Chapter 2 SUMMARY

- A substance that has a fixed chemical composition throughout is called a *pure substance.*

- A pure substance exists in different phases depending on its energy level. In the liquid phase, a substance that is not about to vaporize is called a *compressed* or *subcooled liquid.*

- In the gas phase, a substance that is not about to condense is called a *superheated vapor.*

- During a phase-change process, the temperature and pressure of a pure substance are dependent properties. At a given pressure, a substance changes phase at a fixed temperature called the *saturation temperature.* At a given temperature, the pressure at which a substance changes phase is called the *saturation pressure.* During a boiling process, both the liquid and the vapor phases coexist in equilibrium, and under this condition the liquid is called *saturated liquid* and the vapor *saturated vapor.*

- In a saturated liquid-vapor mixture, the mass fraction of the vapor phase is called the *quality* and is defined as

$$x = \frac{mass_{vapor}}{m_{total}}$$

- The quality may have values between 0 (saturated liquid) and 1 (saturated vapor). It has no meaning in the compressed liquid or superheated vapor regions.

- In the saturated mixture region, the average value of any intensive property y is determined from

$$y = y_f + x\,y_{fg}$$

  where *f* stands for saturated liquid and *g* for saturated vapor.

- In the absence of compressed liquid data, a general approximation is to treat a compressed liquid as a saturated liquid at the given *temperature,* that is,

$$y = y_{f\,@T}$$

  where *y* stands for *v*, *u*, or *h.*

- The state beyond which there is no distinct vaporization process is called the *critical point.* At supercritical pressures, a substance gradually and uniformly expands from the liquid to vapor phase.

- All three phases of a substance coexist in equilibrium at states along the *triple line* characterized by triple-line temperature and pressure.

- Various properties of some pure substances are listed in the appendix. As can be noticed from these tables, the compressed liquid has lower v, u, and h values than the saturated liquid at the same T or P. Likewise, superheated vapor has higher v, u, and h values than the saturated vapor at the same T or P.

- Any relation among the pressure, temperature, and specific volume of a substance is called an *equation of state.* The simplest and best-known equation of state is the *ideal-gas equation of state,* given as

  $$Pv = RT$$

  where R is the gas constant. Caution should be exercised in using this relation since an ideal gas is a fictitious substance. Real gases exhibit ideal-gas behavior at relatively low pressures and high temperatures.

- The deviation from ideal-gas behavior can be properly accounted for by using the *compressibility factor* Z, defined as

  $$Z = \frac{Pv}{RT} \quad \text{or} \quad Z = \frac{v_{actual}}{v_{ideal}}$$

- The Z factor is approximately the same for all gases at the same *reduced temperature* and *reduced pressure*, which are defined as

  $$T_R = \frac{T}{T_{cr}} \quad \text{and} \quad P_R = \frac{P}{P_{cr}}$$

  where P_{cr} and T_{cr} are the critical pressure and temperature, respectively. This is known as the *principle of corresponding states.* When either P or T is unknown, Z can be determined from the compressibility chart with the help of the *pseudo-reduced specific volume,* defined as

$$v_R = \frac{v_{actual}}{RT_{cr} / P_{cr}}$$

- The P-v-T behavior of substances can be represented more accurately by the more complex equations of state. Three of the best known are

  van der Waals:

$$(P + \frac{a}{v^2})(v - b) = RT$$

  where

$$a = \frac{27R^2T_{cr}^2}{64P_{cr}} \quad \text{and} \quad b = \frac{RT_{cr}}{8P_{cr}}$$

  Beattie-Bridgeman:

$$P = \frac{R_uT}{\overline{v}^2}\left(1 - \frac{c}{\overline{v}T^3}\right)(\overline{v} + B) - \frac{A}{\overline{v}^2}$$

  where

$$A = A_o\left(1 - \frac{a}{\overline{v}}\right) \quad \text{and} \quad B = B_o\left(1 - \frac{b}{\overline{v}}\right)$$

  Benedict-Webb-Rubin:

$$P = \frac{R_uT}{\overline{v}} + \left(B_oR_uT - A_o - \frac{C_o}{T^2}\right)\frac{1}{\overline{v}^2} + \frac{bR_uT - a}{\overline{v}^3} + \frac{a\alpha}{\overline{v}^6} + \frac{c}{\overline{v}^3T^2}\left(1 + \frac{\gamma}{\overline{v}^2}\right)e^{-\gamma/\overline{v}^2}$$

  The constants appearing in the Beattie-Bridgeman and Benedict-Webb-Rubin equations are given in Table A-29 for various substances.

- The amount of energy needed to raise the temperature of a unit of mass of a substance by one degree is called the *specific heat at constant volume* C_v for a constant-volume process and the *specific heat at constant pressure* C_p for a constant pressure process. They are defined as

$$C_v = \left(\frac{\partial u}{\partial T} \right)_v \quad and \quad C_p = \left(\frac{\partial h}{\partial T} \right)_p$$

- For ideal gases u, h, C_v, and C_p are functions of temperature alone. The Δu and Δh of ideal gases can be expressed as

$$\Delta u = u_2 - u_1 = \int_1^2 C_v(T)\, dT \cong C_{v,av}(T_2 - T_1)$$
$$\Delta h = h_2 - h_1 = \int_1^2 C_p(T)\, dT \cong C_{p,av}(T_2 - T_1)$$

- For ideal gases C_v, and C_p are related by

$$C_p = C_v + R \qquad [kJ/(kg \cdot K)]$$

- The *specific heat ratio* k is defined as

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

- For *incompressible substances* (liquids and solids), both the constant-pressure and constant-volume specific heats are identical and denoted by C:

$$C_p = C_v = C \qquad [kJ/(kg \cdot K)]$$

- The Δu and Δh of incompressible substances are given by

$$\Delta u = \int_1^2 C(T)\, dT \cong C_{av}(T_2 - T_1) \qquad (kJ/kg)$$
$$\Delta h = \Delta u + v\Delta P \qquad (kJ/kg)$$