversibly to a heat engine (Fig. 6.2). Taking an elementary cycle, if dQ_1 is the heat received by the engine reversibly at T_1,

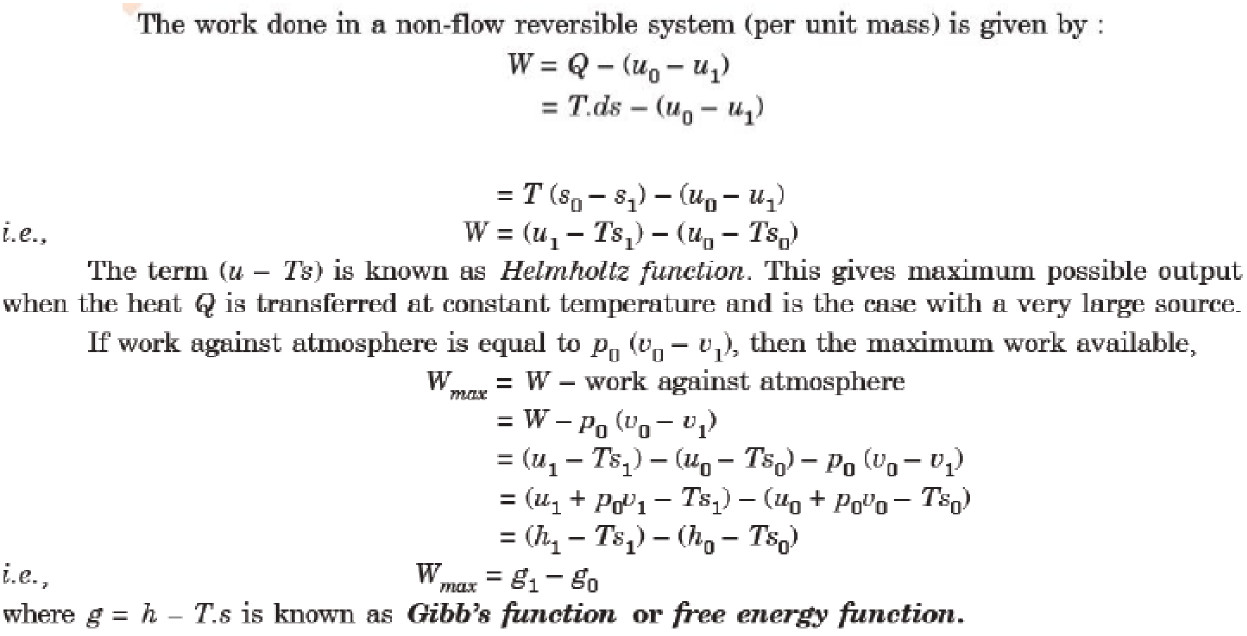
Then

$$dW_{max} = \frac{T_1 - T_0}{T_1} dQ_1$$

$$= dQ_1 - \frac{T_0}{T_1} dQ_1 = \text{A.E.}$$

Availability functions:
HELMHOLTZ AND GIBB'S FUNCTIONS

Availability functions:
HELMHOLTZ AND GIBB'S FUNCTIONS



The work done in a non-flow reversible system (per unit mass) is given by :

$$W = Q - (u_0 - u_1)$$
$$= T.ds - (u_0 - u_1)$$
$$= T\,(s_0 - s_1) - (u_0 - u_1)$$

i.e.,
$$W = (u_1 - Ts_1) - (u_0 - Ts_0)$$

The term $(u - Ts)$ is known as *Helmholtz function*. This gives maximum possible output when the heat Q is transferred at constant temperature and is the case with a very large source.

If work against atmosphere is equal to $p_0\,(v_0 - v_1)$, then the maximum work available,

$$W_{max} = W - \text{work against atmosphere}$$
$$= W - p_0\,(v_0 - v_1)$$
$$= (u_1 - Ts_1) - (u_0 - Ts_0) - p_0\,(v_0 - v_1)$$
$$= (u_1 + p_0v_1 - Ts_1) - (u_0 + p_0v_0 - Ts_0)$$
$$= (h_1 - Ts_1) - (h_0 - Ts_0)$$

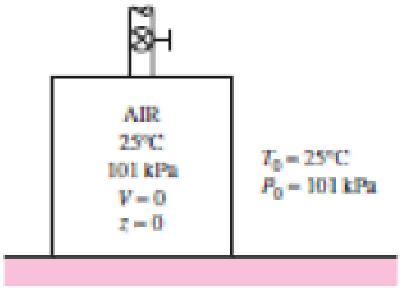
i.e.,
$$W_{max} = g_1 - g_0$$

where $g = h - T.s$ is known as ***Gibb's function*** **or** ***free energy function.***

Dead State

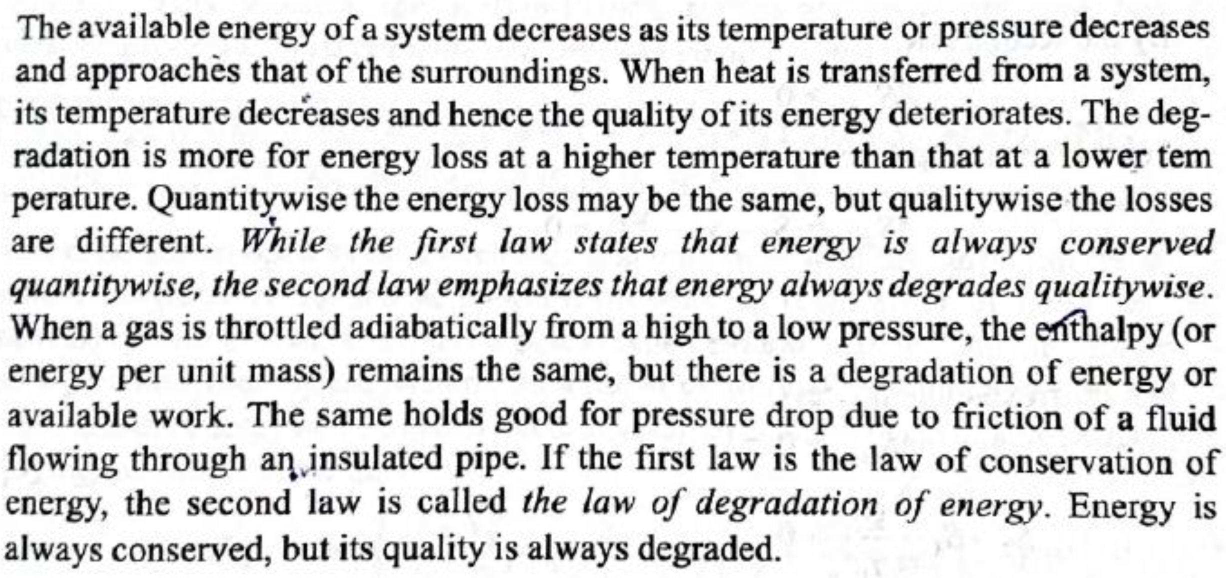
- A system is said to be in the **dead state when it is in thermodynamic equilibrium** with the environment it is in
- At the dead state, a system is at the temperature and pressure of its environment (in thermal and mechanical equilibrium); it has no kinetic or potential energy relative to the environment (zero velocity and zero elevation above a reference level); and it does not react with the environment (chemically inert).

- No work can be produced from a system that is initially at the dead state.
- A system delivers the maximum possible work as it undergoes a reversible process from the specified initial state to the state of its environment, that is, the dead state.



A system that is in equilibrium with its environment is said to be at the dead state.

LAW OF DEGRADATION OF ENERGY

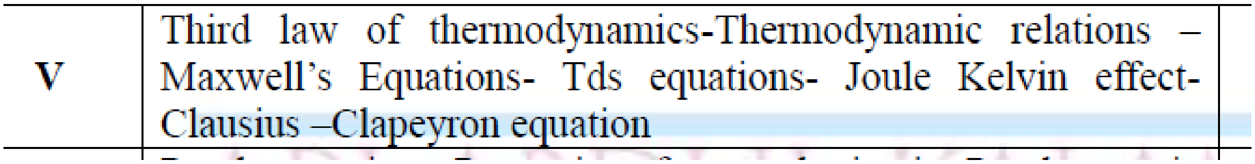

The available energy of a system decreases as its temperature or pressure decreases and approaches that of the surroundings. When heat is transferred from a system, its temperature decreases and hence the quality of its energy deteriorates. The degradation is more for energy loss at a higher temperature than that at a lower temperature. Quantitywise the energy loss may be the same, but qualitywise the losses are different. *While the first law states that energy is always conserved quantitywise, the second law emphasizes that energy always degrades qualitywise.* When a gas is throttled adiabatically from a high to a low pressure, the enthalpy (or energy per unit mass) remains the same, but there is a degradation of energy or available work. The same holds good for pressure drop due to friction of a fluid flowing through an insulated pipe. If the first law is the law of conservation of energy, the second law is called *the law of degradation of energy*. Energy is always conserved, but its quality is always degraded.

SECOND LAW EFFICIENCY

- It is defined as the ratio of minimum exergy intake to perform the given task to actual exergy intake to perform the same task

MODULE – 5

ENTROPY



V	Third law of thermodynamics-Thermodynamic relations – Maxwell’s Equations- Tds equations- Joule Kelvin effect- Clausius –Clapeyron equation

THIRD LAW OF THERMODYNAMICS

- The molecules of a substance in solid phase continually oscillate, creating an uncertainty about their position.
- These oscillations, however, fade as the temperature is decreased, and the molecules supposedly become motionless at absolute zero.
- This represents a state of ultimate molecular order (and minimum energy).
- Therefore, *the entropy of a pure crystalline substance at absolute zero temperature is zero since there is* no uncertainty about the state of the molecules at that instant.

The third law of thermodynamics provides an absolute reference point for the determination of entropy.

- The entropy determined relative to this point is called absolute entropy, and it is extremely useful in the thermodynamic analysis of chemical reactions.
- Notice that the entropy of a substance that is not pure crystalline (such as a solid solution) is not zero at absolute zero temperature.
- This is because more than one molecular configuration exists for such substances, which introduces some uncertainty about the microscopic state of the substance.

A LITTLE MATH—PARTIAL DERIVATIVES AND ASSOCIATED RELATIONS

$$dz = \left(\frac{\partial z}{\partial x}\right)_y dx + \left(\frac{\partial z}{\partial y}\right)_x dy$$

$$dz = M\,dx + N\,dy$$

where

$$M = \left(\frac{\partial z}{\partial x}\right)_y \quad \text{and} \quad N = \left(\frac{\partial z}{\partial y}\right)_x$$

- Taking the partial derivative of *M with respect to y and of N with respect to x yields*

$$\left(\frac{\partial M}{\partial y}\right)_x = \frac{\partial^2 z}{\partial x\,\partial y} \quad \text{and} \quad \left(\frac{\partial N}{\partial x}\right)_y = \frac{\partial^2 z}{\partial y\,\partial x}$$

$$\left(\frac{\partial M}{\partial y}\right)_x = \left(\frac{\partial N}{\partial x}\right)_y$$

THE MAXWELL RELATIONS

- The equations that relate the partial derivatives of properties P, v, T, and s of a simple compressible system to each other are called the Maxwell relations.

$$du = T\,ds - P\,dv$$

$$dh = T\,ds + v\,dP$$

- The other two Gibbs relations are based on two new combination properties— the Helmholtz function a and the Gibbs function g, defined as

$$a = u - Ts$$

$$g = h - Ts$$

Differentiating, we get

$$da = du - T\,ds - s\,dT$$

$$dg = dh - T\,ds - s\,dT$$

$$da = -s\,dT - P\,dv$$

$$dg = -s\,dT + v\,dP$$

- A careful examination of the four Gibbs relations reveals that they are of the form

$$dz = M\,dx + N\,dy$$

$$\left(\frac{\partial M}{\partial y}\right)_x = \left(\frac{\partial N}{\partial x}\right)_y$$

- since *u, h, a, and g are properties and thus have exact differentials.*

$$\left(\frac{\partial T}{\partial v}\right)_s = -\left(\frac{\partial P}{\partial s}\right)_v$$

$$\left(\frac{\partial T}{\partial P}\right)_s = \left(\frac{\partial v}{\partial s}\right)_P$$

$$\left(\frac{\partial s}{\partial v}\right)_T = \left(\frac{\partial P}{\partial T}\right)_v$$

$$\left(\frac{\partial s}{\partial P}\right)_T = -\left(\frac{\partial v}{\partial T}\right)_P$$

- These are called the **Maxwell relations**
- They are extremely valuable in thermodynamics because they provide a means of determining the change in entropy, which cannot be measured directly, by simply measuring the changes in properties *P, v, and T.*
- *Note that the Maxwell relations* given above are limited to simple compressible systems

Tds Equations

$$s = f(T, v)$$

i.e.,
$$ds = \left(\frac{\partial s}{\partial T}\right)_v dT + \left(\frac{\partial s}{\partial v}\right)_T dv$$

or
$$Tds = T\left(\frac{\partial s}{\partial T}\right)_v dT + T\left(\frac{\partial s}{\partial v}\right)_T dv$$

But for a reversible constant volume change

$$dq = c_v\,(dT)_v = T(ds)_v$$

or
$$c_v = T\left(\frac{\partial s}{\partial T}\right)_v$$

But,
$$\left(\frac{\partial s}{\partial v}\right)_T = \left(\frac{\partial p}{\partial T}\right)_v$$
Third maxwell's equation

$$\boxed{Tds = c_v dT + T\left(\frac{\partial p}{\partial T}\right)_v dv}$$

This is known as the *first form of entropy equation* or the **first Tds equation.**

$$s = f(T, p)$$

$$Tds = T\left(\frac{\partial s}{\partial T}\right)_p dT + T\left(\frac{\partial s}{\partial p}\right)_T dp$$

where $c_p = T\left(\frac{\partial s}{\partial T}\right)_p$

Also
$$\left(\frac{\partial s}{\partial p}\right)_T = -\left(\frac{\partial v}{\partial T}\right)_p$$
[Maxwell's eqn.

whence, substituting in eqn. (7.25)

$$\boxed{Tds = c_p dT - T\left(\frac{\partial v}{\partial T}\right)_p dp}$$

This is known as the *second form of entropy equation* or the **second Tds equation.**

The Clapeyron Equation

Clausius-Claperyon equation is a relationship between the *saturation pressure*, *temperature*, the *enthalpy of evaporation*, and the *specific volume* of the *two phases* involved. This equation provides a basis for calculations of properties in a two-phase region. It gives the *slope* of a curve separating the two phases in the p-T diagram.

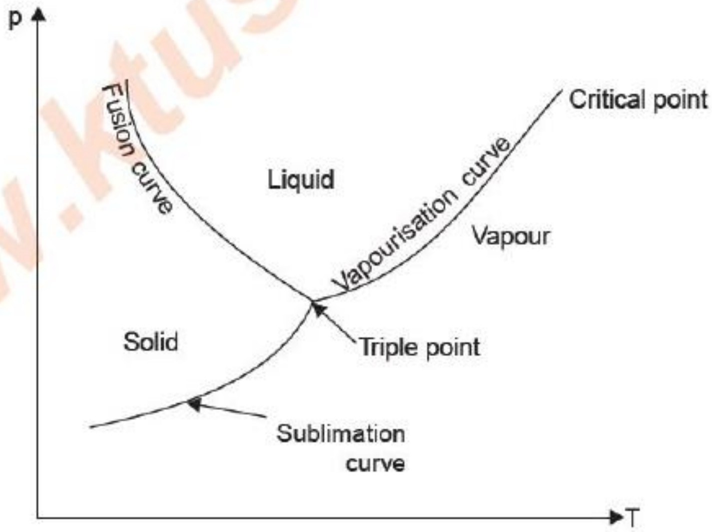
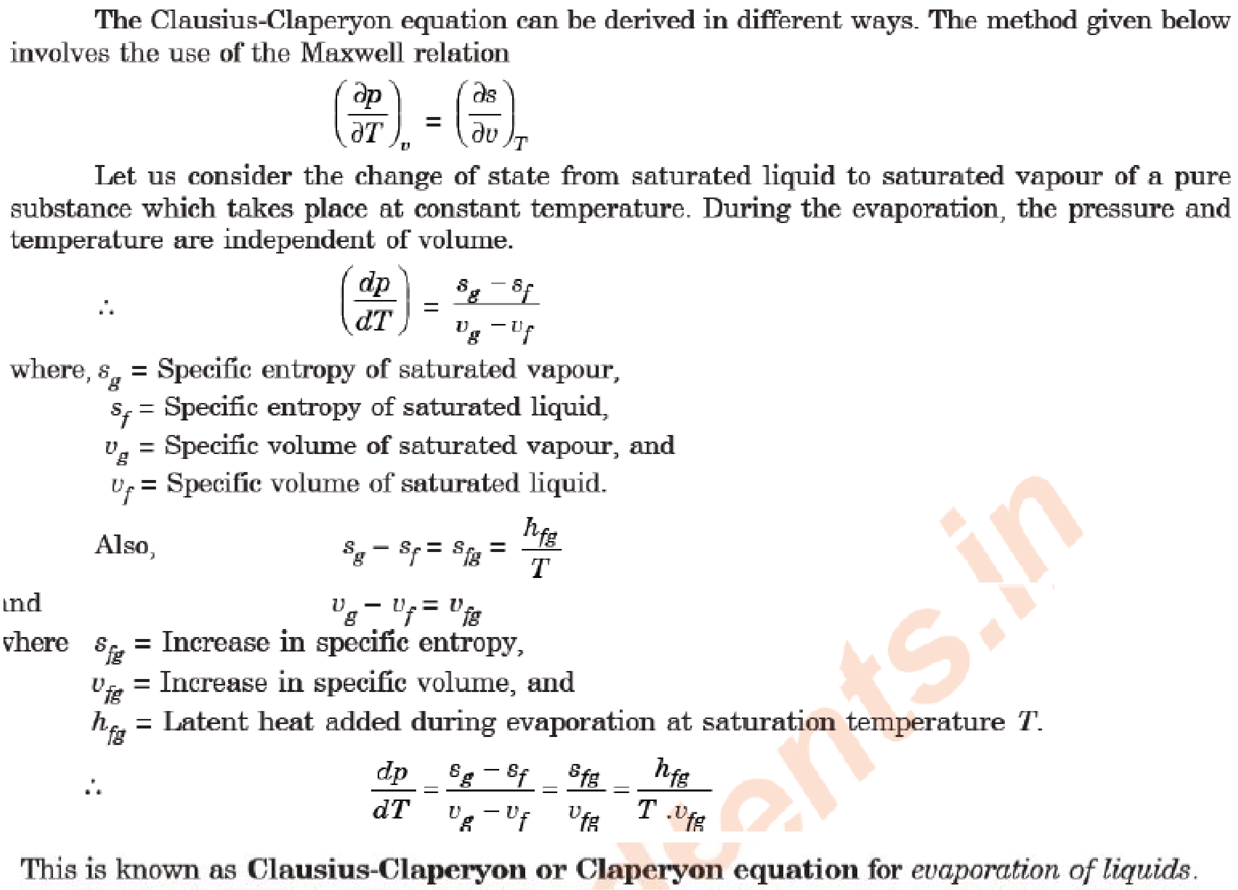


Fig. 7.4. p-T diagram.



The Clausius-Claperyon equation can be derived in different ways. The method given below involves the use of the Maxwell relation

$$\left(\frac{\partial p}{\partial T}\right)_v = \left(\frac{\partial s}{\partial v}\right)_T$$

Let us consider the change of state from saturated liquid to saturated vapour of a pure substance which takes place at constant temperature. During the evaporation, the pressure and temperature are independent of volume.

$$\therefore \qquad \left(\frac{dp}{dT}\right) = \frac{s_g - s_f}{v_g - v_f}$$

where, s_g = Specific entropy of saturated vapour,

s_f = Specific entropy of saturated liquid,

v_g = Specific volume of saturated vapour, and

v_f = Specific volume of saturated liquid.

Also, $$s_g - s_f = s_{fg} = \frac{h_{fg}}{T}$$

nd $$v_g - v_f = v_{fg}$$

here s_{fg} = Increase in specific entropy,

v_{fg} = Increase in specific volume, and

h_{fg} = Latent heat added during evaporation at saturation temperature T.

$$\therefore \qquad \frac{dp}{dT} = \frac{s_g - s_f}{v_g - v_f} = \frac{s_{fg}}{v_{fg}} = \frac{h_{fg}}{T\,.v_{fg}}$$

This is known as **Clausius-Claperyon or Claperyon equation** for *evaporation of liquids*.

JOULE –KELVIN EFFECT (JOULE–THOMSON EFFECT)

In thermodynamics, the Joule–Thomson effect (also known as the Joule–Kelvin effect or Kelvin–Joule effect) describes the temperature change of a *real* gas or liquid (as differentiated from an ideal gas) when it is forced through a valve or porous plug while keeping it insulated so that no heat is exchanged with the environment.This procedure is called a *throttlingprocess* or *Joule–Thomson process*

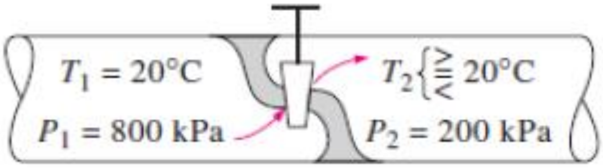
THE JOULE-THOMSON COEFFICIENT

When a fluid passes through a restriction such as a porous plug, a capillary tube, or an ordinary valve, its pressure decreases.

- The enthalpy of the fluid remains approximately constant during such a throttling process.

• The temperature of the fluid may remain unchanged, or it may even increase during a throttling process

$$\mu = \left(\frac{\partial T}{\partial P}\right)_h$$



The temperature of a fluid may increase, decrease, or remain constant during a throttling process.

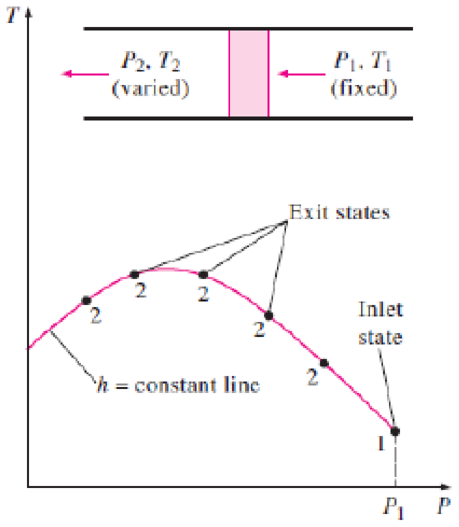
$$\mu_{JT} \begin{cases} < 0 & \text{temperature increases} \\ = 0 & \text{temperature remains constant} \\ > 0 & \text{temperature decreases} \end{cases}$$

Joule-Thomson coefficient represents the slope of *h =constant lines on a T-P diagram.*

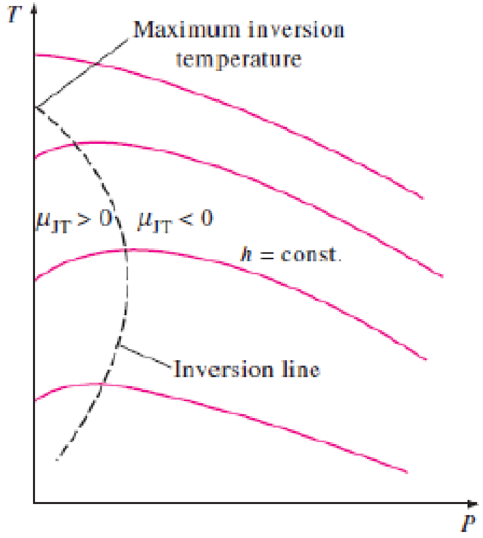
• Such diagrams can be easily constructed from temperature and pressure measurements alone during throttling processes.

• A fluid at a fixed temperature and pressure *T1 and P1 (thus fixed enthalpy) is forced to flow through a* porous plug, and its temperature and pressure downstream (*T2 and P2) are* measured.

Repeating the experiment for different sets of inlet pressure and temperature and plotting the results, we can construct a T-P diagram for a substance with severalh = constant lines



The development of an h = constant line on a P-T diagram.



Constant-enthalpy lines of a substance on a T-P diagram.

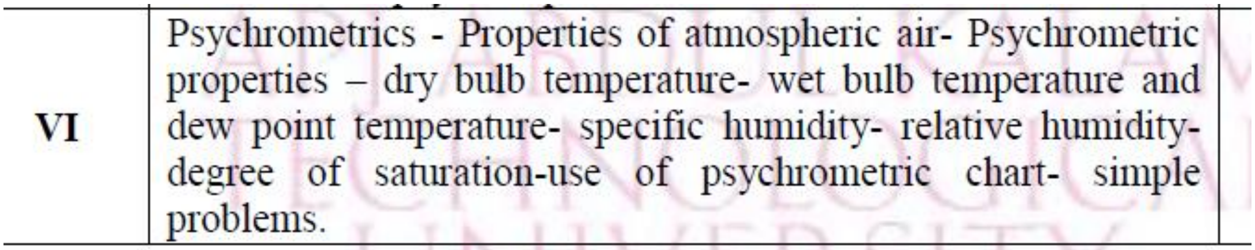
Some constant-enthalpy lines on the *T-P diagram past hrough a point of* zero slope or zero Joule-Thomson coefficient.

- The line that passes through these points is called the **inversion line,** and the temperature at a point where aconstant-enthalpy line intersects the inversion line is called the **inversion temperature.**
- The temperature at the intersection of the $P = 0$ line (ordinate) and the upper part of the inversion line is called the **maximum inversion temperature.**

Notice that the slopes of the *h = constant* lines are negative ($\mu_{JT} < 0$) at states to the right of the inversion line and positive ($\mu_{JT} > 0$) to the left of the inversion line.

- A throttling process proceeds along a constant-enthalpy line in the direction of decreasing pressure, that is, from right to left.

MODULE – 6
PSYCHROMETRICS



VI	Psychrometrics - Properties of atmospheric air- Psychrometric properties – dry bulb temperature- wet bulb temperature and dew point temperature- specific humidity- relative humidity- degree of saturation-use of psychrometric chart- simple problems.

PSYCHROMETRICS

- Air consists of fixed gases principally, nitrogen and oxygen with an admixture of water vapour in varying amounts.
- Water is always present in atmospheric air (1% to 3%) by weight.
- The art of measuring the moisture content of air is termed as psychrometry.
- The science which investigates the thermal properties of moist air, considers the measurement and control of the moisture content of air, and studies the effect of atmospheric moisture on material and human comfort may properly be termed "psychrometrics".

PSYCHROMETRIC PROPERTIES

- **Dry air:** water free contents of air, no moisture.

Dry air never found in practice.

- **Moist air (air):** contain some amount of moisture.
- **Saturated air:** if the water vapour is added to dry air, a limit will be reached when the air is saturated. Beyond this limit the cannot hold any more water vapour.
- **Dry bulb temperature (DBT):** It is the temperature of air as registered by an ordinary thermometer (t_{db}).-when the air is not affected by the moisture present in the air.
- **Wet bulb temperature (WBT):** It is the temperature of air as registered by a thermometer when the bulb is covered by a wetted wick and is exposed to a current of rapidly moving air (t_{wb}).

Adiabatic saturation temperature: It is the temperature at which the water or ice can saturate air by evaporating adiabatically into it. It is numerically equivalent to the measures wet bulb temperature (twb).

- **Wet bulb depression:** It is the difference between dry-bulb and wet-bulb temperature. (tdb- twb).

• **Dew point temperature (DPT):** It is the temperature to which air must be cooled at constant pressure in order to cause condensation of any of its water vapour. It is equal to steam table saturation temperature corresponding to the actual partial pressure of water in the air (t_{dp}).

Dew point depression: It is the difference between dry-bulb and dew point temperature. (tdb- tdp).

• **Specific humidity (Humidity ratio or moisture content):** It is the ratio of the mass of water vapour per unit mass of dry air in the mixture of vapour and air. It is generally expressed as grams of water per kg of dry air. For a given barometric pressure it is a function of dew point temperature alone.

• **Relative humidity (RH) [φ]:** It is the ratio of the partial pressure of water vapour in the mixture to the saturated partial pressure at the dry bulb temperature, expressed as percentage .

• **Sensible heat:** It is the heat that changes the temperature of a substance when added to or abstracted from it.

• **Latent heat:** It is the heat that does not affect the temperature but changes the state of a substance when added to or abstracted from it.

• **Enthalpy:** It is the combination energy which represents the sum of internal and flow energy in a steady flow process. It is determined from an arbitrary datum point for the air.

Degree of saturation (Percentage humidity) μ: ratio of actual mass of water vapour in a unit mass of dry air to the mass of water vapour in the same mass of dry air when it is saturated at the same temperature (dry bulb temperature).

• **Dry air:** molecular mass of dry air is taken as 28, and the gas constant of air (Ra) is equal to 0.287kJ/kg K or 287J/kgK.

• Density of dry air-1.293 kg/m3 at 1.0135 bar and at temperature 0°C.

• Molecular mass of water vapour is taken as 18, and the gas constant for water vapour (Rv) is equal to 0.461kJ/kgK.

• **Unsaturated air:** The water vapour in the air exists in the superheated sate and the moist air containing moisture in such a form (superheated state) is said to be unsaturated air.

PSYCHROMETERIC PROCESSES

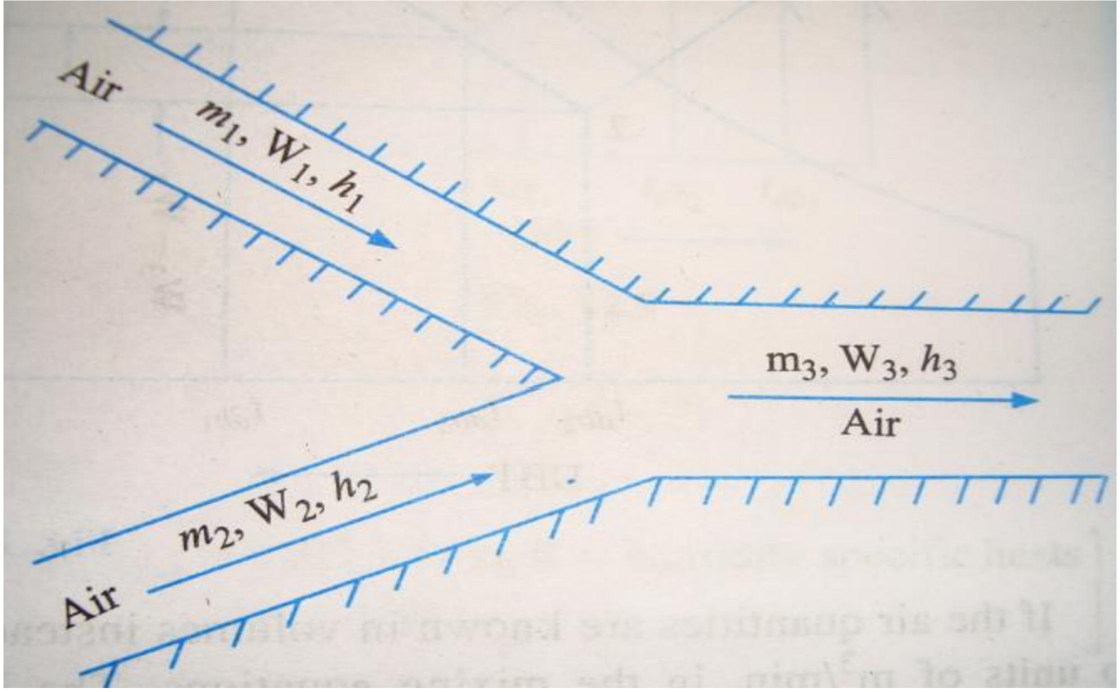
• Inorder to condition air to the conditions of human comfort or of the optimum control of an industrial processes required, certain processes are to be carried out on the outside air available.

• **The processes effecting the psychrometric properties of air are called psychrometric processes.**

• These processes involve mixing of air streams, heating, cooling, humidifying, dehumidifying, adiabatic saturation and mostly the combination of these.
– Mixing of air streams
– Sensible heating
– Sensible cooling
– Cooling and dehumidification
– Cooling and humidification
– Heating and dehumidification
– Heating and humidification

MIXING OF AIR STREAMS

• Frequently used in air-conditioning.
• Mixing normally takes place without the addition or rejection of either heat or moisture.
• Adiabatically and at constant total moisture content.



SENSIBLE HEATING

• When air passes over a dry surface which is at a temperature greater than its (air) dry bulb temperature, it undergoes **sensible heating**.
• Passing air over heating coil like electric resistance heating coils or steam coils.
• Specific humidity remains constant but the dry bulb

temperature rises and approaches that of the surface. The extent to which it approaches at mean effective surface temperature of the coil conveniently expressed in terms of the equivalent “**by-pass factor**”.

- **Bypass factor (BF*)*:** defined as the ratio of the difference between the mean surface temperature of the coil and leaving air temperature to the difference between the mean surface and the entering air temperature.
- The value of the by-pass factor is a function of the coil design and velocity.
- In complete air-conditioning system the preheating and reheating of air are among the familiar examples of sensible heating.

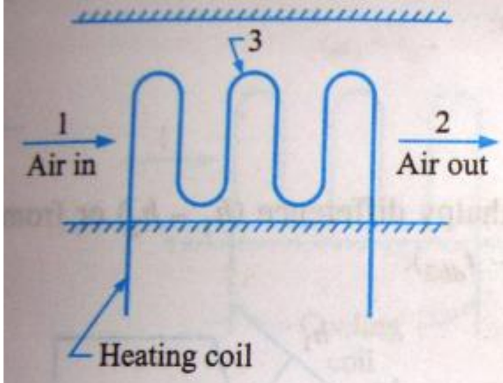


Fig. 12.8. Sensible heating.

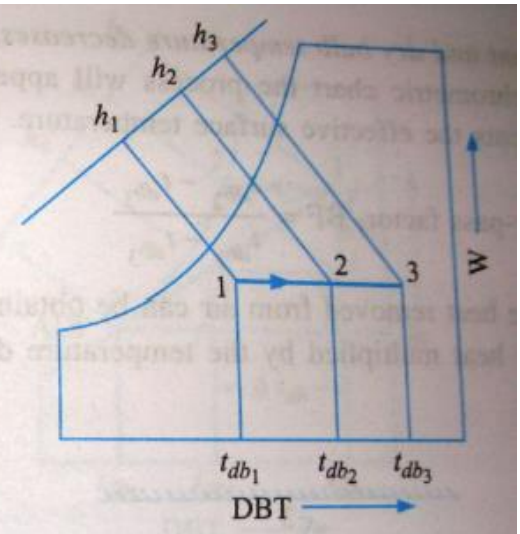


Fig. 12.9

BYPASS FACTOR

- The by-pass process of air is measured in terms of a by-pass factor.
- The amount of air that by-passes or the by-pass factor depends upon thefollowing factors:

– The number of fins provided in a unit length i.e. the pitch of the cooling coil fins.
– The number of rows in a coil in the direction of flow.
– The velocity of flow of air.

- By-pass factor of a cooling coil decreases with decrease in fin spacing and increase in number of rows.
- The performance of a heating or cooling coil is measured in terms of bypass factor. A coil with low by-pass factor has better performance

SENSIBLE COOLING

- Air undergoes sensible cooling whenever it passes over a surface that is at a temperature less than the dry bulb temperature of the air but greater than the dew point temperature.
- It is achieved by passing the air over cooling coil like evaporating coil of the refrigeration cycle or secondary brine coil.
- Specific humidity remains constant and dry bulb temperature decreases, approaching the mean effective surface temperature.
- For sensible cooling, the cooling coil may have refrigerant, cooling water or cool gas flowing through it.
- Sensible cooling can be done only up to the dew point temperature (t_{dp}). The cooling below this temperature will result in the condensation of moisture

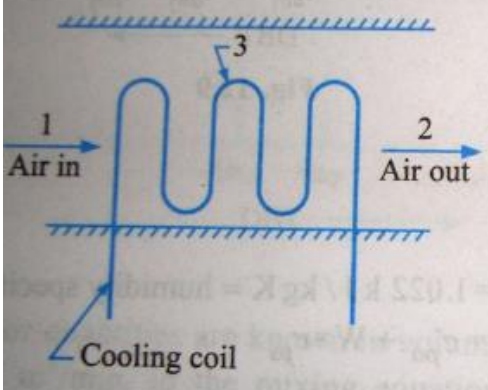


Fig. 12.10. Sensible cooling.

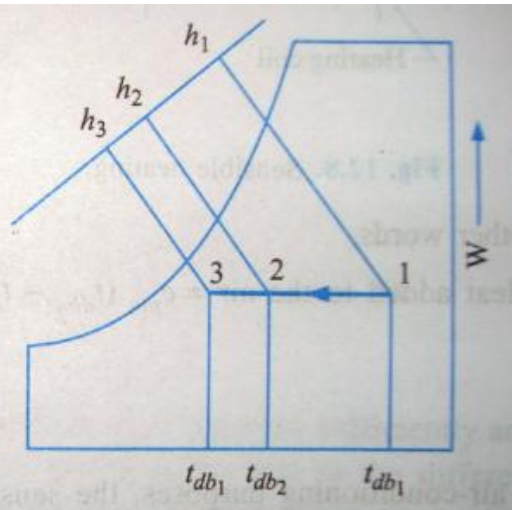


Fig. 12.11

COOLING AND DEHUMIDIFICATION

- Whenever air is made to pass over a surface or through a spray of water that is at a temperature less than the dew point temperature of the air, condensation of some of the water vapour in air will occur simultaneously with the sensible cooling process.

- **Apparatus Dew Point (ADP) :** For process involving condensation, the effective surface temperature, (t_{db3}) is called **Apparatus Dew Point (ADP).**
- **Sensible heat factor (SHF)** or **Sensible heat ratio (SHR):**

– SHF=[(SH)/(SH+LH)]
– SH= Sensible heat removed

– LH= Latent heat removed

– Total Heat (TH) = SH+LH

- Sensible heat factor is shown on the right hand side of the psychrometric chart.
- The dehumidification of air is only possible when the effective surface temperature of the cooling coil (td4) is less than the dew point temperature of the air entering the coil (t_{dp1}) . The effective surface temperature of the coil is known as apparatus dew point(ADP).

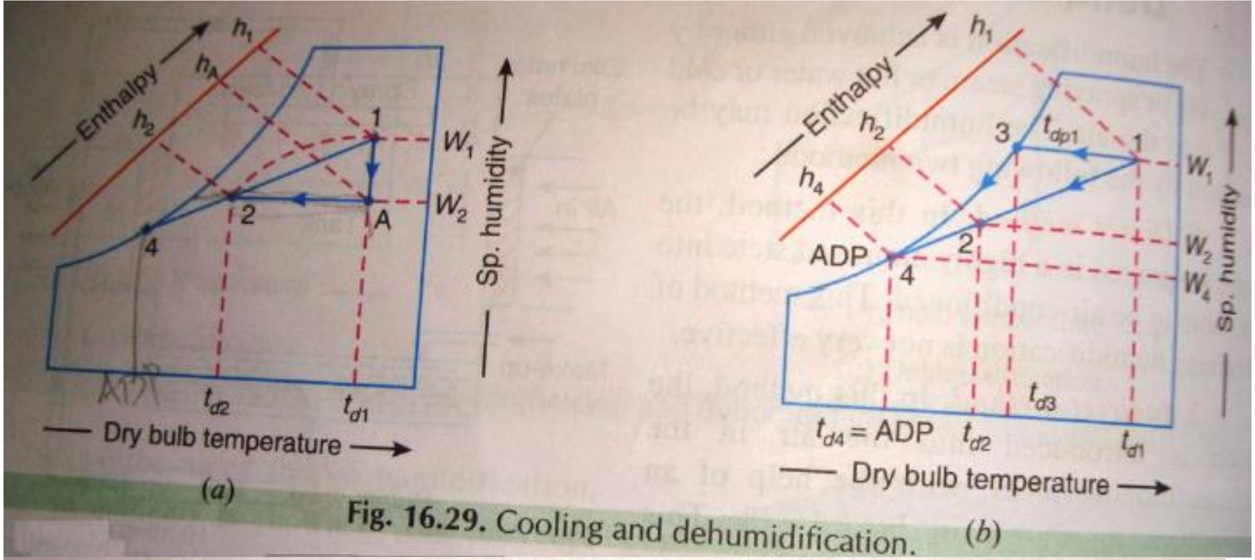
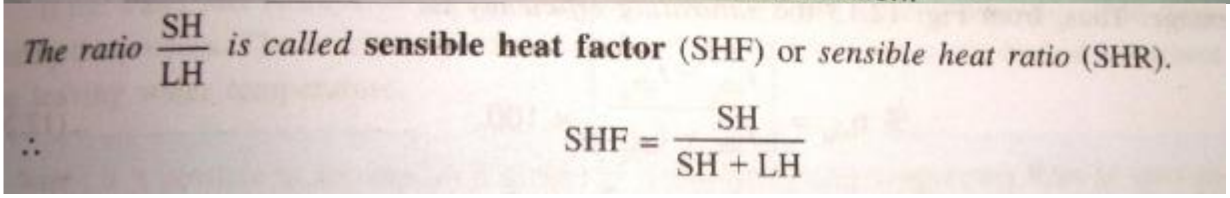


Fig. 16.29. Cooling and dehumidification.



The ratio $\frac{\text{SH}}{\text{LH}}$ *is called* **sensible heat factor** (SHF) or *sensible heat ratio* (SHR).

$$\therefore \quad \text{SHF} = \frac{\text{SH}}{\text{SH} + \text{LH}}$$

HUMIDIFICATION AND DEHUMIDIFICATION

- The addition of moisture to the air, without change in its dry bulbtemperature, is known as **humidification.**
- Removal of moisture from the air, without change in its dry bulb temperature, is known as **dehumidification.**
- Cooling coil
- Cold water spray

APPENDIX 1

CONTENT BEYOND THE SYLLABUS

AIR CONDITIONING

Air conditioning is the process of removing heat and moisture from the interior of an occupied space to improve the comfort of occupants. Air conditioning can be used in both domestic and commercial environments. This process is most commonly used to achieve a more comfortable interior environment, typically for humans and other animals; however, air conditioning is also used to cool and dehumidify rooms filled with heat-producing electronic devices, such as computer servers, power amplifiers, and to display and store some delicate products, such as artwork.

Air conditioners often use a fan to distribute the conditioned air to an enclosed space such as a building or a car to improve thermal comfort and indoor air quality. Electric refrigerant-based AC units range from small units that can cool a small bedroom, which can be carried by a single adult, to massive units installed on the roof of office towers that can cool an entire building. The cooling is typically achieved through a refrigeration cycle, but sometimes evaporation or free cooling is used. Air conditioning systems can also be made based on desiccants (chemicals that remove moisture from the air). Some AC systems reject or store heat in subterranean pipes.

HEATING, VENTILATION, AND AIR CONDITIONING

Heating, ventilation, and air conditioning (HVAC) is the technology of indoor and vehicular environmental comfort. Its goal is to provide thermal comfort and acceptable indoor air quality. HVAC system design is a subdiscipline of mechanical engineering, based on the principles of thermodynamics, fluid mechanics and heat transfer. "Refrigeration" is sometimes added to the field's abbreviation, as HVAC&R or HVACR or "ventilation" is dropped, as in HACR (as in the designation of HACR-rated circuit breakers).

HVAC is an important part of residential structures such as single family homes, apartment buildings, hotels and senior living facilities, medium to large industrial and office buildings such as skyscrapers and hospitals, vehicles such as cars, trains, airplanes, ships and submarines, and in marine environments, where safe and healthy building conditions are regulated with respect to temperature and humidity, using fresh air from outdoors.

Ventilating or ventilation (the "V" in HVAC) is the process of exchanging or replacing air in any space to provide high indoor air quality which involves temperature control, oxygen replenishment, and removal of moisture, odors, smoke, heat, dust, airborne bacteria, carbon dioxide, and other gases. Ventilation removes unpleasant smells and excessive moisture, introduces outside air, keeps interior building air circulating, and prevents stagnation of the interior air.