Course Title:

Fundamental of Thermodynamics and Heat Transfer

Lecture 5 (Week 5):

Properties of Pure Working Substances

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Learning Objective of Lecture:

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

- ✓ Pure Working Substances
- ✓ State Postulate
- ✓ Ideal Gases and Ideal Gas Relations
- ✓ Two Phase (Liquid and Vapor) Systems
- ✓ Saturation Curve on T – v Diagram
- ✓ Saturation Curve P – v Diagram
- ✓ Change of Phase of a Pure Substance on P – T Diagram
- ✓ Phase Change Terminologies and Definitions

3.1. Pure Working Substance and State Postulate

3. 1.1. Pure Working Substance

A substance used in any thermodynamic system (device) with the help of which energy is transported between the system and its surroundings is called a *working substance*. In all thermodynamic applications, the working substance serves as a medium for the transfer of energy to produce certain useful effect. The working substance is the fluid that can be conveniently made to absorb, store and release energy. Moreover, fluids can be easily transported from one place to another through pipes.

Examples of the common working substances generally used in thermodynamic applications are: combustion products obtained by igniting air fuel mixture in petrol and diesel engines, water vapor in steam power plants, atmospheric air in gas power plants, refrigerants (ammonia, hydrocarbons like Freon-12) in refrigerator, heat pump and air conditioning etc.

The property analysis of the common working substances that can be used to transport energy is the one of the tasks in engineering thermodynamics. Working substances may exist in different phases or states, e.g. gas, liquid, vapor or mixture of liquid and vapor for most of the applications. When the working substance is in gaseous phase, it can be modeled as an ideal gas and its analysis is done by using ideal gas equations. When the working substance is in liquid phase, vapor phase (saturated or slightly superheated vapor) or mixture of liquid and vapor (two phase mixture), its properties cannot be calculated by using ideal gas equations and in such a case experimental data (also called empirical data) are used. However, sometimes superheated vapor with high degree of superheat can be modeled as an ideal gas and its property analysis may be done by using ideal gas equations with slight error.

Working substance used in any thermodynamic system is assumed to be a pure substance for most of the thermodynamic analysis. A pure substance can exist in three different states or phases. The three conditions of a pure substance namely solid, liquid and gas (vapor) are known as phases. Thus, a phase is any part of a system that is chemically uniform (homogeneous) and physically distinct. A pure substance is therefore a system which is [1]

(i) homogeneous in composition
(ii) homogeneous in chemical aggregation (combination) and
(iii) invariable in chemical aggregation

Homogeneous in composition means that the relative proportions of the chemical elements of the substance is the same in all parts of the system. It does not matter how these elements are combined. Homogeneous in chemical combination means that the chemical elements of the substance must be combined chemically in the same way in all samples of the system. Invariable in chemical aggregation means that the state of chemical combination of the system does not change with time.

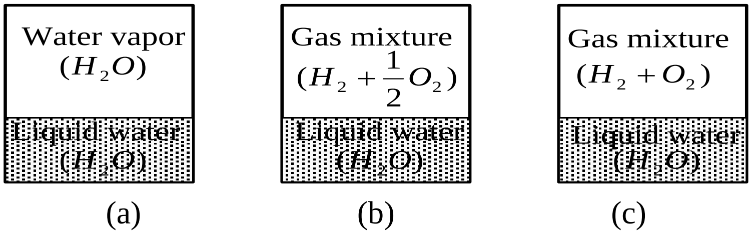

Figure 3.1. Illustrating the definition of a pure substance

Let us consider three systems as shown in figure 3.1, system (a) consisting of homogeneous mixture of liquid water and water vapor, system (b) consisting of a mixture of liquid water and gaseous mixture of hydrogen and oxygen in which H and O present in the proportion of 2:1 and system (c) consisting of a mixture of liquid water and gaseous mixture of hydrogen and oxygen in which H and O present in the proportion of 1:1.

When system (a) is analyzed, it is observed that hydrogen atoms H and oxygen atoms O are present in the ratio 2:1 (i.e. relative proportion of the chemical elements is identical) whether the sample is taken from the liquid water or from the steam. Similarly, chemical combination of H and O is identical for both liquid water and water vapor. Hence, system 1 satisfies both conditions (i) and (ii) and it can be regarded as a pure substance as long as its state of combination does not change with time (i.e. satisfies condition (iii)).

While analyzing system (b), it is observed that the relative proportion of hydrogen atoms H and oxygen atoms O is the same (2:1) whether the sample is taken from the liquid water or from the gaseous mixture of H_2 and O_2. So, it is homogenous in composition. But in the gaseous mixture the hydrogen and oxygen atoms are not combined chemically whereas in case of the liquid water the hydrogen and oxygen atoms are combined to form liquid water. Hence it is not homogeneous in chemical combination and cannot be regarded as a pure substance.

Similarly, if system (c) is analyzed, it can be observed that it is neither homogeneous in composition nor homogeneous in chemical combination. Because the hydrogen and oxygen atoms are present in the ratio 1:1 and chemically uncombined form in the gaseous mixture, but in the ratio 2:1 in the liquid water. Hence, it is not a pure substance.

Thus, a mixture of hydrogen and oxygen gases which changed into water or steam during the time interval would not be a pure substance. However, air is considered as a pure substance although it is a mixture of several gases as long as it is all gas or all liquid. Hence, single elements and compounds are both examples of pure substances. A mixture of various chemical elements or compounds like air also qualifies as a pure substance as long as the mixture is homogeneous. A mixture of two or more phases of a pure substance (e.g. a mixture of ice and liquid water) is still a pure substance as long as the chemical composition of all phases is the same. Pure substances display a sharp melting and boiling point.

3.1.2. State Postulate (Two Property Rule)

Repeated observations and experiments on a pure substance show that if the values of any two independent thermodynamic properties are known, all other thermodynamic properties can be determined for a state of a system composed of a pure substance. This special feature of a pure substance for a state of a system is known as state postulate or two property rule. The observed behavior of a pure substance can be stated as

The values of any two independent thermodynamic properties are sufficient to establish a stable thermodynamic state of a system composed of a pure substance [2].

This statement is applicable for a system which is composed of a pure substance in the absence of electric, magnetic, gravitational motion and surface tension effects and displacement (expansion or compression) work is the only quasi-static mode of work ($W = P.dV$). Such a pure substance is also called a *simple compressible substance*. The state postulate gives the information needed to specify the state of a particular system completely. For example, if the values of temperature T and volume V are known for a pure substance for any state, then all other properties are uniquely specified i.e. properties like pressure P and internal energy U of a pure substance can be determined for that state as

$$P = P(T, V) \text{ and } U = U(T, V) \qquad \text{……….} (3.1)$$

Stable thermodynamic state implies an equilibrium state, so that state postulate does not provide any information concerning non-equilibrium states. The state postulate is applicable for a closed system or a control mass. Therefore, it can be stated on a per unit mass basis or in terms of intensive properties as:

The values of two independent intensive thermodynamic properties are sufficient to establish a stable thermodynamic state of a system composed of a pure substance [2].

For example, the pressure P and specific internal energy u of a pure substance for any state can be determined if its temperature T and specific volume v are independent and their values at that state are known. Mathematically,

$$P = P(T, v) \text{ and } u = u(T, v) \qquad \text{……….} (3.2)$$

Two properties are considered independent if one can be varied while the other is held constant. For example, temperature and specific volume are always independent. However, temperature and pressure are independent only for a single phase system and they are no more independent for a system consisting of two phase mixture.

3.2. Ideal Gas and Ideal Gas Relations

A gas having no forces of molecular attraction is known as an *ideal gas* or a perfect gas. A gas which obeys the gas laws (ideal gas relations) at all ranges of pressures and temperatures can be considered as an ideal gas. But there is no such gas existing in nature. However, real gases tend to obey these laws at low pressures or high temperature or at both. This is because the molecules are far apart to each other at low pressures and high temperatures due to which the forces of molecular attraction tend to be negligibly small. The range of this low pressure and high temperature at which any gas behaves as an ideal gas is different for different gases [3].

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3.2.1. Boyle's Law

This law was enunciated by Robert Boyle on the basis of his experimental results. Boyle's law states that

The volume of a given mass of an ideal gas is inversely proportional to the pressure if the temperature remains constant. i.e.

$$V \propto \left(\frac{1}{P}\right)_T \qquad \text{……… (3.3)}$$

Eliminating the sign of proportionality, the equation (3.3) can be written as

$$PV = \text{Constant} \qquad \text{…………(3.4)}$$

For any intermediate state (e.g. 1 and 2) of a thermodynamic process, equation (3.4) is expressed as

$$PV = P_1V_1 = P_2V_2 = ... \qquad \text{…………(3.5)}$$

3.2.2. Charle’s Law

This law is also called Gay-Lussac's law, and it can be stated in two parts:

(a) *The volume of a given mass of an ideal gas varies directly with the absolute temperature if the pressure is held constant. i.e.*

$$V \propto (T)_P \qquad \text{……… (3.6)}$$

Eliminating the sign of proportionality, the equation (3.6) can be written as

$$V/T = Constant \qquad \text{……….(3.7)}$$

For any intermediate state (e.g. 1 and 2) of a thermodynamic process, equation (3.7) is expressed as

$$V/T = V_1/T_1 = V_2/T_2 = \cdots \qquad \text{…………..(3.8)}$$

(b) *The pressure of a given mass of an ideal gas varies directly with the absolute temperature if the volume is held constant. i.e.*

$$P \propto (T)_V \qquad \text{……… ....(3.9)}$$

Eliminating the sign of proportionality, the equation (3.9) can be written as

$$P/T = Constant \qquad \text{………...(3.10)}$$

For any intermediate state (e.g. 1 and 2) of a thermodynamic process, equation (3.10) is expressed as

$$P/T = P_1/T_1 = P_2/T_2 = \cdots \qquad \text{………...(3.11)}$$

3.2.3. Equation of state

Equation of state is a relationship between all three common parameters or thermodynamic properties namely pressure (P), specific volume(v), and temperature (T) for a particular state of a system (an ideal gas). The equation of state can be derived by combining Boyle's law and Charles' law. For this, let us consider unit mass of an ideal gas undergoing constant pressure or isobaric process $1-A$ and constant temperature or isothermal process $A-2$ in series as shown in figure 3.2.

Applying Charle's law for the isobaric process: $1-A$

$$\frac{v_1}{v_A} = \frac{T_1}{T_A}$$

$$v_A = \frac{T_A}{T_1}v_1 = \frac{T_2}{T_1}v_1 \qquad \text{……… (3.12)}$$

Here, $T_A = T_2$ for isothermal process $A - 2$.

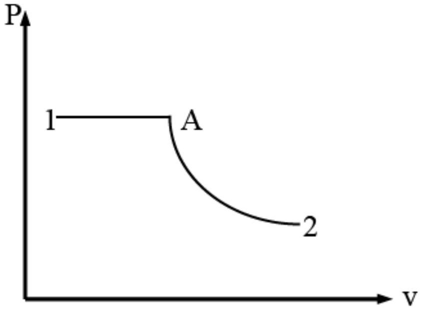

Figure 3.2: Isobaric and isothermal processes in series

Using Boyle's law for the isothermal process: $A - 2$,

$$\frac{v_2}{v_A} = \frac{P_A}{P_2}$$

$$v_A = \frac{P_2}{P_A} v_2 = \frac{P_2}{P_1} v_2 \qquad \text{……… (3.13)}$$

Here, $P_A = P_1$ for isobaric process $1 - A$. Equating the two relations (3.12) and (3.13)

$$\frac{P_1 v_1}{T_1} = \frac{P_2 v_2}{T_2} \qquad \text{……… (3.14)}$$

Since states *1* and *2* are selected arbitrarily, it follows that

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

$$Pv = RT \qquad \text{……… (3.15)}$$

where R is a constant for a given ideal gas. The constant R is known as the particular or characteristic gas constant and has the unit of J/kg.K. It has a particular value for a particular gas. Characteristic gas constants for different gases are given in Appendix 1.

For a system consisting of a mass of m kg of an ideal gas, equation (3.15) becomes

$$PV = mRT \qquad \text{……… (3.16)}$$

Equation (3.16) is known as equation of state of a system and it provides the relationship between common thermodynamic properties *P, V, T* of a state of a system.

Similarly, for a system consisting of n number of moles of an ideal gas, equation (3.16) can be expressed as

$$PV = n\bar{R}T \qquad \text{…………(3.17)}$$

where $\bar{R}$ is called a universal gas constant and its value is 8.3144 J/mol.K and same for all gases. Equating equations (3.16) and (3.17), the relationship between two gas constants will be

$$\bar{R} = \frac{m}{n}R = MR \qquad \text{…………(3.18)}$$

where M is molecular weight of a gas and has a unit of kg/mol.

3.3. Two Phase (Liquid and Vapor Mixture) Systems

There are many practical situations where two phases of a pure substance coexist in equilibrium. For example, water exists as a mixture of liquid and vapor phase in the boiler and the condenser of a steam power plant. The refrigerant converts from liquid to vapor phase in the freezer of a refrigerator. Even though freezing of water is an important phase-change process, attention in this section is focused on the liquid and vapor phases and their mixture. As a familiar and easily available substance, pure water is used to demonstrate the basic principles involved. However, all pure substances exhibit the same general behavior as water.

A system is said to be in two phases when it exists as a mixture of two phases during phase change, e.g. a mixture of solid and vapor, solid and liquid, liquid and vapor. Working substance in the form of mixture of liquid and vapor are used in many thermodynamic systems to transport energy. For example, liquid water or steam (water vapor) as the working substance is used in power plants, refrigerators, heat pumps, air conditionings and many thermal distribution systems.

Water may exist in the compressed or sub-cooled liquid region, two phase or saturation region where saturated liquid water and saturated water vapor are in equilibrium and superheated vapor region.

The variations of properties during phase change processes are best studied and understood with the help of property diagrams. In the next section, saturation curve on T–v, P–v, and P–T diagrams for pure substances like water are developed and discussed.

3.3.1. Saturation Curve on $T - v$ Diagram

Let us consider the results of constant pressure heating of some amount of liquid water in a piston-cylinder device (see figure 3.3), from atmospheric condition i.e. temperature 20°C and pressure 1 atm as specified by state 1 shown in figure 3.4. If the piston is assumed to be frictionless and have negligible weight, the system pressure remains constant at $1\,atm$ throughout the process. Under these conditions water exists in the liquid phase and it is called a *compressed liquid* or a *subcooled liquid* which is not about to vaporize. When the liquid water is heated by keeping the pressure constant, the following events occurs:

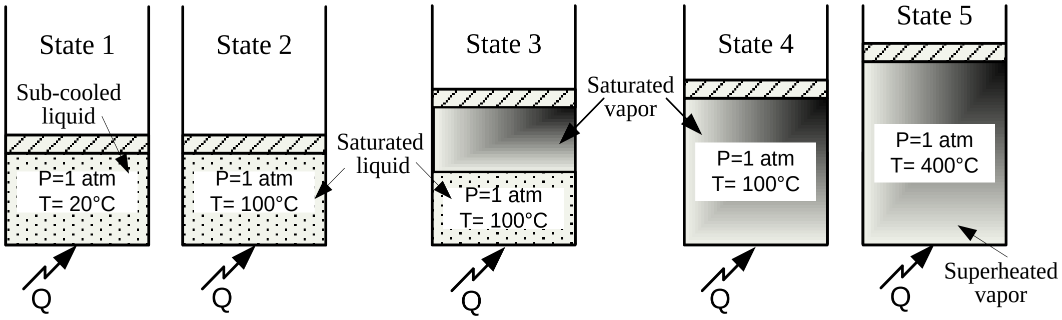

Figure 3.3. Arrangements for constant pressure heating of water

Process 1 – 2

The temperature of the compressed or subcooled liquid from state 1 increases with slight increment of its specific volume when heat is supplied to the system from surroundings. This process is represented by process line 1–2 in figure 3.4. This trend continues until its temperature reaches to $100\,℃$, state 2. In the compressed liquid region, the specific volume of the liquid is approximately same to the specific volume of the saturated liquid state as the water in the liquid phase is almost incompressible. Hence, the process line 1–2 is almost a vertical line.

Process 2 – 4

At this particular state of water (state 2) the liquid has reached the temperature at which it begins to boil called the saturation temperature and is said to exist as a saturated liquid state. Hence, a liquid that is about to vaporize is called a *saturated liquid*. Further heat addition causes it to vaporize. Vaporization process is represented by phase change process line 2–4. During the phase change both the pressure and temperature remains constant and specific volume of the substance increases due to intermolecular expansion. At state 3 on the process line 2–4, the liquid phase and vapor phase are in equilibrium. After the absorption of complete latent heat, water gets completely converted into vapor, state 4 in figure 3.4, which is known as saturated vapor state. At this state the entire cylinder is filled with saturated vapor. Any heat loss from this vapor causes some of the vapor to condense. Hence, a vapor that is about to condense is called a *saturated vapor*.

A substance at states between 2 and 4 is referred to as a *saturated liquid - vapor mixture* since the liquid and vapor phases coexist in equilibrium at any state between them.

Process 4 - 5

If the constant pressure heating is continued, the temperature of the saturated vapor increases above the saturation temperature (100°C in this example) with the increase in specific volume as shown by process line 4–5 in figure 3.4 and the state 5 is called a superheated vapor state. Because the temperature at state 5 (400°C) is greater than the saturation temperature (100°C) for the given pressure (1 atm), the vapor at state 5 is not about to condense. Hence, a vapor that is not about to condense is called a *superheated vapor*.

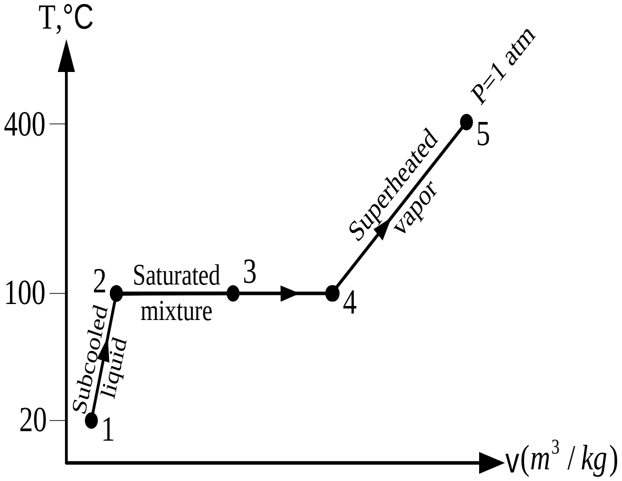

Figure 3.4. $T - v$ diagram for constant pressure heating of water

Similar constant pressure heating process at different higher pressures is repeated, by adding weights on top of the piston gradually until the pressure inside the cylinder reaches a required value to develop the T–v diagram at various pressures as shown in figure 3.5, but there are some noticeable characteristics. For example, the vaporization of water takes place relatively at higher temperatures at higher pressures. Likewise, the specific volume of the saturated liquid is larger and the specific volume of the saturated vapor is smaller than the corresponding values at lower pressures. Thus, as the pressure is increased further, the length of the saturation line 2–4 goes on decreasing. At some particular pressure it becomes a point, i.e. saturated liquid state and saturated vapor state coincide. This particular point or state of the substance is called the *critical point* at which both saturated liquid and saturated vapor states are identical. At pressure above the critical pressure there is no a distinct phase change process i.e. the phase transition from liquid to vapor is no longer discrete.

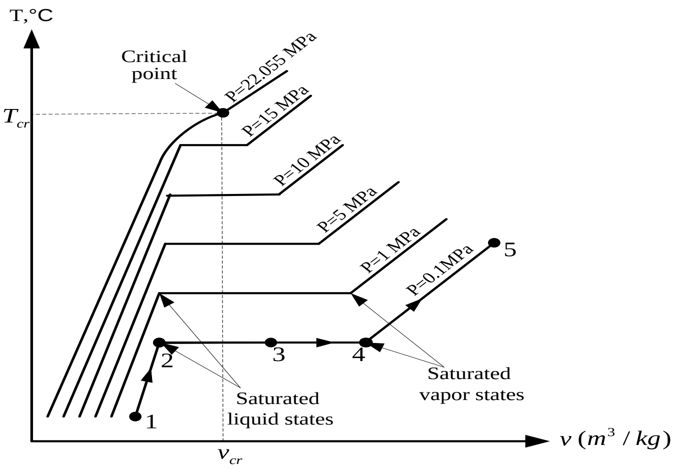

Figure 3.5. $T - v$ diagram for constant pressure heating of a pure substance (e.g. water) at various pressures

The curve that joins all of the saturated liquid states is defined as saturated liquid line and the curve that joins all of the saturated vapor states is known as saturated vapor line. These two lines intersect at the critical point and form what is often called a *steam dome*. The entire curve so formed as shown in figure 3.6 is called *saturation curve* on T–v diagram. The region to the left of the saturated liquid line below the critical temperature is known as the *compressed liquid or subcooled liquid region*, region to the right of the saturated vapor line below the critical temperature is known as the *superheated vapor region* and the region within the curve is defined as the *two-phase mixture region* or saturated liquid - vapor mixture region and just the saturation region or wet region.

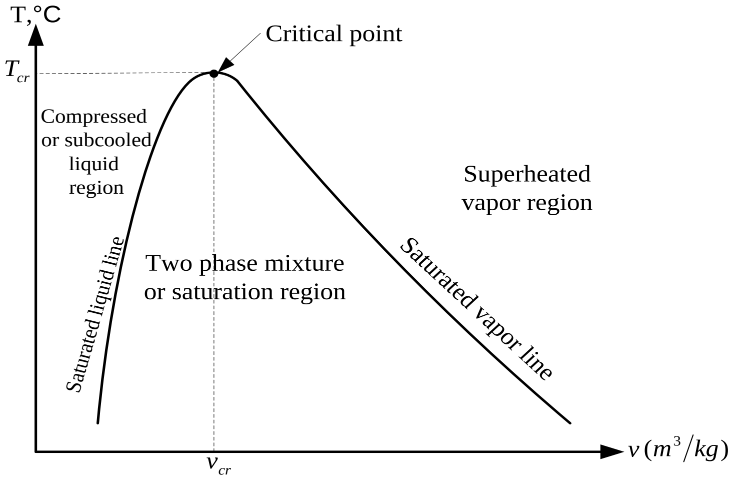

Figure 3.6. Saturation curve for two-phase mixture on T–v diagram

3.3.2. Saturation Curve on $P - v$ Diagram

To draw a P–v diagram of a pure substance, let us consider again a piston cylinder device that contains liquid water at 250 atm (25.33 MPa) pressure and 100°C temperature, at which state water exists as a compressed or subcooled liquid as shown in figure 3.7. The system is placed in contact with a source also called a thermal reservoir which has a temperature same as that of the water inside the cylinder.

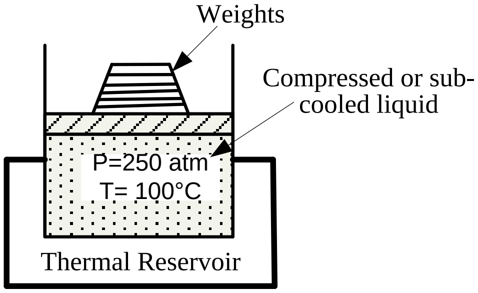

Figure 3.7. Arrangement for constant temperature expansion of water

If the weights on top of the piston are removed one by one such that temperature is maintained constant at 100^0C (by allowing heat exchange between the water and the thermal reservoir), the pressure inside the cylinder decreases gradually and there will be slight increment in the specific volume of the water, as shown by process line 1–2 in figure 3.8. This trend continues until its pressure reduces to the saturation pressure at the specified temperature (e.g. 1 atm at 100°C), i.e. the water starts to boil after taking heat energy from the thermal reservoir when it reaches to the saturated liquid state 2. During the vaporization process, the specific volume increases until it reaches to saturated vapor state 4 via state 3 at which the saturated liquid and saturated vapor are in equilibrium, while the pressure and temperature remain constant as shown by process line 2–4 in figure 3.8. Further reduction in the pressure of the water vapor from the state 4 will result in a further increase in its specific volume as shown by process line 4–5 in figure 3.8.

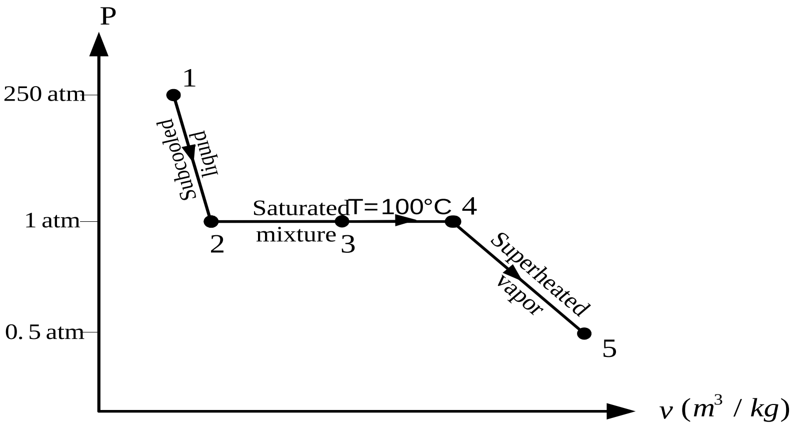

Figure 3.8. $P - v$ diagram for constant temperature expansion of water

If the similar expansion process is repeated for other higher temperatures, vaporization takes place correspondingly at higher pressure and similar paths are obtained for the phase change processes as shown in figure 3.9. After connecting the saturated liquid states and saturated vapor states by a curve, the saturation curve for two phase mixture on P–v diagram of a pure substance can be obtained as shown in figure 3.10.

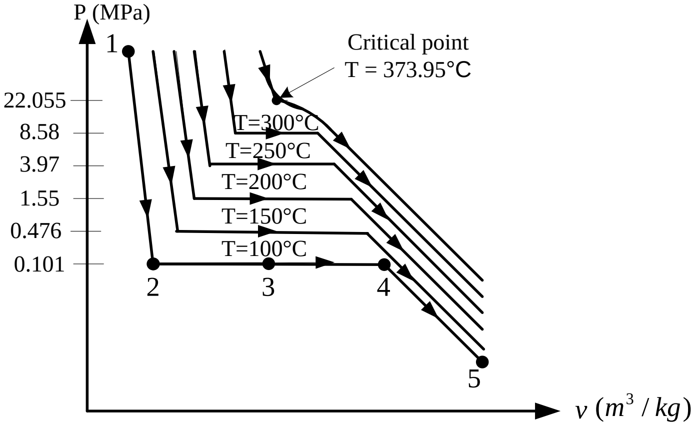

Figure 3.9. $P - v$ diagram for expansion of water at different temperatures

The general shape of the P–v diagram of a pure substance (figure 3.10) is very much similar to the T–v diagram (figure 3.6), except the T = constant lines (isotherms) on P–v diagram have a downward trend as shown in figure 3.9.

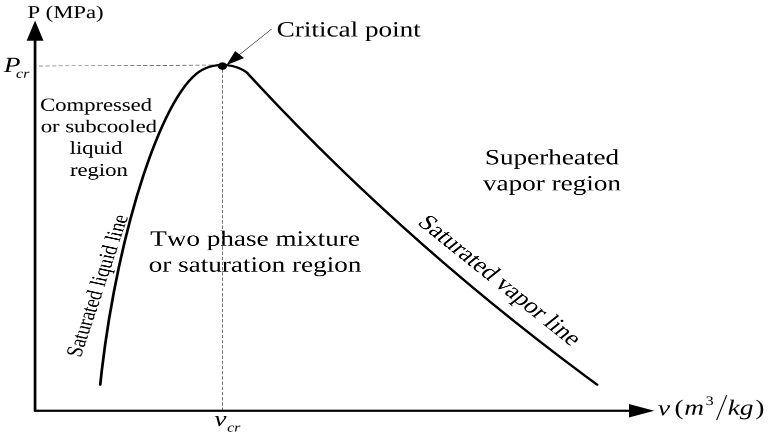

Figure 3.10. Saturation curve for two-phase mixture on $P - v$ diagram

3.3.3. Change of Phase of a Pure Substance on *P–T* Diagram

The P–T diagram of a pure substance is shown in figure 3.11. This diagram is often called the *phase diagram* since all three phases are separated from each other by three lines. The sublimation line separates the solid and vapor regions; the vaporization line separates the liquid and vapor regions, and the melting (or fusion) line separates the solid and liquid regions. These three lines meet at the triple point, where all three phases coexist in equilibrium. The vaporization line ends at the critical point because no distinction can be made between liquid and vapor phases above the critical point. Substances that expand and contract on freezing differ only in the melting line on the *P*–*T* diagram.

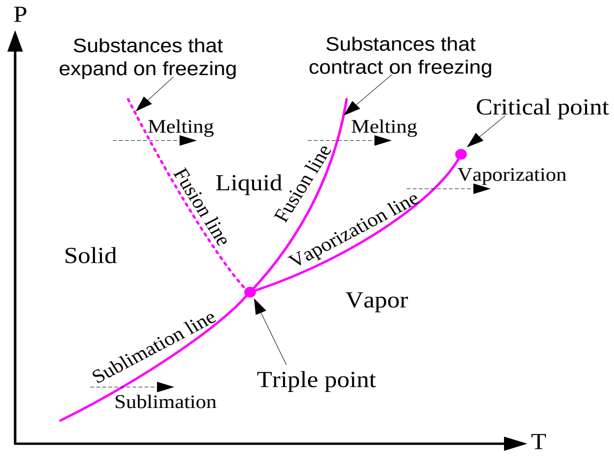

Figure 3.11. *P*–*T* diagram of pure substances [4]

3.3.3. Phase Change Terminologies and Definitions

Saturation Temperature and Saturation Pressure

We have observed that the temperature at which pure substance e.g. water starts boiling depends on the pressure. Therefore, if the pressure is fixed, so is the boiling temperature. Hence, we can define saturation temperature and saturation pressure as below:

At a given pressure, the temperature at which a pure substance changes its phase (e.g., liquid to vapor or vapor to liquid) is known as a saturation temperature. For example, if the liquid water is heated at the pressure of 1 atm (101.325 kPa), vaporization takes place at the temperature of 100^0C, i.e., saturation temperature of the water for 1 atm pressure is 100^0C.

At a given temperature, the pressure at which a pure substance changes its phase (e.g., liquid to vapor or vapor to liquid) is known as a saturation pressure. For example, if the liquid water is expanded at the temperature of 100^0C, vaporization takes place at the pressure of 1 atm (101.325 kPa), i.e., saturation pressure of the water for 100^0C temperature is 1 atm.

Saturated Liquid and Saturated Vapor

A liquid which is about to vaporize is called a saturated liquid. In other words, the state of a pure substance at which vaporization just starts during heating process or condensation just completes during cooling process is known as a saturated liquid.

A vapor which is about to condense is called a saturated vapor. In other words, the state of a pure substance at which vaporization just completes during heating process or condensation just starts during cooling process is known as a saturated vapor.

Subcooled or Compressed Liquid

A liquid which is not about to vaporize is called a subcooled or compressed vapor. In other words, when saturated liquid is further cooled or compressed, it is known as a subcooled or compressed liquid.

Superheated Vapor

A vapor which is not about to condense is called a superheated vapor. In other words, when saturated vapor is further heated or expanded, it is known as a superheated vapor.

Degree of Superheat

The difference in temperature of a superheated vapor and the corresponding saturation temperature is known as a degree of superheat as shown in figure 3.17. Here, the degree of superheat of a superheated vapor at state A is $(T_A - T_{sat})$.

Degree of Subcooling

The difference in the corresponding saturation temperature and the temperature of a sub-cooled or compressed liquid is known as a degree of sub-cooling as shown in figure 3.17. Here, the degree of sub-cooling of a sub-cooled or compressed liquid at state B is $(T_{sat} - T_B)$.

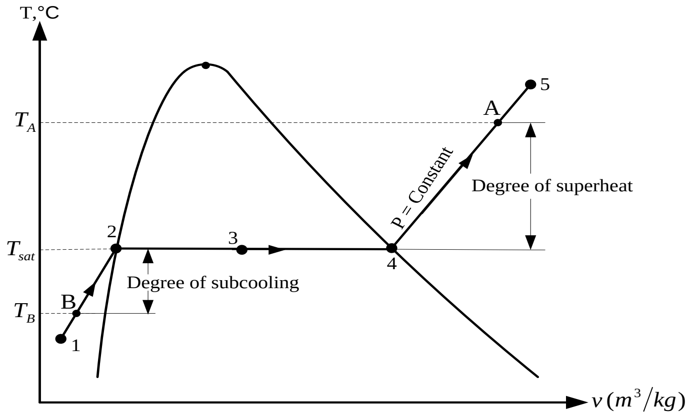

Figure 3.17. Degree of superheat and degree of subcooling

Critical Point

Critical point is defined as a point or a state of a pure substance at which saturated liquid and saturated vapor states are identical (see figure 3.11). The pressure, temperature and specific volume of a pure substance at the critical point are called the critical pressure, the critical temperature and the critical specific volume respectively. For water, critical pressure, critical temperature and critical specific volume are 22.055MPa, 373.95^{0}C and 0.003106 m^3/kg respectively. If the water is heated at the pressure of 22.055 MPa, it suddenly gets converted into saturated vapor from saturated liquid when temperature reaches 373.95^{0}C. Therefore, when a pure substance is heated at its critical pressure, there will be sudden phase change from a saturated liquid state to a saturated vapor state and hence no phase transition interval is observed.

Triple Point

Triple point of a pure substance is defined as the point at which all the three phases of the pure substance namely solid, liquid and vapor coexist in equilibrium. In other words, a point at which the sublimation, fusion and vaporization lines meet to each other is called the triple point as shown in figure 3.11. The triple point of a pure substance exists at a definite pressure and temperature. For water, the pressure and the temperature at the triple point are 0.6117 kPa and 0.01°C respectively.

Sensible Heat

It is the quantity of heat (or enthalpy) required to raise the temperature of 1 kg of water from 0°C to its boiling point (saturation temperature) at a given pressure and is denoted by h_l . Hence, there is a change in the temperature in sensible heating process. Sensible heat increases as the pressure increases. When saturated vapor is further heated at a constant pressure, there is also an increase of temperature of the vapor, but this process is called *superheating* and the additional heat required to convert saturated vapor to superheated vapor is called *heat of superheat* $[c_{pv}(T_{\sup}-T_{sat})]$.

Latent (Hidden) Heat

There are two types of latent heat namely latent heat of fusion and latent heat of vaporization. *Latent heat of fusion* is defined as the quantity of heat required to convert 1 kg of ice into water at constant temperature of 0°C. Whereas *latent het of vaporization* is defined as the quantity of heat required to convert 1 kg of water at saturation temperature for a given pressure into its vapor at the same temperature and pressure and is denoted by h_{lg}. Hence, there is no change in the temperature for a given pressure in latent heating process. Latent heat decreases as the pressure increases.

Lecture Highlights

- *Working substance:* A substance used in any thermodynamic system (device) with the help of which energy is transported between the system and its surroundings is called a working substance. For examples, water, air, refrigerants etc.
- *Pure substance*: A pure substance is of homogeneous in composition, homogenous in chemical aggregation and invariable in chemical aggregation and is also called simple compressible substance.
- *Two property rule (state postulate):* It states that "Any two independent intensive properties are sufficient to specify a thermodynamic equilibrium state of a system composed of a pure substance".
- *Ideal gas:* An ideal gas has very negligible or no force of intermolecular attraction. It follows ideal gas equation or gas laws at all ranges of pressure and temperature. Most of real gases tend to behave an ideal gas at low pressure and high temperature.
- *Boyle's law*: It states that "At a constant temperature, the volume of a given mass of a gas is inversely proportional to the pressure, i.e. $PV = constant$.
- *Charle's law or Gay Lussac's law*: (i) It states that "At constant pressure, the volume of a given mass of a gas is directly proportional to the absolute temperature", i.e., $V/T = constant$. (ii) It can also be stated as "At constant volume, the pressure of a given mass of a gas is directly proportional to the absolute temperature", i.e., $P/T = constant$.
- *Equation of state:* It is the combined gas equation and given by $PV = mRT$ or, $Pv = RT$.
- *Saturation pressure:* For a given temperature, the pressure at which phase change takes place is called saturation pressure, e.g. saturation pressure for 100°C is 1 atm.
- *Saturation temperature:* For a given pressure, the temperature at which phase change takes place is called saturation temperature, e.g., saturation temperature for 1 atm is 100°C.
- *Saturated liquid:* It is the state of a substance at which evaporation just starts during heating or condensation just completes during cooling.
- *Saturated vapor:* It is the state of a substance at which evaporation just completes during heating and condensation just starts during cooling.
- *Sub-cooled or compressed liquid:* When saturated liquid is further cooled below its saturation temperature or compressed above its saturation pressure, it is called sub-cooled or compressed liquid.
- *Superheated vapor:* When saturated vapor is further heated or expanded, it is called a superheated vapor.
- *Two phase mixture:* It is the mixture of saturated liquid and saturated vapor.
- *Degree of superheat:* It is the difference in temperature of superheated vapor and its corresponding saturation temperature.
- *Degree of sub-cooling:* It is the difference in the corresponding saturation temperature and the temperature of a sub-cooled or compressed liquid.
- *Critical point:* It is the meeting point of saturated liquid line and saturated vapor line at which saturated liquid and vapor have identical properties. The pressure, temperature and volume

corresponding to critical point are called critical pressure, critical temperature and critical volume respectively.

- *Triple point:* It is the state of a substance at which the solid, liquid and vapor phase co-exist.
- *Sensible heat:* It is defined as the quantity of heat required to raise the temperature of certain mass of a substance through certain rise in temperature.
- *Latent heat:* The amount of energy absorbed or released during a phase-change process is called the *latent heat***.** More specifically, the amount of energy absorbed during melting is called the *latent heat of fusion* and is equivalent to the amount of energy released during freezing. Similarly, the amount of energy absorbed during vaporization is called the *latent heat of vaporization* and is equivalent to the energy released during condensation.

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