



Higher Technological Institute

Mechanical Engineering Dep.

Principles of Thermodynamics (Lectures)

Dr. Ahmed Abd EL badie

Eng. Mahmoud Osama

Eng. Mohammed Mosaad

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Chapter (1)

Introduction and Basic Concepts

1.1 Introduction:

Thermodynamics can be defined as the science of energy. Although everybody has a feeling of what energy is, it is difficult to give a precise definition for it. Energy can be viewed as the ability to cause changes. The name thermodynamics stems from the Greek words therme (heat) and dynamis (power), which is most descriptive of the early efforts to convert heat into power. Today the same name is broadly interpreted to include all aspects of energy and energy transformations including power generation, refrigeration, and relationships among the properties of matter. One of the most fundamental laws of nature is the **conservation of energy principle** simply states that during an interaction, energy can change from one form to another but the total amount of energy remains constant That is, **energy cannot be created or destroyed**. The first law of thermodynamics is simply an expression of the conservation of energy principle, and it asserts that energy is a thermodynamic property. The second law of thermodynamics asserts that energy has quality as well as quantity, and actual processes occur in the direction of decreasing quality of energy.

1.2 Application Areas of Thermodynamics



Human Body



Power Plant

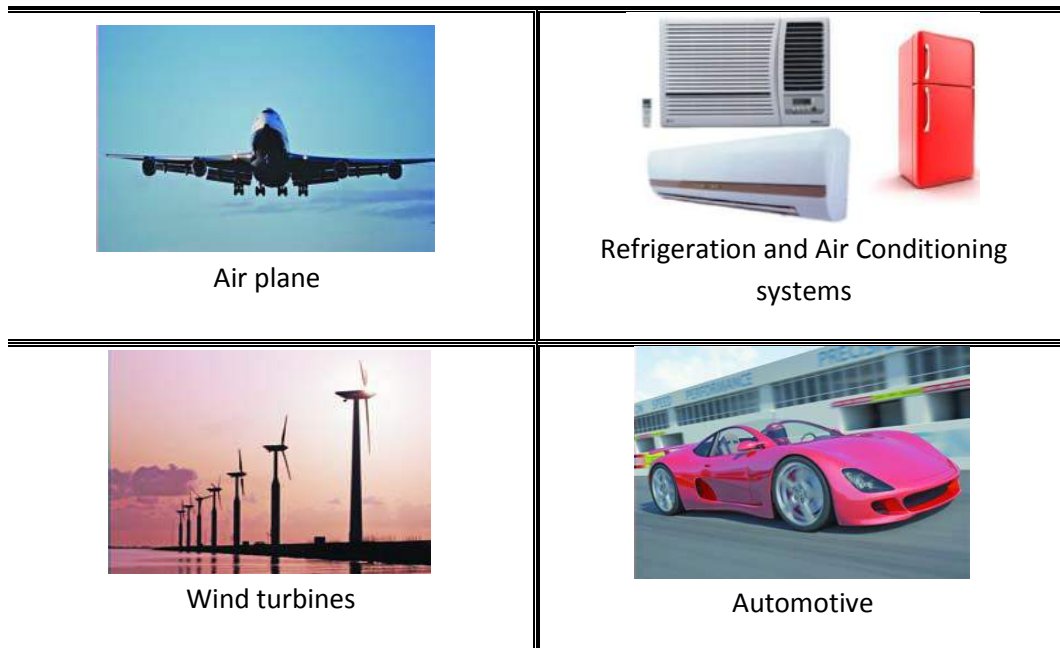


Fig. 1.1

1.3 Systems and Control Volume

A system is defined as a quantity of matter or a region in space chosen for study. The mass or region outside the system is called the **surroundings**. The real or imaginary surface that separates the system from its surroundings is called the **boundary** (Fig. 1.2). The boundary of a system can be fixed or movable.

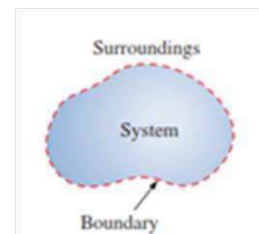


Fig. 1.2

A **closed system** (also known as a **control mass** or **just system when the context makes it clear**) consists of a fixed amount of mass, and no mass can cross its boundary, as shown in Fig. 1.3. But energy, in the form of heat or work, can cross the boundary; and the volume of a closed system does not have to be fixed. If, as a special case, even energy is not allowed to cross the boundary, that system is called an **isolated system**.

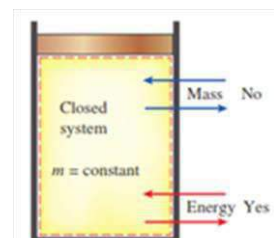


Fig. 1.3

An **open system, or a control volume, as it is often called, is a properly** selected region in space. It usually encloses a device that involves Mass flow such as a compressor, turbine, or nozzle. Flow through these devices is best studied by selecting the region within the device as the Control volume. Both mass and energy can cross the boundary of a control volume (Fig. 1.4).

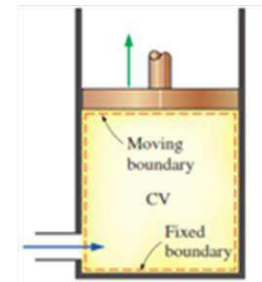


Fig. 1.4

1.4 Properties of A System

Any characteristic of a system is called a **property**. Some familiar properties are pressure P , temperature T , volume V , and mass m . viscosity, thermal conductivity, modulus of elasticity, thermal expansion coefficient, electric resistivity, and even velocity and elevation. Properties are considered to be either intensive or extensive.

Intensive properties are those that are independent of the mass of a system, such as temperature, pressure, and density. And **Extensive properties** are those whose values depend on the size or extent of the system. Total mass, total volume and total momentum are some examples of extensive properties.

An easy way to determine whether a property is intensive or extensive is to divide the system into two equal parts with an imaginary partition, as shown in Fig. 1.5. Each part will have the same value of intensive properties as the original system, but half the value of the extensive properties.

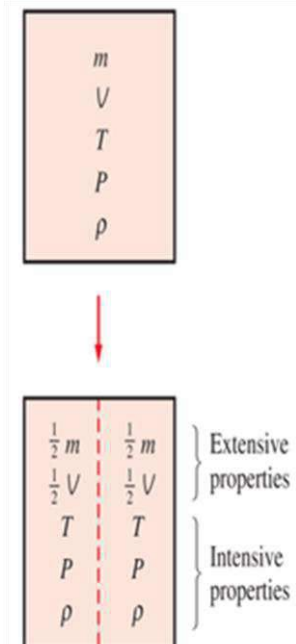


Fig. 1.5

Extensive properties per unit mass are called **specific properties**. Some examples of specific properties are specific volume ($v=V/m$) and specific total energy ($e = E/m$).

The number of properties required to fix the state of a system is given by the **state postulate**. The state of a simple compressible system is completely specified by **two independent, intensive properties**. Two properties are **independent** if one property can be varied while the other one is held constant.

1.5 State and Equilibrium

At a given **State**, All the properties of a system have fixed values. If the value of even one property changes, the state will change to a different one. The word **Equilibrium** implies a state of balance. A system in equilibrium experiences no changes when it is isolated from its surroundings. There are many types of equilibrium, and a system is not in thermodynamic equilibrium unless the conditions of all the relevant types of equilibrium are satisfied. **Thermal equilibrium** occurs when the system involves no temperature differential, which is the driving force for heat flow. **Mechanical equilibrium** happens when there is no change in pressure at any point of the system with time. **Phase equilibrium** occurs if a system involves two phases. And **Chemical equilibrium** occurs if its chemical composition does not change with time.

1.6 Processes and Cycles

Any change that a system undergoes from one equilibrium state to another is called a **process** (Fig. 1.6). The series of states through which a system passes during a process is called the **path of the process**. To describe a process

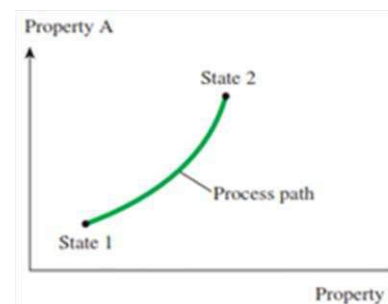


Fig. 1.6

completely, one should specify the initial and final states of the process, as well as the path it follows, and the interactions with the surroundings.

When a process proceeds in such a manner that the system remains infinitesimally close to an equilibrium state at all times, it is called a **quasi-static**, or **quasi-equilibrium, process** (Fig. 1.7).

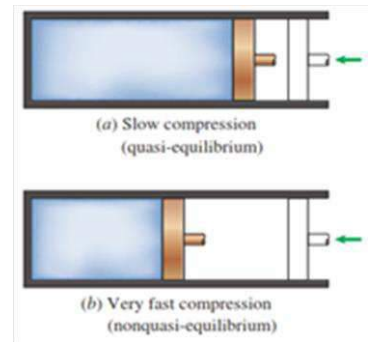


Fig. 1.7

A **quasi-equilibrium process** can be viewed as a sufficiently slow process that 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.

A system is said to have undergone a cycle if it returns to its initial state at the end of the process as shown in Fig. 1.8. That is, for a cycle the initial and final states are identical.

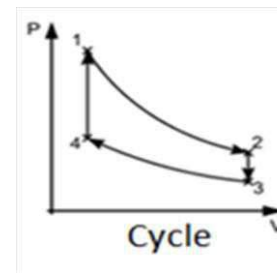


Fig. 1.8

1.7 The Steady-Flow Process

The term steady implies no change with time. The opposite of steady is unsteady, or transient. The term uniform, however, implies no change with location over a specified region.

A large number of engineering devices operate for long periods of time under the same conditions, and they are classified as steady-flow devices.

Processes involving such devices can be represented

reasonably well by a somewhat idealized process, called the **steady-flow process, which can be** defined as a process during which a fluid flows through a control volume steadily (Fig. 1.9).

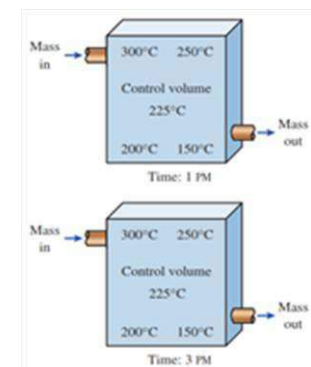


Fig. 1.9

1.8 Units

In 1960, the General Conference of Weights and Measures (CGPM) produced the SI, which was based on six fundamental quantities, and their units were adopted in 1954 at the Tenth General Conference of Weights and Measures: meter (m) for length, kilogram (kg) for mass, second (s) for time, ampere

(A) for electric current, degree Kelvin ($^{\circ}\text{K}$) for temperature, and candela (cd).

TABLE 1-1

The seven fundamental (or primary) dimensions and their units in SI

Dimension	Unit
Length	meter (m)
Mass	kilogram (kg)
Time	second (s)
Temperature	kelvin (K)
Electric current	ampere (A)
Amount of light	candela (cd)
Amount of matter	mole (mol)

TABLE 1-2

Term	Law	Unit
Area	Length(m) * Length(m)	m^2
Volume	Length (m) Width (m) Height (m)	m^3
Velocity	Length (m) / Time (s)	m/s
Acceleration	Velocity (m/s) /Time (s)	m/s^2
Force	Mass (kg) * Acceleration (m/s^2)	Newton (N)
Work	Force (N) * Distance (m)	Joule (J)
Power	Work(J)/Time (s)	Watt (W)
Pressure	Force (N) / Area (m^2)	Pascal (Pa)

In 1971, the CGPM added a seventh fundamental quantity and unit: mole (mol) for the amount of matter. Based on the notational scheme, the degree symbol was officially dropped from the absolute temperature unit, and all unit names were to be written without capitalization even if they were derived from proper names. The SI is based on a decimal relationship between units. The prefixes used to express the multiples of the various units are listed in Table 1-3.

TABLE 1 - 3			
Standard prefixes in SI units			
Multiple	Prefix	Multiple	Prefix
10^{24}	yotta, Y	10^{-1}	deci, d
10^{21}	zetta, Z	10^{-2}	centi, c
10^{18}	exa, E	10^{-3}	milli, m
10^{15}	peta, P	10^{-6}	micro, μ
10^{12}	tera, T	10^{-9}	nano, n
10^9	giga, G	10^{-12}	pico, p
10^6	mega, M	10^{-15}	femto, f
10^3	kilo, k	10^{-18}	atto, a
10^2	hecto, h	10^{-21}	zepto, z
10^1	deka, da	10^{-24}	yocto, y

1.9 Temperature and The Zeroth Law of Thermodynamics

1.9.1 Thermal Equilibrium

Heat is transferred from the body at higher temperature to the one at lower temperature until both bodies attain the same temperature (Fig. 1.10). At that point, the heat transfer stops, and the two bodies are said to have reached **thermal equilibrium**.

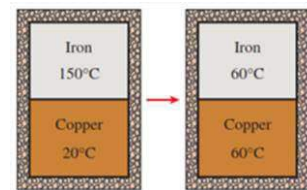


Fig. 1.10

The **zeroth law of thermodynamics** states that two bodies are in thermal equilibrium if both have the same temperature reading even if they are not in contact.

1.9.2 Temperature Scales

The temperature scales used in the SI and in the English system today are:

1.9.2.1 Two-Point Scales

Temperature values are assigned at two different points (Ice point and Steam point), The temperature scales used in the SI and in the English system today are the **Celsius scale (or centigrade scale)**; the ice point and

steam point values are 0 and 100°C respectively in Celsius scale; while in **Fahrenheit scale** which is an **English Scale**. The ice and steam points values is 32 and 212°F respectively.

1.9.2.2 Thermodynamics Temperature Scales

A temperature scale that is independent of the properties of any substance or substances. The thermodynamic temperature scale in the SI is **Kelvin scale**, the temperature unit on this scale is the **kelvin** which is designated by K (not °K; the degree symbol was officially dropped from kelvin); the lowest temperature on the Kelvin scale is absolute zero, or 0 K. Then it follows that only one nonzero reference point needs to be assigned to establish the slope of this linear scale. Using nonconventional refrigeration techniques, scientists have approached absolute zero kelvin so they achieved 0.000000002 K.

The thermodynamic temperature scale in the English system is the **Rankine scale**; the temperature unit on this scale is the **Rankine**, which is designated by R.

1.9.2.3 Ideal Gas Temperature Scales

The temperatures on this scale measured using a constant-volume gas thermometer which is based on the principle that at low pressures, the temperature of a gas is proportional to its pressure at constant volume (Fig. 1.11).

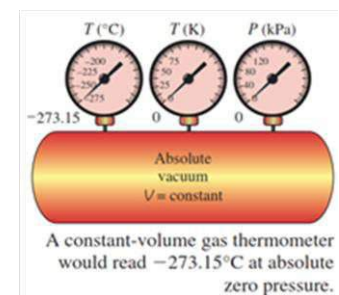


Fig. 1.11

1.9.2.4 Relation Between Temperature Scales

The Kelvin scale is related to the Celsius scale by

$$T(\text{K}) = T(^{\circ}\text{C}) + 273.15 \quad (1-1)$$

The Rankine scale is related to the Fahrenheit scale by

$$T(R) = T(^{\circ}F) + 459.67 \quad (1-2)$$

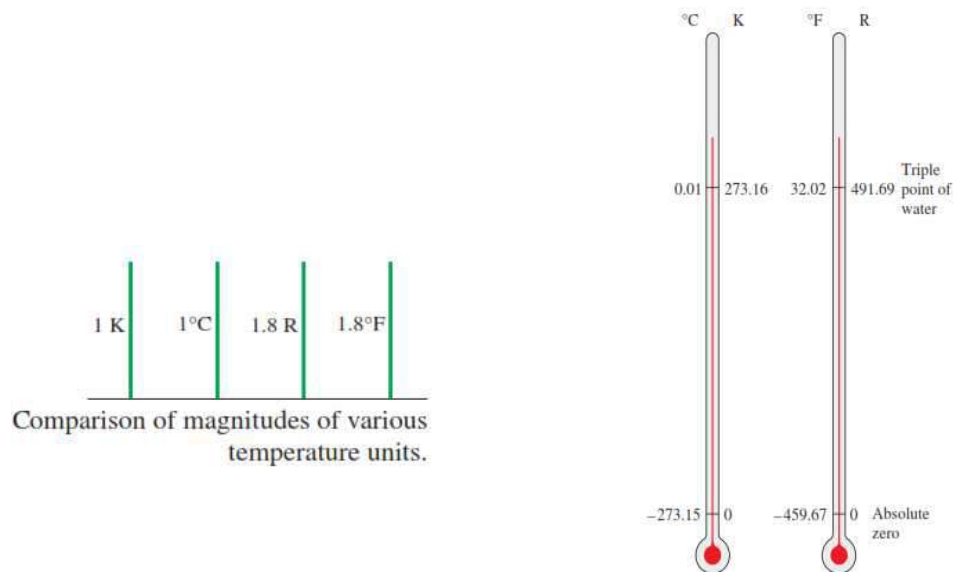


Fig. 1.12

The temperature scales in the two unit systems are related by

$$T(R) = 1.8T(K) \quad (1-3)$$

$$T(^{\circ}F) = 1.8T(^{\circ}C)+32 \quad (1-4)$$

The temperature difference in SI scales are equal and also in English scales

$$\Delta T(K) = \Delta T(^{\circ}C) \quad (1-5)$$

$$\Delta T(R) = \Delta T(^{\circ}F) \quad (1-6)$$

A comparison of various temperature scales is given in Fig. 1.12.

1.10 Pressure

Pressure is defined as a normal force exerted by a fluid per unit area.

1.10.1 Pressure Units

The Pressure unit is Pascal (Pa)= N/m². Pascal is too small for most pressures encountered in practice. Therefore, its multiples kilopascal

(kPa= 10^3 Pa) and Megapascal (MPa= 10^6 Pa). Three other pressure units commonly used in practice, especially in Europe, are Bar, Standard atmosphere (atm) and Kilogram-force per square centimeter (kgf/cm²).

Where

$$1 \text{ bar} = 10^5 \text{ Pa} = 0.1 \text{ MPa} = 100 \text{ kPa}$$

$$1 \text{ atm} = 101.325 \text{ kPa} = 1.01325 \text{ bars}$$

$$1 \text{ kgf/cm}^2 = 9.807 \times 10^4 \text{ Pa} = 0.9807 \text{ bar} = 0.9679 \text{ atm}$$

In the English system, the pressure unit is pound-force per square inch (lbf/in², or psi), and 1 atm = 14.696 psi.

1.10.2 Pressure Scales

Absolute pressure, it is measured relative to absolute vacuum (i.e., absolute zero pressure). Most pressure-measuring devices, however, are calibrated to read zero in the atmosphere, **Gage pressure** (P_{gage}) which is difference between the absolute pressure and the local atmospheric pressure. P_{gage} can be positive or negative, and **Vacuum a pressure** which is pressures below atmospheric pressure and are measured by vacuum gages that indicate the difference between the atmospheric pressure and the absolute pressure.

Absolute, gage, and vacuum pressures are related to each other by

$$P_{\text{abs}} = P_{\text{atm}} + P_{\text{gage}} \quad (1-7)$$

$$P_{\text{abs}} = P_{\text{atm}} - P_{\text{vac}} \quad (1-8)$$

This is illustrated in Fig. 1.13.

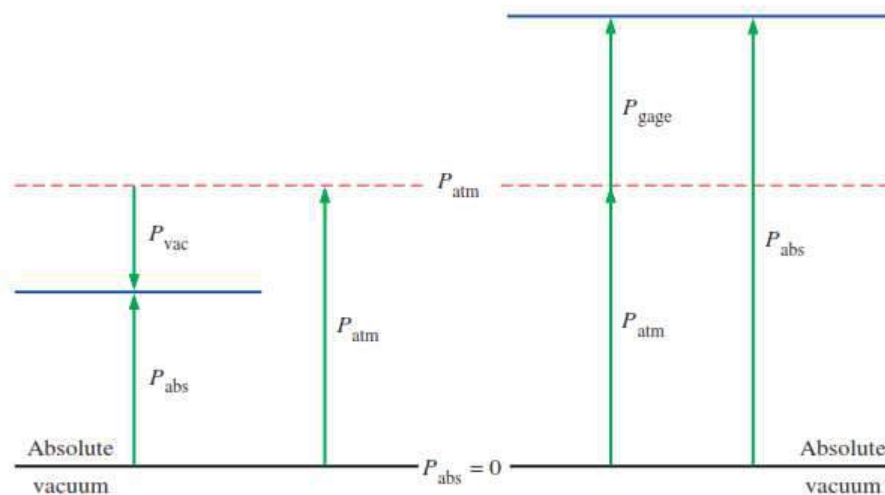


Fig. 1.13

1.10.3 Pressure Measurement Device

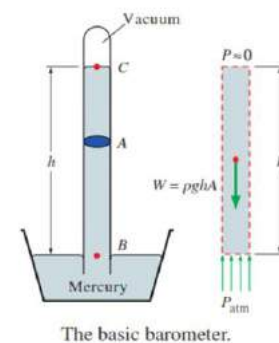
1.10.3.1 The Barometer

Atmospheric pressure is measured by a device called a barometer (Fig.1.14); thus, the atmospheric pressure is often referred to as the *barometric pressure*.

$$P_{\text{atm}} = \rho gh \quad (1-9)$$

g: gravitational acceleration,

ρ : density & h: head

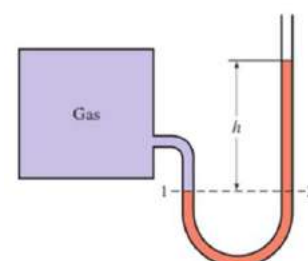


The basic barometer.

Fig. 1.14

1.10.3.2 The Manometer

An elevation change of Δz in a fluid at rest corresponds to $\Delta P/\rho g$, which suggests that a fluid column can be used to measure pressure differences. A device based on this principle is called a manometer (Fig. 1.15), and it is commonly used to measure small and moderate pressure differences.



The basic manometer.

Fig. 1.15

$$\Delta P = P_2 - P_1 = -\rho g \Delta z = -\gamma_s \Delta z \quad (1-10)$$

1.10.3.3 Bourdon Tube

Which consists of a bent, coiled, or twisted hollow metal tube whose end is closed and connected to a dial indicator needle as shown in Fig. 1.16, when the tube is open to the atmosphere, the tube is undeflected, and the needle on the dial at this state is calibrated to read zero (gage pressure). When the fluid inside the tube is pressurized, the tube stretches and moves the needle in proportion to the applied pressure.

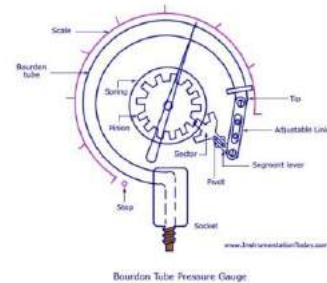
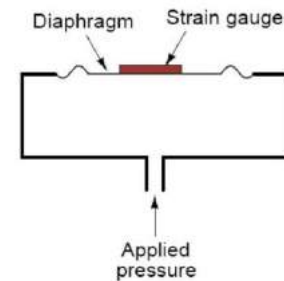


Fig. 1.16

1.10.3.4 Pressure Transducers

Modern pressure sensors which use various techniques to convert the pressure effect to an electrical effect such as a change in voltage, resistance, or capacitance. Strain-gage pressure transducers (Fig. 1.17a) Works by having a diaphragm deflect between two chambers open to the pressure inputs. As the diaphragm stretches in response to a change in pressure difference across it, the strain gage stretches (change in resistance) and a Wheatstone bridge circuit amplifies the output, Piezoelectric transducers (Fig. 1.17b) also called solid-state pressure transducers, work on the principle that an electric potential is generated



(a)



(b)

Fig. 1.17

in a crystalline substance when it is subjected to mechanical pressure. This phenomenon is called the piezoelectric (or press-electric) effect.

1.10.3.5 Deadweight Tester

Is used primarily for *calibration* and can measure extremely high pressures. As its name implies, a deadweight tester (Fig. 1.18) measures pressure directly through application of a weight that provides a force per unit area (the fundamental definition of pressure).

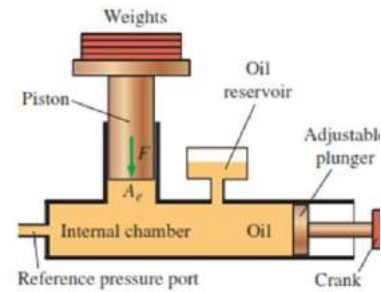


Fig. 1.18

Chapter (2)

Energy, Energy Transfer, and General Energy Analysis

2.1 Form of Energy

2.1.1 Energies of The System

a) Macroscopic:

Forms of energy are those a system possesses as a whole with respect to some outside reference frame, such as **Potential Energy (P.E)** which is The energy that a system possesses as a result of its elevation in a gravitational field, and **Kinetic Energy (K.E)** which is The energy that a system possesses as a result of its motion relative to some reference frame.

$$\Delta KE = \frac{1}{2} m \Delta C^2 \text{ (J)} \quad (2-1)$$

$$\Delta PE = mg \Delta z \text{ (J)} \quad (2-2)$$

Where **m** is the mass, **C** is the velocity and **g** is the gravitational acceleration possesses as a result of its motion relative to some reference frame.

b) Microscopic (Internal Energy (U)):

The sum of all the microscopic forms of energy which are those related to the molecular structure of a system and the degree of the molecular activity, and they are independent of outside reference frames as shown in Fig.2.1. Forms of internal energy are:

- 1) **Sensible energy:** the internal energy of a system associated with the kinetic energies of the molecules.
- 2) **Latent energy:** The internal energy associated with the phase of a system (Keeps the system in a certain Phase).

3) Chemical energy: The internal energy associated with the atomic bonds in a molecule.

$$\text{Total system energy (E)} = U + \text{K.E} + \text{P.E}$$

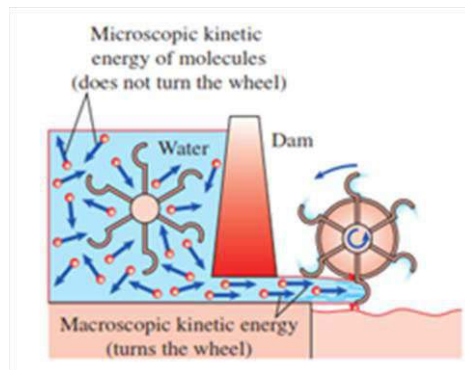


Fig. 2.1

2.1.2 The Dynamic Forms of Energy

This is the form of energies which are recognized at the system boundary as they cross it, and they represent the energy gained or lost by a system during a process. The only two forms of energy interactions associated with a closed system are heat transfer and work as shown in Fig. 2.2. Hot body contains energy, but this

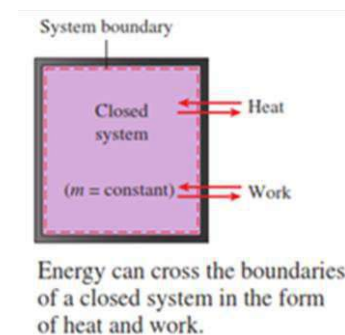


Fig. 2.2

energy is heat transfer only as it passes through the skin of the body (the system boundary) to reach the air. Once in the surroundings, the transferred heat becomes part of the internal energy of the surroundings.

2.1.3 Mechanical Energy

It is the form of energy that can be converted to mechanical work completely and directly by an ideal mechanical device such as an ideal turbine. Kinetic and potential energies are the familiar forms of mechanical energy. Thermal energy is not mechanical energy, however, since it cannot

be converted to work directly and completely (the second law of thermodynamics).

A pump transfer's mechanical energy to a fluid by raising its pressure, and a turbine extracts mechanical energy from a fluid by dropping its pressure. Therefore, the pressure of a flowing fluid is also associated with its mechanical energy.

In fact, the pressure unit Pa is equivalent to $\text{Pa} = \text{N}/\text{m}^2 = \text{N} \cdot \text{m}/\text{m}^3 = \text{J}/\text{m}^3$ which is energy per unit volume, and the product Pv or its equivalent P/ρ has the unit J/kg. Note that pressure itself is not a form of energy but a pressure force acting on a fluid through a distance produces work, called *flow work*; it is convenient to view it as part of the energy of a flowing fluid and call it *flow energy*. Therefore, the mechanical energy of a flowing fluid can be expressed on a unit mass basis as

$$E_{\text{mech}} = \frac{P}{\rho} + \frac{C^2}{2} + gZ \quad (\text{per unit mass}) \quad (2-3)$$

The mechanical energy of a fluid does not change during flow if its pressure, density, velocity, and elevation remain constant, in the absence of any irreversible losses.

2.2 Energy Transfer by Heat (Q)

Heat is defined as the form of energy that is transferred between two systems (or a system and its surroundings) by virtue of a temperature difference. Then it follows that there cannot be any heat transfer between two systems that are at the same temperature. In thermodynamics, the term heat simply means heat transfer. Heat (Q) has energy units, kJ (or Btu).

Heat transferred per unit mass of a system is denoted q

where

$$q = \frac{Q}{m} \quad (\text{kJ}/\text{kg}) \quad (2-4)$$

The rate of heat transfer(\dot{Q}): The amount of heat transferred per unit time (unit kJ/s = kW)

Where

$$\dot{Q} = Q/\Delta t \text{ (kJ/s} \rightarrow \text{kW)}$$

2.2.1 Some heat related definitions

- **Body heat:** is the thermal energy content of a body.
- **Heat flow:** is the transfer of thermal energy.
- **Heat addition:** the transfer of heat into a system.
- **Heat rejection:** the transfer of heat out of a system.

- **Adiabatic Process:**

A process during which there is no heat transfer. There are two ways a process can be adiabatic; either the system is well insulated so that only a negligible amount of heat can pass through the boundary, or both the system and the surroundings are at the same temperature and therefore there is no driving force (temperature difference) for heat transfer.

An adiabatic process should not be confused with an isothermal process. Even though there is no heat transfer during an adiabatic process, the energy content and thus the temperature of a system can still be changed by other means such as work.

2.2.2 Heat transfer mechanisms

Heat is transferred by three mechanisms: conduction, convection, and radiation.

Conduction: is the transfer of energy from the more energetic particles of a substance to the adjacent less energetic ones as a result of interaction between particles.

Convection: is the transfer of energy between a solid Surface and the adjacent fluid that is in motion, and it involves the combined Effects of conduction and fluid motion.

Radiation: is the transfer of energy Due to the emission of electromagnetic waves (or photons).

2.3 Energy Transfer by Work (W)

If the energy crossing the boundary of a closed system is not heat, it must be work. Work is the energy transfer associated with force acting through a distance. Work has energy units such kJ. The work done per unit mass of a system is denoted by w and is expressed as

$$w = \frac{W}{m} \text{ (kJ/kg)} \quad (2-5)$$

The work done per unit time is called power and is denoted \dot{W} (kJ/s = kW).

2.3.1 Mechanical Forms of Work

a) Work done by a force F on a body displaced a distance s (Fig. 2.3).

$$W = \int_1^2 F ds \text{ (kJ)} \quad (2-6)$$

if F is a constant force then

$$W = F \cdot s \text{ (kJ)} \quad (2-7)$$

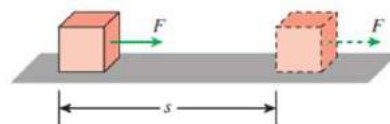


Fig. 2.3

b) **Shaft Work:** It is the energy transmission with a rotating shaft

When there is a force (F) acting through a moment arm (r) as shown in Fig. 2.4, it will generate a torque (T) so that

$$T = F \cdot r \rightarrow F = \frac{T}{r} \quad (2-8)$$

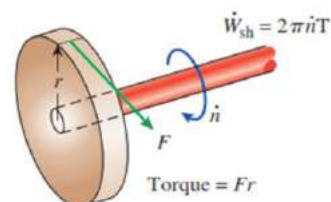


Fig. 2.4

This force acts through a distance (s), which is related to the radius (r) by

$s = (2\pi r)n$ where n is the number of revolutions Then the shaft work is determined from

$$W_{sh} = Fs = \left(\frac{T}{r}\right) (2\pi rn) = 2\pi nT \quad (\text{kJ}) \quad (2-9)$$

The power transmitted through the shaft

$$\dot{W}_{sh} = 2\pi \dot{n}T \quad (\text{kW}) \quad (2-10)$$

Where \dot{n} is the number of revolutions per unit time

c) Spring Work: When a force is applied on a spring changing the spring length. When the length of the spring changes by a differential amount dx under the influence of a force F , the work done is

$$\delta W_{spring} = Fdx \quad \text{and} \quad F = kx \quad (2-11)$$

The displacement x is measured from the undisturbed position of the spring.

$$W = \int_1^2 kx \, dx \quad \rightarrow \quad W = 0.5 k(x_2^2 - x_1^2) \quad (2-12)$$

2.3.2 Non Mechanical Forms of Work

Non mechanical work modes encountered in practice. However, these non-mechanical work modes can be treated in a similar manner by identifying a *generalized force* F acting in the direction of a *generalized displacement* x . Then the work associated with the differential displacement under the influence of this force is determined from $\delta W = F \delta x$.

a) Electric Work:

In an electric field, electrons in a wire move under the effect of electromotive forces, doing work (Fig. 2.5).

$$W_e = V I \Delta t \quad (\text{kJ}) \quad (2-13)$$

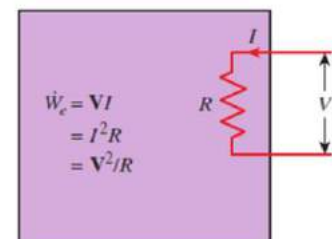


Fig. 2.5

Where V is the potential difference, I is the current and Δt is Time interval.

The electrical work done per unit time is **Electrical Power** (\dot{W}_e)

$$\dot{W}_e = VI = I^2R = V^2/R \quad (W) \quad (2-14)$$

b) Magnetic Work:

Where the generalized force is the magnetic field strength and the generalized displacement is the total magnetic dipole moment.

c) Electrical polarization work:

Where the generalized force is the electric field strength and the generalized displacement is the polarization of the medium (the sum of the electric dipole rotation moments of the molecules).

2.4 The Ideal-Gas Equation of State

Any equation that relates the pressure, temperature, and specific volume of a substance is called an equation of state. Property relations that involve other properties of a substance at equilibrium states are also referred to as equations of state.

The simplest and best-known equation of state for substances in the gas phase is the ideal-gas equation of state. This equation predicts the P - v - T behavior of a gas quite accurately within some properly selected region. It was experimentally determined that at low pressures the volume of a gas is proportional to its temperature. That is,

$$P \propto \frac{1}{v} \quad \text{and} \quad P \propto T \quad ; \text{ so } P \propto \frac{T}{v}$$
$$P = \frac{RT}{v} \quad \rightarrow \quad Pv = RT \quad (2-15)$$

Where

R: is called the **gas constant**,

The Previous equation is called ideal-gas equation of state; a gas that obeys this relation is called an **ideal gas**.

The gas constant R is different for each gas and is determined from

$$R = \frac{R_u}{M} \quad (\text{kJ/kg} \cdot \text{K}) \quad (2-16)$$

R_u (kJ/kMole) is the **Universal gas constant** and M is the molar mass (molecular weight).

R , R_u and M can be found from **Table A-1**

The mass of a system is equal to the product of its molar mass M and the mole number N :

$$m = MN \quad (\text{kg}) \quad (2-17)$$

$$V = mv \rightarrow PV = mRT \quad (2-18)$$

$$mR = (MN)R = NR_u \rightarrow PV = NR_u T \quad (2-19)$$

$$V = N\bar{v} \rightarrow P\bar{v} = R_u T \quad (2-20)$$

\bar{v} is the molar specific volume (the volume per unit mole) as shown in Fig. 2.6.

The properties of an ideal gas at two different states are related to each other by

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad (2-21)$$

Per unit mass	Per unit mole
$v, \text{m}^3/\text{kg}$	$\bar{v}, \text{m}^3/\text{kmol}$
$u, \text{kJ}/\text{kg}$	$\bar{u}, \text{kJ}/\text{kmol}$
$h, \text{kJ}/\text{kg}$	$\bar{h}, \text{kJ}/\text{kmol}$

Fig. 2.6

An **ideal gas** is an imaginary substance that obeys the relation $Pv = RT$. It has been experimentally observed that the ideal-gas relation given closely approximates the P-v-T behavior of **real gases at low densities**.

At low pressures and high temperatures, the density of a gas decreases, and the gas behaves as an ideal gas under these conditions. What constitutes low pressure and high temperature is explained later. In the range of practical interest, many familiar gases such as **air, nitrogen, oxygen, hydrogen, helium, argon, neon, krypton**, and even heavier gases such as **carbon dioxide** can be treated as ideal gases with negligible error (often less than 1 percent). Dense gases such as **water vapor** in steam power plants and

refrigerant vapor in refrigerators, however, **should not be treated as ideal gases**. Instead, the property tables should be used for these substances.

Important Note:

For Pressures below 10 kPa, Water vapor can be treated as an ideal gas, regardless of its temperature, with negligible error (less than 0.1 percent). At higher pressures, however, the ideal gas assumption yields unacceptable errors, particularly in the vicinity of the critical point and the saturated vapor line (over 100 percent).

Therefore, in air-conditioning applications, the water vapor in the air can be treated as an ideal gas with essentially no error since the pressure of the water vapor is very low.

Example 2.1

The gage pressure of an automobile tire is measured to be 210 kPa before a trip and 220 kPa after the trip at a location where the atmospheric pressure is 95 kPa. Assuming the volume of the tire remains constant and the air temperature before the trip is 25°C, determine air temperature in the tire after the trip.

Solution

The pressure in an automobile tire is measured before and after a trip. The temperature of air in the tire after the trip is to be determined.

$$\begin{aligned} P_1 &= P_{\text{gage},1} + P_{\text{atm}} = 210 + 95 = 305 \text{ kPa} \\ P_2 &= P_{\text{gage},2} + P_{\text{atm}} = 220 + 95 = 315 \text{ kPa} \end{aligned}$$

Note that air is an ideal gas and the volume is constant, the air temperatures after the trip is determined to be

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \longrightarrow T_2 = \frac{P_2}{P_1} T_1 = \frac{315 \text{ kPa}}{305 \text{ kPa}} (25 + 273 \text{ K}) = 307.8 \text{ K} = 34.8^\circ\text{C}$$

Therefore, the absolute temperature of air in the tire will increase by 6.9% during this trip.

2.5 Compressibility Factor

The ideal-gas equation is very simple and thus very convenient to use. However, the gases deviate from ideal-gas behavior significantly at states near the saturation region and the critical point.

This deviation from ideal-gas behavior at a given temperature and pressure can accurately be

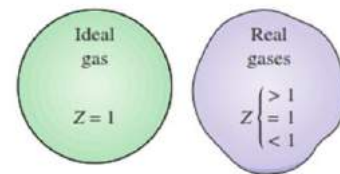


Fig. 2.7

accounted for by the introduction of a correction factor called the compressibility factor Z defined as

$$Z = \frac{Pv}{RT} \quad (2-22)$$

Or

$$Pv = ZRT \quad (2-23)$$

It can be also expressed as

$$Z = \frac{V_{actual}}{V_{ideal}} \quad (2-24)$$

where $v_{ideal} = RT/P$. Obviously, $Z = 1$ for ideal gases. For real gases Z can be greater than or less than unity (Fig. 2.7). The farther away Z is from unity, the more the gas deviates from ideal-gas behavior.

Problems (Equation of State)

- 2.1 Determine the mass of the air in a room whose dimensions are $4\text{ m} \times 5\text{ m} \times 6\text{ m}$ at 100 kPa and 25°C .
- 2.2 A spherical balloon with a diameter of 6 m is filled with helium at 20°C and 200 kPa . Determine the mole number and the mass of the helium in the balloon.
- 2.3 A 1-m^3 tank containing air at 25°C and 500 kPa is connected through a valve to another tank containing 5 kg of air at 35°C and 200 kPa . Now the valve is opened, and the entire system is allowed to reach thermal equilibrium with the surroundings, which are at 20°C . Determine the volume of the second tank and the final equilibrium pressure of air.
- 2.4 The pressure in an automobile tire depends on the temperature of the air in the tire. When the air temperature is 25°C , the pressure gage reads 210 kPa . If the volume of the tire is 0.025 m^3 , determine the pressure rise in the tire when the air temperature in the tire rises to 50°C . Also, determine the amount of air that must be bled off to restore pressure to its original value at this temperature. Assume the atmospheric pressure is 101 kPa .
- 2.5 The pressure gage on a 2.5 m^3 oxygen tank reads 500 kPa . Determine the amount of oxygen in the tank if the temperature is 28°C and the atmospheric pressure is 98 kPa .
- 2.6 A 0.5-m^3 rigid tank containing hydrogen at 20°C and 600 kPa is connected by a valve to another 0.5-m^3 rigid tank that holds hydrogen at 30°C and 150 kPa . Now the valve is opened and the system is allowed to reach thermal equilibrium with the surroundings, which are at 15°C . Determine the final pressure in the tank.

2.6 The Moving boundary work (Work Boundary):

One form of mechanical work frequently encountered in practice is associated with the expansion or compression of a gas in a piston–cylinder device. During this process, part of the boundary (the inner face of the piston) moves back and forth. Therefore, the expansion and compression work is often called **moving boundary work**, or simply **boundary work**. Consider the gas enclosed in the piston–cylinder device shown in Fig. 2.8. The initial pressure of the gas is P , the total volume is V , and the cross sectional area of the piston is A . If the piston is allowed to move a distance ds in a quasi-equilibrium manner, the differential work done during this process is

$$\delta W_b = F ds = PA ds = P dV \quad (2-25)$$

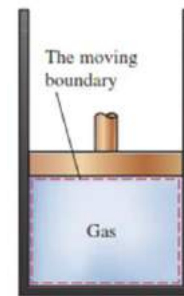
The total boundary work done during the entire process as the piston moves is obtained by adding all the differential works from the initial state to the final state:

$$W_b = \int_1^2 PA ds = \int_1^2 P dV \quad (2-26)$$

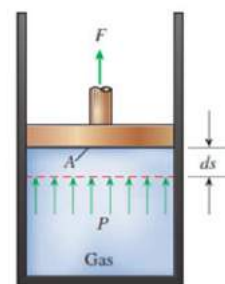
This equation reveals that the area under the process curve on a P-V diagram (Fig. 2.9) is equal, in magnitude, to the work done during a quasi-equilibrium expansion or compression process of a closed system.

a) Boundary Work for a Constant-Volume Process:

Since a rigid tank has a constant volume and $dV = 0$ in this equation. Therefore, there is no boundary work done during this process. That is, the



(a)



(b)

Fig. 2.8

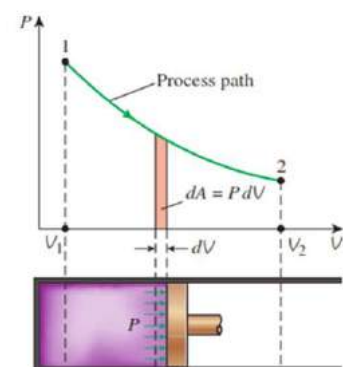


Fig. 2.9

boundary work done during a constant-volume process is always zero. This is also evident from the P - V diagram of the process as shown in Fig. 2.10 (the area under the process curve is zero).

$$W_b = \int_1^2 P \, dV = 0 \quad \because dV = 0$$

$$W_b = 0 \quad (2-27)$$

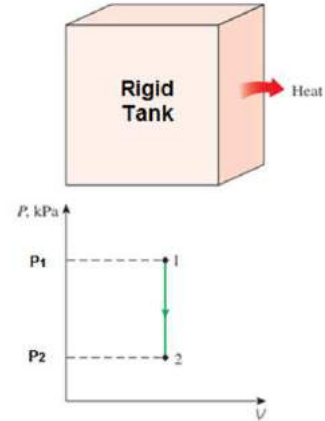


Fig. 2.10

b) Boundary Work for a Constant-Pressure Process:

The fluid in a piston cylinder device is heated /cooled and the temperature rises/decreases at constant pressure which leads to increase/decrease in the volume as shown in Fig. 2.11.

$$W_b = \int_1^2 P \, dV = P_0 \int_1^2 dV$$

$$W_b = P_0(V_2 - V_1) = mP_0(v_2 - v_1) \quad (2-28)$$

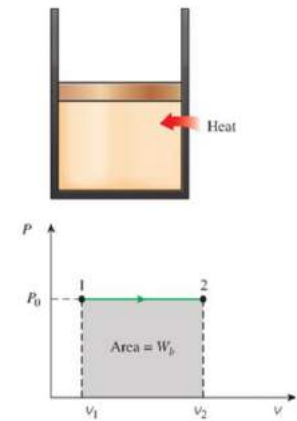


Fig. 2.11

c) Boundary Work for a Constant-Temperature (Isothermal) Process:

Air in a piston-cylinder device is compressed isothermally. For an ideal gas at constant temperature T_0 as shown in Fig. 2.12

$$PV = mRT_0 = C \quad \text{or} \quad P = \frac{C}{V} \quad (2-29)$$

$$\frac{P_1}{P_2} = \frac{V_2}{V_1} \quad (2-30)$$

$$C = P_1V_1 = P_2V_2 = mRT_0 \quad (2-31)$$

Where C is a constant.

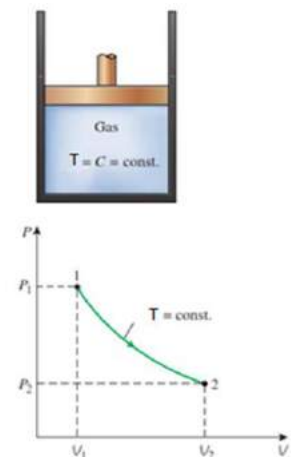


Fig. 2.12

$$\begin{aligned}
 W_b &= \int_1^2 P \, dV = \int_1^2 \frac{C}{V} \, dV = C \int_1^2 \frac{dV}{V} \\
 &= C \ln \left(\frac{V_2}{V_1} \right) = C \ln \left(\frac{P_1}{P_2} \right) \quad (2-32)
 \end{aligned}$$

d) Boundary Work for Polytropic Process:

During actual expansion and compression processes of gases, pressure and volume are often related by $PV^n = C$, where n and C are constants. A process of this kind is called a **Polytropic process**. Below we develop a general expression for the work done during a Polytropic process. The pressure for a Polytropic process can be expressed as $P = CV^{-n}$, Where $C = P_1 V_1^n = P_2 V_2^n$.

$$\begin{aligned}
 W_b &= \int_1^2 P \, dV = \int_1^2 CV^{-n} dV \\
 &= C \frac{V_2^{-n+1} - V_1^{-n+1}}{-n+1} = \frac{P_2 V_2 - P_1 V_1}{1-n} \\
 &= \frac{mR(T_2 - T_1)}{1-n} \quad n \neq 1 \quad (\text{kJ}) \quad (2-33)
 \end{aligned}$$

For the special case of $n = 1$ the boundary work becomes

$$W_b = \int_1^2 P \, dV = \int_1^2 CV^{-1} dV = C \ln \left(\frac{V_2}{V_1} \right)$$

For an ideal gas this result is equivalent to the isothermal process discussed before.

Problems (Work Boundary)

- 2.7 A piston–cylinder device initially contains 0.4 m^3 of air at 100 kPa and 80°C . The air is now compressed to 0.1 m^3 in such a way that the temperature inside the cylinder remains constant. Determine the work done during this process.
- 2.8 A frictionless piston–cylinder device contains 2 kg of nitrogen at 100 kPa and 300 K . Nitrogen is now compressed slowly according to the relation $PV^{1.4} = \text{constant}$ until it reaches a final temperature of 360 K . Calculate the work input during this process.
- 2.9 A gas is compressed from an initial volume of 0.42 m^3 to a final volume of 0.12 m^3 . During the quasi-equilibrium process, the pressure changes with volume according to the relation $P = aV + b$, where $a = -1200 \text{ kPa/m}^3$ and $b = 600 \text{ kPa}$. Calculate the work done during this process (a) by plotting the process on a P - V diagram and finding the area under the process curve and (b) by performing the necessary integrations.
- 2.10 Nitrogen at an initial state of 300 K , 150 kPa , and 0.2 m^3 is compressed slowly in an isothermal process to a final pressure of 800 kPa . Determine the work done during this process.
- 2.11 A piston–cylinder device contains 0.05 m^3 of a gas initially at 200 kPa . At this state, a linear spring that has a spring constant of 150 kN/m is touching the piston but exerting no force on it. Now heat is transferred to the gas, causing the piston to rise and to compress the spring until the volume inside the cylinder doubles. If the cross-sectional area of the piston is 0.25 m^2 , determine (a) the final pressure inside the cylinder, (b) the total work done by the gas, and (c) the fraction of this work done against the spring to compress it.

2.7 The First Law of Thermodynamics

The first law of thermodynamics is often viewed as a statement of **the conservation of energy principle**, provides a sound basis for studying the relationships among the various forms of energy and energy interactions.

Based on experimental observations, the first law of thermodynamics states that **energy can be neither created nor destroyed during a process; it can only change forms**. Therefore, every bit of energy should be accounted for during a process.

For all adiabatic processes between two specified states of a closed system, the net work done is the same regardless of the nature of the closed system and the details of the process.

The first law of thermodynamics simply states that the change in the total energy during an adiabatic process must be equal to the net work done.

Consider a well-insulated (i.e., adiabatic) room heated by an electric heater as our system as shown in Fig. 2.13, the conservation of energy principle dictates that the electrical work done on the adiabatic system must equal the increase in energy of the system.

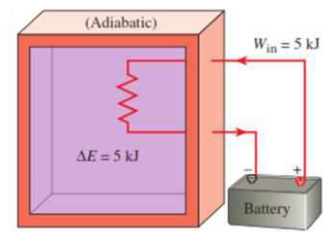


Fig. 2.13

Next, let us replace the electric heater with a paddle wheel. As a result of the stirring process (Fig. 2.14); the energy of the system will increase. Again, since there is no heat interaction between the system and its surroundings ($Q = 0$), the shaft work done on the system must show up as an increase in the energy of the system.

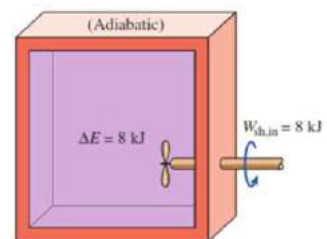


Fig. 2.14

For an Adiabatic Piston cylinder device (Fig. 2.15); the temperature of air rises when it is compressed. This is because energy is transferred to the

air in the form of boundary work. In the absence of any heat transfer ($Q = 0$), the entire boundary work will be stored in the air as part of its total energy. The conservation of energy principle again requires that the increase in the energy of the system be equal to the boundary work done on the system.

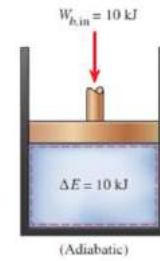


Fig. 2.15

If a system involves various heat and work interactions simultaneously as shown in Fig. 2.16. For example, if a system gains 12 kJ of heat during a process while 6 kJ of work is done on it, the increase in the energy of the system

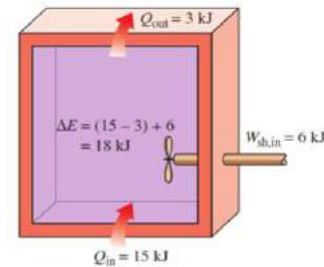


Fig. 2.16

during that process is 18 kJ. That is, the change in the energy of a system during a process is simply equal to the net energy transfer to (or from) the system.

2.7.1 Energy Balance

The conservation of energy principle can be expressed as follows: The net change (increase or decrease) in the total energy of the system during a process is equal to the difference between the total energy entering and the total energy leaving the system during that process.

That is,

$$\left(\begin{array}{c} \text{Total Energy} \\ \text{entering the system} \end{array} \right) - \left(\begin{array}{c} \text{Total Energy} \\ \text{leaving the system} \end{array} \right) = \left(\begin{array}{c} \text{Change in the total} \\ \text{energy of the system} \end{array} \right)$$

$$E_{in} - E_{out} = \Delta E_{system} \quad (2-34)$$

2.7.2 Energy Change of a System, ΔE system

The determination of the energy change of a system during a process involves the evaluation of the energy of the system at the beginning and at the end of the process, and taking their difference. That is,

Energy Change = Energy at Final State – Energy at Initial State

$$\Delta E_{\text{system}} = E_{\text{final}} - E_{\text{initial}} = E_2 - E_1 \quad (2-35)$$

The change in the total energy of a system during a process is the sum of the changes in its internal, kinetic, and potential energies and can be expressed as

$$\begin{aligned} \Delta E_{\text{system}} &= E_2 - E_1 = (U_2 + KE_2 + PE_2) - (U_1 + KE_1 + PE_1) \\ &= \Delta U + \Delta KE + \Delta PE \end{aligned} \quad (2-36)$$

Where

$$\Delta U = m(u_2 - u_1) \quad (\text{kJ}) \quad (2-37a)$$

$$\Delta KE = \frac{1}{2000} m(C_2^2 - C_1^2) \quad (\text{kJ}) \quad (2-37b)$$

$$\Delta PE = mg \left(\frac{Z_2 - Z_1}{1000} \right) \quad (\text{kJ}) \quad (2-37c)$$

The values of the specific internal energies u_1 and u_2 can be determined directly from the property tables or thermodynamic property relations.

For stationary systems (do not involve any changes in their velocity or elevation during a process), the changes in kinetic and potential energies are zero (that is, $\Delta KE + \Delta PE = 0$), and the total energy change relation in Thus:

$$\Delta E_{\text{system}} = \Delta U \quad (2-38)$$

2.7.3 Mechanisms of Energy Transfer, E_{in} and E_{out}

Energy can be transferred to or from a system in three forms: **heat**, **work**, and **mass flow** (Fig. 2.17).

The only two forms of energy interactions associated with a fixed mass or closed system are heat transfer and work.

Heat and work are directional quantities, and thus the complete description of a heat or work interaction requires the specification of both the magnitude and direction. One way of doing that is to adopt a sign convention.

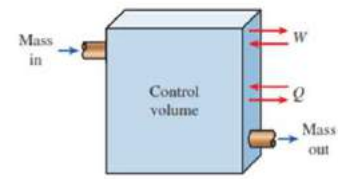


Fig. 2.17

The generally accepted **formal sign convention** for heat and work interactions is as follows (Fig. 2.18): **heat transfer to a system** and **work done by a system** are **positive**; **heat transfer from a system** and **work done on a system** are **negative**.

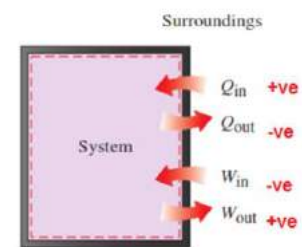


Fig. 2.18

The work done on the system is increasing the systems energy; the Negative sign is only a formal convention and it doesn't mean that the energy decrease, while work done by the system decreases the systems energy; and it has positive sign as a formal convention also.

Heat and work are energy transfer mechanisms between a system and its surroundings, and there are many similarities between them:

1. Both are recognized at the boundaries of a system as they cross the boundaries. That is, both heat and work are boundary phenomena.
2. Systems possess energy, but not heat or work.
3. Both are associated with a process, not a state. Unlike properties, heat or work has no meaning at a state.
4. Both are path functions (i.e., their magnitudes depend on the path followed during a process as well as the end states).

Mass flow in and out of the system serves as an additional mechanism of energy transfer. When mass enters a system, the energy of the system increases because mass carries energy with it (in fact, mass is energy). Likewise, when some mass leaves the system, the energy contained within the system decreases because the leaving mass takes out some energy with it

$$E_{in} - E_{out} = (Q_{in} - Q_{out}) + (W_{in} - W_{out}) + (E_{mass,in} - E_{mass,out})$$

$$= \Delta E_{system} \quad (2-39a)$$

, in the **rate form**

$$E_{in} - E_{out} = (\dot{Q}_{in} - \dot{Q}_{out}) + (\dot{W}_{in} - \dot{W}_{out}) + (\dot{E}_{mass,in} - \dot{E}_{mass,out})$$

$$= dE_{system}/dt \quad (2-39b)$$

For steady operating conditions $dE_{system}/dt = 0$

For constant rates, the total quantities during a time interval Δt are related to the quantities per unit time as

$$Q = \dot{Q}\Delta t, \quad W = \dot{W}\Delta t, \quad \text{and} \quad E = (dE/dt)\Delta t \quad (2-40)$$

The energy balance can be expressed on a per unit mass basis as

$$\Delta e_{system} = e_{in} - e_{out} \quad (\text{kJ/kg}) \quad (2-41)$$

For a closed system undergoing a cycle, the initial and final states are identical, and thus system $\Delta E_{system} = E_2 - E_1 = 0$. Then the energy balance for a cycle simplifies to

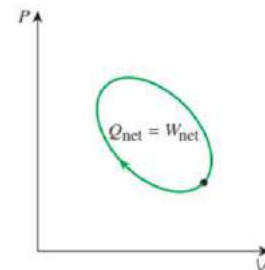


Fig. 2.19

$$E_{in} - E_{out} = 0, \text{ or } E_{in} = E_{out} \quad (2-42)$$

Noting that a closed system does not involve any mass flow across its boundaries, the energy balance for a cycle can be expressed in terms of heat and work interactions as out

$$W_{net,out} = Q_{net,in} \quad \text{or} \quad \dot{W}_{net,out} = \dot{Q}_{net,in} \quad (\text{For a cycle}) \quad (2-43)$$

That is, the **network output** during a cycle is equal to **net heat input** as shown in Fig. 2.19.

To Use the formal convention Signs of the heat and work the first law is restated as

$$\begin{aligned} \sum Q - \sum W &= \Delta E_{\text{system}} \\ (\mathbf{Q}_{\text{in}} + (-\mathbf{Q}_{\text{out}})) - ((-\mathbf{W}_{\text{in}}) + \mathbf{W}_{\text{out}}) &= \Delta E_{\text{system}} \end{aligned} \quad (2-44)$$

2.7.4 Enthalpy

The combination of properties $u + Pv$ are frequently encountered so for the sake of simplicity and convenience, this combination is defined as a new property, **enthalpy**, and given the symbol h :

$$h = u + Pv \quad (\text{kJ/kg}) \quad (2-45a)$$

$$H = U + PV \quad (\text{kJ}) \quad (2-45b)$$

Both H and h are called enthalpy since the context clarifies which one is meant. The term Pv are the amount of energy required to shove the surrounding out of the way and make room for the systems Pressure and volume. And sometimes it called flow energy.

2.7.5 First law of Thermodynamics for a closed system

$$\sum Q - \sum W = \Delta E_{\text{system}} = \Delta U + \Delta KE + \Delta PE \quad (2-46a)$$

Closed system is always stationary system so $\Delta KE \ \& \ \Delta PE = 0$

$$\sum Q - \sum W = \Delta U = m \Delta u \quad (2-46b)$$

Example 2.2

A rigid tank contains a hot fluid that is cooled while being stirred by a paddle wheel. Initially, the internal energy of the fluid is 800 kJ. During the cooling process, the fluid loses 500 kJ of heat, and the paddle wheel does 100 kJ of

work on the fluid. Determine the final internal energy of the fluid. Neglect the energy stored in the paddle wheel.

Solution

$$\sum Q - \sum W = \Delta E_{\text{system}} = \Delta U + \Delta KE + \Delta PE$$

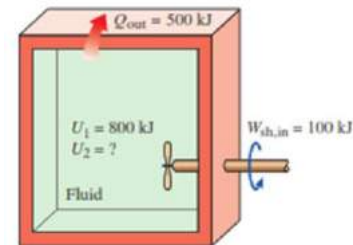
Closed system is always stationary system
so ΔKE & $\Delta PE = 0$

$$\sum Q - \sum W = \Delta U$$

$$- Q_{\text{out}} - (-W_P) = (U_2 - U_1)$$

$$-500 - (-100) = (U_2 - 800)$$

$$U_2 = 400 \quad (\text{KJ})$$



Note: For Piston cylinder at constant pressure:

Piston cylinder devices have moving boundaries (due to The expansion or compression of the gas) which described previously as a boundary work W_b that equals $mP\Delta v$ under a constant pressure process; this amount of work can be added to the change in internal energy $m \Delta u$; so that:

$$\sum Q - \sum W = \Delta U = m \Delta h \quad (2-47a)$$

Hence

$$\Delta h = \Delta u + P\Delta v = \Delta u + W_b \quad (2-47b)$$

2.7.6 First law of Thermodynamics for an open system

Since the open system is a dynamic system so that the system is flowing; the rate form of the first law is more efficient for open systems; and also since the boundary of the system are always moving so that enthalpy h are used instead of internal energy u since;

$$\Delta h = \Delta u + \int Pdv ; \text{ then}$$

$$\begin{aligned}\Sigma \dot{Q} - \Sigma \dot{W} &= \Delta H + \Delta KE + \Delta PE \\ &= \dot{m} \left(\Delta h + \frac{C_2^2 - C_1^2}{2000} + \frac{g\Delta z}{1000} \right) \quad (\text{kJ/s})\end{aligned} \quad (2-48)$$

Example 2.3

A fan that consumes 20 W of electric power when operating is claimed to discharge air from a ventilated room at a rate of 1.0 kg/s at a discharge velocity of 8 m/s. Determine if this claim is reasonable.

Solution

$$\Sigma \dot{Q} - \Sigma \dot{W} = \Delta H + \Delta KE + \Delta PE$$

$$\dot{Q} \& C_1 \& \Delta KE \& \Delta PE = 0 \quad \text{so}$$

$$-(-\dot{W}_{in}) = \dot{m} \left(\frac{C_2^2 - C_1^2}{2000} \right)$$

$$-\left(-\frac{20}{1000}\right) = 1 \left(\frac{C_2^2}{2000} \right)$$

$$C_2 = 6.3 \text{ m/s}$$

which is less than 8 m/s. Therefore, the claim is false.

Problems (First Law of Thermodynamics)

(1st law Closed system)

- 2-12 A rigid tank contains a hot fluid that is cooled while being stirred by a paddle wheel. Initially, the internal energy of the fluid is 800 kJ. During the cooling process, the fluid loses 500 kJ of heat, and the paddle wheel does 100 kJ of work on the fluid. Determine the final internal energy of the fluid.
- 2-13 A mass of 15kg of air in a piston-cylinder device is heated from 25 °C to 77 °C so that the internal energy is changed from $u_1 = 299$ kJ/kg to $u_2 = 250.02$ kJ/kg. The pressure inside the cylinder is held constant at 300 kPa during the process, and a heat loss of 60 kJ occurs if the device was heated by an electric heater. Determine the electric energy supplied in the heater.
- 2.14 A $4 \times 5 \times 7$ -m room is heated by the radiator of a steam heating system. The steam radiator transfers heat at a rate of 10000 kJ/hr and a 100 W fan is used to distribute the warm air in the room. The rate of heat loss from the room is estimated to be about 5000 kJ/hr. If the initial temperature of the room air is 10°C ($u_1 = 203.33$ kJ/kg). Determine how long it will take for the air temperature to rise to 20°C ($u_2 = 210$ kJ/kg).

(1st law open system)

- 2-15 Steam at enthalpy ($h_1 = 3433.8$ kJ/kg) enter an adiabatic nozzle steadily with a velocity of 80 m/s and leaves at enthalpy ($h_2 = 3247.6$ kJ/kg). Determine the exit velocity.

- 2-16 Air at enthalpy ($h_1 = 400.98 \text{ kJ/kg}$) enter a diffuser at velocity of $C_1 = 230 \text{ m/s}$ and leaves at enthalpy ($h_2 = 426.8 \text{ kJ/kg}$). The air is estimated to lose heat at a rate of 18 kJ/s during this process. Determine the exit velocity if the mass flow rate is 2.5 kg/s .
- 2-17 Steam flows steadily through an adiabatic turbine. The inlet condition of steam is ($h_1 = 3240.9 \text{ kJ/kg}$), $C_1 = 80 \text{ m/s}$. the exit condition is ($h_2 = 2393.2 \text{ kJ/kg}$) and $C_2 = 50 \text{ m/s}$. The mass flow rate of steam is 3 kg/s . determines the change in K.E and the power output.
- 2-18 Air enters the compressor of a gas turbine plant at enthalpy of ($h_1 = 298.18 \text{ kJ/kg}$) with a low velocity and exit at ($h_2 = 628.07 \text{ kJ/kg}$) with a velocity of 90 m/s . the compressor is cooled at a rate of 1500 kJ/min , and the power input to the compressor is 250 kW . Determine the mass flow rate of air through the compressor.
- 2-19 A hot water stream at enthalpy 334.9 kJ/kg enters a mixing chamber with a mass flow rate of 0.5 kg/s . where it is mixed with a stream of cold water at enthalpy of 83.96 kJ/kg . If it desired that the mixture leaves the chamber at enthalpy of 175.92 kJ/kg . Determine the mass flow rate of the cooled water steam.
- 2-20 Water is heated in an insulated constant diameter tube by a 7-kW electric resistance heater. If the water enters the heater steadily at enthalpy of ($h_1 = 60 \text{ kJ/kg}$) and leaves at ($h_2 = 280 \text{ kJ/kg}$), Determine the mass flow rate of the water.

2.8 Specific Heats

The **specific heat** is defined as the energy required to raise the temperature of a unit mass of a substance by one degree. In general, this energy depends on how the process is executed.

In thermodynamics, we are interested in two kinds of specific heats: **specific heat at constant volume** c_v and **specific heat at constant pressure** c_p .

Physically, the specific heat at constant volume c_v can be viewed as the energy required to raise the temperature of the unit mass of a substance by one degree as the volume is maintained constant. The energy required to do the same as the pressure is maintained constant is the specific heat at constant pressure c_p . The specific heat at constant pressure c_p is always greater than c_v because at constant pressure the system is allowed to expand and the energy for this expansion work must also be supplied to the system.

The first law of a **constant-volume process** can be expressed in the differential form as

$$\delta q - \delta w = du \quad (2-49)$$

$$\delta w = 0 \quad (\text{No change in volume})$$

From the definition of c_v , this energy must be equal to $c_v dT$, where dT is the differential change in temperature. Thus $du = c_v dT$ at constant volume.

$$c_v = \left(\frac{du}{dT} \right)_{v=c} \quad (\text{kJ/kg.K or kJ/kg.}^\circ\text{C}) \quad (2-50)$$

$$\Delta u = \int_1^2 c_v dT \quad (2-51)$$

Similarly, an expression for the specific heat at constant pressure c_p can be obtained by considering a constant-pressure expansion or compression process. By substituting in eqn. 2-49

$$c_p dT - Pdv = du \quad (2-52)$$

$$c_p dT = du + Pdv = dh \quad (2-53)$$

$$c_p = \left(\frac{dh}{dT} \right)_{P=c} \quad (\text{kJ/kg.K or kJ/kg.}^\circ\text{C})$$

$$\Delta h = \int_1^2 c_p dT \quad (2-54)$$

a measure of the variation of enthalpy of a substance with temperature. Both the internal energy and enthalpy of a substance can be changed by the transfer of energy in any form, with heat being only one of them.

The term specific energy is probably more appropriate than the term specific heat, which implies that energy is transferred (and stored) in the form of heat.

2.9 Internal Energy, Enthalpy, and Specific Heats of Ideal Gases

Using the definition of enthalpy and the equation of state of an ideal gas, we have

$$\left. \begin{array}{l} h = u + Pv \\ Pv = RT \end{array} \right\} h = u + RT \quad (2-55)$$

$$dh = du + RdT \quad (2-56)$$

$$c_p dT = c_v dT + RdT \quad (2-57)$$

$$\therefore c_p = c_v + R \quad (2-58)$$

Another ideal-gas property called the **specific heat ratio** k

$$k = \frac{c_p}{c_v} \quad (2-59)$$

2.10 Evaluation of Δu and Δh of an Ideal Gas:

2.10.1 Using the empirical data from the gas table:

$$\Delta h = (h_2 - h_1) \quad (\text{kJ/kg}) \quad (2-60)$$

$$\Delta u = (u_2 - u_1) \quad (\text{kJ/kg}) \quad (2-61)$$

Where **h** and **u** can be found from the gas data table (Table A17 – A-22)

2.10.2 Using the function form of Specific heat:

From equation 2-54

$$\Delta h = \int_{T_1}^{T_2} c_p \, dT$$

from table (A-2c)

$$\bar{c}_p = a + b T + c T^2 + d T^3 \quad \text{kJ/kmole.K}$$

$$c_p = \frac{1}{M} (a + b T + c T^2 + d T^3) \quad \text{kJ/kg.K} \quad (2-62)$$

Substitute equation 2-62 into equation 2-54

$$\Delta h = \int_{T_1}^{T_2} \frac{1}{M} (a + b T + c T^2 + d T^3) \, dT \quad (2-63)$$

$$\Delta u = \int_{T_1}^{T_2} c_v \, dT, \quad R = c_p - c_v, \quad \text{and} \quad R = \frac{R_u}{M} \quad (2-64)$$

$$\Delta u = \int_{T_1}^{T_2} \frac{1}{M} (a + b T + c T^2 + d T^3 - R_u) \, dT \quad (2-65)$$

Where **a**, **b**, **c** and **d** are constant can be found from **Table (A-2c)**.

2.10.3 Using Average Specific heat:

$$\Delta h = \int_{T_1}^{T_2} c_p dT = c_{p \text{ avg}} (T_2 - T_1) \quad (2-66)$$

$$\Delta u = \int_{T_1}^{T_2} c_v dT = c_{v \text{ avg}} (T_2 - T_1) \quad (2-67)$$

Where $c_{p \text{ avg}}$ and $c_{v \text{ avg}}$ can be found from **Table (A-2b)** at T_{avg} (average temperature).

2.10.4 Using specific heat value at Room temperature:

$$\Delta h = \int_{T_1}^{T_2} c_p dT = c_p (T_2 - T_1) \quad (2-68)$$

$$\Delta u = \int_{T_1}^{T_2} c_v dT = c_v (T_2 - T_1) \quad (2-69)$$

Where c_p and c_v can be found from **Table (A-2a)**

Problems (Internal energy and enthalpy change of ideal gases)

- 2-21 Air at 300 K and 200 kPa is heated at constant pressure to 600 K. Determine the change in internal energy Δu of air per unit mass, using (a) data from the air table (Table A–17), (b) the functional form of the specific heat (Table A–2c), (c) the average specific heat value (Table A–2b), and (d) the specific heat value at room temperature (Table A–2a). Also, determine the percentage error involved in each case.
- 2-22 Determine the enthalpy change Δh of air, in kJ/kg, as it is heated from 600 to 1000 K, using (a) the empirical data from the air table (Table A–17), (b) the empirical specific heat equation as a function of temperature (Table A–2c), (c) the C_p value at the average temperature (Table A–2b), and (d) the C_p value at room temperature (Table A–2a). Also, determine the percentage error involved in each case.
- 2-23 Determine the internal energy change Δu of hydrogen, in kJ/kg, as it is heated from 300 to 800 K, using (a) the empirical data from the hydrogen table (Table A–22), (b) the empirical specific heat equation as a function of temperature (Table A–2c), (c) the C_v value at the average temperature (Table A–2b), and (d) the C_v value at room temperature (Table A–2a). Also, determine the percentage error involved in each case.
- 2-24 Determine the enthalpy change Δh of nitrogen, in kJ/kg, as it is heated from 600 to 1000 K, using (a) the empirical data from the nitrogen table (Table A–18), (b) the empirical specific heat equation as a function of temperature (Table A–2c), (c) the C_p value at the average temperature (Table A–2b), and (d) the C_p value at room temperature (Table A–2a). Also, determine the percentage error involved in each case.

Chapter (3)

Properties of Pure substances

3.1 Pure substances

A substance that has a fixed chemical composition throughout is called a **pure substance**. Water, nitrogen, helium, and carbon dioxide, for example, are all pure substances.

A pure substance does not have to be of a **single chemical element or compound**, however. A mixture of **various chemical elements or compounds** also qualifies as a pure substance as long as **the mixture is homogeneous**. Air, for example, is a mixture of several gases, but it is often considered to be a pure substance because it

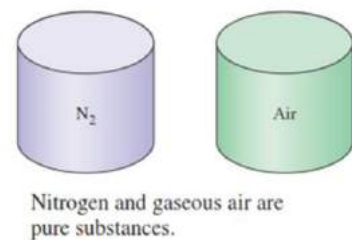


Fig. 3.1

has a uniform chemical composition as shown in Fig. 3.1. A mixture of oil and water is not a pure substance. Since oil is not soluble in water, it will collect on top of the water, forming two chemically dissimilar regions.

A **mixture of two or more phases** of a pure substance is still a pure substance as long as the chemical composition of all phases is the same. A mixture of ice and liquid water, for example, is a pure substance because both phases have the same chemical composition. A mixture of liquid air and gaseous air, however, is not a pure substance

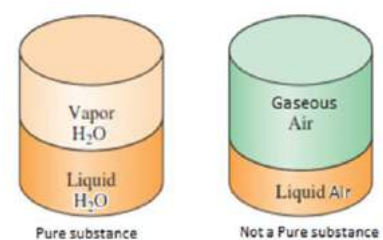


Fig. 3.2

since **the composition of liquid air is different from the composition of gaseous air**, and thus the mixture is no longer chemically homogeneous as shown in Fig. 3.2. This is due to different components in air condensing at different temperatures at a specified pressure.

3.2 Phases of A Pure substances

Substances exist in different phases at room temperature and pressure, copper is a solid, mercury is a liquid, and nitrogen is a gas. Under different conditions, each may appear in a different phase.

There are three principal phases (solid, liquid, and gas) a substance may have several phases within a principal phase, A phase is identified as having a distinct molecular arrangement that is homogeneous throughout and separated from the others by easily identifiable boundary surfaces. When studying phases or phase changes in thermodynamics, one does not need to be concerned with the molecular structure and behavior of different phases.

Molecules in a **solid**, the attractive forces of molecules on each other are large and keep the molecules at fixed positions. At sufficiently high temperatures, the velocity (and thus the momentum) of the molecules may reach a point where the intermolecular forces are partially overcome and groups of molecules break away. This is the beginning of the melting process (Fig. 3.3). The molecular spacing in the **liquid** phase is not much different from that of the solid phase, except the molecules are not at fixed positions relative to each other and they can rotate and translate freely. In the **gas** phase, the molecules are far apart from each other, and a molecular order is nonexistent. Gas molecules move about at random, continually colliding with each other and the walls of the container they are in.

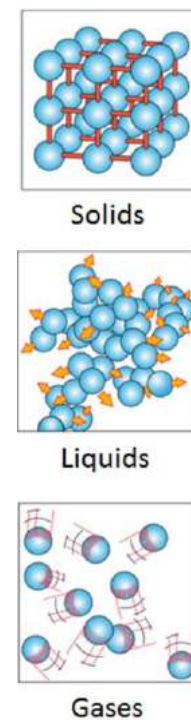


Fig. 3.3

Molecules in the gas phase are at a considerably higher energy level than they are in the liquid or solid phases. Therefore, the gas must release a large amount of its energy before it can condense or freeze.

3.3 Phase-Change of Pure substances

There are many practical situations where two phases of a pure substance coexist in equilibrium. Water exists as a mixture of liquid and vapor in the boiler and the condenser of a steam power plant. The refrigerant turns from liquid to vapor in the freezer of a refrigerator. Even though many home owners consider the freezing of water in underground pipes as the most important phase-change process, attention in this section is **focused on the liquid and vapor phases and their mixture**. As a familiar substance, water is used to demonstrate the basic principles involved. Remember, however, **that all pure substances exhibit the same general behavior**.

Consider a piston–cylinder device containing liquid water at 20°C and 1 atm pressure as shown in Fig. 3.4. Under these conditions, water exists in the liquid phase, and it is called a **Compressed liquid**, or a **Subcooled liquid**, meaning that the liquid is not about to vaporize (liquid below vaporization temperature).

As more heat is transferred, the temperature keeps rising until it reaches 100°C. At this point water is still a liquid, but any heat addition will cause some of the liquid to vaporize. That is, a phase-change process from liquid to vapor is about to take place. A liquid that is *about to vaporize* is called a **saturated liquid**. Therefore, state 2 is a saturated liquid state.

Once boiling starts, the temperature stops rising until the liquid is completely vaporized. That is, the temperature will remain constant during the entire phase-change process if the pressure is held constant. During a boiling process, the only change we will observe is a large increase in the volume and a steady decline in the liquid level as a result of more liquid turning to vapor. Midway about the vaporization line (state 3), the cylinder contains equal amounts of liquid and vapor, once a small amount of the liquid turned to a vapor substance at this state is referred to as a **saturated**

liquid–vapor mixture since the *liquid and vapor phases coexist* in equilibrium at these states.

As we continue transferring heat, the vaporization process continues until the last drop of liquid is vaporized (state 4). At this point, the entire cylinder is filled with vapor that is on the borderline of the liquid phase. Any heat loss from this vapor will cause some of the vapor to condense (phase change from vapor to liquid). A vapor that is *about to condense* is called a **saturated vapor**.

Once the phase-change process is completed, we are back to a single phase region again (vapor), and further transfer of heat results in an increase in both the temperature and the specific volume. At state 5, the temperature of the vapor is, higher than 100°C; A vapor that is *not about to condense* (i.e., not a saturated vapor) is called a **superheated vapor**. Therefore, water at state 5 is a superheated vapor.

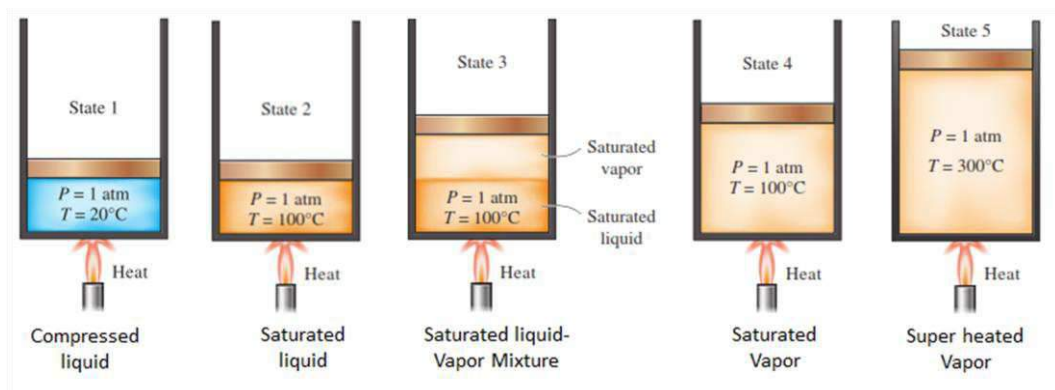


Fig. 3.4

3.4 Saturation Temperature and Pressure

The temperature at which water starts boiling depends on the pressure; therefore, if the pressure is fixed, so is the boiling temperature.

At a given pressure, the temperature at which a pure substance changes phase is called the **Saturation temperature** T_{sat} . Likewise, at a given temperature,

the pressure at which a pure substance changes phase is called the **Saturation pressure** P_{sat}

There are Saturation tables that list the saturation pressure against the temperature (or the saturation temperature against the pressure) and they are available for practically all substances.

It takes a large amount of energy to melt a solid or vaporize a liquid. 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.

The magnitudes of the latent heats depend on the temperature or pressure at which the phase change occurs.

During a phase-change process, pressure and temperature are obviously dependent properties, and there is a definite relation between them, that is, $T_{\text{sat}} = f(P_{\text{sat}})$. A plot of T_{sat} versus P_{sat} , such as the one given for water in Fig. 3.5, is called a **liquid–vapor saturation curve**. A curve of this kind is characteristic of all pure substances.

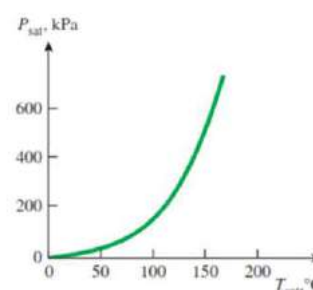


Fig. 3.5

3.5 Property Diagrams for Phase-change Process

3.5.1 The T - v Diagram

By plotting The phase-change process of water at 1 atm pressure inside a piston cylinder device on a **T - v diagram** as shown in Fig. 3.6. and then repeat this process at different pressures to develop **the T - v diagram** as shown in Fig. 3.7. Let us add weights on top of the piston until the pressure

inside the cylinder reaches 1 MPa (for example). At this pressure, water has a somewhat smaller specific volume than it does at 1 atm pressure. As heat is transferred to the water at this new pressure, the process follows a path that looks very much like the process path at 1 atm pressure, but there are some noticeable differences.

First, water starts boiling at a much higher temperature at this pressure. Second, the specific volume of the saturated liquid is larger and the specific volume of the saturated vapor is smaller than the corresponding values at 1 atm pressure. That is, the horizontal line that connects the saturated liquid and saturated vapor states is much shorter.

As the pressure is increased further, this saturation line continues to shrink, and it becomes a point when the pressure reaches 22.06 MPa for the case of water.

This point is called the critical point (Fig. 3.8), and it is defined as the point at which the saturated liquid and saturated vapor states are identical. The temperature, pressure, and specific volume of a substance at the critical point are called, respectively, the critical temperature T_{cr} , critical pressure P_{cr} , and

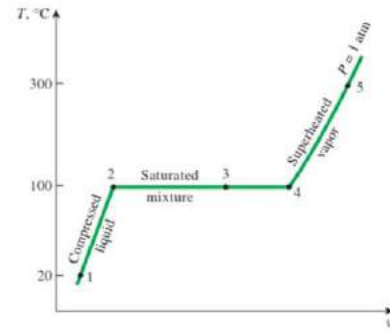


Fig. 3.6

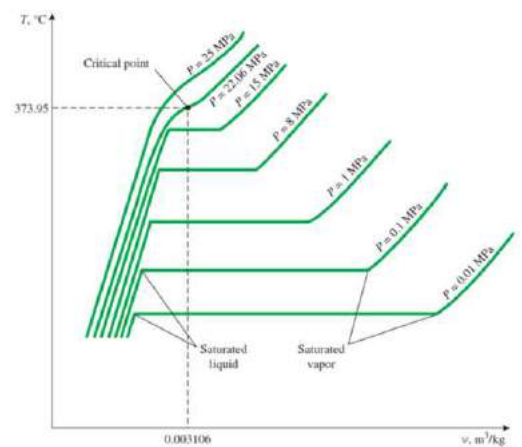


Fig. 3.7

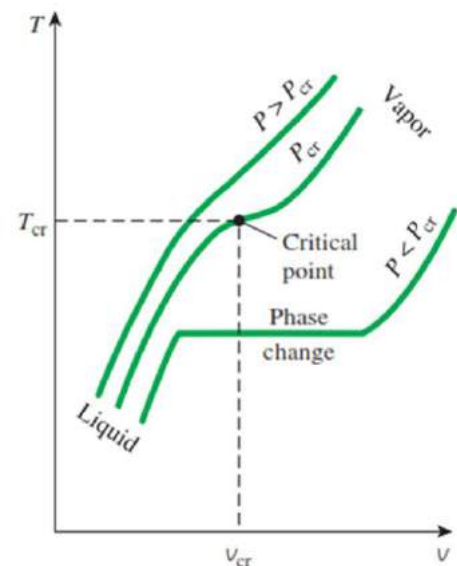


Fig. 3.8

critical specific volume v_{cr} . The critical properties for various substances are given in Table A–1.

Above the critical state, there is no line that separates the compressed liquid region and the superheated vapor region.

The saturated liquid states for different pressures can be connected by a line called the saturated liquid line, and saturated vapor states in the same figure can be connected by another line, called the saturated vapor line. These two lines meet at the critical point, forming a dome as shown in Fig. 3.9. All the compressed liquid states are located in the region to the left of the saturated liquid line, called the compressed liquid region. All the superheated vapor states are located to the right of the saturated vapor line, called the superheated vapor region. In these two regions, the substance exists in a single phase, a liquid or a vapor. All the states that involve both phases in equilibrium are located under the dome, called the **saturated liquid–vapor mixture region**, or the **wet region**.

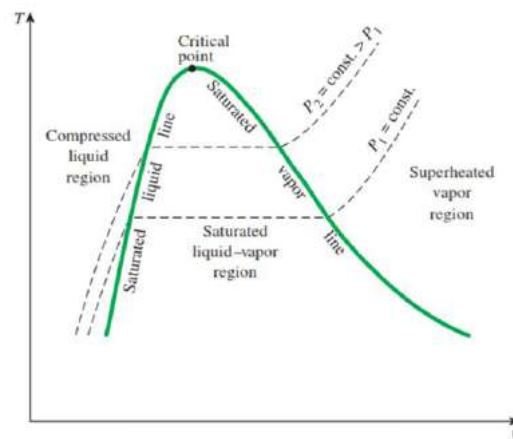


Fig. 3.9

3.5.2 The P - v Diagram

The general shape of the P - v diagram of a pure substance is very much like the T - v diagram (Fig. 3.9), but the $T = \text{constant}$ lines on this diagram have a downward trend. Consider again a piston–cylinder device that contains liquid water at 1 MPa and 150°C. Water at this state exists as a compressed liquid. Now the weights on top of the piston are removed one by one so that the pressure inside the cylinder decreases gradually. The water is allowed to exchange heat with the surroundings so its temperature remains

constant. As the pressure decreases, the volume of the water increases slightly.

When the pressure reaches the saturation-pressure value at the specified temperature, the water starts to boil. During this vaporization process, both the temperature and the pressure remain constant, but the specific volume increases. Once the last drop of liquid is vaporized, further reduction in pressure results

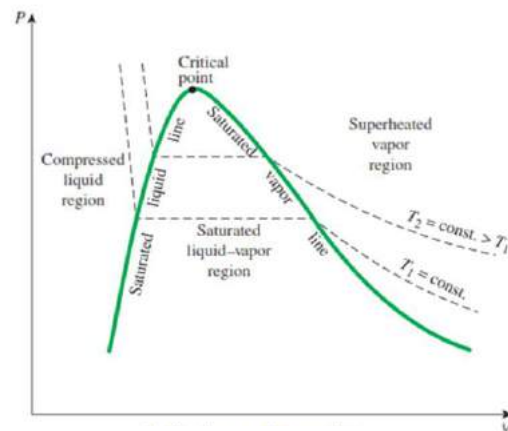


Fig. 3.10

in a further increase in specific volume. Notice that during the phase-change process, we did not remove any weights.

Doing so would cause the pressure and therefore the temperature to drop [since $T_{\text{sat}} = f(P_{\text{sat}})$], and the process would no longer be isothermal. When the process is repeated for other temperatures, similar paths are obtained for the phase-change processes. Connecting the saturated liquid and the saturated vapor states by a curve, we obtain the P - v diagram of a pure substance (Fig. 3.10).

The two equilibrium diagrams developed so far represent the equilibrium states involving the liquid and the vapor phases only. However, these diagrams can easily be extended to include the solid phase as well as the solid–liquid and the solid–vapor saturation regions as shown in Fig. 3.11. The basic principles discussed in conjunction with the liquid–vapor phase-change process apply equally to the solid–liquid and solid vapor phase-change processes. Most substances contract during a solidification (i.e., freezing) process. Others, like water, expand as they freeze. These two diagrams differ only in the solid–liquid saturation region. The T - v diagrams

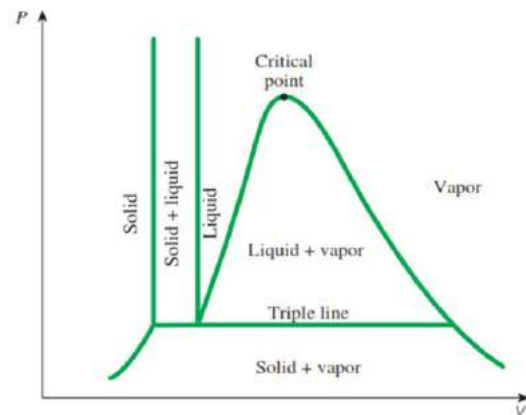
look very much like the P - v diagrams, especially for substances that contract on freezing.

The fact that water expands upon freezing has vital consequences in nature. If water contracted on freezing as most other substances do, the ice formed would be heavier than the liquid water, and it would settle to the bottom of rivers, lakes, and oceans instead of floating at the top.

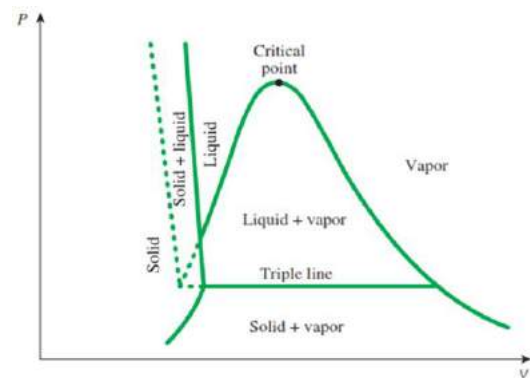
The sun's rays would never reach these ice layers, and the bottoms of many rivers, lakes, and oceans would be covered with ice at times, seriously disrupting marine life.

We are all familiar with two phases being in equilibrium, but under some conditions all three phases of a pure substance coexist in equilibrium.

On P - v or T - v diagrams, these triple-phase states form a line called the triple line. The states on the triple line of a substance have the same pressure and temperature but different specific volumes. The triple line appears as a point on the P - T diagrams and, therefore, is often called the triple point. That is, all three phases of water coexist in equilibrium only if the temperature and pressure have precisely these values. No substance can exist in the liquid phase in stable equilibrium at pressures below the triple-point pressure. The same can be said for temperature for substances that contract on freezing. However, substances at high pressures can exist in the liquid phase at temperatures below the triple-point temperature. For example, water cannot exist in liquid form in



(a) P - v diagram of a substance that contracts on freezing



(b) P - v diagram of a substance that expands on freezing (such as water)

Fig. 3.11

equilibrium at atmospheric pressure at temperatures below 0°C, but it can exist as a liquid at -20°C at 200 MPa pressure. Also, ice exists at seven different solid phases at pressures above 100 MPa. There are two ways a substance can pass from the solid to vapor phase: either it melts first into a liquid and subsequently evaporates, or it evaporates directly without melting first (sublimation) which occurs at pressures below the triple-point value, since a pure substance cannot exist in the liquid phase at those pressures. For substances that have a triple-point pressure above the atmospheric pressure such as solid CO (dry ice), sublimation is the only way to change from the solid to vapor phase at atmospheric conditions.

3.5.3 The P-T Diagram

The P - T diagram of a pure substance is often called the **phase diagram** (Fig. 3.12) since all three phases are separated from each other by three lines. The **sublimation**

line separates the solid and the 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

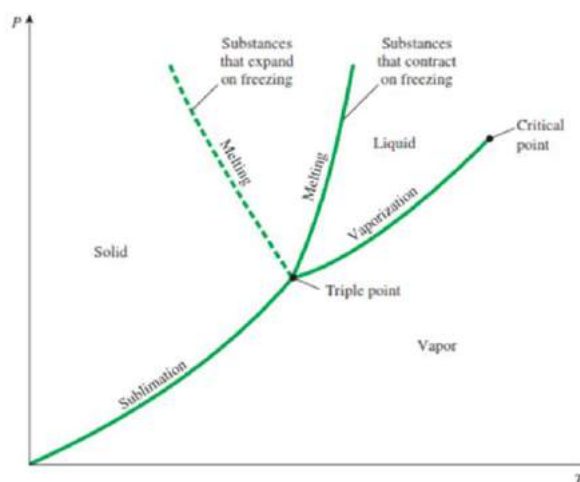


Fig. 3.12

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.

3.6 Property Tables

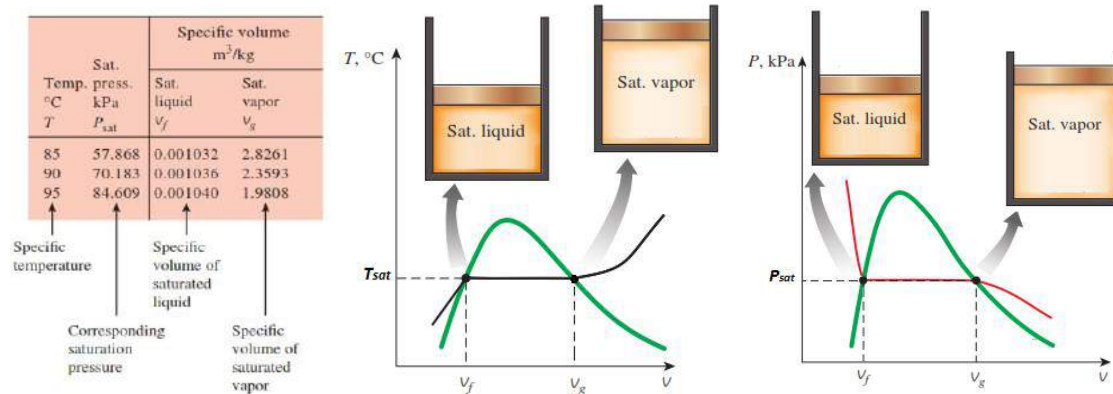
For most substances, the relationships among thermodynamic properties are too complex to be expressed by simple equations. Therefore, properties are frequently presented in the form of tables. Some thermodynamic properties can be measured easily, but others cannot and are calculated by using the relations between them and measurable properties. The results of these measurements and calculations are presented in tables in a convenient format.

For each substance, the thermodynamic properties are listed in more than one table. In fact, a separate table is prepared for each region of interest such as the superheated vapor, compressed liquid, and saturated (mixture) regions. Property tables are given in the appendix in both SI and English units.

3.6.1 Saturated Liquid and Saturated Vapor States

The properties of saturated liquid and saturated vapor for water are listed in **Tables A–4** and **A–5**. Both tables give the same information. The only difference is that in **Table A–4** properties are listed under temperature and in **Table A–5** under pressure. Therefore, it is more convenient to use **Table A–4** when *temperature* is given and **Table A–5** when *pressure* is given. The subscript *f* is used to denote **properties of a saturated liquid**, and the subscript *g* to denote the **properties of saturated vapor**.

For example the specific volume:



v_f : specific volume of saturated liquid.

v_g : specific volume of saturated vapor.

v_{fg} : difference between v_g and v_f (that is $v_{fg} = v_g - v_f$).

As well as the specific volume; the other properties are treated by the same way; so that there are u_f , u_g , u_{fg} , h_f , h_g and h_{fg} .

Example 3.1

A rigid tank contains 50 kg of saturated liquid water at 90°C. Determine the pressure in the tank and the volume of the tank.

Solution

The state of the saturated liquid water is shown on a T-v diagram in Fig. 3.13. Since saturation conditions exist in the tank, the pressure must be the saturation pressure at 90°C:

$$P = P_{sat @ 90^\circ C} = 70.183 \text{ kPa} \quad (\text{Table A-4})$$

The specific volume of the saturated liquid at 90°C is

$$v = v_f @ 90^\circ C = 0.001036 \text{ m}^3/\text{kg} \quad (\text{Table A-4})$$

Then the total volume of the tank becomes

$$V = mv = (50 \text{ kg})(0.001036 \text{ m}^3/\text{kg}) = 0.0518 \text{ m}^3$$

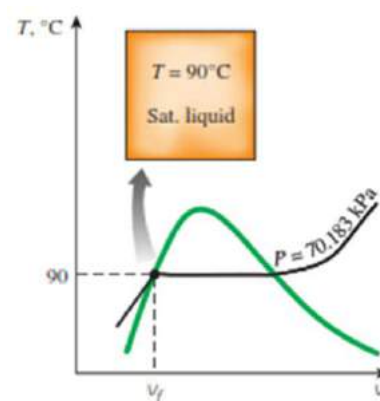


Fig. 3.13

Example 3.2

A piston–cylinder device contains 2 ft³ of saturated water vapor at 50-psia pressure. Determine the temperature and the mass of the vapor inside the cylinder.

Solution

The state of the saturated water vapor is shown on a P - v diagram in Fig. 3.14. Since the cylinder contains saturated vapor at 50 psia, the temperature inside must be the saturation temperature at this pressure:

$$T = T_{\text{sat @ 50 psia}} = 280.99^\circ\text{F} \quad (\text{Table A-5E})$$

The specific volume of the saturated vapor at 50 psia is

$$v = v_g @ 50 \text{ psia} = 8.5175 \text{ ft}^3/\text{lbm} \quad (\text{Table A-5E})$$

Then the mass of water vapor inside the cylinder becomes

$$m = \frac{V}{v} = \frac{2 \text{ ft}^3}{8.5175 \text{ ft}^3/\text{lbm}} = 0.235 \text{ lbm}$$

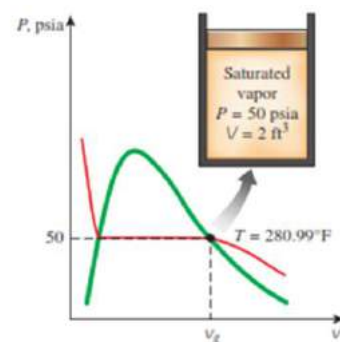


Fig. 3.14

3.6.2 Saturated Liquid–Vapor Mixture

During a vaporization process, a substance exists as part liquid and part vapor. That is, it is a mixture of saturated liquid and saturated vapor as shown in Fig. 3.15. To analyze this mixture properly, we need to know the proportions of the liquid and vapor phases in the mixture. This is done by defining a new property called the **quality** x as the ratio of the mass of vapor to the total mass of the mixture:

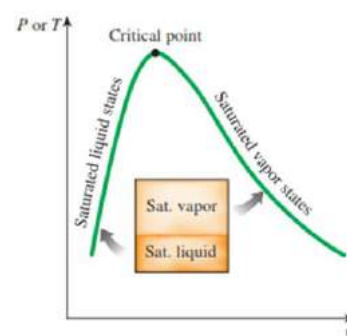


Fig. 3.15

$$x = \frac{m_{\text{vapor}}}{m_{\text{total}}} = \frac{m_g}{m_t} \quad (3-1)$$

Where: $m_{\text{Total}} = m_{\text{vapor}} + m_{\text{liquid}}$

Quality has significance for saturated mixtures only. It has no meaning in the compressed liquid or superheated vapor regions. Its value is between 0 and 1. The quality of a system that consists of saturated liquid is 0 (or 0 percent), and the quality of a system consisting of saturated vapor is 1 (or 100 percent). In saturated mixtures, quality can serve as one of the two independent intensive properties needed to describe a state. Note that the properties of the saturated liquid are the same whether it exists alone or in a mixture with saturated vapor. During the vaporization process, only the amount of saturated liquid changes, not its properties. The same can be said about a saturated vapor.

A saturated mixture can be treated as a combination of two subsystems: the saturated liquid and the saturated vapor. However, the amount of mass for each phase is usually not known. Therefore, it is often more convenient to imagine that the two phases are mixed well, forming a homogeneous mixture. Then the properties of this “mixture” will simply be the average properties of the saturated liquid–vapor mixture under consideration. Here is how it is done.

Consider a tank that contains a saturated liquid–vapor mixture (Fig. 3.16).

The volume occupied by saturated liquid is V_f , and the volume occupied by saturated vapor is V_g . The total volume V is the sum of the two:

$$V = V_f + V_g \quad (3-2)$$

$$V = m v \rightarrow m_t v = m_f v_f + m_g v_g \quad (3-3)$$

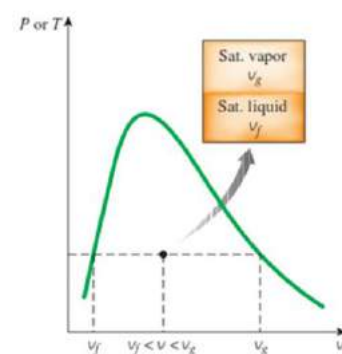


Fig. 3.16

$$m_f = m_t - m_g \rightarrow m_t v = (m_t - m_g) v_f + m_g v_g \quad (3-4)$$

Dividing by the m_t yields

$$v = \frac{(m_t - m_g)}{m_t} v_f + \frac{m_g}{m_t} v_g \quad (3-5)$$

$$v = \left(1 - \frac{m_g}{m_t}\right) v_f + \frac{m_g}{m_t} v_g \quad (3-6)$$

$$v = (1 - x)v_f + x v_g = v_f - xv_f + x v_g \quad (3-7)$$

$$v = v_f + x(v_g - v_f) = v_f + x v_{fg} \quad (3-8)$$

$$x = \frac{v - v_f}{v_g - v_f} = \frac{v - v_f}{v_{fg}}, \quad \because v_{fg} = v_g - v_f \quad (3-9)$$

The analysis given above can be repeated for internal energy and enthalpy with the following results:

$$u = u_f + x u_{fg} \text{ and } h = h_f + x h_{fg} \quad (3-10)$$

All the results are of the same format, and they can be summarized in a single equation as

$$y = y_f + x y_{fg} \quad (3-11)$$

where y is v, u or h

$$y_f \leq y \leq y_g \quad (3-12)$$

Example 3.3

A rigid tank contains 10 kg of water at 90°C. If 8 kg of the water is in the liquid form and the rest is in the vapor form, determine (a) the pressure in the tank and (b) the volume of the tank.

Solution

(a) The state of the saturated liquid–vapor mixture is shown in Fig. 3.17.

Since the two phases coexist in equilibrium, we have a saturated mixture, and the pressure must be the saturation pressure at the given temperature:

$$P = P_{\text{sat @ } 90^\circ\text{C}} = 70.183 \text{ kPa} \quad (\text{Table A-4})$$

(b) At 90°C, we have $v_f = 0.0010365 \text{ m}^3/\text{kg}$ and $v_g = 2.3593 \text{ m}^3/\text{kg}$ (Table A-4). One way of finding the volume of the tank is to determine the volume occupied by each phase and then add them:

$$\begin{aligned}
 V &= V_f + V_g = m_f v_f + m_g v_g \\
 &= (8 \text{ kg})(0.001036 \text{ m}^3/\text{kg}) + (2 \text{ kg})(2.3593 \text{ m}^3/\text{kg}) \\
 &= 4.73 \text{ m}^3
 \end{aligned}$$

Another way is to first determine the quality x , then the average specific volume v , and finally the total volume:

$$x = \frac{m_g}{m_t} = \frac{2 \text{ kg}}{10 \text{ kg}} = 0.2$$

$$\begin{aligned}
 v &= v_f + x v_{fg} \\
 &= 0.001036 \text{ m}^3/\text{kg} + (0.2)[(2.3593 - 0.001036) \text{ m}^3/\text{kg}] \\
 &= 0.473 \text{ m}^3/\text{kg}
 \end{aligned}$$

And

$$V = m v = (10 \text{ kg})(0.473 \text{ m}^3/\text{kg}) = 4.73 \text{ m}^3$$

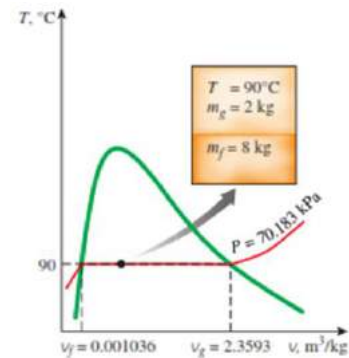


Fig. 3.17

3.6.3 Superheated vapor

In the region to the right of the saturated vapor line and at temperatures above the critical point temperature, a substance exists as superheated vapor. Since the superheated region is a single-phase region (vapor phase only), temperature and pressure are no longer dependent properties and they can conveniently be used as the two independent properties in the tables (**Table A-6**)

Compared to saturated vapor, superheated vapor is characterized by:

Higher temperatures ($T > T_{\text{sat}}$ at a given P)

Lower pressures ($P < P_{\text{sat}}$ at a given T)

Higher specific volumes ($v > v_g$ at a given P or T)

Higher internal energies ($u > u_g$ at a given P or T)

Higher enthalpies ($h > h_g$ at a given P or T)

3.6.4 Compressed Liquid

Compressed liquid tables are not as commonly available, and **Table (A-7)** is the only compressed liquid table in this text. The format of **Table (A-7)** is very much like the format of the superheated vapor tables. One reason for the lack of compressed liquid data is the relative independence of compressed liquid properties from pressure. Variation of properties of compressed liquid with pressure is very mild. Increasing the pressure 100 times often causes properties to change less than 1 percent.

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

This is because the compressed liquid properties depend on temperature much more strongly than they do on pressure. Thus,

$$y \cong y_{f@T} \text{ (From Table (A - 4))}$$

for compressed liquids, where y is v , u , or h . so that

$$v \cong v_{f@T}, \quad u \cong u_{f@T}, \quad \text{and } h \cong h_{f@T}$$

In general, a compressed liquid is characterized by:

Higher pressures ($P > P_{\text{sat}}$ at a given T)

Lower temperatures ($T < T_{\text{sat}}$ at a given P)

Lower specific volumes ($v < v_f$ at a given P or T)

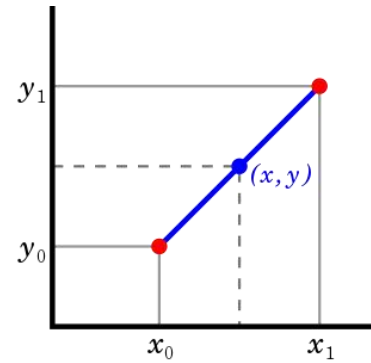
Lower internal energies ($u < u_f$ at a given P or T)

Lower enthalpies ($h < h_f$ at a given P or T)

Region	T-y (y is v, u or h)	P-y (y is v, u or h)	Aspect	How to determine Properties
Compressed Liquid (c)			$P > P_{sat}$ at a given T $T < T_{sat}$ at a given P $v < v_f$ at a given P or T $u < u_f$ at a given P or T $h < h_f$ at a given P or T	$y \cong y_{f@T}$ (From Table (A - 4) <u>only</u>) where y is v, u, or h. so that $v \cong v_{f@T}$, $u \cong u_{f@T}$, and $h \cong h_{f@T}$
Saturated Liquid and Saturated Vapor (f and g)			$P = P_{sat}$ at a given T $T = T_{sat}$ at a given P At a given P or T	From Saturation table (A-4) at given T, or table (A-5) at given P ; beside other independent property x for sat. liquid = 0 x for sat. vapor = 1
			Sat. liquid $v = v_f$ $u = u_f$ $h = h_f$	
Saturated Mixture (X)			$P = P_{sat}$ at a given T $T = T_{sat}$ at a given P $v_f > v > v_g$, $u_f > u > u_g$, or $h_f > h > h_g$	$v = v_f + x v_{fg}$ $u = u_f + x u_{fg}$ $h = h_f + x h_{fg}$ $x = \frac{y - y_f}{y_{fg}}$, (y is v, u or h) $0 < x < 1$
Superheated Vapor (s)			$T > T_{sat}$ at a given P $P < P_{sat}$ at a given T $v > v_g$ at a given P or T $u > u_g$ at a given P or T $h > h_g$ at a given P or T	From table (A-6) using two independent properties

3.7 Linear interpolation

If the two known points are given by the coordinates (x_0, y_0) and (x_1, y_1) , the **linear interpolant** is the straight line between these points. For a value x in the interval (x_0, x_1) , the value y along the straight line is given from the equation of slopes



$$\frac{y - y_0}{y_1 - y_0} = \frac{x - x_0}{x_1 - x_0}$$

This interpolation can be applied on the empirical data from the tables of ideal gases and pure substances; for example:

If we want to determine the properties from the following table at $T = 12^\circ\text{C}$

Temp., T °C	Sat. press., P_{sat} kPa	Sat. liquid, v_f	Sat. vapor, v_g	Sat. liquid, u_f	Evap., u_{fg}	Sat. vapor, u_g	Sat. liquid, h_f	Evap., h_{fg}	Sat. vapor, h_g
0.01	0.6117	0.001000	206.00	0.000	2374.9	2374.9	0.001	2500.9	2500.9
5	0.8725	0.001000	147.03	21.019	2360.8	2381.8	21.020	2489.1	2510.1
10	1.2281	0.001000	106.32	42.020	2346.6	2388.7	42.022	2477.2	2519.2
15	1.7057	0.001001	77.885	62.980	2332.5	2395.5	62.982	2465.4	2528.3
20	2.3392	0.001002	57.762	83.913	2318.4	2402.3	83.915	2453.5	2537.4

It is clear from this table that there are no given properties at 12°C ; but we have the properties for both 10°C and 15°C ; then we can find the properties at 12°C using interpolation.

Temp., T °C	Sat. press., P_{sat} kPa	Sat. liquid, v_f	Sat. vapor, v_g	Sat. liquid, u_f	Evap., u_{fg}	Sat. vapor, u_g	Sat. liquid, h_f	Evap., h_{fg}	Sat. vapor, h_g
0.01	0.6117	0.001000	206.00	0.000	2374.9	2374.9	0.001	2500.9	2500.9
5	0.8725	0.001000	147.03	21.019	2360.8	2381.8	21.020	2489.1	2510.1
10	1.2281	0.001000	106.32	42.020	2346.6	2388.7	42.022	2477.2	2519.2
12	?????	???????	?????	???????	?????	?????	?????	?????	?????
15	1.7057	0.001001	77.885	62.980	2332.5	2395.5	62.982	2465.4	2528.3
20	2.3392	0.001002	57.762	83.913	2318.4	2402.3	83.915	2453.5	2537.4

To find P_{sat} at 12°C:

$$\frac{12 - 10}{15 - 10} = \frac{P_{sat@12} - P_{sat@10}}{P_{sat@15} - P_{sat@10}}$$

$$\frac{12 - 10}{15 - 10} = \frac{P_{sat@12} - 1.2281}{1.7057 - 1.2281}$$

$$P_{sat@12} = 1.41914 \text{ kPa}$$

To find v_f at 12°C:

$$\frac{12 - 10}{15 - 10} = \frac{v_f@12 - v_f@10}{v_f@15 - v_f@10}$$

$$\frac{12 - 10}{15 - 10} = \frac{v_f@12 - 0.001}{0.001001 - 0.001}$$

$$v_f@12 = 0.0010004 \text{ m}^3/\text{kg}$$

By the same method we can determine all other properties.

Problems (Properties of A pure Substance)

3.1 Complete this table for H₂O, and then illustrate each point on both of *P-v* and *T-v*.

point	T(°C)	P(kPa)	v(m ³ /kg)	u(kj/kg)	h(kj/kg)	X	Phase description
A	100	110					
B		200					Saturated liquid
C	300	400					
D	110	600					
E	130			2200			
F		500					Saturated vapor
G	140				1800		
H	50		4.16				
I		325				0.8	
J	160				1215		

3.2 Complete this table for refrigerant-134a.

point	T(°C)	P(kPa)	v(m ³ /kg)	u(kj/kg)	h(kj/kg)	X	Phase description
A	-8	320					
B	-12						Saturated liquid
C	20			95			
D	8	600					
E		400		300			
F		180					Saturated vapor

3.3 A mass of 200 g of saturated liquid water is completely vaporized at a constant pressure of 100 kPa. Determine (a) the volume change and (b) the amount of energy transferred to the water.

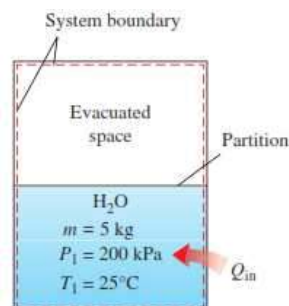
3.4 A rigid vessel contains 2 kg of refrigerant-134a at 800 kPa and 120°C. Determine the volume of the vessel and the total internal energy.

- 3.5 A 0.5-m^3 vessel contains 10 kg of refrigerant-134a at -20°C . Determine (a) the pressure, (b) the total internal energy, and (c) the volume occupied by the liquid phase.
- 3.6 A rigid tank contains 10 kg of water at 90°C . If 8 kg of the water is in the liquid form and the rest is in the vapor form, determine (a) the pressure in the tank and (b) the volume of the tank.
- 3.7 A cooking pan whose inner diameter is 20 cm is filled with water and covered with a 4-kg lid. If the local atmospheric pressure is 101 kPa, determine the temperature at which the water starts boiling when it is heated.
- 3.8 Water is being heated in a vertical piston–cylinder device. The piston has a mass of 20 kg and a cross-sectional area of 100 cm^2 . If the local atmospheric pressure is 100 kPa, determine the temperature at which the water starts boiling.
- 3.9 A rigid tank with a volume of 2.5 m^3 contains 15 kg of saturated liquid–vapor mixture of water at 75°C . Now the water is slowly heated. Determine the temperature at which the liquid in the tank is completely vaporized. Also, show the process on a T - v diagram with respect to saturation lines.
- 3.10 A piston–cylinder device contains 0.1 m^3 of liquid water and 0.9 m^3 of water vapor in equilibrium at 800 kPa. Heat is transferred at constant pressure until the temperature reaches 350°C .
- (a) What is the initial temperature of the water?
- (b) Determine the total mass of the water.
- (c) Calculate the final volume.
- (d) Show the process on a P - v diagram with respect to saturation lines.

- 3.11 A piston–cylinder device contains 0.8 kg of steam at 300°C and 1 MPa. Steam is cooled at constant pressure until one-half of the mass condenses. (a) Show the process on a T - v diagram; (b) Find the final temperature; (c) Determine the volume change.
- 3.12 Superheated water vapor at 1MPa and 300°C is allowed to cool at constant volume until the temperature drops to 150°C. At the final state, determine (a) the pressure, (b) the quality, and the enthalpy. Also show the process on T - V and P - V diagrams.
- 3.13 A piston–cylinder device initially contains 50 L of liquid water at 40°C and 200 kPa. Heat is transferred to the water at constant pressure until the entire liquid is vaporized.
- (a) What is the mass of the water?
 - (b) What is the final temperature?
 - (c) Determine the total enthalpy change
 - (d) Show the process on a T - v diagram with respect to saturation lines.

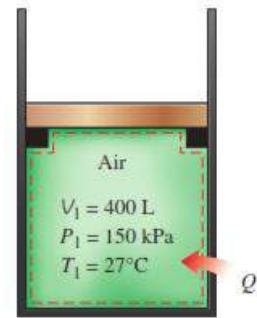
Problems (General Problems on the First Law of thermodynamics)

1. A piston–cylinder device contains 25 g of saturated water vapor that is maintained at a constant pressure of 300 kPa. A resistance heater within the cylinder is turned on and passes a current of 0.2 A for 5 min from a 120-V source. At the same time, a heat loss of 3.7 kJ occurs. (a) Show that for a closed system the boundary work W_b and the change in internal energy ΔU in the first-law relation can be combined into one term, ΔH , for a constant pressure process. (b) Determine the final temperature of the steam.
2. A rigid tank is divided into two equal parts by a partition. Initially, one side of the tank contains 5 kg of water at 200 kPa and 25°C, and the other side is evacuated. The partition is then removed, and the water expands into the entire tank. The water is allowed to exchange heat with its surroundings until the temperature in the tank returns to the initial value of 25°C. Determine (a) the volume of the tank, (b) the final pressure, and (c) the heat transfer for this process.

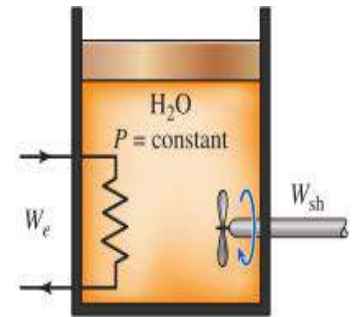


3. A piston–cylinder device initially contains 0.5 m³ of nitrogen gas at 400 kPa and 27°C. An electric heater within the device is turned on and is allowed to pass a current of 2 A for 5 min from a 120-V source. Nitrogen expands at constant pressure, and a heat loss of 2800 J occurs during the process. Determine the final temperature of nitrogen.

4. A piston–cylinder device initially contains air at 150 kPa and 27°C. At this state, the piston is resting on a pair of stops, and the enclosed volume is 400 L. The mass of the piston is such that a 350-kPa pressure is required to move it. The air is now heated until its volume has doubled. Determine (a) the final temperature, (b) the work done by the air, and (c) the total heat transferred to the air.

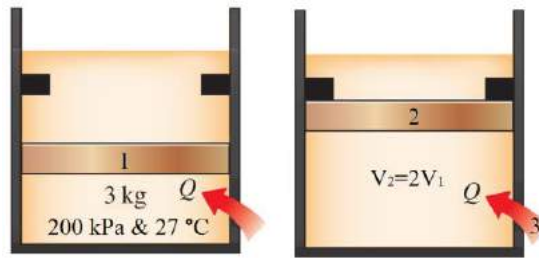


5. An insulated piston–cylinder device contains 5 L of saturated liquid water at a constant pressure of 175 kPa. Water is stirred by a paddle wheel while a current of 8 A flows for 45 min through a resistor placed in the water. If one-half of the liquid is evaporated during this constant pressure process and the paddle-wheel work amounts to 400 kJ, determine the voltage of the source. Also, show the process on a P - v diagram with respect to saturation lines.



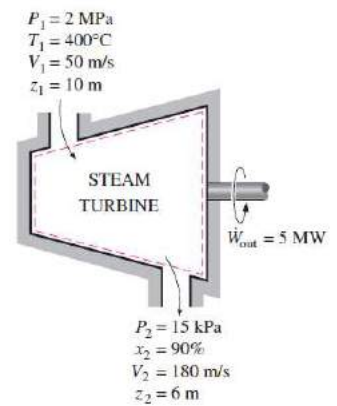
6. A 4-m x 5-m x 6-m room is to be heated by a baseboard resistance heater. It is desired that the resistance heater be able to raise the air temperature in the room from 7 to 23°C within 15 min. Assuming no heat losses from the room and an atmospheric pressure of 100 kPa, determine the required power of the resistance heater. Assume constant specific heats at room temperature.
7. A piston–cylinder device, with a set of stops on the top, initially contains 3 kg of air at 200 kPa and 27°C. Heat is now transferred to the air, and the piston rises until it hits the stops, at which point the volume is twice the initial volume. More heat is transferred until the pressure inside the

cylinder also doubles. Determine the work done and the amount of heat transfer for this process. Also, show the process on a P - v diagram.



8. Air at 10°C and 80 kPa enters the diffuser of a jet engine steadily with a velocity of 200 m/s. The inlet area of the diffuser is 0.4 m². The air leaves the diffuser with a velocity that is very small compared with the inlet velocity. Determine (a) the mass flow rate of the air and (b) the temperature of the air leaving the diffuser.
9. Air enters an adiabatic nozzle steadily at 300 kPa, 200°C, and 30 m/s and leaves at 100 kPa and 180 m/s. The inlet area of the nozzle is 80 cm². Determine (a) the mass flow rate through the nozzle, (b) the exit temperature of the air, and (c) the exit area of the nozzle.
10. Air at 100 kPa and 280 K is compressed steadily to 600 kPa and 400 K. The mass flow rate of the air is 0.02 kg/s, and a heat loss of 16 kJ/kg occurs during the process. Assuming the changes in kinetic and potential energies are negligible, determine the necessary power input to the compressor.
11. The electric heating systems used in many houses consist of a simple duct with resistance heaters. Air is heated as it flows over resistance wires. Consider a 15-kW electric heating system. Air enters the heating section at 100 kPa and 17°C with a volume flow rate of 150 m³/min. If heat is lost from the air in the duct to the surroundings at a rate of 200 W, determine the exit temperature of air.

12. The power output of an adiabatic steam turbine is 5 MW, and the inlet and the exit conditions of the steam are as indicated in Fig.

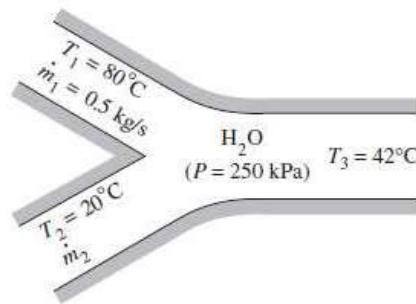


- (a) Compare the magnitudes of Δh , ΔKE , and ΔPE .
 (b) Determine the work done per unit mass of the steam flowing through the turbine.
 (c) Calculate the mass flow rate of the steam.

13. Nitrogen gas at 60 kPa and 7°C enters an adiabatic diffuser steadily with a velocity of 200 m/s and leaves at 85 kPa and 22°C . Determine (a) the exit velocity of the nitrogen and (b) the ratio of the inlet to exit area A_1/A_2 .
14. Steam flows steadily through an adiabatic turbine. The inlet conditions of the steam are 10 MPa, 450°C , and 80 m/s, and the exit conditions are 10 kPa, 92 percent quality, and 50 m/s. The mass flow rate of the steam is 12 kg/s. Determine (a) the change in kinetic energy, (b) the power output, and (c) the turbine inlet area.
15. Steam enters an adiabatic turbine at 10 MPa and 500°C and leaves at 10 kPa with a quality of 90 percent. Neglecting the changes in kinetic and potential energies, determine the mass flow rate required for a power output of 5 MW.
16. Argon gas enters an adiabatic turbine steadily at 900 kPa and 450°C with a velocity of 80 m/s and leaves at 150 kPa with a velocity of 150 m/s. The inlet area of the turbine is 60 cm^2 . If the power output of the turbine is 250 kW, determine the exit temperature of the argon.
17. Air enters the compressor of a gas-turbine plant at ambient conditions of 100 kPa and 25°C with a low velocity and exits at 1 MPa and 347°C with a velocity of 90 m/s. The compressor is cooled at a rate of 1500

kJ/min, and the power input to the compressor is 250 kW. Determine the mass flow rate of air through the compressor.

18. Helium is to be compressed from 120 kPa and 310 K to 700 kPa and 430 K. A heat loss of 20 kJ/kg occurs during the compression process. Neglecting kinetic energy changes, determine the power input required for a mass flow rate of 90 kg/min.
19. A hot-water stream at 80°C enters a mixing chamber with a mass flow rate of 0.5 kg/s where it is mixed with a stream of cold water at 20°C. If it is desired that the mixture leave the chamber at 42°C, determine the mass flow rate of the cold-water stream. Assume all the streams are at a pressure of 250 kPa.



20. Liquid water at 300 kPa and 20°C is heated in a chamber by mixing it with superheated steam at 300 kPa and 300°C. Cold water enters the chamber at a rate of 1.8 kg/s. If the mixture leaves the mixing chamber at 60°C, determine the mass flow rate of the superheated steam required.

Chapter (4)

The Second Law of Thermodynamics

4.1 Introduction:

In the preceding chapters we applied the first law of thermodynamics, or the conservation of energy principle, to processes involving closed and open systems. As pointed out repeatedly in the preceding chapters, energy is a conserved property, and no process is known to have taken place in violation of the first law of thermodynamics. Therefore, it is reasonable to conclude that a process must satisfy the first law to occur. However, as explained here, satisfying the first law alone does not ensure that the process will actually take place.

It is common experience that a cup of hot coffee left in a cooler room eventually cools off (Fig. 4.1). This process satisfies the first law of thermodynamics since the amount of energy lost by the coffee is equal to the amount gained by the surrounding air. Now let us consider the reverse process the hot coffee getting even hotter in a cooler room as a result of heat transfer from the room air. We all know that this process never takes place. Yet, doing so would not violate the first law as long as the amount of energy lost by the air is equal to the amount gained by the coffee.

As another familiar example, consider the heating of a room by the passage of electric current through a resistor (Fig. 4.2). Again, the first law dictates that the amount of electric energy supplied to the resistance wires be equal to the amount of energy transferred to the room air as heat. Now let us attempt to reverse this process. It will come as



Fig.4.1

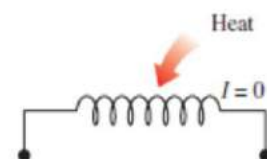


Fig. 4.2

no surprise that transferring some heat to the wires does not cause an equivalent amount of electric energy to be generated in the wires.

Finally, consider a paddle-wheel mechanism that is operated by the fall of a mass (Fig. 4.3). The paddle wheel rotates as the mass falls and stirs a fluid within an insulated container. As a result, the potential energy of the mass decreases, and the internal energy of the fluid increases in accordance with the conservation of energy principle.

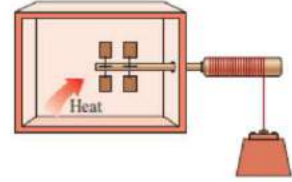


Fig. 4.3

However, the reverse process, raising the mass by transferring heat from the fluid to the paddle wheel, does not occur in nature, although doing so would not violate the first law of thermodynamics.

It is clear from these arguments that processes proceed in a certain direction and not in the reverse direction. The first law places no restriction on the direction of a process, but satisfying the first law does not ensure that the process can actually occur. This inadequacy of the first law to identify whether a process can take place is remedied by introducing another general principle, the second law of thermodynamics. We show later in this chapter that the reverse processes discussed above violate the second law of thermodynamics. This violation is easily detected with the help of a property, called entropy. A process cannot occur unless it satisfies both the first and the second laws of thermodynamics (Fig.

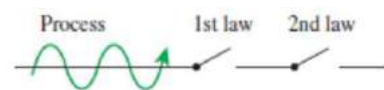


Fig. 4.4

4.4). There are numerous valid statements of the second law of thermodynamics. Two

such statements are presented and discussed later in this chapter in relation to some engineering devices that operate on cycles.

The use of the second law of thermodynamics is not limited to identifying the direction of processes. The second law also asserts that energy has *quality* as well as quantity. The first law is concerned with the quantity of energy and the transformations of energy from one form to

another with no regard to its quality. Preserving the quality of energy is a major concern to engineers, and the second law provides the necessary means to determine the quality as well as the degree of degradation of energy during a process. As discussed later in this chapter, more of high-temperature energy can be converted to work, and thus it has a higher quality than the same amount of energy at a lower temperature.

The second law of thermodynamics is also used in determining the theoretical limits for the performance of commonly used engineering systems, such as heat engines and refrigerators, as well as predicting the degree of completion of chemical reactions. The second law is also closely associated with the concept of perfection. In fact, the second law defines perfection for thermodynamic processes. It can be used to quantify the level of perfection of a process, and point the direction to eliminate imperfections effectively.

4.2 Thermal Energy Reservoirs:

In the development of the second law of thermodynamics, it is very convenient to have a hypothetical body with a relatively large thermal energy capacity (mass \times specific heat) 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. In practice, large bodies of water such as oceans, lakes, and rivers as well as the atmospheric air can be

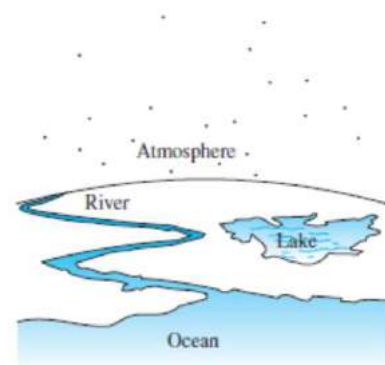


Fig. 4.5

modeled accurately as thermal energy reservoirs because of their large thermal energy storage capabilities or thermal masses (Fig. 4.5). The atmosphere, for example, does not warm up as a result of heat losses from residential buildings in winter. Likewise, mega joules of waste energy

dumped in large rivers by power plants do not cause any significant change in water temperature.

A two-phase system can also be modeled as a reservoir since it can absorb and release large quantities of heat while remaining at constant temperature. Another familiar example of a thermal energy reservoir is the industrial furnace. The temperatures of most furnaces are carefully controlled, and they are capable of supplying large quantities of thermal energy as heat in an essentially isothermal manner. Therefore, they can be modeled as reservoirs.

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 (Fig. 4.6). Thermal energy reservoirs are often referred to as heat reservoirs since they supply or absorb energy in the form of heat.



Fig. 4.6

4.3 Heat Engines:

As pointed out earlier, 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 shown in Fig. 4.7, for example, is first converted to the internal energy to the water.

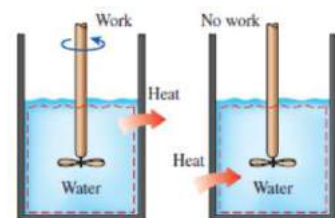


Fig. 4.7

This energy may then leave the water as heat. We know 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.

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. Heat engines differ considerably from one another, but all can be characterized by the following (Fig. 4.8):

1. They receive heat from a high-temperature source (solar energy, oil furnace, nuclear reactor, etc.).
2. They convert part of this heat to work (usually in the form of a rotating shaft).
3. They reject the remaining waste heat to a low-temperature sink (the atmosphere, rivers, etc.).
4. They operate on a cycle.

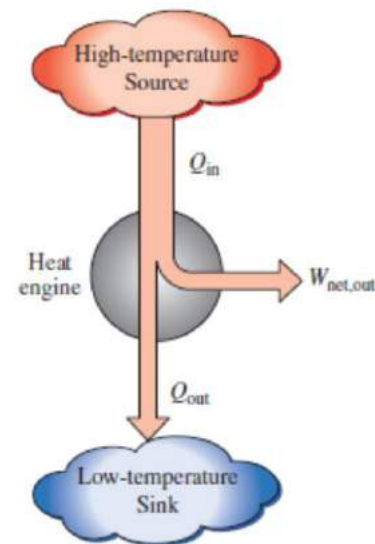


Fig. 4.8

Heat engines and other cyclic devices usually involve a fluid to and from which heat is transferred while undergoing a cycle. This fluid is called the working fluid.

The term *heat engine* is often used in a broader sense to include work producing devices that do not operate in a thermodynamic cycle. Engines that involve internal combustion such as gas turbines and car engines fall into this category. These devices operate in a mechanical cycle but not in a thermodynamic cycle since the working fluid (the combustion gases) does not undergo a complete cycle. Instead of being cooled to the initial temperature, the exhaust gases are purged and replaced by fresh air-and-fuel mixture at the end of the cycle.

The work-producing device that best fits into the definition of a heat engine is the *steam power plant*, which is an external-combustion engine. That is, combustion takes place outside the engine, and the thermal energy released

during this process is transferred to the steam as heat. The schematic of a basic steam power plant is shown in Fig. 4.9. This is a rather simplified diagram. The various quantities shown on this figure are as follows:

Q_{in} = amount of heat supplied to steam in boiler from a high-temperature source (furnace)

Q_{out} = amount of heat rejected from steam in condenser to a low temperature sink (the atmosphere, a river, etc.)

W_{out} = amount of work delivered by steam as it expands in turbine

W_{in} = amount of work required to compress water to boiler pressure

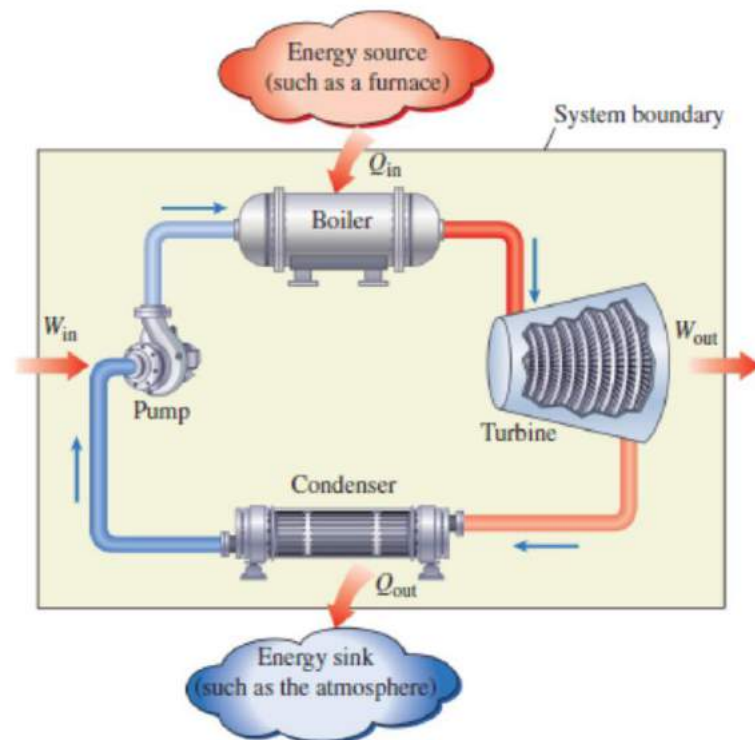


Fig. 4.9

The network output of this power plant is simply the difference between the total work output of the plant and the total work input:

$$W_{net,out} = W_{out} - W_{in} \quad (\text{kJ}) \quad (4-1)$$

The network can also be determined from the heat transfer data alone. The four components of the steam power plant involve mass flow in and out, and

therefore should be treated as open systems. These components, together with the connecting pipes, however, always contain the same fluid (not counting the steam that may leak out, of course). No mass enters or leaves this combination system, which is indicated by the shaded area on Fig. 4.9; thus, it can be analyzed as a closed system. Recall that for a closed system undergoing a cycle, the change in internal energy ΔU is zero, and therefore the network output of the system is also equal to the net heat transfer to the system:

$$W_{\text{net,out}} = Q_{\text{in}} - Q_{\text{out}} \quad (\text{kJ}) \quad (4-2)$$

4.3.1 Thermal Efficiency

In Eq. 4-2, Q_{out} represents the magnitude of the energy wasted in order to complete the cycle. But Q_{out} is never zero; thus, the network output of a heat engine is always less than the amount of heat input. That is, only part of the heat transferred to the heat engine is converted to work. The fraction of the heat input that is converted to network output is a measure of the performance of a heat engine and is called the thermal efficiency η_{th} (Fig. 4.10). For heat engines, the desired output is the network output, and the required input is the amount of heat supplied to the working fluid. Then the thermal efficiency of a heat engine can be expressed as

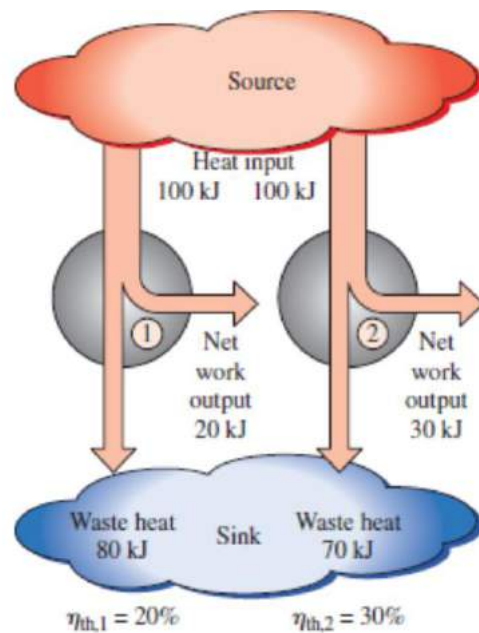


Fig. 4.10

$$\text{Thermal Efficiency} = \frac{\text{Desired Output}}{\text{Required input}} \quad (4-3)$$

Or
$$\eta_{th} = \frac{W_{net,output}}{Q_{in}} \quad (4-4)$$

It can also be expressed as

$$\eta_{th} = 1 - \frac{Q_{out}}{Q_{in}} \quad (4-5)$$

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

Cyclic devices of practical interest such as heat engines, refrigerators, and heat pumps operate between a high-temperature medium (or reservoir) at temperature T_H and a low-temperature medium (or reservoir) at temperature T_L . To bring uniformity to the treatment of heat engines, refrigerators, and heat pumps, we define these two quantities:

Q_H = magnitude of heat transfer between the cyclic device and the high temperature medium at temperature T_H

Q_L = magnitude of heat transfer between the cyclic device and the low temperature medium at temperature T_L .

Notice that both Q_L and Q_H are defined as *magnitudes* and therefore are positive quantities. The direction of Q_H and Q_L is easily determined by inspection. Then, the network output and thermal efficiency relations for any heat engine (shown in Fig. 4.11) can also be expressed as

$$W_{net,out} = Q_H - Q_L$$

and

$$\eta_{th} = \frac{W_{net,output}}{Q_H} \quad \text{or} \quad \eta_{th} = 1 - \frac{Q_L}{Q_H} \quad (4-6)$$

The thermal efficiency of a heat engine is always less than unity since both Q_L and Q_H are defined as positive quantities.

Thermal efficiency is a measure of how efficiently a heat engine converts the heat that it receives to work. Heat engines are built for the purpose of converting heat to work, and engineers are constantly trying to improve the

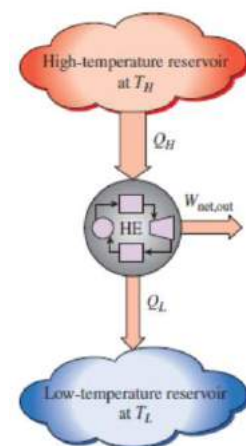


Fig. 4.11

efficiencies of these devices since increased efficiency means less fuel consumption and thus lower fuel bills and less pollution.

The thermal efficiencies of work-producing devices are relatively low. Ordinary spark-ignition automobile engines have a thermal efficiency of about 25 percent. That is, an automobile engine converts about 25 percent of the chemical energy of the gasoline to mechanical work. This number is as high as 40 percent for diesel engines and large gas-turbine plants and as high as 60 percent for large combined gas-steam power plants. Thus, even with the most efficient heat engines available today, almost one-half of the energy supplied ends up in the rivers, lakes, or the atmosphere as waste or useless energy.

Example 4.1

Heat is transferred to a heat engine from a furnace at a rate of 80 MW. If the rate of waste heat rejection to a nearby river is 50 MW, determine the net power output and the thermal efficiency for this heat engine.

Solution

$$\dot{Q}_H = 80 \text{ MW} \quad \text{and} \quad \dot{Q}_L = 50 \text{ MW}$$

The net power output of this heat engine is

$$\dot{W}_{\text{net}} = \dot{Q}_H - \dot{Q}_L = 80 - 50 = 30 \text{ MW}$$

Then the thermal efficiency is easily determined to be

$$\eta_{\text{th}} = \frac{\dot{W}_{\text{net}}}{\dot{Q}_H} = \frac{30}{80} = 0.375 \text{ (or 37.5\%)}$$

4.3.2 The Second Law of Thermodynamics: Kelvin–Planck Statement

Under the ideal conditions, a heat engine must reject some heat to a low-temperature reservoir in order to complete the cycle. That is, no heat

engine can convert all the heat it receives to useful work. This limitation on the thermal efficiency of heat engines forms the basis for the Kelvin–Planck statement of the second law of thermodynamics, which is expressed as follows:

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.

That is, a heat engine must exchange heat with a low-temperature sink as well as a high-temperature source to keep operating. The Kelvin–Planck statement can also be expressed as no heat engine can have a thermal efficiency of 100 percent (Fig. 4.12), or as for a power plant to operate, the working fluid must exchange heat with the environment as well as the furnace.

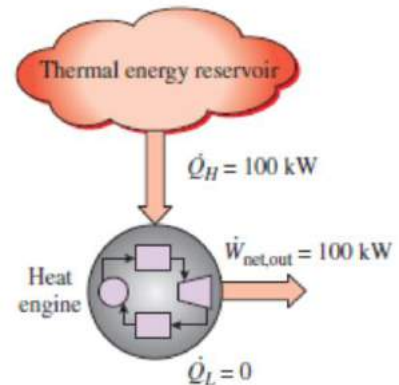


Fig. 4.12

Note that the impossibility of having a 100 percent efficient heat engine is not due to friction or other dissipative effects. It is a limitation that applies to both the idealized and the actual heat engines. Later in this chapter, we develop a relation for the maximum thermal efficiency of a heat engine. We also demonstrate that this maximum value depends on the reservoir temperatures only.

4.4 Refrigerators and Heat Pump:

We all know from experience that heat is transferred in the direction of decreasing temperature, that is, from high-temperature mediums to low temperature ones. This heat transfer process occurs in nature without requiring any devices. The reverse process, however, cannot occur by itself. The transfer of heat from a low-temperature medium to a high-temperature one requires special devices called refrigerators. Refrigerators, like heat

engines, are cyclic devices. The working fluid used in the refrigeration cycle is called a refrigerant. The most frequently used refrigeration cycle is the vapor-compression refrigeration cycle, which involves four main components: a compressor, a condenser, an expansion valve, and an evaporator, as shown in Fig. 4.13.

The refrigerant enters the compressor as a vapor and is compressed to the condenser pressure. It leaves the compressor at a relatively high

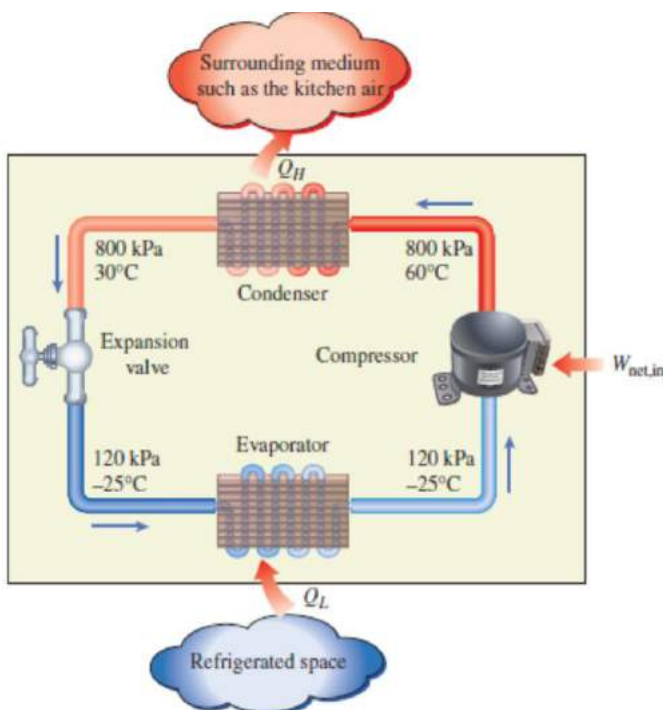


Fig. 4.13

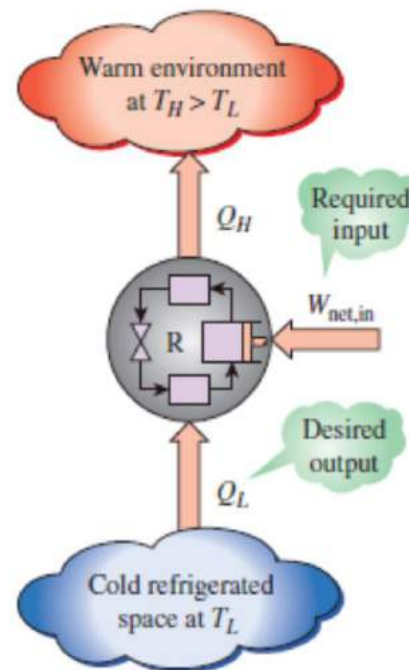


Fig. 4.14

temperature and cools down and condenses as it flows through the coils of the condenser by rejecting heat to the surrounding medium. It then enters a capillary tube where its pressure and temperature drop drastically due to the throttling effect. The low-temperature refrigerant then enters the evaporator, where it evaporates by absorbing heat from the refrigerated space. The cycle is completed as the refrigerant leaves the evaporator and reenters the compressor.

In a household refrigerator, the freezer compartment where heat is absorbed by the refrigerant serves as the evaporator, and the coils, usually behind the refrigerator where heat is dissipated to the kitchen air, serve as the condenser. A refrigerator is shown schematically in Fig. 4.14. Here Q_L is the magnitude of the heat removed from the refrigerated space at temperature T_L , Q_H is the magnitude of the heat rejected to the warm environment at temperature T_H , and $W_{\text{net,in}}$ is the network input to the refrigerator. As discussed before, Q_L and Q_H represent magnitudes and thus are positive quantities.

4.4.1 Coefficient of Performance

The efficiency of a refrigerator is expressed in terms of the coefficient of performance (COP), denoted by COP_R . The objective of a refrigerator is to remove heat (Q_L) from the refrigerated space. To accomplish this objective, it requires a work input of $W_{\text{net,in}}$. Then the COP of a refrigerator can be expressed as

$$\text{COP}_R = \frac{\text{Desired output}}{\text{Required Input}} = \frac{Q_L}{W_{\text{net,in}}} \quad (4-7)$$

This relation can also be expressed in rate form by replacing Q_L by \dot{Q}_L and $W_{\text{net,in}}$ by $\dot{W}_{\text{net,in}}$.

The conservation of energy principle for a cyclic device requires that

$$W_{\text{net,in}} = Q_H - Q_L \quad (\text{KJ}) \quad (4-8)$$

Then the COP_R relation becomes

$$\text{COP}_R = \frac{Q_L}{Q_H - Q_L} = \frac{1}{Q_H/Q_L - 1} \quad (4-9)$$

Notice that the value of COP_R can be *greater than unity*. That is, the amount of heat removed from the refrigerated space can be greater than the amount of work input. This is in contrast to the thermal efficiency, which can never be greater than 1. In fact, one reason for expressing the efficiency of a

refrigerator by another term the coefficient of performance is the desire to avoid the oddity of having efficiencies greater than unity.

4.4.2 Heat Pump

Another device that transfers heat from a low-temperature medium to a high temperature one is the heat pump, shown schematically in Fig. 4.15. 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. Discharging this heat to a higher temperature medium is merely a necessary part of the operation, not the purpose.

The objective of a heat pump, however, is to maintain a heated space at a high temperature. This is accomplished by absorbing heat from a low temperature source, such as well water or cold outside air in winter, and supplying this heat to the high-temperature medium such as a house (Fig. 4.16).

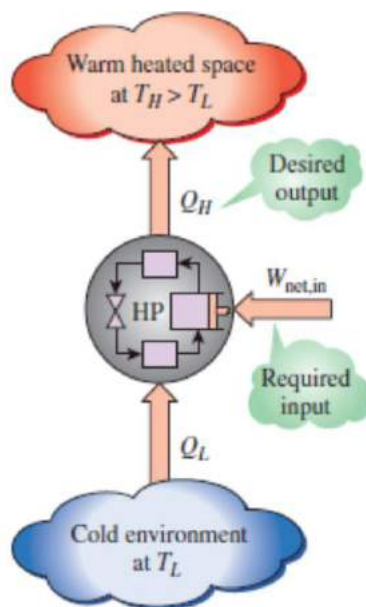


Fig. 4.15

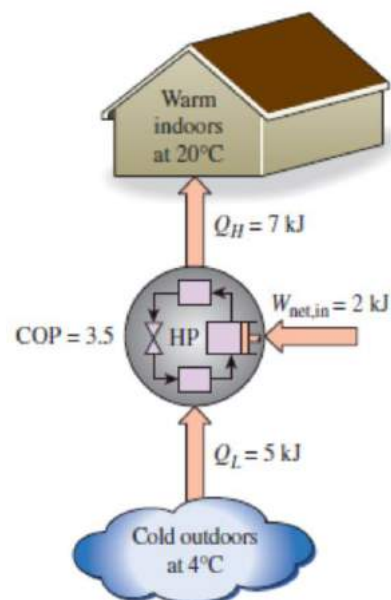


Fig. 4.16

An ordinary refrigerator that is placed in the window of a house with its door open to the cold outside air in winter will function as a heat pump since it

will try to cool the outside by absorbing heat from it and rejecting this heat into the house through the coils behind it.

The measure of performance of a heat pump is also expressed in terms of the coefficient of performance COP_{HP} , defined as

$$COP_{HP} = \frac{\text{Desired output}}{\text{Required Input}} = \frac{Q_H}{W_{\text{net,in}}} \quad (4-10)$$

which can also be expressed as

$$COP_{HP} = \frac{Q_H}{Q_H - Q_L} = \frac{1}{1 - Q_L/Q_H} \quad (4-11)$$

A comparison of Eqs. 4–7 and 4–10 reveals that

$$COP_{HP} = 1 + COP_R \quad (4-12)$$

for fixed values of Q_L and Q_H . This relation implies that the coefficient of performance of a heat pump is always greater than unity since COP_R is a positive quantity. That is, a heat pump will function, at worst, as a resistance heater, supplying as much energy to the house as it consumes. In reality, however, part of Q_H is lost to the outside air through piping and other devices, and COP_{HP} may drop below unity when the outside air temperature is too low. When this happens, the system usually switches to a resistance heating mode. Most heat pumps in operation today have a seasonally averaged COP of 2 to 3.

Most existing heat pumps use the cold outside air as the heat source in winter, and they are referred to as *air-source heat pumps*. The COP of such heat pumps is about 3.0 at design conditions. Air-source heat pumps are not appropriate for cold climates since their efficiency drops considerably when

temperatures are below the freezing point. In such cases, geothermal (also called ground-source) heat pumps that use the ground as the heat source can be used. Geothermal heat pumps require the burial of pipes in the ground 1 to 2 m deep. Such heat pumps are more expensive to install, but they are also more efficient (up to 45 percent more efficient than air-source heat pumps). The COP of ground-source heat pumps can be as high as 6 in the cooling mode.

Air conditioners are basically refrigerators whose refrigerated space is a room or a building instead of the food compartment. A window air-conditioning unit cools a room by absorbing heat from the room air and discharging it to the outside. The same air-conditioning unit can be used as a heat pump in winter by installing it backwards. In this mode, the unit absorbs heat from the cold outside and delivers it to the room. Air-conditioning systems that are equipped with proper controls and a reversing valve operate as air conditioners in summer and as heat pumps in winter.

Example 4.2

The food compartment of a refrigerator, shown in Fig. 4.17, is maintained at 4°C by removing heat from it at a rate of 360 kJ/min. If the required power input to the refrigerator is 2 kW, determine (a) the coefficient of performance of the refrigerator and (b) the rate of heat rejection to the room that houses the refrigerator.

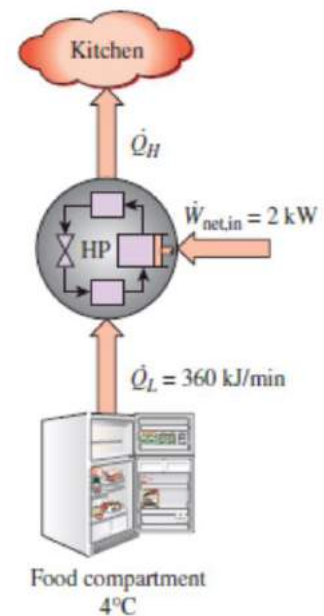


Fig. 4.17

Solution

$$\text{COP}_R = \frac{Q_L}{W_{\text{net,in}}} = \frac{360}{2 \times 60} = 3$$

$$\dot{Q}_H = \dot{Q}_L + \dot{W}_{\text{net,in}} = 360 + 2 \times 60 = 480 \text{ KJ/min}$$

Example 4.3

A heat pump is used to meet the heating requirements of a house and maintained at 20°C. On a day when the outdoor air temperature drops to -2°C, the house is estimated to lose heat at a rate of 80,000 kJ/h. If the heat pump under these conditions has a COP of 2.5, determine (a) the power consumed by the heat pump and (b) the rate at which heat is absorbed from the cold outdoor air.

Solution

$$W_{\text{net,in}} = \frac{\dot{Q}_H}{\text{COP}_R} = \frac{80000}{2.5} = 32000 \text{ KJ/h}$$

$$\dot{Q}_L = \dot{Q}_H - W_{\text{net,in}} = 80000 - 32000 = 48000 \text{ KJ/h}$$

4.4.3 The Second Law of Thermodynamics: Clausius Statement

There are two classical statements of the second law the Kelvin–Planck statement, which is related to heat engines and discussed in the preceding section, and the Clausius statement, which is related to refrigerators or heat pumps. The Clausius statement is expressed as follows:

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.

It is common knowledge that heat does not, of its own volition, transfer from a cold medium to a warmer one. The Clausius statement does not imply that a cyclic device that transfers heat from a cold medium to a warmer one is impossible

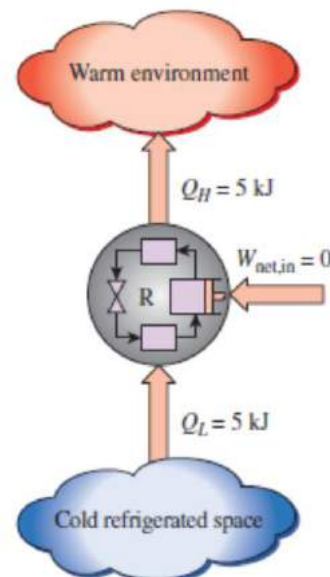


Fig. 4.18

to construct. In fact, this is precisely what a common household refrigerator does. It simply states that a refrigerator cannot operate unless its compressor is driven by an external power source, such as an electric motor (Fig. 4.18). This way, the net effect on the surroundings involves the consumption of some energy in the form of work, in addition to the transfer of heat from a colder body to a warmer one. That is, it leaves a trace in the surroundings. Therefore, a household refrigerator is in complete compliance with the Clausius statement of the second law.

Both the Kelvin–Planck and the Clausius statements of the second law are negative statements, and a negative statement cannot be proved. Like any other physical law, the second law of thermodynamics is based on experimental observations. To date, no experiment has been conducted that contradicts the second law, and this should be taken as sufficient proof of its validity.

4.4.4 Equivalence of the Two Statements

The Kelvin–Planck and the Clausius statements are equivalent in their consequences, and either statement can be used as the expression of the second law of thermodynamics. Any device that violates the Kelvin–Planck statement also violates the Clausius statement, and vice versa. This can be demonstrated as follows.

Consider the heat-engine-refrigerator combination shown in Fig. 4.19*a*, 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 Q_L from the low-temperature 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, as shown in Fig. 4.19b, that transfers heat in an amount of Q_L 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.

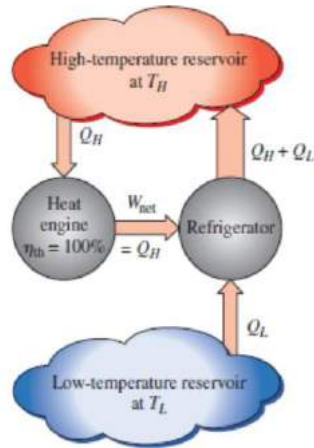


Fig. 4.19a

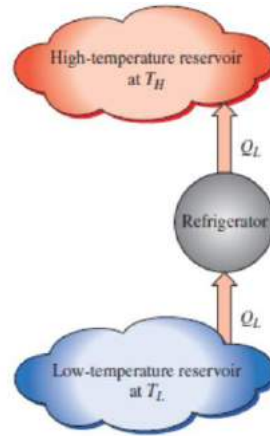


Fig. 4.19b

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 as shown in Fig. 4.20. Therefore, the Clausius and the Kelvin–Planck statements are two equivalent expressions of the second law of thermodynamics.

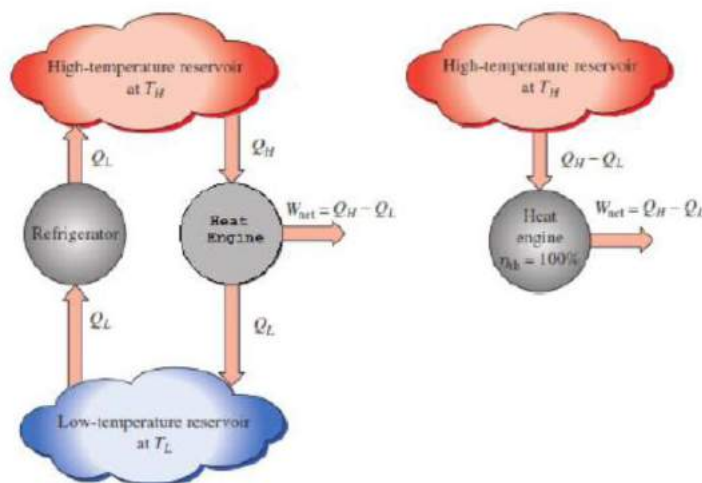


Fig. 4.20

4.5 Perpetual Motion Machines

We have repeatedly stated that a process cannot take place unless it satisfies both the first and second laws of thermodynamics. Any device that violates either law is called a **perpetual-motion machine**, and despite numerous attempts, no perpetual-motion machine is known to have worked. But this has not stopped inventors from trying to create new ones. A device that violates the **first law** of thermodynamics (by *creating* energy) is called a **perpetual-motion machine of the first kind (PMM1)**, and a device that violates the **second law** of thermodynamics is called a **perpetual-motion machine of the second kind (PMM2)**.

Consider the steam power plant shown in Fig. 4.21. It is proposed to heat the steam by resistance heaters placed inside the boiler, instead of by the energy supplied from fossil or nuclear fuels. Part of the electricity generated by the plant is to be used to power the resistors as well as the pump. The rest of the electric energy is to be supplied to the electric network as the network output. The inventor claims that once the system is started, this power plant will produce electricity indefinitely without requiring any energy input from the outside.

Well, here is an invention that could solve the world's energy problem if it works, of course. A careful examination of this invention reveals that the system enclosed by the shaded area is continuously supplying energy to the outside at a rate of $\dot{Q}_{out} + \dot{W}_{net,out}$ without receiving any energy. That is, this system is creating energy at a rate of $\dot{Q}_{out} + \dot{W}_{net,out}$, which is clearly a violation of the first law. Therefore, this wonderful device is nothing more than a PMM1 and does not warrant any further consideration.

Now let us consider another novel idea by the same inventor. Convinced that energy cannot be created, the inventor suggests the following modification

that will greatly improve the thermal efficiency of that power plant without violating the first law.

Aware that more than one-half of the heat transferred to the steam in the furnace is discarded in the condenser to the environment, the inventor suggests getting rid of this wasteful component and sending the steam to the pump as soon as it leaves the turbine, as shown in Fig. 4.22. This way, all the heat transferred to the steam in the boiler will be converted to work, and thus the power plant will have a theoretical efficiency of 100 percent. The inventor realizes that some heat losses and friction between the moving components are unavoidable and that these effects will hurt the efficiency somewhat, but still expects the efficiency to be no less than 80 percent (as opposed to 40 percent in most actual power plants) for a carefully designed system.

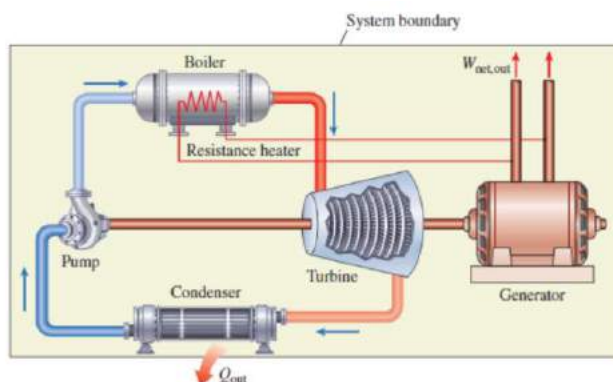


Fig. 4.21

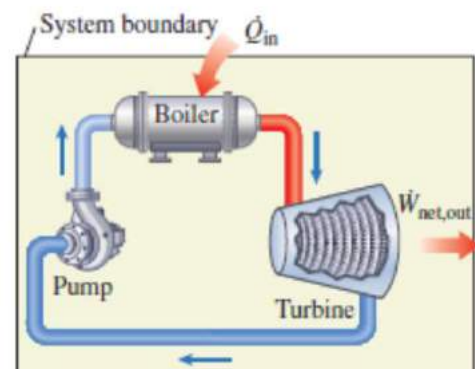


Fig. 4.22

4.6 Reversible and Irreversible Processes

The second law of thermodynamics states that no heat engine can have an efficiency of 100 percent. Then one may ask, what is the highest efficiency that a heat engine can possibly have? Before we can answer this question, we need to define an idealized process first, which is called the **reversible process**.

A reversible process is defined as a process that can be reversed without leaving any trace on the surroundings (Fig. 4.23). That is, both the system and the surroundings are returned to their initial states at the end of the reverse process.

Irreversible process is defined as a process that cannot be reversed without leaving any trace on the surroundings.

It should be pointed out that a system can be restored to its initial state following a process, regardless of whether the process is reversible or irreversible. But for reversible processes, this restoration is made without leaving any net change on the surroundings, whereas for irreversible processes, the surroundings usually do some work on the system and therefore does not return to their original state.

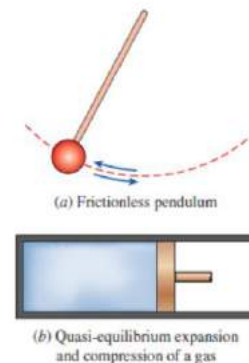


Fig. 4.23

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. You may be wondering, then, why we are bothering with such fictitious processes. There are two reasons. First, they 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.

Engineers are interested in reversible processes because work-producing devices such as car engines and gas or steam turbines deliver the most work, and work-consuming devices such as compressors, fans, and pumps consume the least work when reversible processes are used instead of irreversible ones.

Reversible processes can be viewed as theoretical limits for the corresponding irreversible ones. Some processes are more irreversible than others. We may never be able to have a reversible process, but we can certainly approach it. The more closely we approximate a reversible process, the more work delivered by a work-producing device or the less work required by a work-consuming device.

The concept of reversible processes leads to the definition of the second law efficiency for actual processes, which is the degree of approximation to the corresponding reversible processes. This enables us to compare the performance of different devices that are designed to do the same task on the basis of their efficiencies. The better the design, the lower the irreversibilities and the higher the second-law efficiency.

4.7 Irreversibilities

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.

4.7.1 Friction:

is a familiar form of irreversibility associated with bodies in motion. When two bodies in contact are forced to move relative to each other (a piston in a cylinder, for example, as shown in Fig. 4.24), a friction force that opposes the motion develops at the interface of these two bodies, and some work is needed to overcome this friction force. The energy supplied as work is eventually converted to heat during the process and is transferred to the bodies in contact, as evidenced by a temperature rise at the interface. When

the direction of the motion is reversed, the bodies are restored to their original position, but the interface does not cool, and heat is not converted back to work. Instead, more of the work is converted to heat while overcoming the friction forces that also oppose the reverse motion. Since the system (the moving bodies) and the surroundings cannot be returned to their original states, this process is irreversible. Therefore, any process that involves friction is irreversible. The larger the friction forces involved, the more irreversible the process is.

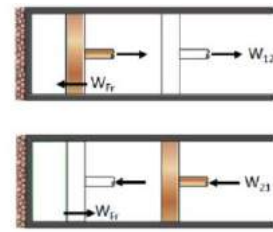


Fig. 4.24

$$W_{12} = W_{\text{ideal}} - W_{\text{Fr}} \quad (4-13)$$

$$W_{21} = -(W_{\text{ideal}} + W_{\text{Fr}}) \quad (4-14)$$

$$W_{12} + W_{21} = -2W_{\text{Fr}} \neq 0 \quad (4-15)$$

4.7.2 Unrestrained expansion:

Unrestrained expansion of a gas separated from a vacuum by a membrane, as shown in Fig. 4.25. When the membrane is ruptured, the gas fills the entire tank. The only way to restore the system to its original state is to compress it to its initial volume, while transferring heat from the gas until it reaches its initial temperature. From the conservation of energy considerations, it can easily be shown that the amount of heat transferred from the gas equals the amount of

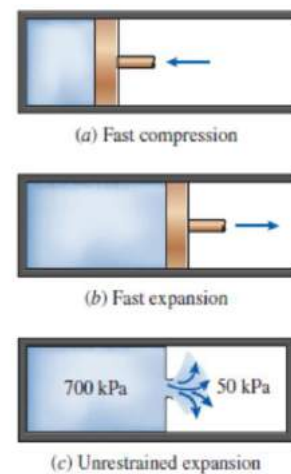


Fig. 4.25

work done on the gas by the surroundings. The restoration of the surroundings involves conversion of this heat completely to work, which would violate the second law. Therefore, unrestrained expansion of a gas is an irreversible process.

4.7.3 Heat transfer:

Heat transfer can occur only when there is a temperature difference between a system and its surroundings. Therefore, it is physically impossible to have a reversible heat transfer process. But a heat transfer process becomes less and less irreversible as the temperature difference between the two bodies approaches zero. Then, heat transfer through a differential temperature difference dT can be considered to be reversible. As dT approaches zero, the process can be reversed in direction (at least theoretically) without requiring any refrigeration. Notice that reversible heat transfer is a conceptual process and cannot be duplicated in the real world.

The smaller the temperature difference between two bodies, the smaller the heat transfer rate will be. Any significant heat transfer through a small temperature difference requires a very large surface area and a very long time. Therefore, even though approaching reversible heat transfer is desirable from a thermodynamic point of view, it is impractical and not economically feasible.

4.8 Internally and Externally Reversible Processes

4.8.1 Internally Reversible process:

is defined as there is no irreversibilities occur within the boundaries of the system during the process.

4.8.2 Externally Reversible process:

is defined as there is no irreversibilities occur outside the system boundaries during the process.

4.8.3 Totally Reversible process:

is defined as there are no irreversibilities within the system or its surroundings.

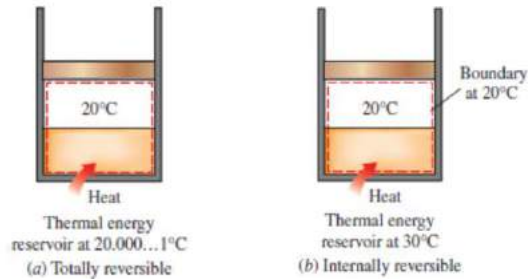


Fig. 4.26

4.9 The Carnot Cycle

We mentioned earlier that heat engines are cyclic devices and that the working fluid of a heat engine returns to its initial state at the end of each cycle. Work is done by the working fluid during one part of the cycle and on the working fluid during another part. The difference between these two is the network delivered by the heat engine. The efficiency of a heat engine cycle greatly depends on how the individual processes that make up the cycle are executed. The network, thus the cycle efficiency, can be maximized by using processes that require the least amount of work and deliver the most, that is, by using reversible processes. Therefore, it is no surprise that the most efficient cycles are reversible cycles, that is, cycles that consist entirely of reversible processes.

Reversible cycles cannot be achieved in practice because the irreversibilities associated with each process cannot be eliminated. However, reversible cycles provide upper limits on the performance of real cycles. Heat engines and refrigerators that work on reversible cycles serve as models to which actual heat engines

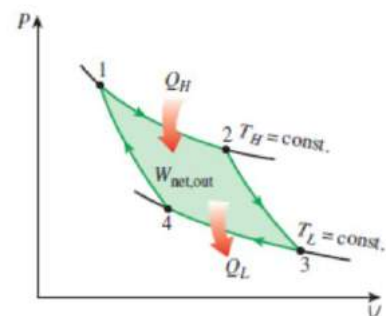


Fig. 4.27

and refrigerators can be compared. Reversible cycles also serve as starting points in the development of actual cycles and are modified as needed to meet certain requirements.

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 steady flow system. The four reversible processes that make up the Carnot cycle are as follows:

1. Reversible Isothermal Expansion (process 1-2, $T_H = \text{constant}$)
2. Reversible Adiabatic Expansion (process 2-3)
3. Reversible Isothermal Compression (process 3-4, $T_L = \text{constant}$)
4. Reversible Adiabatic Compression (process 4-1)

The P - v diagram of this cycle is shown in Fig. 4.27. Remembering that on a P - v diagram the area under the process curve represents the boundary work for quasi-equilibrium (internally reversible) processes, we see that the area under curve 1-2-3 is the work done by the gas during the expansion part of the cycle, and the area under curve 3-4-1 is the work done on the gas during the compression part of the cycle. The area enclosed by the path of the cycle (area 1-2-3-4-1) is the difference between these two and represents the net work done during the cycle.

4.10 The Reversed Carnot Cycle

The Carnot heat engine cycle just described is a totally reversible cycle. Therefore, all the processes that comprise it can be reversed, in which case it becomes the Carnot refrigeration cycle. This time, the cycle remains exactly the same, except that the directions of any heat and work interactions are reversed: Heat in the amount of Q_L is absorbed from the low-temperature

medium, heat in the amount of Q_H is rejected to a high-temperature medium, and a work input of $W_{\text{net,in}}$ is required to accomplish all this.

The P - v diagram of the reversed Carnot cycle is the same as the one given for the Carnot cycle, except that the directions of the processes are reversed, as shown in Fig. 4.28.

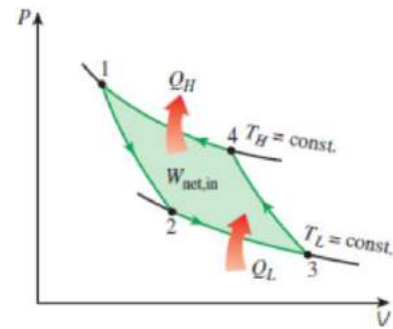


Fig. 4.28

4.11 The Carnot Principles

The second law of thermodynamics puts limits on the operation of cyclic devices as expressed by the Kelvin–Planck and Clausius statements. A heat engine cannot operate by exchanging heat with a single reservoir, and a refrigerator cannot operate without a net energy input from an external source. We can draw valuable conclusions from these statements. Two conclusions pertain to the thermal efficiency of reversible and irreversible (i.e., actual) heat engines and they are known as the Carnot principles (Fig. 4.29), expressed as follows:

4.11.1 The First Carnot Principle:

The efficiency of an irreversible heat engine is always less than the efficiency of a reversible one operating between the same two reservoirs.

$$\eta_{\text{th,irrev}} < \eta_{\text{th,rev}} \quad , \quad W_{\text{irrev}} < W_{\text{rev}} \quad , \quad Q_{\text{L,irrev}} > Q_{\text{L,rev}} \quad (4-16)$$

4.11.2 The Second Carnot Principle:

The efficiencies of all reversible heat engines operating between the same two reservoirs are the same.

$$\eta_{th,1} = \eta_{th,2} \quad , \quad W_1 = W_2 \quad , \quad Q_{L1} = Q_{L2} \quad (4-17)$$

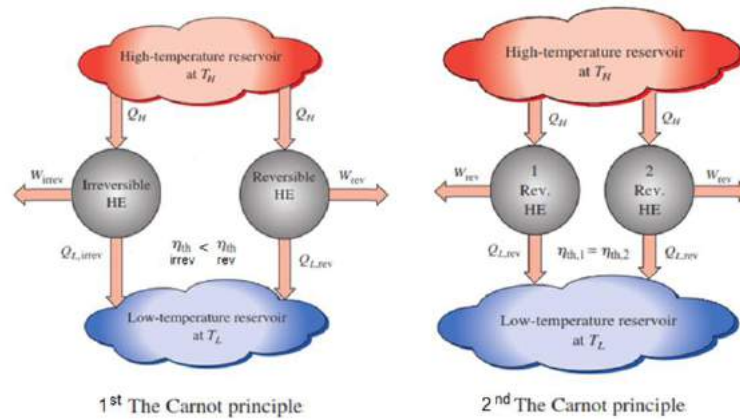


Fig. 4.29

4.12 The Thermodynamic Temperature Scale

A temperature scale that is independent of the properties of the substances that are used to measure temperature is called a thermodynamic temperature scale. Such a temperature scale offers great conveniences in thermodynamic calculations, and its derivation is given below using some reversible heat engines.

The second Carnot principle discussed in previous Section states that all reversible heat engines have the same thermal efficiency when operating between the same two reservoirs. That is, the efficiency of a reversible engine is independent of the working fluid employed and its properties, the way the cycle is executed, or the type of reversible engine used. Since energy reservoirs are characterized by their temperatures, the thermal efficiency of reversible heat engines is a function of the reservoir temperatures only. That is,

$$(\eta_{th})_{rev} = g(T_H, T_L) \quad \text{or} \quad \eta_{th} = 1 - \left(\frac{Q_L}{Q_H}\right)_{rev} = g(T_H, T_L) \quad (4-18)$$

So

$$\left(\frac{Q_L}{Q_H}\right)_{rev} = f(T_H, T_L) \quad (4-19)$$

Kelvin proposed that:

$$Q(T_H) = T_H \quad , \quad Q(T_L) = T_L \quad \text{or} \quad \left(\frac{Q_L}{Q_H}\right)_{rev} = \frac{T_L}{T_H} \quad (4-20)$$

This temperature scale is called the Kelvin scale, and the temperatures on this scale are called absolute temperatures. On the Kelvin scale, the temperature ratios depend on the ratios of heat transfer between a reversible heat engine and the reservoirs and are independent of the physical properties of any substance. On this scale, temperatures vary between zero and infinity.

4.13 The Carnot Heat Engine

The hypothetical heat engine that operates on the reversible Carnot cycle is called the Carnot heat engine. The thermal efficiency of any heat engine, reversible or irreversible, is given by Eq. 4–6 as

$$\eta_{th} = 1 - \frac{Q_L}{Q_H}$$

where Q_H is heat transferred to the heat engine from a high-temperature reservoir at T_H , and Q_L is heat rejected to a low-temperature reservoir at T_L . For reversible heat engines, the heat transfer ratio in the above relation can be replaced by the ratio of the absolute temperatures of the two reservoirs, as given by Eq. 4–20. Then the efficiency of a Carnot engine, or any reversible heat engine, becomes

$$\eta_{th,rev} = 1 - \frac{T_L}{T_H} \quad (4-21)$$

This relation is often referred to as the Carnot efficiency, since the Carnot heat engine is the best known reversible engine. This is the highest efficiency a heat engine operating between the two thermal energy reservoirs at temperatures T_L and T_H can have. All irreversible (i.e., actual) heat engines operating between these temperature limits (T_L and T_H) have lower efficiencies. An actual heat engine cannot reach this maximum theoretical

efficiency value because it is impossible to completely eliminate all the irreversibilities associated with the actual cycle.

Note that T_L and T_H in Eq. 4–21 are *absolute temperatures*. Using °C or °F for temperatures in this relation gives results grossly in error. The thermal efficiencies of actual and reversible heat engines operating between the same temperature limits compare as follows:

$$\eta_{th} \begin{cases} < \eta_{th,rev} & \text{irreversible heat engine} \\ = \eta_{th,rev} & \text{reversible heat engine} \\ > \eta_{th,rev} & \text{impossible heat engine} \end{cases} \quad (4-22)$$

Example 4.4

A Carnot heat engine, receives 500 kJ of heat per cycle from a high-temperature source at 652°C and rejects heat to a low-temperature sink at 30°C. Determine (a) the thermal efficiency of this Carnot engine and (b) the amount of heat rejected to the sink per cycle.

Solution

$$\eta_{th} = 1 - \frac{T_L}{T_H} = 1 - \frac{(30 + 273)K}{(652 + 273)K} = 0.672$$

$$Q_{L,rev} = \frac{T_L}{T_H} Q_{H,rev} = \frac{(30+273)K}{(652+273)K} (500KJ) = 164 \text{ KJ}$$

4.14 The Carnot Refrigerator and Heat Pump

A refrigerator or a heat pump that operates on the reversed Carnot cycle is called a Carnot refrigerator, or a Carnot heat pump. The coefficient of performance of any refrigerator or heat pump, reversible or irreversible, is given by Eqs. 4–9 and 4–11 as

$$COP_R = \frac{Q_L}{Q_H - Q_L} = \frac{1}{Q_H/Q_L - 1} \quad \text{and} \quad COP_{HP} = \frac{Q_H}{Q_H - Q_L} = \frac{1}{1 - Q_L/Q_H}$$

Where Q_L is the amount of heat absorbed from the low-temperature medium and Q_H is the amount of heat rejected to the high-temperature medium. The COP of all reversible refrigerators or heat pumps can be determined by replacing the heat transfer ratios in the above relations by the ratios of the absolute temperatures of the high- and low-temperature reservoirs, as expressed by Eq. 4–21. Then the COP relations for reversible refrigerators and heat pumps become These are the highest coefficients of performance that a refrigerator or a heat pump operating between the temperature limits of T_L and T_H can have. All actual refrigerators or heat pumps operating between these temperature limits (T_L and T_H) have lower coefficients of performance.

$$COP_{R,rev} = \frac{T_L}{T_H - T_L} = \frac{1}{T_H/T_L - 1} \quad (4-23)$$

And

$$COP_{HP,rev} = \frac{T_H}{T_H - T_L} = \frac{1}{1 - T_L/T_H} \quad (4-24)$$

The coefficients of performance of actual and reversible refrigerators operating between the same temperature limits can be compared as follows:

$$COP_R \begin{cases} < COP_{R,rev} & \text{irreversible refrigerator} \\ = COP_{R,rev} & \text{reversible refrigerator} \\ > COP_{R,rev} & \text{impossible refrigerator} \end{cases} \quad (4-25)$$

A similar relation can be obtained for heat pumps by replacing all COP_R 's in Eq. 4–25 by COP_{HP} . The COP of a reversible refrigerator or heat pump is the maximum theoretical value for the specified temperature limits. Actual refrigerators or heat pumps may approach these values as their designs are improved, but they can never reach them.

As a final note, the COP's of both the refrigerators and the heat pumps decrease as T_L decreases. That is, it requires more work to absorb heat from

lower-temperature media. As the temperature of the refrigerated space approaches zero, the amount of work required to produce a finite amount of refrigeration approaches infinity and COP_R approaches zero.

Example 4.5

A Carnot refrigeration cycle is executed in a closed system in the saturated liquid–vapor mixture region using 0.8 kg of refrigerant-134a as the working fluid. The maximum and the minimum temperatures in the cycle are 20 and -8°C , respectively. It is known that the refrigerant is saturated liquid at the end of the heat rejection process, and the network input to the cycle is 15 kJ. Determine the fraction of the mass of the refrigerant that vaporizes during the heat addition process, and the pressure at the end of the heat rejection process.

Solution

$$\text{COP}_R = \frac{1}{T_H/T_L - 1} = \frac{1}{(20 + 273 \text{ K})/(-8 + 273 \text{ K}) - 1} = 9.464$$

$$Q_L = \text{COP}_R \times W_{\text{in}} = (9.464)(15 \text{ kJ}) = 142 \text{ kJ}$$

$$Q_L = m_{\text{evap}} h_{fg@-8^\circ\text{C}} \rightarrow m_{\text{evap}} = \frac{142 \text{ kJ}}{204.59 \text{ kJ/kg}} = 0.694 \text{ kg}$$

$$\text{Mass fraction} = \frac{m_{\text{evap}}}{m_{\text{total}}} = \frac{0.694 \text{ kg}}{0.8 \text{ kg}} = \mathbf{0.868} \text{ or } \mathbf{86.8\%}$$

The pressure at the end of heat rejection process is simply the saturation pressure at heat rejection temperature,

$$P_4 = P_{\text{sat}@20^\circ\text{C}} = \mathbf{572.1 \text{ kPa}}$$

Example 4.6

A heat pump is to be used to heat a house during the winter. The house is to be maintained at 21°C at all times. The house is estimated to be losing heat at a rate of 135,000 kJ/h when the outside temperature drops to -5°C. Determine the minimum power required to drive this heat pump.

Solution

$$\text{COP}_{\text{HP,rev}} = \frac{1}{1 - T_L/T_H} = \frac{1}{1 - (-5 + 273 \text{ K})/(21 + 273 \text{ K})} = 11.3$$

$$\dot{W}_{\text{net,in}} = \frac{\dot{Q}_H}{\text{COP}_{\text{HP}}} = \frac{37.5 \text{ kW}}{11.3} = 3.32 \text{ kW}$$

Problems (Second Law of Thermodynamics)

- 4.1. a) What is a thermal energy reservoir? Give some examples.
b) Is it possible for a heat engine to operate without rejecting any waste heat to a low-temperature reservoir? Explain.
c) What are the characteristics of all heat engines?
- 4.2 Heat is transferred to a heat engine from a furnace at a rate of 90 MW. If the rate of waste heat rejection to a nearby river is 60 MW, determine the net power output and the thermal efficiency for this heat engine.
- 4.3 A 600-MW steam power plant, which is cooled by a nearby river, has a thermal efficiency of 40 percent. Determine the rate of heat transfer to the river water. Will the actual heat transfer rate be higher or lower than this value? Why?
- 4.4 A steam power plant receives heat from a furnace at a rate of 280 GJ/h. Heat losses to the surrounding air from the steam as it passes through the pipes and other components are estimated to be about 8 GJ/h. If the waste heat is transferred to the cooling water at a rate of 145 GJ/h, determine (a) net power output and (b) the thermal efficiency of this power plant.
- 4.5 An automobile engine consumes fuel at a rate of 28 L/h and delivers 60 kW of power to the wheels. If the fuel has a heating value of 44,000 kJ/kg and a density of 0.8 g/cm^3 , determine the efficiency of this engine.
- 4.6. a) What is the difference between a refrigerator and a heat pump?
b) What is the difference between a refrigerator and an air conditioner?
c) In a refrigerator, heat is transferred from a lower temperature medium (the refrigerated space) to a higher temperature one (the kitchen air). Is this a violation of the second law of thermodynamics? Explain.

- d) Show that the Kelvin–Planck and the Clausius expressions of the second law are equivalent.
- 4.7 A household refrigerator with a COP of 1.2 removes heat from the refrigerated space at a rate of 60 kJ/min. Determine (a) the electric power consumed by the refrigerator and (b) the rate of heat transfer to the kitchen air.
- 4.8 Determine the COP of a heat pump that supplies energy to a house at a rate of 8000 kJ/h for each kW of electric power it draws. Also, determine the rate of energy absorption from the outdoor air.
- 4.9 Determine the COP of a refrigerator that removes heat from the food compartment at a rate of 5040 kJ/h for each kW of power it consumes. Also, determine the rate of heat rejection to the outside air.
- 4.10 A heat pump with a COP of 2.5 supplies energy to a house at a rate of 60,000 kJ/h. Determine (a) the electric power drawn by the heat pump and (b) the rate of heat absorption from the outside air.
- 4.11 A Carnot heat engine operates between a source at 1000 K and a sink at 300 K. If the heat engine is supplied with heat at a rate of 800 kJ/min, determine (a) the thermal efficiency and (b) the power output of this heat engine.
- 4.12 A heat engine is operating on a Carnot cycle and has a thermal efficiency of 55 percent. The waste heat from this engine is rejected to a nearby lake at 40°C at a rate of 800 kJ/min. Determine (a) the power output of the engine and (b) the temperature of the source.
- 4.13 An inventor claims to have developed a heat engine that receives 700 kJ of heat from a source at 500 K and produces 300 kJ of network while

rejecting the waste heat to a sink at 290 K. Is this a reasonable claim? Why?

- 4.14 An experimentalist claims that, based on his measurements, a heat engine receives 300 kJ of heat from a source of 900 k, converts 160 kJ of it to work, and rejects the rest as waste heat to a sink at 540 k. Are these measurements reasonable? Why?
- 4.15 A Carnot refrigerator operates in a room in which the temperature is 22°C and consumes 2 kW of power when operating. If the food compartment of the refrigerator is to be maintained at 3°C, determine the rate of heat removal from the food compartment.
- 4.16 An inventor claims to have developed a refrigeration system that removes heat from the closed region at -12°C and transfers it to the surrounding air at 25°C while maintaining a COP of 6.5. Is this claim reasonable? Why?
- 4.17 During an experiment conducted in a room at 25°C, a laboratory assistant measures that a refrigerator that draws 2 kW of power has removed 30,000 kJ of heat from the refrigerated space, which is maintained at -30°C. The running time of the refrigerator during the experiment was 20 min. Determine if these measurements are reasonable.
- 4.18 A Carnot heat engine receives heat from a reservoir at 900°C at a rate of 800 kJ/min and rejects the waste heat to the ambient air at 27°C. The entire work output of the heat engine is used to drive a refrigerator that removes heat from the refrigerated space at -5°C and transfers it to the same ambient air at 27°C. Determine (a) the maximum rate of heat removal from the refrigerated space and (b) the total rate of heat rejection to the ambient air.

References

Yunus A. Cengel, Michael A. Boles, (2015) "Thermodynamics an engineering approach", eighth edition .