

# CHAPTER ONE

## FUNDAMENTALS OF ENERGY CONVERSION

### 1.1 Introduction

Energy conversion engineering (or heat-power engineering, as it was called prior to the Second World War), has been one of the central themes in the development of the engineering profession. It is concerned with the transformation of energy from sources such as fossil and nuclear fuels and the sun into conveniently used forms such as electrical energy, rotational and propulsive energy, and heating and cooling.

A multitude of choices and challenges face the modern energy conversion engineer. A few years ago major segments of the energy conversion industry were settled into a pattern of slow innovation. Most automobile manufacturers were satisfied to manufacture engines that had evolved from those produced twenty years earlier, some of which boasted 400 horsepower and consumed a gallon of leaded gasoline every eight or nine miles. Many electric power utilities were content with state-of-the-art, reliable, fossil-fuel-consuming steam power plants, except for a few forward-looking, and in several cases unfortunate, exceptions that risked the nuclear alternative.

Then came the oil embargo of the 1970s, high fuel prices, and threatened shortages. Also, the public and legislatures began to recognize that air pollution produced by factories, power plants, and automobiles and other forms of environmental pollution were harmful. International competitors, producing quality automobiles with smaller, lower-pollution engines, exceptional gas mileage, and lower prices shook the automobile industry. The limitations of the Earth's resources and environment started to come into clearer focus. These and other influences have been helping to create a more favorable climate for consideration, if not total acceptance, of energy conversion alternatives and new concepts.

There are opposing factors, however. Among them are limited research and development funding due to budgetary constraints, emphasis on short-term rather than long-term goals because of entrepreneurial insistence on rapid payback on investment, and managerial obsession with the bottom line. But more open attitudes have become established. New as well as previously shelved ideas are now being considered or reconsidered, tested, and sometimes implemented. A few examples are combined steam and gas turbine cycles, rotary combustion engines, solar and windmill power farms, stationary and vehicular gas turbine power plants, cogeneration, photovoltaic solar power, refuse-derived fuel, stratified charge engines, turbocharged engines, fluidized-bed combustors, and coal-gasification power plants. We are living in a

rapidly changing world that requires continuing adaptation of old technologies and the development of new ones. Energy conversion engineering is a more stimulating, complex, and viable field today because of this altered climate.

## A Look Backward

How did we get where we are today? A good answer requires a study of the history of science and engineering worthy of many volumes. Table 1.1 identifies a few pivotal ideas and inventions, some of them landmarks to energy conversion engineers, and the names of the thinkers and movers associated with them. Of course, the table cannot present the entire history of energy conversion engineering. Omitted are the contributions of Newton and Euler, the first rocket engine, the V-8 engine, the ramjet and fanjet. The reader could easily come up with many other glaring omissions and extend the table indefinitely.

While the names of one or two persons are associated with each landmark achievement, most of these landmarks were the products of teams of unheralded individuals whose talents were crucial to success. Moreover, the successes did not occur in a vacuum, but benefited and followed from the advances and failures of others. Unknown or renowned, each engineer and his or her associates can make a contribution to the progress of mankind.

Table 1.1 can only hint at how the persons, ideas, and events listed there relied on their predecessors and on a host of less well-known scientific and technological advances. A brief bibliography of historical sources is given at the end of the chapter. These works chronicle the efforts of famous and unsung heroes, and a few villains, of energy conversion and their struggles with ideas and limiting tools and resources to produce machines for man and industry.

The historical progress of industry and technology was slow until the fundamentals of thermodynamics and electromagnetism were established in the nineteenth century. The blossoming of energy technology and its central role in the industrial revolution is well known to all students of history. It is also abundantly clear that the development of nuclear power in the second half of the twentieth century grew from theoretical and experimental scientific advances of the first half century. After a little reflection on Table 1.1, there should be no further need to justify a fundamental scientific and mathematical approach to energy conversion engineering.

**TABLE 1-1** Some Significant Events in the History of Energy Conversion

Giovanni Branca	Impulse steam turbine proposal	1629
Thomas Newcomen	Atmospheric engine using steam (first widely used Heat engine)	1700
James Watt	Separate steam condenser idea; and first Boulton and Watt condensing steam engine	1765 1775

**Table 1.1** (continued)

John Barber	Gas turbine ideas and patent	1791
Benjamin Thompson (Count Rumford)	Observed conversion of mechanical energy to heat while boring cannon	1798
Robert Fulton	First commercial steamboat	1807
Robert Stirling	Stirling engine	1816
N. L. Sadi Carnot	Principles for an ideal heat engine (foundations of thermodynamics)	1824
Michael Faraday	First electric current generator	1831
Robert Mayer	Equivalence of heat and work	1842
James Joule	Basic ideas of the First Law of Thermodynamics; and measured the mechanical equivalent of heat	1847 1849
Rudolph Clausius	Second Law of Thermodynamics	1850
William Thompson (Lord Kelvin)	Alternate form of the Second Law of Thermodynamics	1851
Etienne Lenoir	Internal combustion engine without without mechanical compression	1860
A. Beau de Rochas	Four-stroke cycle internal combustion engine concept	1862
James C. Maxwell	Mathematical principles of electromagnetics	1865
Niklaus Otto	Four-stroke cycle internal combustion engine	1876
Charles Parsons	Multistage, axial-flow reaction steam turbine	1884
Thomas Edison	Pearl Street steam-engine-driven electrical power plant	1884
C.G.P. de Laval	Impulse steam turbine with convergent-divergent nozzle	1889
Rudolph Diesel	Compression ignition engine	1892
—	First hydroelectric power at Niagara Falls	1895
Albert Einstein	Mass-energy equivalence	1905
Ernst Schrodinger	Quantum wave mechanics	1926
Frank Whittle	Turbojet engine patent application; and first jet engine static test	1930 1937

**Table 1.1** (concluded)

Otto Hahn	Discovery of nuclear fission	1938
Hans von Ohain	First turbojet engine flight	1939
J. Ackeret, C. Keller	Closed-cycle gas turbine electric power generation	1939
Enrico Fermi	Nuclear fission demonstration at the University of Chicago	1942
Felix Wankel	Rotary internal combustion engine	1954
	Production of electricity via nuclear fission by a utility at Shippingport, Pennsylvania	1957
NASA	Rocket-powered landing of man on the moon	1969
Electricité de France	Superphénix 1200-MW fast breeder reactor – first grid power	1986

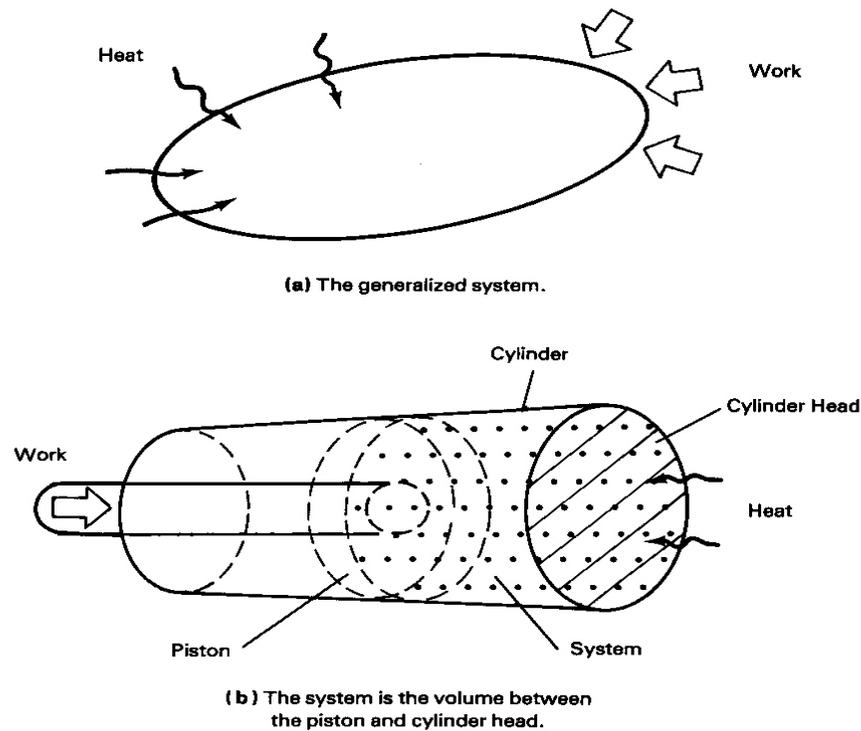
Since energy conversion engineering is deeply rooted in thermodynamics, fluid mechanics, and heat transfer, this chapter briefly reviews those aspects of these disciplines that are necessary for understanding, analysis, and design in the field of energy conversion.

## 1.2 Fundamentals of Thermodynamics

The subject of thermodynamics stems from the notions of temperature, heat, and work. Although, the following discussion makes occasional reference to molecules and particles, useful in clarifying and motivating concepts in thermodynamics, thermodynamics as a science deals with matter as continuous rather than as discrete or granular.

### System, Surroundings, and Universe

We define a *pure substance* as a homogeneous collection of matter. Consider a fixed mass of a pure substance bounded by a closed, impenetrable, flexible surface. Such a mass, called a *system*, is depicted schematically in Figure 1.1(a). For example, the system could be a collection of molecules of water, air, refrigerant, or combustion gas confined in a closed container such as the boundary formed by a cylinder and a fitted piston, Figure 1.1 (b). A system should always be defined carefully, to ensure that the same particles are in the system at all times. All other matter which can interact with the system is called the *surroundings*. The combination of the system and the surroundings is termed the *universe*, used here not in a cosmological sense, but to include only the



**FIGURE 1.1** System concept.

system and all matter which could interact with the system. Thermodynamics and energy conversion are concerned with changes in the system and in its interactions with the surroundings.

### State

The mass contained within a system can exist in a variety of conditions call *states*. Qualitatively, the concept of state is familiar. For example, the system state of a gas might be described qualitatively by saying that the system is at a high temperature and a low pressure. Values of temperature and pressure are characteristics that identify a particular condition of the system. Thus a unique condition of the system is called a state.

### Thermodynamic Equilibrium

A system is said to be in *thermodynamic equilibrium* if, over a long period of time, no change in the character or state of the system is observed.

## Thermodynamic Properties

It is a fundamental assumption of thermodynamics that a state of thermodynamic equilibrium of a given system may be described by a few observable characteristics called *thermodynamic properties*, such as pressure, temperature, and volume. Obviously, this approach excludes the possibility of description of the condition of the molecules of the system, a concern that is left to the fields of statistical and quantum mechanics and kinetic theory. Nevertheless, it is frequently useful to think of thermodynamic phenomena in molecular terms.

**The Temperature Property.** Temperature is a measure of the vigor of the molecular activity of a system. How can it be observed? A thermometer measures a system property called *temperature* when it is in intimate and prolonged contact (thermodynamic equilibrium) with the system. A mercury-in-glass thermometer, for instance, functions by thermal expansion or contraction of mercury within a glass bulb. The bulb must be in intimate thermal contact with the observed system so that the temperatures of the bulb and the system are the same. As a result of the equilibrium, elongation or contraction of a narrow column of mercury connected to the bulb indicates the temperature change of the system with which it is in contact.

**The Pressure Property.** Another way to observe changes in the state of a liquid or gaseous system is to connect a manometer to the system and observe the level of the free surface of the manometer fluid. The manometer free surface rises or drops as the force per unit area or *pressure* acting on the manometer-system interface changes.

## Defining a State

It has been empirically observed that an equilibrium state of a system containing a single phase of a pure substance is defined by two thermodynamic properties. Thus, if we observe the temperature and pressure of such a system, we can identify when the system is in a particular thermodynamic state.

## Extensive and Intensive properties

Properties that are dependent on mass are known as *extensive properties*. For these properties that indicate quantity, a given property is the sum of the the corresponding properties of the subsystems comprising the system. Examples are internal energy and volume. Thus, adding the internal energies and volumes of subsystems yields the internal energy and the volume of the system, respectively.

In contrast, properties that may vary from point to point and that do not change with the mass of the system are called *intensive properties*. Temperature and pressure are well-known examples. For instance, thermometers at different locations in a system may indicate differing temperatures. But if a system is in equilibrium, the temperatures

of all its subsystems must be identical and equal to the temperature of the system. Thus, a system has a single, unique temperature only when it is at equilibrium.

## Work

From basic mechanics, *work*,  $W$ , is defined as the energy provided by an entity that exerts a force,  $F$ , in moving one or more particles through a distance,  $x$ . Thus work must be done by an external agent to decrease the volume,  $V$ , of a system of molecules. In the familiar piston-cylinder arrangement shown in Figure 1.1(b), an infinitesimal volume change of the system due to the motion of the piston is related to the differential work through the force-distance product:

$$dW = Fdx = pAdx = pdV \quad [\text{ft}\cdot\text{lb}_f \mid \text{n}\cdot\text{m}] \quad (1.1a)$$

or

$$dw = pdv \quad [\text{Btu}/\text{lb}_m \mid \text{kJ}/\text{kg}] \quad (1.1b)$$

where  $p$  is the system pressure, and  $A$  is the piston cross-sectional area.

Note that in Equation (1.1b), the lower case letters  $w$  and  $v$  denote work and volume on a unit mass basis. All extensive properties, i.e., those properties of state that are proportional to mass, are denoted by lowercase characters when on a unit mass basis. These are called *specific properties*. Thus, if  $V$  represents volume, then  $v$  denotes *specific volume*. Although work is not a property of state, it is dealt with in the same way.

Also note that the English units of energy in Equation (1.1a) are given in mechanical units. Alternately, the British Thermal unit [Btu] may be used, as in Equation (1.1b). The two sets of units are related by the famous conversion factor known as the *mechanical equivalent of heat*, 778 ft·lb<sub>f</sub>/Btu. The student should pay close attention to the consistency of units in all calculations. Conversion factors are frequently required and are not explicitly included in many equations. For the convenience of the reader, Appendix A lists physical constants and conversion factors.

When work decreases the volume of a system, the molecules of the system move closer together. The moving molecules then collide more frequently with each other and with the walls of their container. As a result, the average forces (and hence pressures) on the system boundaries increase. Thus the state of the system may be changed by work done on the system.

## Heat

Given a system immersed in a container of hot fluid, by virtue of a difference in temperature between the system and the surrounding fluid, energy passes from the fluid to the system. We say that *heat*,  $Q$  [Btu | kJ], is transferred to the system. The system is

observed to increase in temperature or to change phase or both. Thus heat transfer to or from the system, like work, can also change the state of the matter within the system.

When the system and the surrounding fluid are at the same temperature, no heat is transferred. In this case the system and surroundings are said to be in *thermal equilibrium*. The term *adiabatic* is used to designate a system in which no heat crosses the system boundaries. A system is often approximated as an adiabatic system if it is well insulated.

### Heat and Work Are Not Properties

Mechanics teaches that work can change the kinetic energy of mass and can change the elevation or potential energy of mass in a gravitational field. Thus work performed by an outside agent on the system boundary can change the energy associated with the particles that make up the system. Likewise, heat is energy crossing the boundary of a system, increasing or decreasing the energy of the molecules within. Thus heat and work are not properties of state but forms of energy that are transported across system boundaries to or from the environment. They are sometimes referred to as *energy in transit*. Energy conversion engineering is vitally concerned with devices that use and create energy in transit.

### Internal Energy and The First Law of Thermodynamics

A property of a system that reflects the energy of the molecules of the system is called the *internal energy*,  $U$ . The Law of Conservation of Energy states that energy can be neither created nor destroyed. Thus the internal energy of a system can change only when energy crosses a boundary of the system, i.e., when heat and/or work interact with the system. This is expressed in an equation known as the *First Law of Thermodynamics*. In differential form the First Law is:

$$du = dq - dw \quad [\text{Btu/lb}_m \mid \text{kJ/kg}] \quad (1.2)$$

Here,  $u$  is the internal energy per unit mass, a property of state, and  $q$  and  $w$  are, respectively, heat and work per unit mass. The differentials indicate infinitesimal changes in quantity of each energy form. Here, we adopt the common sign convention of thermodynamics that both the heat entering the system and work done by the system are positive. This convention will be maintained throughout the text. Thus Equation (1.2) shows that heat into the system (positive) and work done on the system (negative) both increase the system's internal energy.

### Cyclic Process

A special and important form of the First Law of Thermodynamics is obtained by

integration of Equation (1.2) for a cyclic process. If a system, after undergoing arbitrary change due to heat and work, returns to its initial state, it is said to have participated in a *cyclic process*. The key points are: (1) the integral of any state property differential is the difference of its limits, and (2) the final state is the same as the initial state (hence there is no change in internal energy of the system)

$$\oint du = u_f - u_i = 0$$

where the special integral sign indicates integration over a single cycle and subscripts  $i$  and  $f$  designate, respectively, initial and final states. As a consequence, the integration of Equation (1.2) for a cycle yields:

$$\oint dq = \oint dw \quad [\text{Btu/lb}_m \mid \text{kJ/kg}] \quad (1.3)$$

This states that the integral of all transfers of heat into the system, taking into account the sign convention, is the integral of all work done by the system. The latter is the *net work* of the system. The integrals in Equation (1.3) may be replaced by summations for a cyclic process that involves a finite number of heat and work terms. Because many heat engines operate in cyclic processes, it is sometimes convenient to evaluate the net work of a cycle using Equation (1.3) with heat additions and losses rather than using work directly.

### Arbitrary Process of a System

Another important form of the First Law of Thermodynamics is the integral of Equation (1.2) for an arbitrary process involving a system:

$$q = u_f - u_i + w \quad [\text{Btu/lb}_m \mid \text{kJ/kg}] \quad (1.4)$$

where  $q$  and  $w$  are, respectively, the net heat transferred and net work for the process, and  $u_f$  and  $u_i$  are the final and initial values of the internal energy. Equation (1.4), like Equation (1.2), shows that a system that is rigid ( $w = 0$ ) and adiabatic ( $q = 0$ ) has an unchanging internal energy. It also shows, like Equation (1.3), that for a cyclic process the heat transferred must equal the work done.

### Reversibility and Irreversibility

If a system undergoes a process in which temperature and pressure gradients are always small, the process may be thought of as a sequence of near-equilibrium states. If each of the states can be restored in reverse sequence, the process is said to be *internally reversible*. If the environmental changes accompanying the process can also be reversed in sequence, the process is called *externally reversible*. Thus, a *reversible process* is

one that is both internally and externally reversible. The reversible process becomes both a standard by which we measure the success of real processes in avoiding losses and a tool that we can use to derive thermodynamic relations that approximate reality.

All real processes fail to satisfy the requirements for reversibility and are therefore irreversible. *Irreversibility* occurs due to temperature, pressure, composition, and velocity gradients caused by heat transfer, solid and fluid friction, chemical reaction, and high rates of work applied to the system. An engineer's job frequently entails efforts to reduce irreversibility in machines and processes.

### Entropy and Enthalpy

*Entropy* and *enthalpy* are thermodynamic properties that, like internal energy, usually appear in the form of differences between initial and final values. The entropy change of a system,  $\Delta s$  [Btu/lb<sub>m</sub>-R | kJ/kg-K], is defined as the integral of the ratio of the system differential heat transfer to the absolute temperature for a reversible thermodynamic path, that is, a path consisting of a sequence of well-defined thermodynamic states. In differential form this is equivalent to:

$$ds = dq_{rev}/T \quad [\text{Btu/lb}_m\text{-R} \mid \text{kJ/kg-K}] \quad (1.5)$$

where the subscript *rev* denotes that the heat transfer must be evaluated along a reversible path made up of a sequence of neighboring thermodynamic states. It is implied that, for such a path, the system may be returned to its condition before the process took place by traversing the states in the reverse order.

An important example of the use of Equation (1.5) considers a thermodynamic cycle composed of reversible processes. The cyclic integral, Equation (1.3), may then be used to show that the net work of the cycle is:

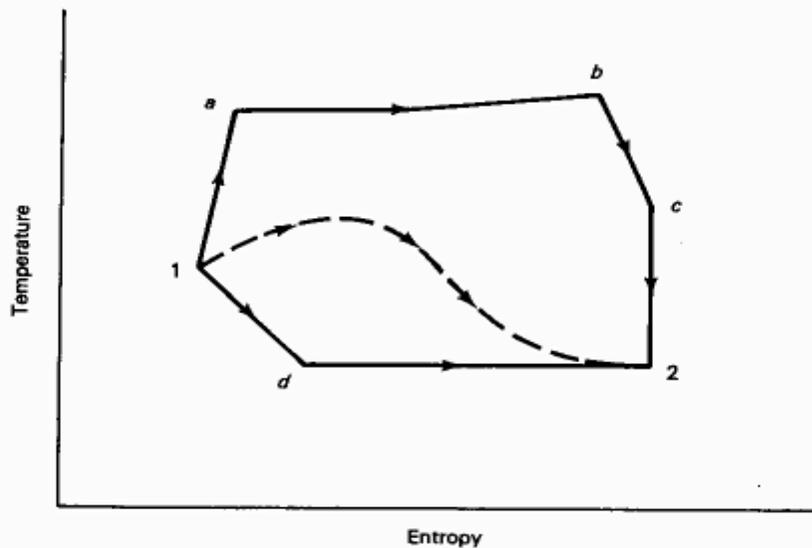
$$w_n = \oint dq = \oint Tds \quad [\text{Btu/lb}_m \mid \text{kJ/kg}]$$

This shows that the area enclosed by a plot of a reversible cyclic process on a temperature-entropy diagram is the net work of the cycle.

The *enthalpy*,  $h$ , is a property of state defined in terms of other properties:

$$h = u + pv \quad [\text{Btu/lb}_m \mid \text{kJ/kg}] \quad (1.6)$$

where  $h$ ,  $u$  and  $v$  are, respectively, the system specific enthalpy, specific internal energy, and specific volume, and  $p$  is the pressure.



**FIGURE 1.2** Two reversible paths and an irreversible path from state 1 to state 2.

Two other important forms of the First Law make use of these properties. Substitution of Equations (1.1) and (1.5) in Equation (1.2) yields, for a reversible process

$$Tds = du + pdv \quad [\text{Btu/lb}_m \mid \text{kJ/kg}] \quad (1.7)$$

and differentiation of Equation (1.6), combined with elimination of  $du$  in Equation (1.7), gives

$$Tds = dh - vdp \quad [\text{Btu/lb}_m \mid \text{kJ/kg}] \quad (1.8)$$

Equations (1.7) and (1.8) may be regarded as relating changes in entropy for reversible processes to changes in internal energy and volume in the former and to changes in enthalpy and pressure in the latter. The fact that all quantities in these equations are properties of state implies that entropy must also be a thermodynamic property.

Because entropy is a state property, the entropy change between two equilibrium states of a system is the same for all processes connecting them, reversible or irreversible. Figure 1.2 depicts several such processes 1-*a*-*b*-*c*-2, 1-*d*-2, and a sequence of nonequilibrium states not describable in thermodynamic terms indicated by the dashed line (an irreversible path). To use Equation (1.5) directly or as in Equations (1.7) and (1.8), a reversible path must be employed. Because of the path independence of state property changes, any reversible path will do. Thus the entropy change,  $s_2 - s_1$ ,

may be evaluated by application of Equations (1.5), (1.7), or (1.8) to either of the reversible paths shown in Figure 1.2 or to any other reversible path connecting states 1 and 2.

### The Second Law of Thermodynamics

While Equation (1.5) may be used to determine the entropy change of a system, the *Second Law of Thermodynamics*, is concerned with the entropy change of the universe, i.e., of both the system and the surroundings. Because entropy is an extensive property, the entropy of a system is the sum of the entropy of its parts. Applying this to the universe, the entropy of the universe is the sum of the entropy of the system and its surroundings. The Second Law may be stated as "The entropy change of the universe is non-negative":

$$\Delta S_{univ} \geq 0 \quad [\text{Btu/R} \mid \text{kJ/K}] \quad (1.9)$$

Note that the entropy change of a system may be negative (entropy decrease) if the entropy change of its environment is positive (entropy increase) and sufficiently large that inequality (1.9) is satisfied.

As an example: if the system is cooled, heat is transferred from the system. The heat flow is therefore negative, according to sign convention. Then, according to Equation (1.5), the system entropy change will also be negative; that is, the system entropy will decrease. The associated heat flow, however, is into the environment, hence positive with respect to the environment (considered as a system). Then Equation (1.5) requires that the environmental entropy change must be positive. The Second Law implies that, for the combined process to be possible, the environmental entropy change must exceed the magnitude of the system entropy change.

The First Law of Thermodynamics deals with how the transfer of heat influences the system internal energy but says nothing about the nature of the heat transfer, i.e., whether the heat is transferred from hotter or colder surroundings. Experience tells us that the environment must be hotter to transfer heat to a cooler object, but the First Law is indifferent to the condition of the heat source. However, calculation of the entropy change for heat transfer from a cold body to a hot body yields a negative universe entropy change, violates the Second Law, and is therefore impossible. Thus the Second Law provides a way to distinguish between real and impossible processes. This is demonstrated in the following example:

#### EXAMPLE 1.1

- (a) Calculate the entropy change of an infinite sink at 27°C temperature due to heat transfer into the sink of 1000 kJ.
- (b) Calculate the entropy change of an infinite source at 127°C losing the same amount of heat.

- (c) What is the entropy change of the universe if the aforementioned source supplies 1000 kJ to the sink with no other exchanges?  
 (d) What are the entropy changes if the direction of heat flow is reversed and the source becomes the sink?

**Solution**

(a) Because the sink temperature is constant, Equation (1.5) shows that the entropy change of the sink is the heat transferred reversibly divided by the absolute temperature of the sink. This reversible process may be visualized as one in which heat is transferred from a source which is infinitesimally hotter than the system:

$$\Delta S_{sink} = 1000/(273 + 27) = + 3.333 \text{ kJ/K.}$$

(b) Treating the source in the same way:

$$\Delta S_{source} = - 1000/(273 + 127) = - 2.5 \text{ kJ/K.}$$

(c) Because the entropy change of the universe is the sum of the entropy changes of source and sink, the two acting together to transfer 1000kJ irreversibly give:

$$\Delta S_{universe} = 3.333 - 2.5 = +0.833 \text{ kJ/K} > 0$$

which satisfies the Second Law inequality (1.9).

(d) A similar approach with the direction of heat flow reversed, taking care to observe the sign convention, gives

$$\Delta S_{sink} = (- 1000)/(273 + 27) = - 3.333 \text{ kJ/K}$$

$$\Delta S_{source} = (1000)/(273 + 127) = + 2.555 \text{ kJ}$$

$$\Delta S_{univ} = - 3.333 + 2.5 = - 0.833 \text{ kJ/K.}$$

Thus we see that heat flow from a low to a high temperature reduces the entropy of the universe, violates the Second Law, and therefore is not possible.

---

Parts a, b, and c of Example 1.1 show that the entropy change of the universe depends on the temperature difference driving the heat transfer process:

$$\Delta S_{univ} = Q(1/T_{sink} - 1/T_{source}) = Q(T_{source} - T_{sink}) / T_{source} T_{sink}$$

Note that if the temperature difference is zero, the universe entropy change is also zero and the heat transfer is reversible. For finite positive temperature differences,  $\Delta S_{univ}$

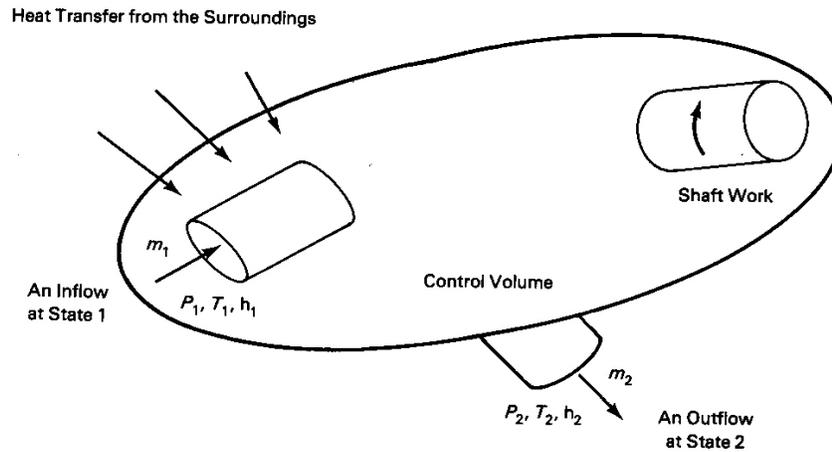


FIGURE 1.3 Control volume with an inlet and an outlet.

exceeds zero and the process is irreversible. As the temperature difference increases,  $\Delta S_{\text{univ}}$  increases. This exemplifies the fact that the entropy change of the universe produced by a process is a measure of the irreversibility of the process.

For an isolated system, there is no change in the entropy of the surroundings. Hence the system entropy change is the entropy change of the universe and therefore must be non-negative. In other words, the entropy of an isolated system can only increase or at best stay constant.

### 1.3 Control Volumes and Steady Flows

In many engineering problems it is preferable to deal with a flow of fluid particles as they pass through a given region of space rather than following the flow of a fixed collection of particles. Thus, putting aside the system concept (fixed collection) for the moment, consider a volume with well-defined spatial boundaries as shown in Figure 1-3. This is called a *control volume*. Mass at state 1 enters at a rate  $m_1$  and leaves at state 2 with mass flow  $m_2$ . If one mass flow rate exceeds the other, mass either accumulates in the volume or is depleted. The important special case of *steady flow*, in which no accumulation or depletion of mass occurs in the control volume, is considered here. In steady flow, the conservation of mass requires equal mass flows in and out, i.e.,  $m_1 = m_2$ , [lb<sub>m</sub>/s, | kg /s].

If  $\dot{Q}$  is the rate of heat flow into the control volume and  $\dot{W}$  is the rate at which shaft work is delivered from the control volume to the surroundings, conservation of energy requires that the excess of inflowing heat over outgoing work equal the net excess of the energy (enthalpy) flowing out of the ports, i.e.,

$$\dot{Q} = \sum_o m_o h_o - \sum_i m_i h_i + \dot{W} = m_2 h_2 - m_1 h_1 + \dot{W} \quad [\text{Btu/s} \mid \text{kJ/s}] \quad (1.10)$$

where summations apply to inflows  $i$  and outflows  $o$ , and where other types of energy terms, such as kinetic and potential energy flows, are assumed negligible. For clarity, the figure shows only one port in and one port out. Kinetic and potential energy terms may be added analogous to the enthalpy terms at each port, if needed.

Equation (1.10) may be the most important and frequently used equation in this book. Mastery of its use is therefore essential. It is known as the *steady flow form of the First Law of Thermodynamics*. It may be thought of as a bookkeeping relation for keeping track of energy crossing the boundaries of the control volume.

The Second Law of Thermodynamics applied to steady flow through an adiabatic control volume requires that  $m_2 s_2 \geq m_1 s_1$ , or by virtue of mass conservation:

$$s_2 \geq s_1 \quad [\text{Btu/lb}_m\text{-R} \mid \text{kJ/kg-K}] \quad (1.11)$$

That is, because entropy cannot accumulate within the control volume in a steady flow, the exit entropy must equal or exceed the inlet entropy. In steady flows, heat transfer can increase or decrease the entropy of the flow, depending on the direction of heat transfer, as long as the entropy change of the surroundings is such that the net effect is to increase the entropy of the universe.

We will often be concerned with adiabatic flows. In the presence of fluid friction and other irreversibilities, the exit entropy of an adiabatic flow exceeds its inlet entropy. Adiabatic flows that have no irreversibilities also have no entropy change and therefore are called *isentropic flows*.

#### 1.4 Properties of Vapors: Mollier and T-s Diagrams

When heated, liquids are transformed into vapors. The much different physical character of liquids and vapors makes engines in which phase change takes place possible. The Newcomen atmospheric engine, for instance condensed steam to liquid water in a piston-cylinder enclosure to create a partial vacuum. The excess of atmospheric pressure over the low pressure of the condensed steam, acting on the opposite face of the piston, provided the actuating force that drove the first successful engines in the early eighteenth century. In the latter half of the eighteenth century, engines in which work was done by steam pressure on the piston rather than by the atmosphere, replaced Newcomen-type engines. Steam under pressure in reciprocating engines was a driving force for the industrial revolution for about two centuries. By the middle of the twentieth century, steam turbines and diesel engines had largely replaced the steam engine in electric power generation, marine propulsion, and railroad locomotives.

Figure 1.4 shows typical saturation curves for a pure substance plotted in temperature and entropy coordinates. A line of constant pressure (an isobar) is shown in which the *subcooled* liquid at state 1 is heated, producing increases in entropy, temperature, and enthalpy, until the liquid is saturated at state 2. Isobars in the

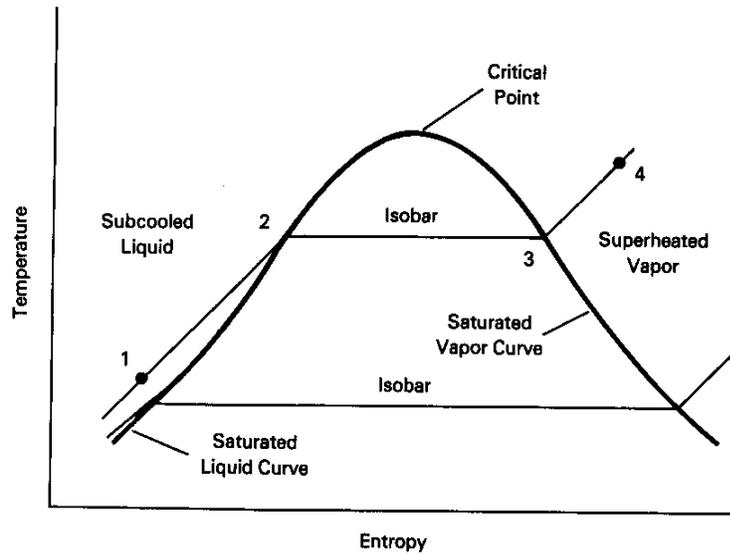


FIGURE 1.4 Liquid, vapor, and mixed phases.

subcooled region of the diagram lie very close to the saturated liquid curve. The separation of the two is exaggerated for clarity.

Once the substance has reached state 2, further transfer of heat fails to increase the system temperature but is reflected in increased enthalpy and entropy in a vaporization or boiling process. During this process the substance is converted from a saturated liquid at state 2 to a mixture of liquid and vapor, and finally to a saturated vapor at state 3. The enthalpy difference between the saturation values,  $h_3 - h_2$ , is called the *enthalpy of vaporization* or *heat of vaporization*.

Continued addition of heat to the system, starting at state 3, *superheats* the steam to state 4, again increasing temperature, enthalpy, and entropy.

Several observations about the isobaric process may be made here. Equation (1.5) and Figure 1.4 show that the effect of adding heat is to always increase system entropy and that of cooling to always decrease it. A similar conclusion can be drawn from Equation (1.10) regarding heat additions acting to increase enthalpy flow through a control volume in the absence of shaft work.

A measure of the proximity of a superheated state (state 4 in the figure) to the saturated vapor line is the *degree of superheat*. This is the difference between the temperature  $T_4$  and the saturated vapor temperature  $T_3$ , at the same pressure. Thus the degree of superheat of superheated state 4 is  $T_4 - T_3$ .

In the phase change from state 2 to state 3, the temperature and pressure give no indication of the relative quantities of liquid and vapor in the system. The *quality*  $x$  is defined as the ratio of the mass of vapor to the mass of the mixture of liquid and vapor at any point between the saturation curves at a given pressure. By virtue of this definition, the quality varies from 0 for a saturated liquid to 1 for a saturated vapor.

Because extensive properties are proportional to mass, they vary directly with the vapor quality in the mixed region. The entropy, for example, varies from the entropy of the saturated liquid  $s_l$  at state 2 to the saturated vapor entropy  $s_v$  at state 3 in accordance with the following quality equation:

$$s = s_l + x(s_v - s_l) \quad [\text{Btu/lb}_m\text{-R} \mid \text{kJ/kg-K}] \quad (1.12)$$

where  $s$  is the entropy per unit mass. Other extensive properties such as enthalpy and volume vary with quality in the same way.

A variable closely related to the quality is moisture fraction (both quality and moisture fraction can be expressed as percentages). Moisture fraction,  $M$ , is defined as the ratio of the mass of liquid to the total mass of liquid and vapor. It can be easily shown that the sum of the quality and the moisture fraction of a mixture is one.

A *Mollier chart*, a diagram with enthalpy as ordinate and entropy as abscissa, is much like the temperature-entropy diagram. A Mollier diagram for steam is included in Appendix B. An isobar on a Mollier chart, unlike that on a T-s diagram, has a continuous slope. It shows both enthalpy and entropy increasing monotonically with heat addition. Such a diagram is frequently used in energy conversion and other areas because of the importance of enthalpy in applying the steady-flow First Law.

### 1.5 Ideal Gas Basics

Under normal ambient conditions, the average distance between molecules in gases is large, resulting in negligible influences of intermolecular forces. In this case, molecular behavior and, therefore, system thermodynamics are governed primarily by molecular translational and rotational kinetic energy. Kinetic theory or statistical thermodynamics may be used to derive the *ideal gas* or *perfect gas* law:

$$pv = RT \quad [\text{ft-lb}_f/\text{lb}_m \mid \text{kJ/kg}] \quad (1.13)$$

where  $p$  [ $\text{lb}_f/\text{ft}^2 \mid \text{kN/m}^2$ ],  $v$  [ $\text{ft}^3/\text{lb}_m \mid \text{m}^3/\text{kg}$ ] and  $T$  [ $^\circ\text{R} \mid ^\circ\text{K}$ ] are pressure, specific volume, and temperature respectively and  $R$  [ $\text{ft-lb}_f/\text{lb}_m\text{-}^\circ\text{R} \mid \text{kJ/kg-}^\circ\text{K}$ ] is the ideal gas constant. The gas constant  $R$  for a specific gas is the *universal gas constant*  $\mathcal{R}$  divided by the molecular weight of the gas.

Thus, the gas constant for air is  $(1545 \text{ ft-lb}_f/\text{lb-mole-}^\circ\text{R}) / (29 \text{ lb}_m/\text{lb-mole}) = 53.3 \text{ ft-lb}_f/\text{lb}_m\text{-}^\circ\text{R}$  in the English system and  $(8.31 \text{ kJ/kg-mole-}^\circ\text{K}) / (29 \text{ kg/kg-mole}) = 0.287 \text{ kJ/kg-}^\circ\text{K}$  in SI units.

The *specific heats* or *heat capacities* at constant volume and at constant pressure, respectively, are:

$$c_v = (\partial u / \partial T)_v \quad [\text{Btu/lb}_m\text{-}^\circ\text{R} \mid \text{kJ/kg-}^\circ\text{K}] \quad (1.14)$$

and

$$c_p = (\partial h / \partial T)_p \quad [\text{Btu/lb}_m\text{-}^\circ\text{R} \mid \text{kJ/kg-}^\circ\text{K}] \quad (1.15)$$

As thermodynamic properties, the heat capacities are, in general, functions of two other thermodynamic properties. For solids and liquids, pressure change has little influence on volume and internal energy, so that to a very good approximation:  $c_v = c_p$ .

A gas is said to be *thermally perfect* if it obeys Equation (1.13) and its internal energy, enthalpy, and heat capacities are functions of temperature only. Then

$$du = c_v(T) dT \quad [\text{Btu/lb}_m \mid \text{kJ/kg}] \quad (1.16)$$

and

$$dh = c_p(T) dT \quad [\text{Btu/lb}_m \mid \text{kJ/kg}] \quad (1.17)$$

A gas is said to be *calorically perfect* if in addition to being thermally perfect it also has constant heat capacities. This is reasonably accurate at low and moderate pressures and at temperatures high enough that intermolecular forces are negligible but low enough that molecular vibrations are not excited and dissociation does not occur. For air, vibrational modes are not significantly excited below about 600K, and dissociation of oxygen does not occur until the temperature is above about 1500K. Nitrogen does not dissociate until still higher temperatures. Excitation of molecular vibrations causes specific heat to increase with temperature increase. Dissociation creates further increases in heat capacities, causing them to become functions of pressure.

It can be shown (see Exercise 1.4) that for a thermally perfect gas the heat capacities are related by the following equation:

$$c_p = c_v + R \quad [\text{Btu/lb}_m\text{-R} \mid \text{kJ/kg-K}] \quad (1.18)$$

This relation does not apply for a dissociating gas, because the molecular weight of the gas changes as molecular bonds are broken. Note the importance of assuring that  $R$  and the heat capacities are in consistent units in this equation.

Another important gas property is the *ratio of heat capacities* defined by  $k = c_p/c_v$ . It is constant for gases at room temperatures but decreases as vibrational modes become excited. The importance of  $k$  will be seen in the following example.

### EXAMPLE 1.2

- Derive an expression for the entropy change of a system in terms of pressure and temperature for a calorically perfect gas.
- Derive a relation between  $p$  and  $T$  for an isentropic process in a calorically perfect gas.

### Solution

(a) For a reversible process, Equation (1.8) gives  $Tds = dh - vdp$ . Dividing by  $T$  and applying the perfect gas law gives  $ds = c_p dT/T - Rdp/p$ .

Then integration between states 1 and 2 yields

$$s_2 - s_1 = c_p \ln(T_2/T_1) - R \ln(p_2/p_1)$$

(b) For an isentropic process,  $s_2 = s_1$ . Then the above equation gives

$$T_2/T_1 = (p_2/p_1)^{(R/c_p)}$$

But  $R/c_p = (c_p - c_v)/c_p = (k - 1)/k$ . Hence  $T_2/T_1 = (p_2/p_1)^{(k-1)/k}$ .

---

This and other important relations for an isentropic process in a calorically perfect gas are summarized as follows

$$T_2/T_1 = (p_2/p_1)^{(k-1)/k} \quad [\text{dl}] \quad (1.19)$$

$$T_2/T_1 = (v_2/v_1)^{(k-1)} \quad [\text{dl}] \quad (1.20)$$

$$p_2/p_1 = (v_1/v_2)^k \quad [\text{dl}] \quad (1.21)$$

These relations show that the ratio of heat capacities governs the variation of thermodynamic properties in an isentropic process. For this reason the ratio of heat capacities is sometimes called the *isentropic exponent*.

## 1.6 Fundamentals of Fluid Flow

Almost all energy conversion devices involve the flow of some form of fluid. Air, liquid water, steam, and combustion gases are commonly found in some of these devices. Here we review a few of the frequently used elementary principles of fluid flow.

The volume flow rate,  $Q$  [ $\text{ft}^3/\text{s}$  |  $\text{m}^3/\text{s}$ ] at which a fluid flows across a surface is the product of the area,  $A$  [ $\text{ft}^2$  |  $\text{m}^2$ ], of the surface and the component of velocity normal to the area,  $V$  [ $\text{ft}/\text{s}$  |  $\text{m}/\text{s}$ ]. The corresponding *mass flow rate* is the ratio of the volume rate and the specific volume,  $v$  [ $\text{ft}^3/\text{lb}_m$  |  $\text{m}^3/\text{kg}$ ]:

$$m = AV/v = Q/v \quad [\text{lb}_m/\text{s} \text{ | } \text{kg}/\text{s}] \quad (1.22)$$

Alternatively the flow rate can be expressed in terms of the reciprocal of the specific volume, the density,  $\rho$  [ $\text{lb}_m/\text{ft}^3$  |  $\text{kg}/\text{m}^3$ ]:

$$m = AV\rho = Q\rho \quad [\text{lb}_m/\text{s} \text{ | } \text{kg}/\text{s}] \quad (1.23)$$

The first important principle of fluid mechanics is the *conservation of mass*, a principle that we have already used in Section 1.3. For a steady flow, the net inflow to a control volume must equal the net outflow. Any imbalance between the inflow and outflow implies an accumulation or a reduction of mass within the control volume, i.e., an unsteady flow. Given a control volume with  $n$  ports, the conservation of mass provides an equation that may be used to solve for the  $n$ th port flow rate, given the other  $n-1$  flow rates. These flows may be (1) given, (2) calculated from data at the ports using Equation (1.22) or (1.23), (3) obtained by solving  $n-1$  other equations, or (4) a combination of the preceding three.

For isentropic flow of an *incompressible* (constant density,  $\rho$ ) fluid, the *Bernoulli equation* applies:

$$p_1/\rho + V_1^2/2 = p_2/\rho + V_2^2/2 \quad [\text{ft}\cdot\text{lb}_f/\text{lb}_m \mid \text{kJ/kg}] \quad (1.24)$$

This is an invariant form, i.e. an equation with the same terms on both sides,  $p/\rho + V^2/2$ . The subscripts identify the locations in the flow where the invariants are evaluated. The first term of the invariant is sometimes called the *pressure head*, and the second the *velocity head*. The equation applies only in regions where there are no irreversibilities such as viscous losses or heat transfer.

The invariant sum of the two terms on either side of Equation (1.24) may be called the *total head* or *stagnation head*. It is the head that would be observed at a point where the velocity approaches zero. The pressure associated with the total head is therefore called the *total pressure* or *stagnation pressure*,  $p_o = p + \rho V^2/2$ . Each point in the flow may be thought of as having its own stagnation pressure resulting from an imaginary isentropic deceleration.

In the event of significant irreversibilities, there is a loss in total head and the Bernoulli equation may be generalized to:

$$p_1/\rho + V_1^2/2 = p_2/\rho + V_2^2/2 + \text{loss} \quad [\text{ft}\cdot\text{lb}_f/\text{lb}_m \mid \text{kJ/kg}] \quad (1.25a)$$

or

$$p_{o1}/\rho = p_{o2}/\rho + \text{loss} \quad [\text{ft}\cdot\text{lb}_f/\text{lb}_m \mid \text{kJ/kg}] \quad (1.25b)$$

Stagnation pressure or head losses in ducts, such as due to flow turning or sudden area change, are tabulated in reference books as fractions of the upstream velocity head for a variety of geometries. Another example is the famous Darcy-Weisbach equation which gives the head loss resulting from fluid friction in a pipe of constant cross-section.

## 1.7 Compressible Flow

While many engineering analyses may reasonably employ incompressible flow principles, there are cases where the compressibility of gases and vapors must be considered. These are situations where the magnitude of the kinetic energy of the flow

is comparable to its enthalpy such as in supersonic nozzles and diffusers, in turbines and compressors, and in supersonic flight. In these cases the steady-flow First Law must be generalized to include kinetic energy per unit mass terms. For two ports:

$$\dot{Q} = m_2(h_2 + V_2^2 / 2) - m_1(h_1 + V_1^2 / 2) + \dot{W} \quad [\text{Btu/s} \mid \text{kJ/s}] \quad (1.26a)$$

Care should be taken to assure consistency of units, because enthalpy is usually stated in thermal units [Btu /lb<sub>m</sub> | kJ/kg] and velocity in mechanical units [ft /s | m /s].

Another invariant of significance appears in Equation (1.26a). The form

$$h_o = h + V^2/2 \quad [\text{Btu/lb}_m \mid \text{kJ/kg}] \quad (1.27)$$

is seen to be invariant in applications where heat transfer and shaft work are insignificant. The invariant,  $h_o$ , is usually given the name *stagnation enthalpy* because it is the enthalpy at a point in the flow (real or imagined) where velocity approaches zero. In terms of stagnation enthalpy, Equation (1.26a) may be rewritten as

$$\dot{Q} = m(h_{o2} - h_{o1}) + \dot{W} \quad [\text{Btu/s} \mid \text{kJ/s}] \quad (1.26b)$$

where conservation of mass with steady flow through two ports has been assumed.

Writing  $dh_o = c_p dT_o$  with  $c_p$  constant, we get

$$h_{o2} - h_{o1} = c_p(T_{o2} - T_{o1})$$

Combining this with Equation (1.27), we are led to define another invariant, the *stagnation temperature* for a calorically perfect gas:

$$T_o = T + V^2/2c_p \quad [\text{R} \mid \text{K}] \quad (1.28)$$

The stagnation temperature may be regarded as the temperature at a real or imaginary point where the gas velocity has been brought to zero adiabatically. For this special case of a constant heat capacity, Equation (1.26b) may be written as

$$\dot{Q} = mc_p(T_{o2} - T_{o1}) + \dot{W} \quad [\text{Btu/s} \mid \text{kJ/s}] \quad (1.26c)$$

In both incompressible and compressible flows, the mass flow rates at all stations in a streamtube are the same. Because the specific volume and density are constant in incompressible flow, Equation (1.22) shows that the volume flow rates are the same at all stations also. However for compressible flow, Equation (1.23) shows that density change along a streamtube implies volume flow rate variation. Thus, while it is frequently convenient to think and talk in terms of volume flow rate when dealing with incompressible flows, mass flow rate is more meaningful in compressible flows and in general.

A measure of the compressibility of a flow is often indicated by a Mach number, defined as the dimensionless ratio of a flow velocity to the local speed of sound in the fluid. For ideal gases the speed of sound is given by

$$a = (kp/\rho)^{1/2} = (kRT)^{1/2} \quad [\text{ft /s} \mid \text{m /s}] \quad (1.29)$$

Compressible flows are frequently classified according to their Mach number:

$M = 0$	Incompressible
$0 < M < 1$	Subsonic
$M = 1$	Sonic
$M > 1$	Supersonic

Studies of compressible flows show that supersonic flows have a significantly different physical character than subsonic and incompressible flows. For example, the velocity fields in subsonic flows are continuous, whereas discontinuities known as shock waves are common in supersonic flows. Thus the student should not be surprised to find that different relations hold in supersonic flows than in subsonic flows.

## 1.8 Energy Classification

Energy exists in a variety of forms. All human activities involve conversion of energy from one form to another. Indeed, life itself depends on energy conversion processes. The human body, through complex processes, transforms the chemical energy stored in food into external motion and work produced by muscles as well as electrical impulses that control and activate internal functions.

It is instructive to examine some of the processes for transformation between technically important forms of energy. Table 1.2 shows a matrix of energy forms and the names of some associated energy converters.

Table 1.2 Energy Transformation Matrix

From:	To:	Thermal Energy	Mechanical Energy	Electrical Energy
Chemical Energy	Furnace	Diesel engine	Fuel cell	
Thermal Energy	Heat exchanger	Steam turbine	Thermocouple	
Mechanical Energy	Refrigerator, heat pump	Gearbox	Electrical generator	
Nuclear Energy	Fission reactor	Nuclear steam turbine	Nuclear power plant	

The table is far from complete, and other energy forms and energy converters could readily be added. However, it does include the major energy converters of interest to mechanical engineers. It is a goal of this book to present important aspects of the design, analysis, performance, and operation of most of these devices.

One of the major criteria guiding the design of energy conversion systems is

efficiency. Each of the conversions in Table 1.2 is executed by a device that operates with one or more relevant efficiencies. The following section explores some of the variety of definitions of efficiency used in design and performance studies of energy conversion devices.

## 1.9 Efficiencies

*Efficiency* is a measure of the quality of an operation or of a characteristic of a device. Several types of efficiencies are widely used. It is important to clearly distinguish among them. Note that the terms *work* and *power* are equally applicable here.

The efficiency of a machine that transmits mechanical power is measured by its *mechanical efficiency*, the fraction of the power supplied to the transmission device that is delivered to another machine attached to its output, Figure 1.5(a). Thus a gearbox for converting rotational motion from a power source to a device driven at another speed dissipates some mechanical energy by fluid and/or dry friction, with a consequent loss in power transmitted to the second machine. The efficiency of the gearbox is the ratio of its power output to the power input, a value less than one. For example, a turboprop engine with a gearbox efficiency of 0.95 will transmit only 95% of its power output to its propeller.

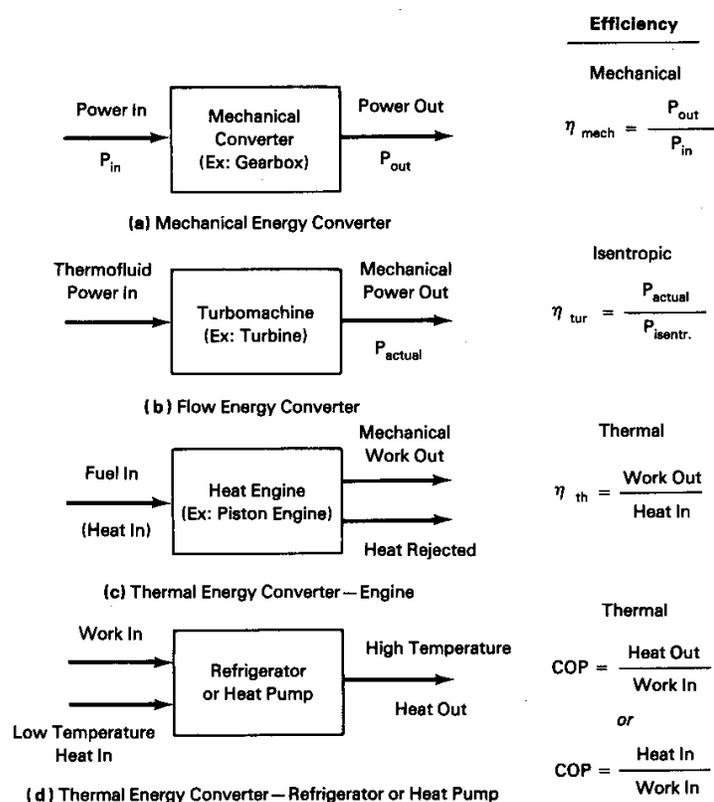


FIGURE 1.5 Types of efficiencies.

Another type of efficiency that measures internal losses of power is used to indicate the quality of performance of turbomachines such as pumps, compressors and turbines. These devices convert flow energy to work (power), or vice versa. Here the efficiency compares the output with a theoretical ideal in a ratio [Figure 1.5(b)]. The resulting efficiency ranges from 0 to 1 as a measure of how closely the process approaches a relevant isentropic process. A turbine with an efficiency of 0.9, will, for example, deliver 90% of the power of a perfect (isentropic) turbine operating under the same conditions. This efficiency, sometimes referred to as *isentropic efficiency* or *turbine efficiency*, will be considered in more detail in the next chapter.

Another form of isentropic efficiency, sometimes called *compressor efficiency* (or pump efficiency), is defined for compressors (or pumps). It is the ratio of the isentropic work to drive the compressor (or pump) to the actual work required. Because the actual work required exceeds the isentropic work, this efficiency is also less than to 1.

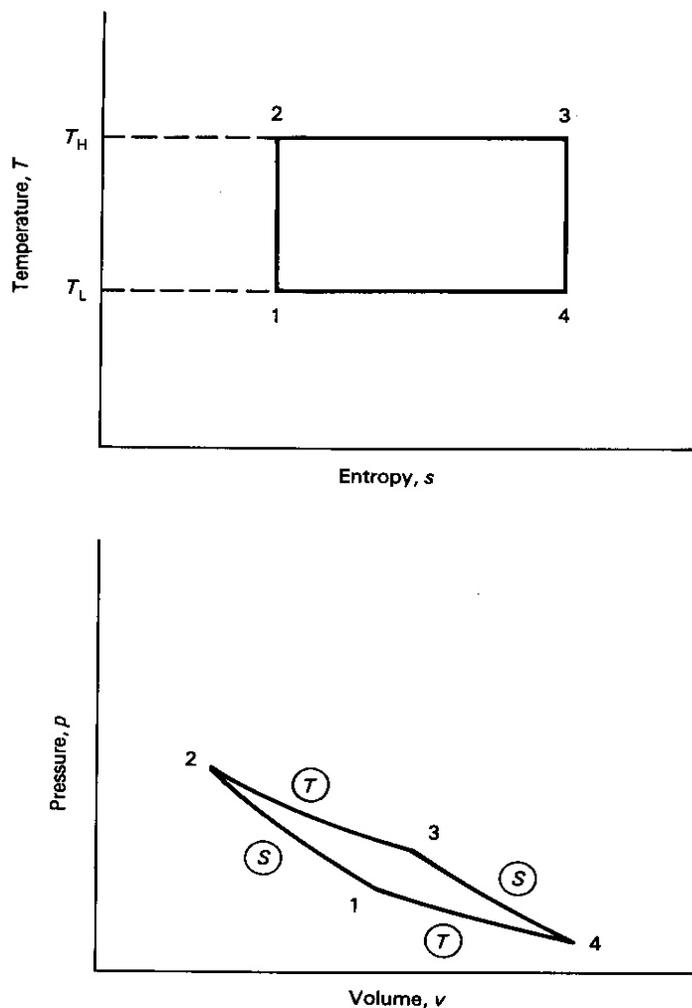
A third type of efficiency compares the magnitude of a useful effect to the cost of producing the effect, measured in comparable units. An example of this type of efficiency compares the net work output,  $w_n$ , of a heat engine to the heat supplied,  $q_a$ , to operate the engine. This is called the  $\eta_{th} = w_n/q_a$  [Figure 1.5(c)]. For example, the flow of natural gas to an electrical power plant provides a chemical energy flow rate or heat flow rate to the plant that leads to useful electric power output. It is known from basic thermodynamics that this efficiency is limited by the Carnot efficiency, as will be discussed in the next section.

Another example of this type of efficiency as applied to refrigerators and heat pumps [Figure 1.5(d)] is called the *coefficient of performance*, COP. In this case the useful effect is the rate of cooling or heating, and the cost to produce the effect is the power supplied to the device. The term "coefficient of performance" is used instead of efficiency for this measure of quality because the useful effect usually exceeds the cost in comparable units of measure. Hence, unlike other efficiencies, the COP can exceed unity. As seen in Figure 1.5(d), there are two definitions for COP, one for a refrigerator and another for a heat pump. It may be shown using the First Law of Thermodynamics, that a simple relationship exists between the two definitions:  $COP_{hp} = COP_{refr} + 1$ .

### 1.10 The Carnot Engine

On beginning the study of the energy conversion ideas and devices that will serve us in the twenty-first century, it is appropriate to review the theoretical cycle that stands as the ideal for a heat engine. The ideas put forth by Sadi Carnot in 1824 in his "Reflections on the Motive Power of Heat" (see Historical Bibliography) expressed the content of the Second Law of Thermodynamics relevant to heat engines, which, in modern form (attributed to Kelvin and Planck) is: "It is impossible for a device which operates in a cycle to receive heat from a single source and convert the heat *completely* to work." Carnot's great work also described the cycle that today bears his name and provides the theoretical limit for efficiency of heat engine cycles that operate between two given temperature levels: the *Carnot cycle*.

The *Carnot cycle* consists of two reversible, isothermal processes separated by two



**FIGURE 1.6** Carnot cycle, in T-s and p-v coordinates.

reversible adiabatic or isentropic processes, as shown in Figure 1.6. All of the heat transferred to the working fluid is supplied isothermally at the high temperature  $T_H = T_3$ , and all heat rejected is transferred from the working medium at the low temperature  $T_L = T_1$ . No heat transfer takes place, of course, in the isentropic processes. It is evident from Equation (1.5) and the T-s diagram that the heat added is  $T_3(s_3 - s_2)$ , the heat rejected is  $T_1(s_1 - s_4)$ , and, by the cyclic integral relation, the net work is  $T_3(s_3 - s_2) + T_1(s_1 - s_4)$ . The thermal efficiency of the Carnot cycle, like that of other cycles, is given by  $w_n / q_a$  and can be expressed in terms of the high and low cycle temperatures as :

$$\begin{aligned}
 \eta_{Carnot} &= \frac{T_3(s_3 - s_2) + T_1(s_1 - s_4)}{T_3(s_3 - s_2)} \\
 &= 1 + \frac{T_1(s_1 - s_4)}{T_3(s_3 - s_2)} \\
 &= 1 - T_1/T_3 = 1 - T_L/T_H
 \end{aligned}$$

because both isothermal processes operate between the same entropy limits.

The Carnot efficiency equation shows that efficiency rises as  $T_L$  drops and as  $T_H$  increases. The message is clear: a heat engine should operate between the widest possible temperature limits. Thus the efficiency of a heat engine will be limited by the maximum attainable energy-source temperature and the lowest available heat-sink temperature.

Students are sometimes troubled by the idea of isothermal heat transfer processes because they associate heat transfer with temperature rise. A moments reflection, however, on the existence of latent heats—e.g., the teakettle steaming on the stove at constant temperature—makes it clear that one should not always associate heat transfer with temperature change.

It is important here, as we start to consider energy conversion devices, to recall the famous *Carnot Theorem*, the proof of which is given the most thermodynamics texts. It states that it is impossible for any engine operating in a cycle between two reservoirs at different temperatures to have an efficiency that exceeds the Carnot efficiency corresponding to those temperatures. It can also be shown that all reversible engines operating between two given reservoirs have the same efficiency and that all irreversible engines must have lower efficiencies. Thus the Carnot efficiency sets an upper limit on the performance of heat engines and therefore serves as a criterion by which other engines may be judged.

### 1.11 Additional Second-Law Considerations

The qualitative relationship between the irreversibility of a process and the entropy increase of the universe associated with it was considered in section 1.2. Let us now consider a quantitative approach to irreversibility and apply it to a model of a power plant.

#### Reversible Work

Instead of comparing the work output of the power plant with the energy supplied from fuel to run the plant, it is instructive to compare it with the maximum work achievable by a reversible heat engine operating between the appropriate temperature limits, the *reversible work*. It has been established that any reversible engine would have the same efficiency as a Carnot engine. The Carnot engine provides a device for determining the reversible work associated with a given source temperature,  $T_H$ , and a lower sink temperature,  $T_L$ .

## Irreversibility

The *irreversibility*,  $I$ , of a process is defined as the difference between the reversible work and the actual work of a process:

$$I = W_{rev} - W_{act} \quad [\text{Btu} \mid \text{kJ}]$$

It is seen that the irreversibility of a process vanishes when the actual work is the same as that produced by an appropriate Carnot engine. Moreover, the irreversibility of a non-work-producing engine is equal to the reversible work. It is clear that  $I \geq 0$ , because no real engine can produce more work than a Carnot engine operating between the same limiting temperatures.

## Second-Law Efficiency

We can also define a “second-law efficiency,”  $\eta_{II}$ , as the ratio of the actual work of a process to the reversible work:

$$\eta_{II} = \frac{W_{act}}{W_{rev}} \quad [dl]$$

This is an efficiency that is limited to 100%, as opposed to the thermal efficiency of a heat engine, sometimes referred to as a “first-law efficiency,” which may not exceed that of the appropriate Carnot engine. Note that an engine which has no irreversibility is a reversible engine and has a second-law efficiency of 100%.

## A Power Plant Model

Let us consider a model of a power plant in which a fuel is burned at a high temperature,  $T_H$ , in order to transfer heat to a working fluid at an intermediate temperature,  $T_{INT}$ . The working fluid, in turn, is used in an engine to produce work and reject heat to a sink at the low temperature,  $T_L$ . Figure 1.7 presents a diagram of the model that shows explicitly the combustion temperature drop from the source temperature,  $T_H$ , to the intermediate temperature,  $T_{INT}$ , the actual work-producing engine, and a Carnot engine used to determine the reversible work for the situation. The Carnot engine has an efficiency of  $\eta_C = 1 - T_L/T_H$  and develops work in the following amount:

$$W_{rev} = W_C = Q_{IN} - Q_C = \eta_C Q_{IN} \quad [\text{Btu} \mid \text{kJ}]$$

Suppose that the engine we are considering is a Carnot engine that operates from the intermediate source at  $T_{INT}$ . Its efficiency and work output are, respectively,  $\eta_I = 1 - T_L/T_{INT}$  and  $W_{act} = Q_{IN} - Q_I = \eta_I Q_{IN}$ . The irreversibility,  $I$ , of the power plant is then

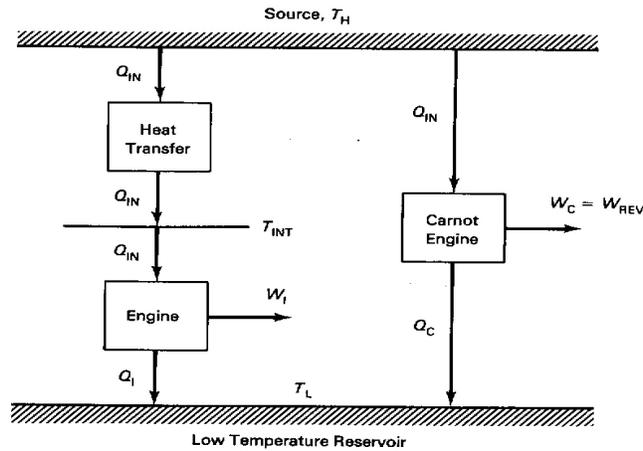


FIGURE 1.7 An engine model and a Carnot engine to determine the reversible work for the model.

$$\begin{aligned}
 I &= W_{rev} - W_{act} = (\eta_C - \eta_I)Q_{IN} \\
 &= [1 - T_L/T_H - (1 - T_L/T_{INT})]Q_{IN} \\
 &= T_L Q_{IN} (T_H - T_{INT}) / T_{INT} T_H \quad [\text{Btu|kJ}]
 \end{aligned}$$

and the second-law efficiency is:

$$\eta_{II} = W_{act} / W_{rev} = (1 - T_L/T_{INT}) / (1 - T_L/T_H) \quad [dl]$$

Note that when  $T_{INT} = T_H$ , the irreversibility vanishes and the second-law efficiency becomes 100%. Also, when  $T_{INT} = T_L$ , the irreversibility is equal to the reversible work of the Carnot engine and the second-law efficiency is zero. The latter condition indicates that a pure heat transfer process or any process that produces no useful work causes a loss in the ability to do work in the amount of  $W_{rev}$ . Thus the reversible work associated with the extremes of a given process is a measure of how much capability to do work *can be* lost, and the irreversibility is a measure of how much of that work-producing potential *is actually* lost. The following example illustrates these ideas.

### EXAMPLE 1.3

Through combustion of a fossil fuel at 3500°R, an engine receives energy at a rate of 3000 Btu/s to heat steam to 1500°R. There is no energy loss in the combustion process. The steam, in turn, produces 1000 Btu/s of work and rejects the remaining energy to the surroundings at 500°R.

- What is the thermal efficiency of the plant?
- What are the reversible work and the Carnot efficiency corresponding to the source?

and sink temperatures?

(c) What is the irreversibility?

(d) What is the second-law efficiency?

(e) What would the irreversibility and the second-law efficiency be if the working fluid were processed by a Carnot engine rather than by the real engine?

### Solution

(a) The thermal efficiency, or first-law efficiency, of the plant is  $\eta_I = W_{\text{act}}/Q_{\text{IN}} = 1000/3000 = 0.333$  or 33.3%.

(b) The relevant Carnot efficiency is  $1 - 500/3500 = 0.857$ , or 85.7%. The engine's reversible work is then  $\eta_C Q_{\text{IN}} = 0.857(3000) = 2571$  Btu/s.

(c) The plant irreversibility is  $2571 - 1000 = 1571$  Btu/s.

(d) The second-law efficiency is then

$$\eta_I / \eta_C = 0.333/0.857 = 0.389, \text{ or } 38.9\%$$

or

$$W_{\text{act}} / W_{\text{rev}} = 1000/2571 = 0.389, \text{ or } 38.9\%.$$

(e) The Carnot efficiency corresponding to the maximum temperature of the working fluid is  $1 - 500/1500 = 0.667$  or 66.7%. The second-law efficiency for this system is then  $0.667/0.857 = 0.778$ , compared with 0.389 for the actual engine. The actual work produced by the irreversibly heated Carnot engine is  $0.667(3000) = 2001$  Btu/s, and its irreversibility is then  $I = 2571 - 2001 = 570$  Btu/s.

In summary, the thermal efficiency, or first-law efficiency, of an engine is a measure of how well the engine converts the energy in its fuel to useful work. It says nothing about energy loss, because energy is conserved and cannot be lost: it can only be transformed. The second-law efficiency, on the other hand, recognizes that some of the energy of a fuel is not available for conversion to work in a heat engine and therefore assesses the ability of the engine to convert only the available work into useful work. This is a reason why some regard the second-law efficiency as more significant than the more commonly used first-law efficiency.

### Bibliography and References

1. Van Wylen, Gordon J., and Sonntag, Richard E., *Fundamentals of Classical Thermodynamics*, 3<sup>rd</sup> ed. New York: Wiley, 1986.

2. Balmer, Robert, *Thermodynamics*. Minneapolis: West, 1990.

3. Cengel, Yunus A., and Boles, Michael A., *Thermodynamics*. New York: McGraw-Hill, 1989.

4. Faires, Virgil Moring, *Thermodynamics*, 5<sup>th</sup> ed. New York: Macmillan, 1970.
5. Silver, Howard F., and Nydahl, John E., *Engineering Thermodynamics*. Minneapolis: West, 1977.
6. Bathie, William W., *Fundamentals of Gas Turbines*. New York: Wiley, 1984
7. Wilson, David Gordon, *The Design of High Efficiency Turbomachinery and Gas Turbines*. Boston: MIT Press, 1984.
8. Anderson, John D., *Modern Compressible Flow*. New York: McGraw-Hill, 1982.
9. Anderson, John D., *Introduction to Flight*. New York: McGraw-Hill, 1978.
10. Chapman, Alan J. and Walker, William F., *Introductory Gas Dynamics*. New York: Holt, Rinehart, and Winston, 1971.

### **Historical Bibliography**

1. Barnard, William N., Ellenwood, Frank E., and Hirshfeld, Clarence F., *Heat-Power Engineering*. Wiley, 1926.
2. Bent, Henry, *The Second Law*. New York: Oxford University Press, 1965
3. Cummins, C. Lyle, Jr., *Internal Fire*, rev. ed. Warrendale, Penna.: Society of Automotive Engineers, 1989.
4. Tann, Jennifer, *The Selected Papers of Boulton and Watt*. Boston: MIT Press, 1981.
5. Carnot, Sadi, *Reflexions Sur la Puissance Motrice de Feu*. Paris: Bachelor, 1824.
6. Potter, J. H., "The Gas Turbine Cycle." ASME Paper presented at the Gas Turbine Forum Dinner, ASME Annual Meeting, New York, Nov. 27, 1972.
7. Grosser, Morton, *Diesel: The Man and the Machine*. New York: Atheneum, 1978.
8. Nitske, W. Robert, and Wilson, Charles Morrow, *Rudolph Diesel: Pioneer of the Age of Power*. Norman, Okla: University of Oklahoma Press, 1965.
9. Rolt, L.T. C., and Allen, J. S., *The Steam Engine of Thomas Newcomen*. New York: Moorland Publishing Co., 1977.
10. Briggs, Asa, *The Power of Steam*. Chicago: University of Chicago Press, 1982.

11. Thurston, Robert H., *A History of the Growth of the Steam Engine*. 1878; rpr. Ithaca, N.Y.: Cornell University Press, 1939.

### EXERCISES

- 1.1 Determine the entropy of steam at 1000 psia and a quality of 50%.
- 1.2 Show that the *moisture fraction* for a liquid water-steam mixture, defined as the ratio of liquid mass to mixture mass, can be written as  $1 - x$ , where  $x$  is steam quality.
- 1.3 Write expressions for the specific entropy, specific enthalpy, and specific volume as functions of the moisture fraction. Determine the values of these properties for steam at 500°F and a moisture fraction of 0.4.
- 1.4 Show that for a thermally perfect gas,  $c_p - c_v = R$ .
- 1.5 During a cyclic process, 75kJ of heat flow into a system and 25kJ are rejected from the system later in the cycle. What is the net work of the cycle.
- 1.6 Seventy-five kJ of heat flow into a rigid system and 25 kJ are rejected later. What are the magnitude and sign of the change in internal energy? What does the sign indicate?
- 1.7 The mass contained between an insulated piston and an insulated cylinder decreases in internal energy by 50 Btu. How much work is involved, and what is the sign of the work term? What does the sign indicate?
- 1.8 Derive Equation (1.8) from Equation (1.7).
- 1.9 Use Equation (1.8) to derive an expression for the finite enthalpy change of an incompressible fluid in an isentropic process. If the process is the pressurization of saturated water initially at 250 psia, what is the enthalpy rise, in Btu /lb<sub>m</sub> and in ft-lb<sub>f</sub>/ lb<sub>m</sub>, if the final pressure is 4000 psia? What is the enthalpy rise if the initial pressure is 100 kPa and the final pressure is 950 kPa?
- 1.10 Sixty kg /s of brine flows into a device with an enthalpy of 200 kJ/kg. Brine flows out of the other port at a flow rate of 20 kg /s. What is the net inflow? Is the system in steady flow? Explain.
- 1.11 Use the steam tables in Appendices B and C to compare the heats of vaporization at 0.01, 10, and 1000 psia. Compare the saturated liquid specific volumes at these pressures. What do you conclude about the influence of pressure on these properties?

1.12 Using the heat capacity equation for nitrogen:

$$c_p = 9.47 - 3.47 \times 10^3/T + 1.16 \times 10^6/T^2$$

where  $c_p$  is in Btu / lb-mole and  $T$  is in degrees Rankine (From Gordon J. Van Wylen and Richard E. Sonntag, *Fundamentals of Classical Thermodynamics*, 3rd ed. New York: Wiley, 1986). Compare the enthalpy change per mole of nitrogen between 540°R and 2000°R for nitrogen as a thermally perfect gas and as a calorically perfect gas.

1.13 Use Equation (1.7) to derive Equation (1.20) for a calorically perfect gas.

1.14 Use Equation (1.7) to derive a relation for the entropy change as a function of temperature ratio for a constant-volume process in a calorically perfect gas.

1.15 Use Equation (1.8) to derive a relation for the entropy change as a function of temperature for an isobaric process in a calorically perfect gas.

1.16 A *convergent nozzle* is a flow passage in which area decreases in the streamwise direction (the direction of the flow). It is used for accelerating the flow from a low velocity to a higher velocity. Use the generalized form of the steady-flow First Law given in Equation (1.26a) to derive an equation for the exit velocity for an adiabatic nozzle.

1.17 Derive an equation for the pressure drop for a loss-free incompressible flow in a varying-area duct as a function of area ratio.

1.18 Two units of work are required to transfer 10 units of heat from a refrigerator to the environment. What is the COP of the refrigerator? Suppose that the same amount of heat transfer instead is by a heat pump into a house. What is the heat pump COP?

1.19 A power plant delivers 100 units of work at 30% thermal efficiency. How many heat units are supplied to operate the plant? How many units of heat are rejected to the surroundings?

1.20 A steam turbine has an efficiency of 90% and a theoretical isentropic power of 100 kW. What is the actual power output?

1.21 Thomas Newcomen used the fact that the specific volume of saturated liquid is much smaller than the specific volume of saturated steam at the same pressure in his famous "atmospheric engine." Calculate the work done on the piston by the atmosphere if steam is condensed at an average pressure of 6 psia by cooling in a tightly fitted piston-cylinder enclosure if the piston area is 1 ft<sup>2</sup> and the piston stroke is 1 ft. If the process takes place 10 times a minute, what is the power delivered? Discuss what can

be done to increase the power of the engine. Describe the characteristics of a Newcomen engine that would theoretically deliver 20 horsepower.

1.22 Expand Table 1.2 to include solar and geothermal sources.

1.23 Twenty pounds of compressed air is stored in a tank at 200 psia and 80°F. The tank is heated to bring the temperature to 155°F. What is the final tank pressure, and how much heat was added?

1.24 Ten kilograms of compressed air is stored in a tank at 250 kPa and 50°C. The tank is heated to bring the air temperature to 200 °C. What is the final tank pressure, and how much heat was added?

1.25 An 85-ft<sup>3</sup> tank contains air at 30 psia and 100° F. What mass of air must be added to bring the pressure to 50 psia and the temperature to 150°F?

1.26 A 20-L tank contains air at 2 bar and 300K. What mass of air must be added to bring the pressure to 2 bar and the temperature to 375K?

1.27 Air enters a wind tunnel nozzle at 160°F, 10 atm, and a velocity of 50 ft/s. The entrance area is 5ft<sup>2</sup>. If the heat loss per unit mass is 10 Btu/lb<sub>m</sub> and the exit pressure and velocity are, respectively, 1.5 atm and 675 ft/s, what are the exit temperature and area?

1.28 Air enters a wind tunnel nozzle at 90°C, 250 kPa, and a velocity of 40 m/s. The entrance area is 3 m<sup>2</sup>. If the heat loss per unit mass is 7 kJ/kg and the exit pressure and velocity are, respectively, 105 kPa and 250 m/s, what are the exit temperature and area?

1.29 Air enters a wind tunnel nozzle at 160°F, 10 atm, and a velocity of 50 ft/s. The entrance area is 5ft<sup>2</sup>. If the heat loss per unit mass is 8 Btu/lb<sub>m</sub> and the exit pressure and temperature are, respectively, 1.25 atm and 120°F, what are the exit velocity and area?

1.30 Air enters a wind tunnel nozzle at 90°C, 250 kPa, and a velocity of 40 m/s. The entrance area is 3 m<sup>2</sup>. If the heat loss per unit mass is 5 kJ/kg and the exit pressure and temperature are, respectively, 120 kPa and 43°C, what are the exit velocity and area?

1.31 Sketch a Mollier diagram showing the character of three isotherms and three isobars for a calorically perfect gas. Label each curve with a value in SI units to show the directions of increasing temperature and pressure. Explain how the diagram would differ if the gas were not calorically perfect.