Course Title:

Fundamental of Thermodynamics and Heat Transfer

Chapter 1:

Introduction

Lecture 2:

Basic Concepts of Thermodynamics

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Learning Objective of Lecture:

To impart a great deal of knowledge to undergraduate students on the topics:

- ✓ Thermodynamic State , Path, Process and Cycle
- ✓ State or Point Function and Path Function
- ✓ Thermodynamic Processes: Quasi – Equilibrium Process, Reversible and Irreversible Process
- ✓ Some Common Properties: Volume, Pressure, Temperature
- ✓ Equality of Temperature
- ✓ Zeroth Law of Thermodynamics

1.6. Thermodynamic State, Path, Process and Cycle

1.6.1. Thermodynamic State

Each unique condition of a system at which the system is in thermodynamic equilibrium is known as a *thermodynamic state* of the system. Thus, a thermodynamic state is the condition of a system at an instant of time which can be identified by its thermodynamic properties such as temperature, pressure, volume etc. The number of thermodynamic properties which are required to specify a thermodynamic state of a system depends upon the nature of the system. Thermodynamic state of a system composed of a pure substance can be specified by a set of two independent thermodynamic properties. For example, for a system constituted by gas enclosed in the piston cylinder assembly, the thermodynamic state of the system will be prescribed by any two independent properties such as pressure, temperature and volume of the gas corresponding to position of the piston at any instant. With properties as co-ordinates, the thermodynamic state of the system in thermodynamic equilibrium can be represented by a point such as 1 (initial state) or 2 (final state) as shown in figure 1.5 below.

On the basis of the definition of thermodynamic state, the salient features of a thermodynamic property are:

- it has a single value at each equilibrium state, and
- the change in its value between two prescribed equilibrium states is independent of the path or route the system follows.

Any thermodynamic system cannot have multiple values of any thermodynamic property like temperature, pressure, volume etc at a thermodynamic state when it is in equilibrium condition. Whereas quantities such as heat and work cannot have single value for any equilibrium state because heat and work can only be defined for an interval between two prescribed equilibrium states. Hence pressure, temperature, volume etc satisfy the first feature but heat and work do not satisfy it.

Similarly, let us consider three paths A, B and C traced by a system from initial state 1 to final state 2 as shown in Figure 1.5. When we observe the change in pressure between two end states 1 and 2 by taking reference of any one path A or B or C, then we find $(P_2 - P_1)_A = (P_2 - P_1)_B = (P_2 - P_1)_C$. Hence, any variable having such relation is said to be path independent. If we consider work as a variable, it will have different values for three different paths A, B, C between same initial and final states 1 and 2, e.g. $W_A > W_B > W_C$. This is because work is determined by calculating area covered by each path A, B, C on P-V diagram (as $W = \int_1^2 PdV$). Hence, work is said to be path dependent. Thus, pressure which satisfies both features is a thermodynamic property, but work which does not satisfy both features is not a thermodynamic property.

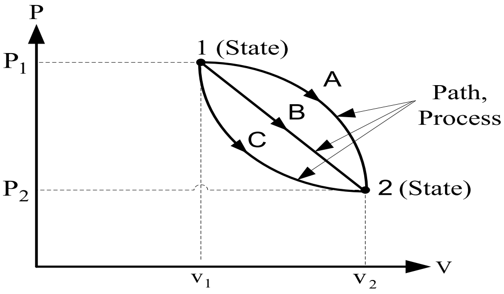


Fig. 1.5. Thermodynamic state, path and process

1.6.2. State or Point Function

Any variable or quantity which depends only on the end states but not on the path is known as a *state or point function*. In other words, if two properties locate a point on the property diagram, then those properties are called point functions. All thermodynamic properties such as temperature, pressure, volume etc are the examples of state functions and are exact differentials because their change can be written as difference between their end states. Thus, for example the change in volume between the states 1 and 2 can be written as $\int_1^2 dV = V_2 - V_1$. The operator *d* is used to denote exact differentials.

1.6.3. Path Function

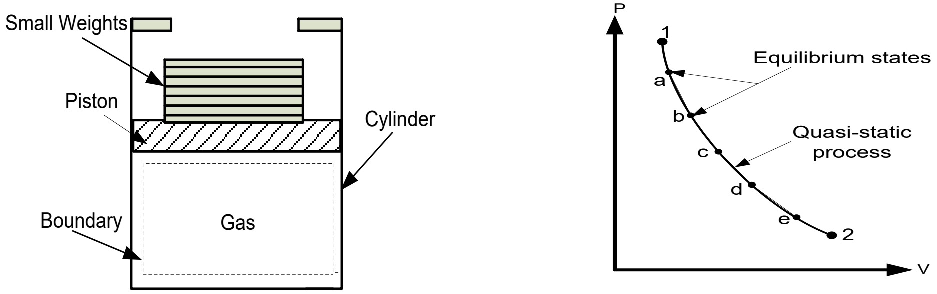
Any variable or quantity which depends not only on the end states but also on the path is known as a *path function*. Such quantities cannot be located on a property diagram by a point but are given by the area under the process curve. Work and heat are the examples of path functions and are inexact differentials because their change cannot be written as difference between their end states. Thus, $\int_1^2 \delta W \neq W_2 - W_1$ and can be denoted as W_{1-2}. Similarly, $\int_1^2 \delta Q \neq Q_2 - Q_1$ and can be denoted as Q_{1-2}. However, their change can be represented by the area under the process curve on property diagrams, e.g. the area under the process curve on *P-V* diagram gives the amount of the work ($W_{1-2} = \int_1^2 \delta W = \int_1^2 PdV$) and the area under the process curve on *T-S* diagram gives the quantity of the heat ($Q_{1-2} = \int_1^2 \delta Q = \int_1^2 TdS$). Since the area under the process curve is different for the different paths (*A, B, C*) undertaken by the processes for the fixed initial state 1 and final state 2 as shown in figure 1.5, both the heat and work are therefore the path functions. The operator δ is used to denote inexact differentials.

1.6.4. Path and Thermodynamic Process

Any operation in which properties of the system change is known as a *change of state*. The locus of the series of intermediate state through which a system passes while it is undergoing from initial state to its final state constitutes the *path*. The path such as *A* or *B* or *C as* in figure 1.5 followed by a system when it is going from one equilibrium state to another state is referred to as a *thermodynamic process*. A thermodynamic process is therefore simply a change in the state of the system and it is described with the help of property diagrams i.e. diagrams drawn with any pair of thermodynamic properties such as *P-V*, *T-V*, *P-T*, *P-h*, *T-s*, *h-s* etc. There are many ways for a system to change from one state to another state and they are identified by special names. For example, an *isochoric process* is a constant volume process, an *isobaric process* is a constant pressure process, an *isothermal process* is constant temperature process, an *isentropic process* is a constant entropy process and an *isenthalpic process* is a constant enthalpy process etc.

1.6.5. Quasi-Static or Quasi-Equilibrium Process

During a thermodynamic process, some unbalanced force must be present either within a system or between a system and its surroundings to cause the change of state. When the process runs slowly, the unbalance force is infinitesimally small and the system would be at almost (*quasi*) close to state of equilibrium or at state of *quasi-equilibrium*. The departure of the state of system from the thermodynamic equilibrium state will be infinitesimally small. So, every state passed by the system will be an almost equilibrium state. The locus of a series of such equilibrium states is defined as a *quasi-static or quasi-equilibrium process* and can be represented graphically as a continuous line on a property diagram. A quasi-static process is thus a succession of equilibrium states. Therefore, a quasi-static process can be viewed as a sufficiently slow process which allows the system to adjust itself internally so that properties in one part of the system do not change any faster than those at other parts of the system.



(a) (b)

Fig. 1.6. Quasi-static process with equilibrium states

To understand the quasi- static or quasi-equilibrium process in a better way let us consider a system of gas contained in a piston cylinder device as shown in figure 1.6(a) with a piston upon which a number of very small pieces of weights are placed. The upward force exerted by the gas just balances the weights above the piston. The system is initially in equilibrium state 1 defined by initial pressure P_1, volume V_1 and temperature T_1. When these small weights are removed slowly, one at a time, the unbalanced force infinitesimally small. The piston will slowly move upwards and at any particular instant of piston position, the system would be almost close to state of equilibrium. At every intermediate state its pressure decreases due to reduced weight and its volume increases and system goes to new state. If we go on removing weight, its pressure further decreases and volume further increase and system passes through intermediate states a, *b*, *c*, *d*, e,… and finally to state 2 defined by final pressure P_2, volume V_2 and temperature T_2 as shown in figure 1.6(b). Thus, this process 1-2 with a series of equilibrium intermediate states (a, b, c, d, e…) is a quasi- static or quasi-equilibrium process.

In contrary to this, if all the weights are removed at a time suddenly, the unbalanced force would be finite. The piston will jump up, strike the stoppers and finally come to equilibrium state 2 defined by final pressure P_2, volume V_2 and temperature T_2 after many oscillations. However, the intermediate states through which the system (gas) passes are indeterminate and cannot be located on the property diagram. Such a process is known as a non-quasistatic or a non-equilibrium process. Most of the real processes run very fast and therefore they are non-quasistatic, even though may be in equilibrium at the initial and final states [1].

1.6.6. Reversible and Irreversible Processes

With reference to intermediate states, thermodynamic processes can be classified as: reversible and irreversible process.

A thermodynamic process is said to be a *reversible process* if the system passes through a continuous series of equilibrium states. An equilibrium state of the system can be located on a property diagram drawn for any pair of thermodynamic properties (e.g. *P-V* diagram) at any instant during the process. The reversible process is then represented by a continuous line (path 1-2) drawn through the intermediate equilibrium states as shown in the figure 1.7(a). If the process is reversed, path 2-1 will be followed and that will restore the system as well as surroundings back to their respective initial states. Any quasi-static process that proceeds at infinitely slow speed without friction is a prerequisite of the reversible process which can be carried out in reverse direction (path 2-1) along the same path (1-2). No actual process in nature is perfectly reversible but some processes may be reversible process to a close approximation. Thus, a reversible process is a quasi-static process, but a quasi-static process need not be a reversible process. Some examples of nearly reversible processes are: frictionless motion, elastic deformation, restricted expansion and compression, heat transfer with infinitely small temperature difference etc.

A thermodynamic process is called an *irreversible process* if the system passes through a sequence of non-equilibrium intermediate states. The thermodynamic properties of the system do not have a unique value during such a process and therefore intermediate states cannot be defined or cannot be located on any property diagram. As shown in figure 1.7(b), an irreversible process is represented by a broken line on the property diagram since the properties are definite only at the initial and final states and not at the intermediate states. When an irreversible process is made to proceed in the backward direction (path 2-1'), the initial state 1 of the system is not restored. All natural processes take place spontaneously at finite speed with friction and hence are irreversible which can be carried out in one direction only. Some examples of irreversible processes are: motion with friction, inelastic deformation, free expansion, heat transfer with finite temperature difference etc.

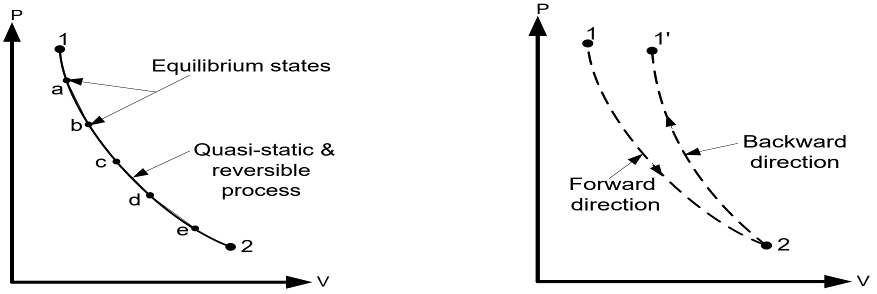


Fig. 1.7. (a) Quasi-static and reversible process (b) Irreversible process

1.6.7. Cyclic Process

A process is said to be a cyclic process, when a system undergoes through a series of processes such that the final state is identical with the initial state. An essential feature of the cyclic process is that the initial condition of the system is restored after a number of different processes in series. The change in the value of any thermodynamic property of the system for a cyclic process is zero, i.e., if P is any thermodynamic property of a system, then $\oint dP = 0$. Figure 1.8 shows a cyclic process 1-2-3-1 which comprises three processes namely 1-2, 2-3 and 3-1.

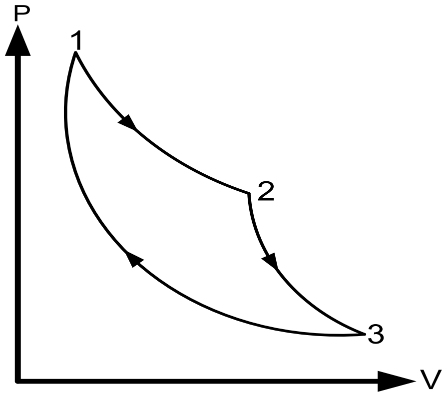


Fig. 1.8. A cyclic process with three thermodynamic processes

1.7. Some Common Properties

Those properties which are commonly or frequently used in the analysis of the thermodynamic systems are known as common properties. Some of the common properties are specific volume, Pressure and temperature.

1.7.1. Specific Volume

The space occupied by a system is called its volume. It is expressed in SI unit by m^3. Specific volume is defined as volume per unit mass of the system. Specific volume of a system is reciprocal of commonly used property density (ρ) of the system. It is denoted by v and it is expressed in m^3/kg.

$$v = \frac{V}{m} = \frac{1}{\rho} \qquad \text{……….} (1.2)$$

In thermodynamics, most of the time we deal with gas and vapor which have very low density. Hence it is appropriate to express property or state relationships of such substances in terms of specific volume.

1.7.2. Pressure

Normal force acting per unit area of a system is called Pressure. If the total normal force F acts uniformly over the entire area A, then the pressure P may be defined as

$$P = F/A \qquad \text{……….} (1.3)$$

It is usually expressed in N/m^2 or Pascal (Pa) (1 Pascal = 1 N/m^2). The unit of Pascal is very small. Hence, Engineers express pressure in Kilopascal (1 kPa = 10^3 Pa) and Megapascal (1 MPa = 10^6 Pa). Some other commonly used units of pressure are bar or atm.

1 bar = 100 kPa = 10^5 Pa = 750 mm of mercury
1 atm = 101.325 kPa = 1.01325×10^5 Pa = 760 mm of mercury

There are generally three pressure terms in thermodynamics which are associated with measurement of pressure: atmospheric pressure (P_{atm}), gauge pressure (P_{gauge}) and absolute pressure (P_{abs}).

1.7.2.1. Atmospheric pressure

This is the pressure exerted by the ambient air on any surface. The standard *atmospheric pressure* is defined as the pressure produced by a column of mercury 760 mm high. Atmospheric pressure is measured by an instrument called barometer. Barometer reads atmospheric pressure in terms of height of the mercury column z_{baro} and it is calculated as

$$P_{atm} = \rho_{Hg} g z_{baro} \qquad \text{……… (1.4)}$$

where $\rho_{Hg} = 13593$ kg/ m^3 is the density of mercury and $g = 9.80665$ m/s^2 is the local gravitational acceleration.

1.7.2.2. Gauge Pressure and Negative Pressure

Pressure of any system measured with reference to surroundings pressure (which is generally atmospheric pressure) is called *gauge pressure*. Gauge pressure is usually measured by instruments like pressure gauge or manometer. Pressure gauge gives pressure reading directly in Pa or bar. Manometer gives gauge pressure of a system in terms of height of certain manometric fluid z_{mano} and it is calculated as

$$P_{gauge} = \rho_{mano}\, g z_{mano} \qquad \text{……… (1.5)}$$

where ρ_{mano} is the density of manometric fluid and g is the local gravitational acceleration.

Instruments like manometer and pressure gauges used to measure fluid pressure generally measure the difference between the unknown real pressure that is to be identified and the existing atmospheric pressure, i.e. $(P_{abs} - P_{atm})$. When the unknown pressure is greater than atmospheric pressure, the pressure measured by the instrument is known as gauge pressure and is positive. If the unknown pressure reading is less than the atmospheric pressure, it is negative and then the pressure reading of the instrument is called vacuum or negative pressure (P_{vac}).

As shown in figure 1.9, a simplest and useful manometer is a transparent U shaped tube consisting two open ends and filled with a manometric liquid whose density is known. One end of the U- tube is open to the atmosphere and the other is connected to a system having pressure higher than the atmospheric pressure which is to be identified.

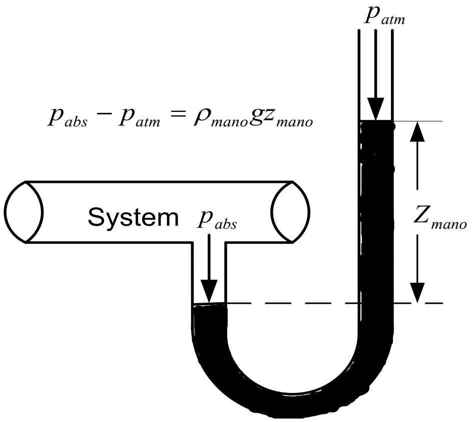


Fig. 1.9. U-tube manometer

As shown in figure 1.9, a simplest and useful manometer is a transparent U shaped tube consisting two open ends and filled with a manometric liquid whose density is known. One end of the U- tube is open to the atmosphere and the other is connected to a system having pressure higher than the atmospheric pressure which is to be identified.

1.7.2.3. Absolute pressure

Pressure of a system expressed relative to the state of perfect vacuum or zero pressure is defined as *absolute pressure.* A state of absolute zero pressure or perfect vacuum will occur only when molecular momentum is zero and it is independent of the changes in atmospheric pressure. Hence actual absolute pressure is the algebraic sum of the gauge pressure and the atmospheric pressure:

$$P_{abs} = P_{atm} + P_{gauge} \qquad \text{..........} (1.6)$$

Sometimes when the pressure of a system may be less than atmospheric pressure, gauge pressure becomes negative. Then the absolute pressure is related as

$$P_{abs} = P_{atm} - P_{vac} \quad \text{…..…….(1.7)}$$

where P_{vac} is the negative gauge pressure or vacuum gauge.

Relationship between these pressure terms P_{atm}, P_{gauge}, P_{abs} and P_{vac} can be illustrated on a pressure scale as shown in Figure 1.10.

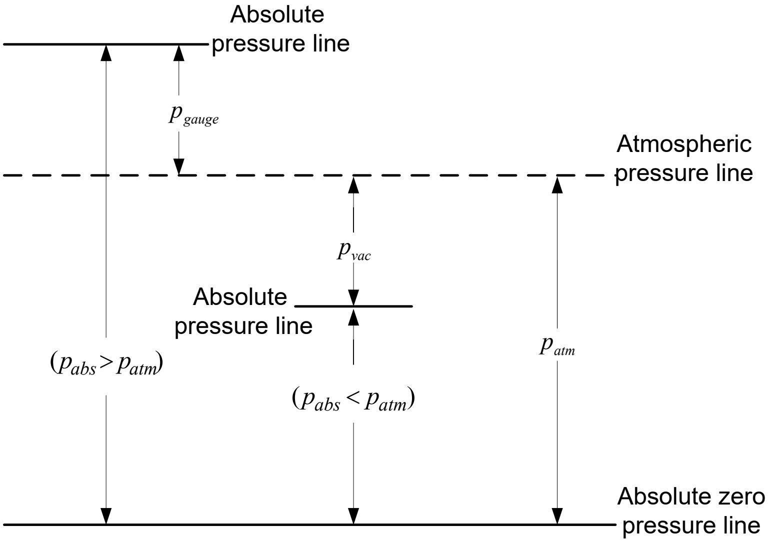


Fig.1.10. Relationship between absolute, gauge, atmospheric pressure and negative pressure or vacuum on a pressure scale

1.7.3. Temperature

A degree of hotness or coldness of a system is known as its *temperature*. But in engineering thermodynamics, it is appropriate to define temperature as a thermodynamic property of a system which dictates the direction and magnitude of heat transfer. It is a driving force which causes the flow of energy as heat from a system to the surroundings and vise versa. If the temperature of the system is higher than that of the surroundings, heat transfer takes place from the system to the surroundings while if the temperature of the system is less than that of the surroundings, heat transfer takes place from surroundings to the system. Moreover, magnitude of the heat transfer is also proportional to the temperature difference between the system and the surroundings.

In other words, temperature may be defined as the measure of the average kinetic energy of the molecules of a system. A change in the temperature of the system promotes the change in the molecular motion and hence the kinetic energy of the system. It is a parameter which determines whether or not a system is in thermal equilibrium with another system. Temperature is an intensive property independent of the size and mass of the system.

An instrument which is used to measure the temperature of a system is called a thermometer. A quantitative measurement of the temperature of a system by a thermometer requires some reference or fixed point and the establishment of a suitable temperature unit. Many temperature scales and reference points have been proposed, but commonly used important temperature scales for temperature measurement are degree Celsius (^0C), degree Fahrenheit (^0F) and Kelvin (K). Celsius and Fahrenheit scales are defined with reference to the ice point ($0\ ^0C\ or\ 32\ ^0F$) and the steam point ($100\ ^0C\ or\ 212\ ^0F$) whereas Kelvin scale has been defined with reference to a single point - the triple point of water ($0.01\ ^0C$). Temperature expressed in Kelvin scale is called Kelvin or absolute temperature. The following correlation can be established between temperatures measured on different scales:

$$\frac{C-FP}{BP-FP}=\frac{F-FP}{BP-FP}=\frac{K-FP}{BP-FP}$$

$$\Rightarrow \quad \frac{C-0}{100-0}=\frac{F-32}{212-32}=\frac{K-273}{373-273}$$

$$\Rightarrow \quad \frac{C}{100}=\frac{F-32}{180}=\frac{K-273}{100} \qquad \text{............(1.8)}$$

where FP and BP are freezing point and boiling point respectively and their values on different temperature scales are given below [2]:

Scale	**Freezing Point (FP)**	**Boiling Point (BP)**
Celsius (^0C)	0	100
Fahrenheit (^0F)	32	212
Kelvin (K)	273	373

1.8. Equality of Temperature and Zeroth Law of Thermodynamics

1.8.1. Equality of Temperature

Two systems are said to have equal temperature if there is no change in any properties of both the systems when they are brought into physical contact.

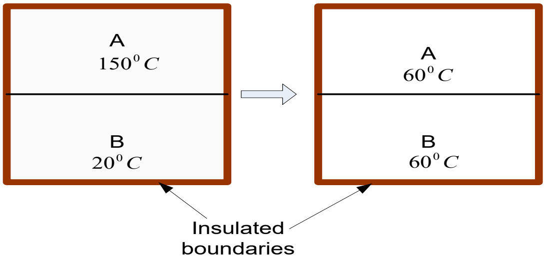


Fig. 1.11. Representation of equality of temperature

Consider two systems A and B which are perfectly insulated from the surroundings and at different temperatures 150 ^0C and 20 ^0C respectively in the initial condition as shown in figure 1.11. When they are brought into physical contact, the energy in the form of heat will flow from the system A at higher temperature to the system B at lower temperature. There would occur changes in any properties of both the system such as temperature, pressure, volume etc. at the beginning. After a sufficient long time of contact, there is no further change in their properties and both systems A and B attain the same temperature 60 ^0C (figure 1.11). At this point, the heat transfer stops, and the two systems are said to have reached thermal equilibrium. The two systems are then equal in temperature. This condition is known as equality of temperature and it is the only requirement for thermal equilibrium.

1.8.2. Zeroth Law of Thermodynamics

The zeroth law of thermodynamics states that

if two systems are separately in thermal equilibrium with a third system, then these two systems are also in thermal equilibrium without bringing them physically in contact with each other.

It may seem silly that such an obvious fact is called one of the basic laws of thermodynamics. However, it cannot be concluded from the other laws of thermodynamics, and it serves as a basis for the validity of temperature measurement by using a thermometer where the thermometer acts as the third system.

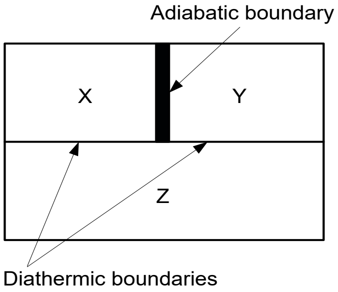


Fig. 1.12. Zeroth law of Thermodynamics

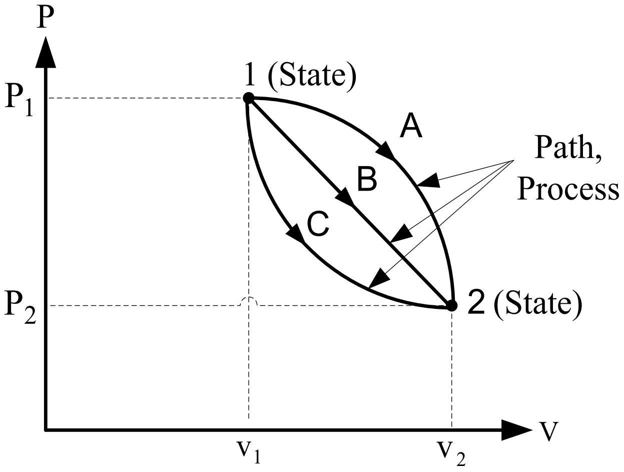
Consider three systems X, Y and Z as shown in figure 1.12. Let us assume that the boundaries between two systems X and Z as well as between Y and Z are diathermic (not adiabatic) whereas the boundary between X and Y is completely insulated (adiabatic). If the systems X and Z interact with each other in form of heat transfer and there is no change in any properties of both the system, then it can be concluded that both X and Z have same temperatures. Again, if Y and Z interact with each other in the form of heat transfer and there is no change in any properties of both the system, then it can be concluded that both Y and Z have same temperatures. Then without performing third experiment, it can be said that X and Y must also be in thermal equilibrium and therefore have same temperatures although their boundary is insulated.

The zeroth law was first formulated by R. H. Fowler in 1931. As the name suggests, its value as a fundamental physical principle was recognized long time after the formulation of the first and the second laws of thermodynamics. It was named the zeroth law since scientists thought that it was so important it should have preceded the first and the second laws of thermodynamics. The zeroth law is applicable in all temperature measuring processes including calibration (setting) of temperature scale as it provides the basis for the measurements of temperature and it therefore enables us to use thermometers to compare the temperatures of multiple objects.

Lecture Highlights:

- *Thermodynamic state:* It is a unique condition at which a thermodynamic system can exist.
- A *thermodynamic property* has two main features: a variable is a thermodynamic property, if and only if ,
  (i) it has a single value at each equilibrium state
  (ii) a change in its value between any two prescribed states is independent of path.
- *Thermodynamic process:* It is the path followed by a system when it undergoes from one equilibrium state to another.
- *Why heat and work are path functions*:

  Any variable or quantity which depends not only on the end states but also on the path is known as a *path function*. Such quantities cannot be located on a property diagram by a point but are given by the area under the process curve. Work and heat are the examples of path functions and their magnitude can be represented by the area under the process curve on property diagrams, e.g. the area under the process curve on *P-V* diagram gives the amount of the work ($W_{1-2} = \int_1^2 \delta W = \int_1^2 PdV$) and the area under the process curve on *T-S* diagram gives the quantity of the heat ($Q_{1-2} = \int_1^2 \delta Q = \int_1^2 TdS$). Since the area under the process curve is different for the different paths (*A, B, C*) undertaken by the processes for the fixed initial state 1 and final state 2 as shown in figure below (e.g. work), both the heat and work are therefore the path functions.



- Differences between *state or point function* and *path function* are:

State/point function	Path function
1. Variable which only depends on the end states is called state or point function.	1. Variable which depends not only on the end states but also on the path is called path function.
2. It is thermodynamic property.	2. It is not thermodynamic property.
3. It has exact differential.	3. It has inexact differential.
4. Its cyclic integral is zero.	4. Its cyclic integral is non-zero.
5. For examples: all thermodynamic properties such as pressure, temperature, volume etc.	5. For example: heat and work.

- With reference to intermediate states, thermodynamic processes can be classified as *reversible* and *irreversible process*.
- Differences between reversible and irreversible processes are:

Reversible process	Irreversible process
1. It is a process which passes through a series of equilibrium states such that each intermediate state can be located on a property diagram. 2. It is denoted by continuous line. 3. It occurs at an infinitely slow rate. 4.The system is in equilibrium state or in quasi-equilibrium state at any instant. 5.It can be carried out in both directions and initial state can be restored. 6.It results in higher efficiency than that of irreversible process. 7.For examples: Frictionless motion, elastic deformation, restricted expansion and compression, heat transfer with infinitely small temperature difference etc.	1. It is a process which passes through a series of intermediate states which cannot be defined or located on a property diagram. 2. It is denoted by broken line indicating that intermediate states are indeterminate. 3. It occurs at a finite rate (fast). 4. The system is never in the equilibrium state at any instant. 5. It can be carried out in one direction only and initial state cannot be restored. 6. It results in lower efficiency than that of reversible process. 7. For examples: Motion with friction, inelastic deformation, free expansion, heat transfer with finite temperature difference etc.

- *Cyclic process:* It is defined as a series of individual processes whose final and initial states are identical. Cyclic integral of a thermodynamic property for a cyclic process is zero.
- *Quasi-static process*: It is a succession of equilibrium states and infinite slowness is its characteristic feature. A reversible process is a quasi-static process, but a quasi-static process need not be a complete reversible process.
- *Specific volume:* Volume per unit mass of the substance is called specific volume and is reciprocal of density ($v = V/m = 1/\rho$).
- *Pressure*: Force per unit area is called pressure. It is generally expressed in Pascal ($Pa = N/m^2$) and sometimes in bar (1 bar = 100 *kPa*) or in atmosphere (1 atm = 101.325 *kPa*).

- *Atmospheric pressure:* It is defined as pressure exerted by ambient (atmospheric) air on any surface and is measured by barometer. Atmospheric pressure measured by barometer with height of mercury Z_{baro} is evaluated as $P_{atm} = \rho g Z_{baro}$.
- *Gauge pressure:* It is the pressure of a system measured with reference to atmospheric pressure by using pressure gauge or manometer. Gauge pressure measured by manometer with height of manometric fluid Z_{mano} is determined as $P_{gauge} = \rho g Z_{mano}$.
- *Absolute pressure:* Pressure of a system relative to perfect vacuum is known as absolute pressure and it is calculated as

$$P_{abs} = P_{gauge} + P_{atm}.$$

- *Vacuum gauge*: When pressure of a system is less than atmospheric pressure, gauge pressure becomes negative and this negative gauge pressure is also called vacuum gauge.
- Differences between gauge pressure and absolute pressure are:

Gauge pressure	Absolute pressure
1. It is the pressure of a system measured with reference to atmospheric pressure or surroundings pressure. 2. It is measured by using pressure gauge or manometer. 3. It may be negative and is called vacuum gauge when pressure of a system is less than atmospheric pressure. 4. Gauge pressure measured by manometer is determined as $P_{gauge} = \rho g Z_{mano}$ whereas pressure gauge gives direct pressure reading.	1. It is the pressure of a system expressed relative to perfect vacuum. 2. It is calculated as the sum of gauge pressure and atmospheric pressure. 3. It is positive and may be zero if the system is perfect vacuum. 4. It is determined as $P_{abs} = P_{gauge} + P_{atm}$

- *Temperature*: It is defined as a property of a system which defines the direction and magnitude of heat transfer. It is also known as a degree of hotness or coldness of a system and is measured by an instrument called thermometer.
- Temperature is measured in different *temperature scales* namely in *degree Celsius, degree Fahrenheit* and *Kelvin* (absolute temperature) scales and they have relation as:

  $$\frac{N-FP}{BP-FP}=\frac{C}{100}=\frac{F-32}{180}=\frac{K-273}{100}$$

  where N is any new scale, FP is *freezing point* of ice and BP is *boiling point* of water.
- *Equality of temperature*: If there is no change in any properties of two systems when they are brought in physical contact, then both the system are said to have equal temperatures.
- *Zeroth law of thermodynamics:* It states that two systems which are separately in thermal equilibrium with a third system are also in thermal equilibrium with each other.
- A*pplications of zeroth law* are: measurement of the temperature of a system and comparison of the temperatures of multiple systems etc.

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[1] *Thermal Science and Engineering*: Kumar D.S., S. K. Kataria & Sons, India, 2009.

[2] *Fundamentals of Thermodynamics & Heat Transfer*: Luintel M.C., Heritage Publishers & Distributors Pvt. Ltd., Kathmandu, Nepal, 2016.