LECTURE NOTES

ON

THERMODYNAMICS

II B. Tech I semester

Prepared By:

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THERMODYNAMICS

III Semester: ME								
Course Code	**Category**	**Hours / Week**			**Credits**	**Maximum Marks**		
AMEB04	**Core**	**L**	**T**	**P**	**C**	**CIA**	**SEE**	**Total**
		3	1	0	4	30	70	100
Contact Classes: 45	**Tutorial Classes: 15**	**Practical Classes: Nil**				**Total Classes: 60**		

COURSE OBJECTIVES:

The course should enable the students to:

I. Understand the laws of thermodynamics and determine thermodynamic properties, gas laws
II. U Knowledge of properties during various phases of pure substances, mixtures, usage of steam tables and Mollier chart, psychometric charts.
III. Understand the direction law and concept of increase in entropy of universe.
IV. Understand the working of ideal air standard, vapor cycles and evaluate their performance in open systems like steam power plants, internal combustion engines, gas turbines and refrigeration systems.
V. Solve problems of different types of cycles and their performance which emphasizes knowledge in IC engines and refrigeration cycles.

COURSE OUTCOMES(COs):

CO 1: Describe the basic concepts and first law of thermodynamics.
CO 2: Describe the second law of thermodynamics and understand the concept of entropy and third law of thermodynamics.
CO 3: Understand the Pure Substances various thermodynamic processes
CO 4: Understand the concept of Mixtures of perfect gases and psychometric properties,
CO 5: Develop the concept power cycle with description and representation on P-V and T-S diagram.

COURSE LEARNING OUTCOMES (CLOs):

1. Understand the concepts of conservation of mass, conservation of energy.
2. Demonstrate knowledge of ability to identify & apply fundamentals to solve problems like system properties, amount of work transfer and heat during various processes.
3. Explore knowledge & ability to design the thermal related components in various fields of energy transfer equipment.
4. Derive the first law of Thermodynamics from the concept of conservation of energy.
5. Discuss the nature of steady and unsteady processes under the influence of time.
6. Determine entropy changes in a wide range of processes and determine the reversibility or irreversibility of a process from such calculations based on Carnot Cycle.
7. Develop the second law of thermodynamics from the limitations of first law.
8. Knowledge of the Gibbs and Helmholtz free energies as equilibrium criteria, and the statement of the equilibrium condition for closed and open systems.
9. Discuss pressure-temperature, volume-temperature, pressure-volume phase diagrams and the steam tables in the analysis of engineering devices and systems.
10. Understand the inter relationship between thermodynamic functions and an ability to use such relationships to solve practical problems.
11. Understand the equation of state, specific and universal gas constants, throttling and free expansion processes.
12. Discuss deviations from perfect gas model, Vander Waals equation of state.
13. Understand mole fraction, mass friction, gravimetric and volumetric analysis, volume fraction.
14. Discuss dalton‘s law of partial pressure, Avogadro‘s laws of additive volumes, and partial pressure, equivalent gas constant.
15. Understand enthalpy, specific heats and entropy of mixture of perfect gases.
16. Understand the process of psychrometry that are used in the analysis of engineering devices like air

conditioning systems.

17. Develop Otto, Diesel, Dual combustion cycles, description and representation on P-V and T-S diagram
18. Discuss thermal efficiency; mean effective pressures on air standard basis.
19. Understand the comparison of various cycles.
20. Understand introduction to Brayton cycle and Bell Coleman cycle.

Module -I | BASIC CONCEPTS AND FIRST LAW OF THERMODYNAMICS | Classes: 09

System, control volume, surrounding, boundaries, universe, types of systems, macroscopic and microscopic viewpoints, concept of continuum, thermodynamic equilibrium, state, property, process, cycle, reversibility, quasi static process, irreversible process, causes of irreversibility, various flow and non-flow processes ,energy in state and in transition, types-work and heat, point and path function, Zeroth law of thermodynamics, concept of quality of temperature, Principles of thermometry, reference points, constant volume gas thermometer, ideal gas scale, PMMI Joule‘s experiments, first law of thermodynamics, corollaries first law applied to a process, applied to a flow system, steady flow energy equation.

Module -II | SECOND LAW OF THERMODYNAMICS | Classes: 09

Thermal reservoir, heat engine, heat pump, parameters of performance, second Law of thermodynamics, Kelvin Planck and Clausius statements and their equivalence, Corollaries, PMM of second kind, Carnot‘s principle, Carnot cycle and its specialties, thermodynamic scale of temperature, Clausius inequality, Entropy, principle of Entropy increase, availability and irreversibility, thermodynamic potentials, Gibbs and Helmholtz functions, Maxwell relations, elementary treatment of the Third Law of thermodynamics.

Module -III | PURE SUBSTANCES | Classes: 09

Phase transformations, T-S and H-S diagrams, P-V-T surfaces, triple point at critical state properties

during change of phase, dryness fraction, Mollier charts, various thermodynamic processes and energy transfer, steam calorimeter.

Equation of state, specific and universal gas constants, throttling and free expansion processes, deviations from perfect gas model, Vander Waals equation of state.

Module -IV | MIXTURES OF PERFECT GASES | Classes: 09

Mole fraction, mass friction, gravimetric and volumetric analysis, volume fraction, Dalton‘s law of partial pressure, Avogadro‘s laws of additive volumes, and partial pressure, equivalent gas constant, internal energy, enthalpy, specific heats and entropy of mixture of perfect gases; psychometric properties, dry bulb temperature, wet bulb temperature, dew point temperature, thermodynamic wet bulb temperature, specific humidity, relative humidity, saturated air, vapour pressure, degree of saturation, adiabatic saturation, Carrier‘s equation, Psychometric chart.

Module-V | AIR CONDITIONING SYSTEMS | Classes: 09

Otto, Diesel, Dual combustion cycles, description and representation on P-V and T-S diagram, thermal efficiency, mean effective pressures on air standard basis, comparison of cycles, introduction to Brayton cycle and Bell Coleman cycle.

Text Books:

1. P. K. Nag, “Engineering Thermodynamics”, Tata McGraw Hill, 4th Edition, 2008.
2. Yunus Cengel, Michael A. Boles, “Thermodynamics-An Engineering Approach”, Tata McGraw Hill, 7th Edition, 2011.

Reference Books:

1. J. B. Jones, R. E. Dugan, “Engineering Thermodynamics”, Prentice Hall of India Learning, 1st Edition,2009.
2. Y. V. C. Rao, “An Introduction to Thermodynamics”, Universities Press, 3rd Edition, 2013.
3. K. Ramakrishna, “Engineering Thermodynamics”, Anuradha Publishers, 2nd Edition, 2011.
4. Holman. J.P, “Thermodynamics”, Tata McGraw Hill, 4th Edition, 2013.

Web References:

1. http://www.livescience.com/50776-thermodynamics.html
2. https://www3.nd.edu/~powers/ame.20231/planckdover.pdf

E-Text Books:

1. https://www3.nd.edu/~powers/ame.20231/planckdover.pdf
2. http://www.ebookdownloadz.net/2014/08/engineering-thermodynamics-by-pknag.html

UNIT I

BASIC CONCEPTS AND FIRST LAW OF THERMODYNAMICS

Thermodynamics is the science that deals with heat and work and those properties of substance that bear a relation to heat and work.

Thermodynamics is the study of the patterns of energy change. Most of this course will be concerned with understanding the patterns of energy change.

More specifically, thermodynamics deals with

(a) Energy conversion and

(b) The direction of change.

Basis of thermodynamics is experimental observation. In that sense it is an empirical science. The principles of thermodynamics are summarized in the form of four laws known as zeroth, first, second, and the third laws of thermodynamics.

The **zeroth law of thermodynamics** deals with thermal equilibrium and provides a means of measuring temperature.

The **first law of thermodynamics** deals with the conservation of energy and introduces the concept of internal energy.

The **second law of thermodynamics** dictates the limits on the conversion of heat into work and provides the yard stick to measure the performance of various processes. It also tells whether a particular process is feasible or not and specifies the direction in which a process will proceed. As a consequence it also introduces the concept of entropy.

The **third law defines** the absolute zero of entropy.

Macroscopic and Microscopic Approaches:

Microscopic approach uses the statistical considerations and probability theory, where we dealwith “average” for all particles under consideration. This is the approach used in the disciplines known as kinetic theory and statistical mechanics.

In the macroscopic point of view, of classical thermodynamics, one is concerned with the time-averaged influence of many molecules that can be perceived by the senses and measured by the instruments.The pressure exerted by a gas is an example of this. It results from the change in momentum of the molecules, as they collide with the wall. Here we are not concerned with the actions of individual molecules but with the time-averaged force on a given area that can be measured by a pressure gage.

From the macroscopic point of view, we are always concerned with volumes that are very large compared to molecular dimensions, and therefore a system (to be defined next) contains many molecules, and this is called continuum. The concept of continuum loses validity when the mean free path of molecules approaches the order of typical system dimensions.

System:

We introduce boundaries in our study called the **system** and **surroundings**.

The boundaries are set up in a way most conducive to understanding the energetics of what we're studying. Defining the system and surroundings is arbitrary, but it becomes important when we consider the exchange of energy between the system and surroundings.

Two types of exchange can occur between system and surroundings: (1) energy exchange (heat, work, friction, radiation, etc.) and (2) matter exchange (movement of molecules across the boundary of the system and surroundings). Based on the types of exchange which take place or don't take place, we will define three types of systems:

- **isolated systems:** no exchange of matter or energy
- **closed systems:** no exchange of matter but some exchange of energy
- **open systems:** exchange of both matter and energy

Control Volume

- Control volume is defined as a volume which encloses the matter and the device inside a control surface.
- Everything external to the control volume is the surroundings with the separation given by the control surface.
- The surface may be open or closed to mass flows and it may have flows from energy in terms of heat transfer and work across it.
- The boundaries may be moveable or stationary.

- In the case of a control surface that is closed to the mass flow, so that no mass can enter or escape the control volume, it is called a **control mass** containing same amount of matter at all times.

Property

- In thermodynamics a property is any characteristic of a system that is associated with the energy and can be quantitatively evaluated.
- The property of a system should have a definite value when the system is in a particular state.
- Thermodynamic property is a point function.
- Properties like volume of a system that depend on the mass of a system are called extensive properties.
- Properties like pressure or temperature which do not depend on the system mass are called intensive properties.
- The ratio of extensive property to the mass of the system are called specific properties and therefore become intensive properties.
- Substance can be found in three states of physical aggregation namely, solid, liquid and vapor which are called its phases.
- If the system consists of mixture of different phases, the phases are separated from each other by phase boundary.
- The thermodynamic properties change abruptly at the phase boundary, even though the intensive properties like temperature and pressure are identical.

Equilibrium:

- When the property of a system is defined, it is understood that the system is in equilibrium.
- If a system is in thermal equilibrium, the temperature will be same throughout the system.
- If a system is in mechanical equilibrium, there is no tendency for the pressure to change. In a single phase system, if the concentration is uniform and there is no tendency for mass transfer or diffusion, the system is said to be in chemical equilibrium.
- A system which is simultaneously in thermal, mechanical, and chemical equilibrium is said to be in thermal equilibrium.

Process

A process is path followed by a system in reaching a given final state of equilibrium state starting from a specified initial state. An actual process occurs only when the equilibrium state does not exist.An ideal process can be defined in which the deviation from thermodynamic equilibrium is

infinitesimal.All the states the system passes through during a quasi-equilibrium process may be considered equilibrium states.

For non-equilibrium processes, we are limited to a description of the system before the process occurs and after the equilibrium is restored.Several processes are described by the fact that one property remains constant.The prefix iso- is used to describe such processes.

A process is said to be reversible if both the system and its surroundings can be restored to their respective initial states by reversing the direction of the process.

- **Reversible**: if the process happens slow enough to be reversed.
- **Irreversible**: if the process cannot be reversed (like most processes).
- **Isobaric:** process done at constant pressure
- **Isochoric:** process done at constant volume
- **Isothermal:** process done at constant temperature
- **Adiabatic:** process where q=0
- **Cyclic:** process where initial state = final state

Internal Energy

- The molecule as a whole can move in x, y and z directions with respective components of velocities and hence possesses kinetic energy.
- There can be rotation of molecule about its center of mass and than the kinetic energy associated with rotation is called rotational energy.
- In addition the bond length undergoes change and the energy associated with it is called vibrational energy.
- The electron move around the nucleus and they possess a certain energy that is called electron energy.
- The microscopic modes of energy are due to the internal structure of the matter and hence sum of all microscopic modes of energy is called the internal energy.

Bulk kinetic energy (KE) and potential energy (PE) are considered separately and the other energy of control mass as a single property (U).

The total energy possessed by the body is given by:

$$E = KE + PE + U$$

Work

Whenever a system interacts with its surroundings, it can exchange energy in two ways- work and heat. In mechanics, work is defined as the product of the force and the displacement in the direction of the force.

Work done when a spring is compressed or extended: According to Hooke's law

$$\text{Spring force} = - k (x – x_0)$$

Where k is the spring constant, x_0 is the equilibrium position, and x is the final position. The negative sign shows that the direction of the spring force is opposite the direction of the displacement from x_0. The external force is equal in magnitude but opposite in sign to the spring force, so

$$\text{External force (force of your hands)} = k (x – x_0).$$

Now, we want to calculate the work done when we stretch the spring from position 1 to position 2.

$$W = \ F\, dx = \ k (x – x_0)\, d(x-x_0) = 1/2\, k\, [(x_2-x_0)^2 - (x_1-x_0)^2]$$

Work done when a volume is increased or decreased

Consider a gas in a container with a movable piston on top. If the gas expands, the piston moves out and work is done by the system on the surroundings.

Alternatively, if the gas inside contracts, the piston moves in and work is done by the surroundings on the system. Why would the gas inside contract or expand?

It would if the external pressure, P_{ex}, and the internal pressure, P_{in}, were different. To calculate the work done in moving the piston, we know that the force = pressure times area and then work equals pressure times area times distance or work equals pressure times the change in volume. So, W = the integral of (P_{ex}) dV.

The differential work done (dW) associated with a differential displacement (dl) is given by

$$dW = F\, dl$$

For a piston cylinder assembly,

$$dW = F\,dl = PA\,(dl) = P\,dV$$

If the gas is allowed to expand reversibly from the initial pressure P to final pressure P, then the work done is given by

$$W = \int p\,dV$$

- The integral represents the area under the curve on a pressure versus volume diagram. Therefore the work depends on the path followed and work is a path function and hence not a property of the system.
- The above expression does not represent work in the case of an irreversible process.
- The thermodynamic definition of work is " Work is said to be done by a system on the surrounding if the sole effect external to the system could be reduced to the raising of a mass through a distance".

Heat

Heat like work, is a form of energy. The energy transfer between a system and its surroundings is called heat if it occurs by virtue of the temperature difference across the boundary.The two modes of energy transfer – work and heat- depend on the choice of the system. Heat energy moves from a hotter body to a colder body upon contact of the two bodies. If two bodies at different temperatures are allowed to remain in contact, the system of two bodies will eventually reach a thermal equilibrium (they will have the same temperature).A body never contains heat. Rather heat is a transient phenomenon and can be identified as it crosses the boundary.

The State Postulate

The state of the system is described by its properties.Once a sufficient number of properties are specified, the rest of the properties assume some values automatically.The number of properties required to fix a 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. The system is calleda **simple compressible system** in the absence of electrical, magnetic, gravitational, motion, and surface tension effects.

The state postulate requires that the two properties specified be independent to fix the state.Two properties are independent if one property can be varied while the other one is held constant. Temperature and specific volume, for example, are always independent properties, and together they can fix the state of a simple compressible system.Thus, temperature and pressure are not sufficient to fix the state of a two-phase system.Otherwise an additional property needs to be specified for each effect that is significant.An additional property needs to be specified for each other effect that is significant.

Zeroth Law of Thermodynamics

We cannot assign numerical values to temperatures based on our sensations alone. Furthermore, our senses may be misleading. Several properties of material changes with temperature in a repeatable and predictable way, and this forms the basis of accurate temperature measurement.The commonly used mercury-in-glass thermometer for example, is based on the expansion of mercury with temperature.Temperature is also measured by using several other temperature dependant properties.Two bodies (eg. Two copper blocks) in contact attain thermal equilibrium when the heat transfer between them stops.

The equality of temperature is the only requirement for thermal equilibrium.

The **Zeroth Law of Thermodynamics**

If two bodies are in thermal equilibrium with a third body, they are also in thermal equilibrium with each other.This obvious fact cannot be concluded from the other laws of thermodynamics, and it serves as a basis of temperature measurement.By replacing the third body with a thermometer, the zeroth law can be restated *two bodies are in thermal equilibrium if both have the same temperature reading even if they are not in contact.* The zeroth law was first formulated and labeled by R.H. Fowler in 1931.

Temperature Scales

All temperature scales are based on some easily reproducible states such as the freezing and boiling point of water, which are also called the ice-point and the steam-point respectively. A mixture of ice and water that is in equilibrium with air saturated with water vapour at 1atm pressure, is said to be at the ice-point, and a mixture of liquid water and water vapour (with no air) in equilibrium at 1atm is said to be at the steam-point.Celsius and Fahrenheit scales are based on these two points (although the value assigned to these two values are different) and are referred as two-point scales. In thermodynamics, it is very desirable to have a temperature scale that is independent of the properties of the substance or substances.

Such a temperature scale is called a **thermodynamic temperature scale**.(Kelvin in SI)

Ideal gas temperature scale

The temperatures on this scale are measured using a constant volume thermometer.Based on the principle that at low pressure, the temperature of the gas is proportional to its pressure at constant volume. The relationship between the temperature and pressure of the gas in the vessel can be expressed as

$$T = a + b.P$$

Where the values of the constants *a* and*b* for a gas thermometer are determined experimentally.

Once *a* and*b* are known, the temperature of a medium can be calculated from the relation above by immersing the rigid vessel of the gas thermometer into the medium and measuring the gas pressure.Ideal gas temperature scale can be developed by measuring the pressures of the gas in the vessel at two reproducible points (such as the ice and steam points) and assigning suitable values to temperatures those two points. Considering that only one straight line passes through two fixed points on a plane, these two measurements are sufficient to determine the constants *a* and*b* in the above equation.

If the ice and the steam points are assigned the values 0 and 100 respectively, then the gas temperature scale will be identical to the Celsius scale. In this case, the value of the constant *a* (that corresponds to an absolute pressure of zero) is determined to be $–273.15^0$C when extrapolated.

The equation reduces to T = bP, and thus we need to specify the temperature at only one point to define an absolute gas temperature scale. Absolute gas temperature is identical to thermodynamic temperature in the temperature range in which the gas thermometer can be used.We can view that thermodynamic temperature scale at this point as an absolute gas temperature scale that utilizes an ideal gas that always acts as a low-pressure gas regardless of the temperature. At the Tenth international conference on weights and measures in 1954, the Celsius scale has been redefined in terms of a single fixed point and the absolute temperature scale.The triple point occurs at a fixed temperature and pressure for a specified substance.

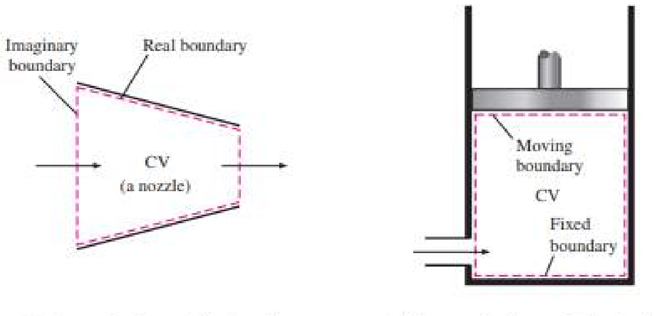
The selected single point is the **triple point** of water (the state in which all three phases of water coexist in equilibrium), which is assigned the value 0.01 C. As before the boiling point of water at 1 atm. Pressure is 100.0 C. Thus the new Celsius scale is essentially the same as the old one.

On the Kelvin scale, the size of Kelvin unit is defined as “ the fraction of 1/273.16 of the thermodynamic temperature of the triple point of water, which is assigned a value of 273.16K”. The ice point on Celsius and Kelvin are respectively 0 and 273.15 K.

SYSTEMS AND CONTROL VOLUMES:

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. These terms are illustrated in Fig. 1–1. The boundary of a system can be fixed or movable. Note that the boundary is the contact surface shared by both the system and the surroundings. Mathematically speaking, the boundary has zero thickness, and thus it can neither contain any mass nor occupy any volume in space.

Systems may be considered to be closed or open, depending on whether a fixed mass or a fixed volume in space is chosen for study. A closed system (also known as a control mass) consists of a fixed amount of mass, and no mass can cross its boundary. That is, no mass can enter or leave a closed system, as shown in Fig. 1–2. 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. Consider the piston-cylinder device shown in Fig. 1–3. Let us say that we would like to find out what happens to the enclosed gas when it is heated. Since we are focusing our attention on the gas, it is our system. The inner surfaces of the piston and the cylinder form the boundary, and since no mass is crossing this boundary, it is a closed system. Notice that energy may cross the boundary, and part of the boundary (the inner surface of the piston, in this case) may move. Everything outside the gas, including the piston and the cylinder, is the surroundings. 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 thesedevices 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. A large number of engineering problems involve mass flow in and out of a system and, therefore, are modeled as control volumes. A water heater, a car radiator, a turbine, and a compressor all involve mass flow and should be analyzed as control volumes (open systems) instead of as control masses (closed systems). In general, any arbitrary region in space can be selected as a control volume. There are no concrete rules for the selection of control volumes, but the proper choice certainly makes the analysis much easier. If we were to analyze the flow of air through a nozzle, for example, a good choice for the control volume would be the region within the nozzle. The boundaries of a control volume are called a control surface, and they can be real or imaginary. In the case of a nozzle, the inner surface of the nozzle forms the real part of the boundary, and the entrance and exit areas form the imaginary part, since there are no physical surfaces there (Fig. 1–4a). A control volume can be fixed in size and shape, as in the case of a nozzle, or it may involve a moving boundary, as shown in Fig. 1–4 b. Most control volumes, however, have fixed boundaries and thus do not involve any moving boundaries. A control volume can also involve heat and work interactions just as a closed system, in addition to mass interaction. As an example of an open system, consider the water heater shown in Fig.Let us say that we would like to determine how much heat we must transfer to the water in the tank in order to supply a steady stream of hot water. Since hot water will leave the tank and be replaced by cold water, it is not convenient to choose a fixed mass as our system for the analysis. Instead, we can concentrate our attention on the volume formed by the interior surfaces of the tank and consider the hot and cold water streams as mass leaving and entering the control volume. The interior surfaces of the tank form the control surface for this case, and mass is crossing the control surface at two locations.



(*a*) A control volume with real and imaginary boundaries

(*b*) A control volume with fixed and moving boundaries

Fig 1.1 A control volume can involve fixed, moving, real, and imaginary boundaries.

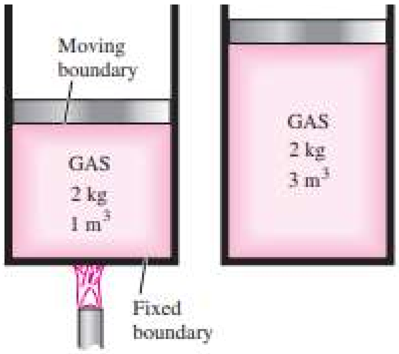


Fig 1.2 :A closed system with a moving boundary

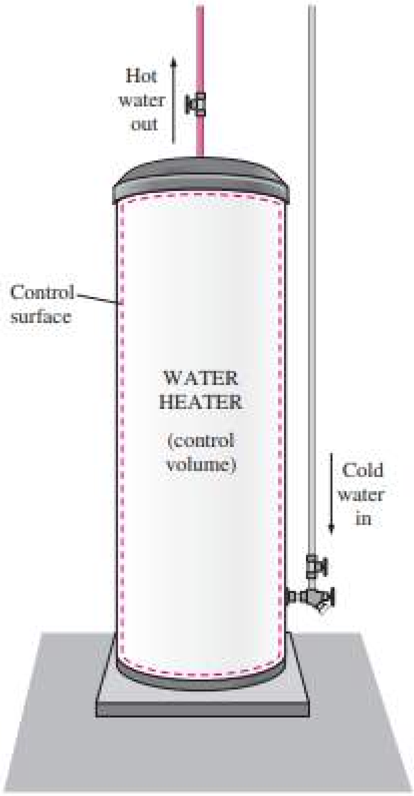


Fig 1.3: An open system (a control volume) with one inlet and one exit.

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. The list can be extended to include less familiar ones such as 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. 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. 1–20. Each part will have the same value of intensive properties as the original system, but half the value of the extensive properties. Generally, uppercase letters are used to denote extensive properties (with mass m being a major exception), and lowercase letters are used for intensive properties (with pressure P and temperature T being the obvious exceptions). 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).

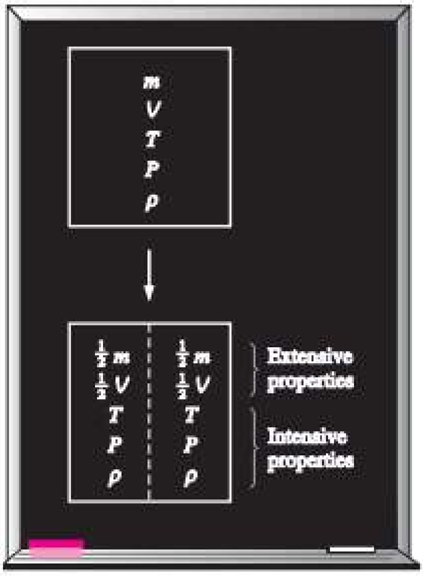


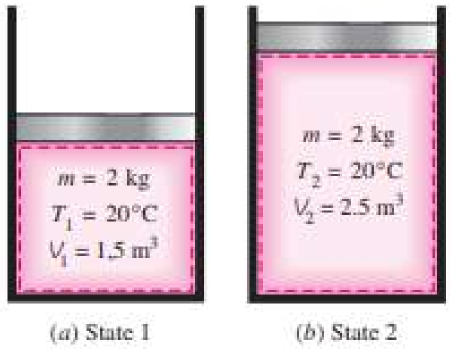
Fig1.4 : Criterion to differentiate intensive and extensive properties.

CONTINUUM:

Continuum Matter is made up of atoms that are widely spaced in the gas phase. Yet it is very convenient to disregard the atomic nature of a substance and view it as a continuous, homogeneous matter with no holes, that is, a continuum. The continuum idealization allows us to treat properties as point functions and to assume the properties vary continually in space with no jump discontinuities. This idealization is valid as long as the size of the system we deal with is large relative to the space between the molecules. This is the case in practically all problems, except some specialized ones. The continuum idealization is implicit in many statements we make, such as “the density of water in a glass is the same at any point.”

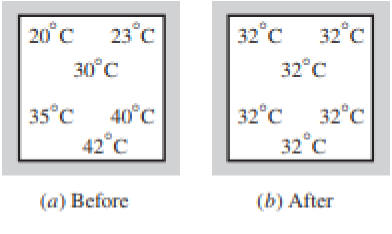
STATE AND EQUILIBRIUM:

Consider a system not undergoing any change. At this point, all the properties can be measured or calculated throughout the entire system, which gives us a set of properties that completely describes the condition, or the state, of the system. 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. In Fig. 1–5 a system is shown at two different states. Thermodynamics deals with equilibrium states. The word equilibrium implies a state of balance. In an equilibrium state there are no unbalanced potentials (or driving forces) within the system. 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. For example, a system is in thermal equilibrium if the temperature is the same throughout the entire system, as shown in Fig. 1–5. That is, the system involves no temperature differential, which is the driving force for heat flow. Mechanical equilibrium is related to pressure, and a system is in mechanical equilibrium if there is no change in pressure at any point of the system with time. However, the pressure may vary within the system with elevation as a result of gravitational effects. For example, the higher pressure at a bottom layer is balanced by the extra weight it must carry, and, therefore, there is no imbalance of forces. The variation of pressure as a result of gravity in most thermodynamic systems is relatively small and usually disregarded. If a system involves two phases, it is in phase equilibrium when the mass of each phase reaches an equilibrium level and stays there. Finally, a system is in chemical equilibrium if its chemical composition does not change with time, that is, no chemical reactions occur. A system will not be in equilibrium unless all the relevant equilibrium criteria are satisfied.



(*a*) State 1 (*b*) State 2

Fig.1.5: A system at two different states.



(*a*) Before (*b*) After

Fig.1.5: A closed system reaching thermal equilibrium.

PROCESSES AND CYCLES

Any change that a system undergoes from one equilibrium state to another is called a process, and the series of states through which a system passes during a process is called the path of the process (Fig. 1–6). To describe a process 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 quasistatic, or quasi-equilibrium, process. 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.

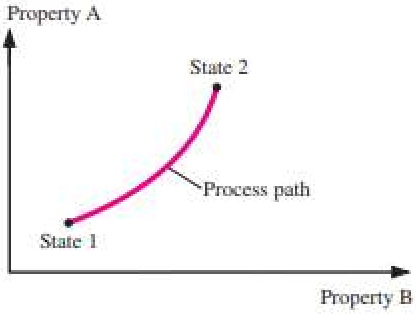
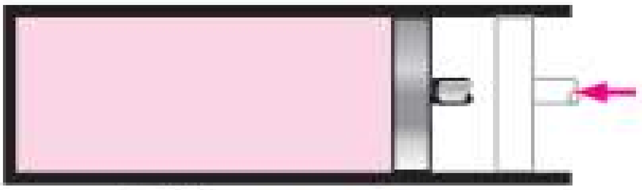
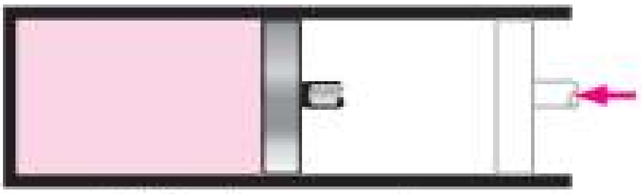


Fig1.6 :A process between states 1 and 2 and the process path.

This is illustrated in Fig. 1–7. When a gas in a piston-cylinder device is compressed suddenly, the molecules near the face of the piston will not have enough time to escape and they will have to pile up in a small region in front of the piston, thus creating a high-pressure region there. Because of this pressure difference, the system can no longer be said to be in equilibrium, and this makes the entire process nonquasi-equilibrium. However, if the piston is moved slowly, the molecules will have sufficient time to redistribute and there will not be a molecule pileup in front of the piston. As a result, the pressure inside the cylinder will always be nearly uniform and will rise at the same rate at all locations. Since equilibrium is maintained at all times, this is a quasi-equilibrium process. It should be pointed out that a quasi-equilibrium process is an idealized process and is not a true representation of an actual process. But many actual processes closely approximate it, and they can be modeled as quasiequilibrium with negligible error. Engineers are interested in quasiequilibrium processes for two reasons. First, they are easy to analyze; second,work-producing devices deliver the most work when they operate on quasiequilibrium processes. Therefore, quasi-equilibrium processes serve as standards to which actual processes can be compared. Process diagrams plotted by employing thermodynamic properties as coordinates are very useful in visualizing the processes. Some common properties that are used as coordinates are temperature T, pressure P, and volume V (or specific volume v). Figure 1–8 shows the P-V diagram of a compression process of a gas. Note that the process path indicates a series of equilibrium states through which the system passes during a process and has significance for quasiequilibrium processes only. For nonquasi-equilibrium processes, we are not able to characterize the entire system by a single state, and thus we cannot speak of a process path for a system as a whole. A nonquasi-equilibrium process is denoted by a dashed line between the initial and final states instead of a solid line. The prefix iso- is often used to designate a process for which a particular property remains constant. An isothermal process, for example, is a process during which the temperature T remains constant; an isobaric process is a process during which the pressure P remains constant; and an isochoric (or isometric) process is a process during which the specific volume v remains constant. A system is said to have undergone a cycle if it returns to its initial state at the end of the process. That is, for a cycle the initial and final states are identical.



(*a*) Slow compression
(quasi-equilibrium)



(*b*) Very fast compression
(nonquasi-equilibrium)

Fig: 1.8 Quasi-equilibrium and nonquasiequilibrium compression processes.

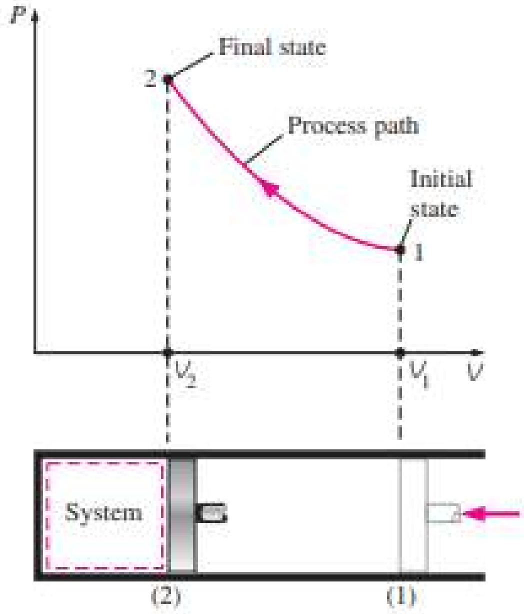


Fig. 1.9: The P-V diagram of a compression process.

The Steady-Flow Process:

The terms steady and uniform are used frequently in engineering, and thus it is important to have a clear understanding of their meanings. 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. These meanings are consistent with their everyday use (steady girlfriend, uniform properties, etc.).

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). That is, the fluid properties can change from point to point within the control volume, but at any fixed point they remain the same during the entire process. Therefore, the volume V, the mass m, and the total energy content E of the control volume remain constant during a steadyflow process.

Steady-flow conditions can be closely approximated by devices that are intended for continuous operation such as turbines, pumps, boilers, condensers, and heat exchangers or power plants or refrigeration systems. Some cyclic devices, such as reciprocating engines or compressors, do not satisfy any of the conditions stated above since the flow at the inlets and the exits will be pulsating and not steady. However, the fluid properties vary with time in a periodic manner, and the flow through these devices can still be analyzed as a steady-flow process by using time-averaged values for the properties.

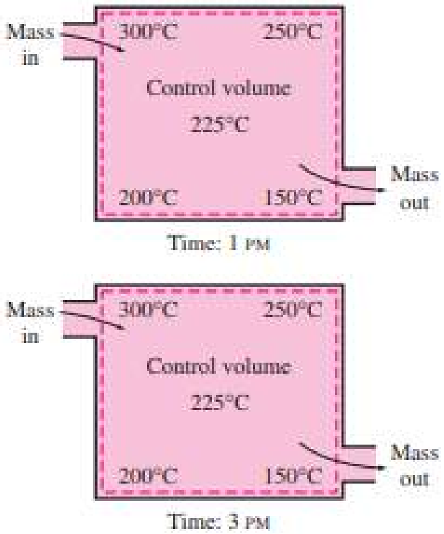


Fig 1– 9: During a steady-flow process, fluid properties within the control volume may change with position but not with time.

TEMPERATURE AND THE ZEROTH LAW OF THERMODYNAMICS

Although we are familiar with temperature as a measure of "hotness" or "coldness," it is not easy to give an exact definition for it. Based on our physiological sensations, we express the level of temperature qualitatively with words like freezing cold, cold, warm, hot, and red-hot. However, we cannot assign numerical values to temperatures based on our sensations alone. Furthermore, our senses may be misleading. A metal chair, for example, will feel much colder than a wooden one even when both are at the same temperature.

Fortunately, several properties of materials change with temperature in a repeatable and predictable way, and this forms the basis for accurate temperature measurement. The commonly used mercury-in-glass thermometer, for example, is based on the expansion of mercury with temperature. Temperature is also measured by using several other temperature-dependent properties.

It is a common experience that a cup of hot coffee left on the table eventually cools off and a cold drink eventually warms up. That is, when a body is brought into contact with another body that is at a different temperature, 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. The equality of temperature is the only requirement for thermal equilibrium.

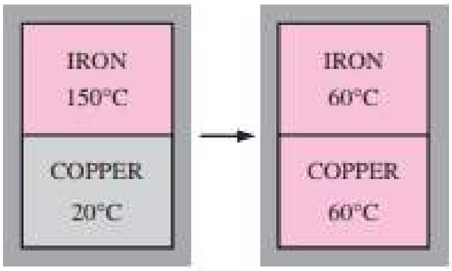


Fig. 1–10: Both bodies attain the same temperature

Fig:Two bodies reaching thermal equilibrium after being brought into contact in an isolated enclosure
The zeroth law of thermodynamics states that if two bodies are in thermal equilibrium with a third body, they are also in thermal equilibrium with each other. It may seem silly that such an obvious fact is called one of the basic laws of thermodynamics. However, it cannot be concluded from the other laws of thermodynamics, and it serves as a basis for the validity of temperature measurement. By replacing the third body with a thermometer, the zeroth law can be restated as two bodies are in thermal equilibrium if both have the same temperature reading even if they are not in contact.

The zeroth law was first formulated and labeled by R. H. Fowler in 1931. As the name suggests, its value as a fundamental physical principle was recognized more than half a century after the formulation of the first and the second laws of thermodynamics. It was named the zeroth law since it should have preceded the first and the second laws of thermodynamics.

ENERGY TRANSFER BY HEAT

Energy can cross the boundary of a closed system in two distinct forms: heat and work (Fig. 1–11). It is important to distinguish between these two forms of energy. Therefore, they will be discussed first, to form a sound basis for the development of the laws of thermodynamics. We know from experience that a can of cold soda left on a table eventually warms up and that a hot baked potato on the same table cools down. When a body is left in a medium that is at a different temperature, energy transfer takes place between the body and the surrounding medium until thermal equilibrium is established, that is, the body and the medium reach the same temperature. The direction of energy transfer is always from the higher temperature body to the lower temperature one. Once the temperature equality is established, energy transfer stops. In the processes described above, energy is said to be transferred in the form of heat.

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 (Fig. 1–12). That is, an energy interaction is heat only if it takes place because of a temperature difference. Then it follows that there cannot be any heat transfer between two systems that are at the same temperature.

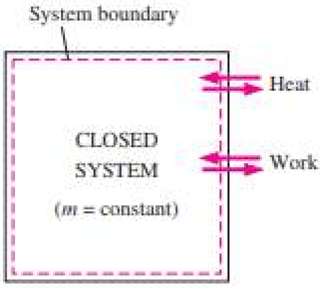


FIGURE 1–11 Energy can cross the boundaries of a closed system in the form of heat and work.

Several phrases in common use today—such as heat flow, heat addition, heat rejection, heat absorption, heat removal, heat gain, heat loss, heat storage, heat generation, electrical heating, resistance heating, frictional heating, gas heating, heat of reaction, liberation of heat, specific heat, sensible heat, latent heat, waste heat, body heat, process heat, heat sink, and heat source—are not consistent with the strict thermodynamic meaning of the term heat, which limits its use to the transfer of thermal energy during a process. However, these phrases are deeply rooted in our vocabulary, and they are used by both ordinary people and scientists without causing any misunderstanding since they are usually interpreted properly instead of being taken literally. (Besides, no acceptable alternatives exist for some of these phrases.) For example, the phrase body heat is understood to mean the thermal energy content of a body. Likewise, heat flow is understood to mean the transfer of thermal energy, not the flow of a fluidlike substance called heat, although the latter incorrect interpretation, which is based on the caloric theory, is the origin of this phrase. Also, the transfer of heat into a system is frequently referred to as heat addition and the transfer of heat out of a system as heat rejection. Perhaps there are thermodynamic reasons for being so reluctant to replace heat by thermal energy: It takes less time and energy to say, write, and comprehend heat than it does thermal energy.

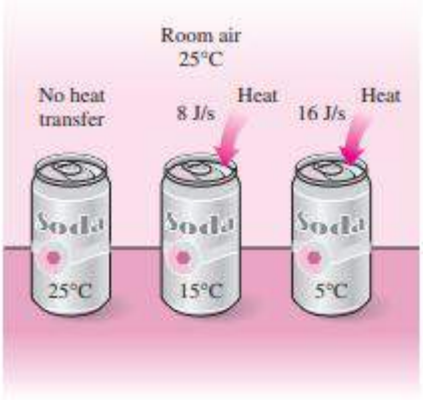


Fig 1–12 Temperature difference is the driving force for heat transfer. The larger the temperature difference, the higher is the rate of heat transfer.

Heat is energy in transition. It is recognized only as it crosses the boundary of a system. Consider the hot baked potato one more time. The potato contains energy, but this energy is heat transfer only as it passes through the skin of the potato (the system boundary) to reach the air, as shown in Fig. 1–13. Once in the surroundings, the transferred heat becomes part of the internal energy of the surroundings. Thus, in thermodynamics, the term heat simply means heat transfer.

A process during which there is no heat transfer is called an adiabatic process (Fig. 1–14). The word adiabatic comes from the Greek word adiabatos, which means not to be passed. 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.

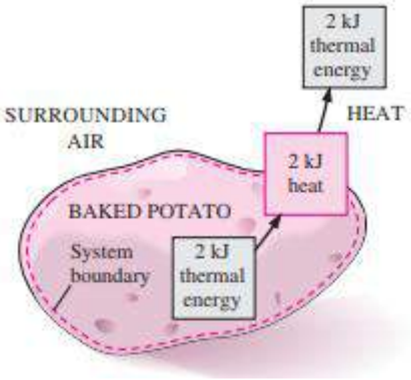


FIGURE 1–13 Energy is recognized as heat transfer only as it crosses the system boundary.

As a form of energy, heat has energy units, kJ (or Btu) being the most common one. The amount of heat transferred during the process between two states (states 1 and 2) is denoted by Q12, or just Q. Heat transfer per unit mass of a system is denoted q and is determined from

$$q = \frac{Q}{m} \qquad (\mathrm{kJ/kg})$$

Sometimes it is desirable to know the rate of heat transfer (the amount of heat transferred per unit time) instead of the total heat transferred over some time interval (Fig. 2–15). The heat transfer rate is denoted Q ., where the overdot stands for the time derivative, or “per unit time.” The heat transfer rate Q .has the unit kJ/s, which is equivalent to kW. When Q .varies with time, the amount of heat transfer during a process is determined by integrating Q . over the time interval of the process:

$$Q = \int_{t_1}^{t_2} \dot{Q}\, dt \qquad \text{(kJ)}$$

When Q .remains constant during a process, this relation reduces to

$$Q = \dot{Q}\, \Delta t \qquad \text{(kJ)}$$

where t t2 t1 is the time interval during which the process takes place.

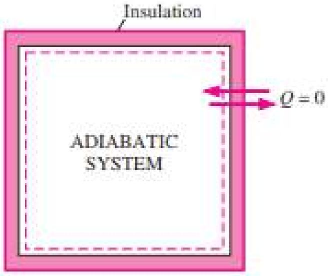


FIGURE 1–14 During an adiabatic process, a system exchanges no heat with its surroundings.

ENERGY TRANSFER BY WORK:

Work, like heat, is an energy interaction between a system and its surroundings. As mentioned earlier, energy can cross the boundary of a closed system in the form of heat or work. Therefore, if the energy crossing the boundary of a closed system is not heat, it must be work. Heat is easy to recognize: Its driving force is a temperature difference between the system and its surroundings. Then we can simply say that an energy interaction that is not caused by a temperature difference between a system and its surroundings is work. More specifically, work is the energy transfer associated with a force acting through a distance. A rising piston, a rotating shaft, and an electric wire crossing the system boundaries are all associated with work interactions.

Work is also a form of energy transferred like heat and, therefore, has energy units such as kJ. The work done during a process between states 1 and 2 is denoted by W12, or simply W. The work done per unit mass of a system is denoted by w and is expressed as

$$w = \frac{W}{m} \qquad \text{(kJ/kg)}$$

The work done per unit time is called power and is denoted W .(Fig. 2–16). The unit of power is kJ/s, or kW.

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. The generally accepted formal sign convention for heat and work interactions is as follows: 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. Another way is to use the subscripts in and out to indicate direction (Fig. 1–

18). For example, a work input of 5 kJ can be expressed as Win 5 kJ, while a heat loss of 3 kJ can be expressed as Qout 3 kJ. When the direction of a heat or work interaction is not known, we can simply assume a direction for the interaction (using the subscript in or out) and solve for it. A positive result indicates the assumed direction is right. A negative result, on the other hand, indicates that the direction of the interaction is the opposite of the assumed direction. This is just like assuming a direction for an unknown force when solving a statics problem, and reversing the direction when a negative result is obtained for the force. We will use this intuitive approach in this book as it eliminates the need to adopt a formal sign convention and the need to carefully assign negative values to some interactions.

FIGURE 2–8 Specifying the directions of heat and work.

Note that a quantity that is transferred to or from a system during an interaction is not a property since the amount of such a quantity depends on more than just the state of the system. 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).

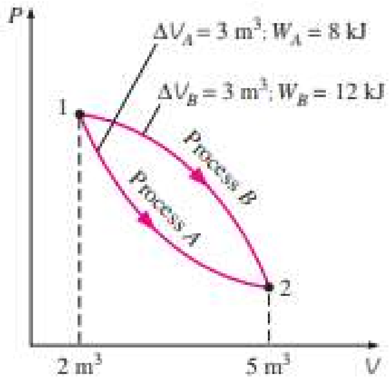
Path functions have inexact differentials designated by the symbol d. Therefore, a differential amount of heat or work is represented by dQ or dW, respectively, instead of dQ or dW. Properties, however, are point functions (i.e., they depend on the state only, and not on how a system reaches that state), and they have exact differentials designated by the symbol d. A small change in volume, for example, is represented by dV, and the total volume change during a process between states 1 and 2 is

$$\int_1^2 dV = V_2 - V_1 = \Delta V$$

That is, the volume change during process 1–2 is always the volume at state 2 minus the volume at state 1, regardless of the path followed (Fig. 2–19). The total work done during process 1–2, however, is

$$\int_1^2 \delta W = W_{12} \qquad (not\ \Delta W)$$

That is, the total work is obtained by following the process path and adding the differential amounts of work (dW) done along the way. The integral of dW is not W2 W1 (i.e., the work at state 2 minus work at state 1), which is meaningless since work is not a property and systems do not possess work at a state.



THE FIRST LAW OF THERMODYNAMICS

So far, we have considered various forms of energy such as heat Q, work W, and total energy E individually, and no attempt is made to relate them to each other during a process. The first law of thermodynamics, also known as 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.

We all know that a rock at some elevation possesses some potential energy, and part of this potential energy is converted to kinetic energy as the rock falls (Fig. 1–8). Experimental data show that the decrease in potential energy (mg z) exactly equals the increase in kinetic energy when the air resistance is negligible, thus confirming the conservation of energy principle for mechanical energy.

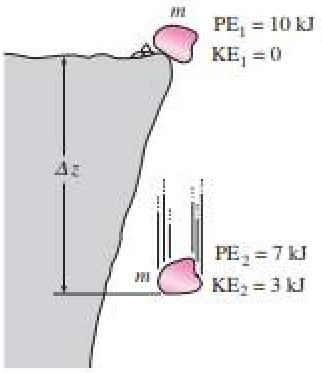


FIGURE 1–8 Energy cannot be created or destroyed; it can only change forms.

Consider a system undergoing a series of adiabatic processes from a specified state 1 to another specified state 2. Being adiabatic, these processes obviously cannot involve any heat transfer, but they may involve several kinds of work interactions. Careful measurements during these experiments indicate the following: 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. Considering that there are an infinite number of ways to perform work interactions under adiabatic conditions, this statement appears to be very powerful, with a potential for far-reaching implications. This statement, which is largely based on the experiments of Joule in the first half of the nineteenth century, cannot be drawn from any other known physical principle and is recognized as a fundamental principle. This principle is called the first law of thermodynamics or just the first law.

A major consequence of the first law is the existence and the definition of the property total energy E. Considering that the net work is the same for all adiabatic processes of a closed system between two specified states, the value of the net work must depend on the end states of the system only, and thus it must correspond to a change in a property of the system. This property is the total energy. Note that the first law makes no reference to the value of the total energy of a closed system at a state. It simply states that the change in the total energy during an adiabatic process must be equal to the net work done. Therefore, any convenient arbitrary value can be assigned to total energy at a specified state to serve as a reference point.

Implicit in the first law statement is the conservation of energy. Although the essence of the first law is the existence of the property total energy, the first law is often viewed as a statement of the conservation of energy principle. Next we develop the first law or the conservation of energy relation with the help of some familiar examples using intuitive arguments.

First, we consider some processes that involve heat transfer but no work interactions. The potato baked in the oven is a good example for this case (Fig. 2–38). As a result of heat transfer to the potato, the energy of the potato will increase. If we disregard any mass transfer (moisture loss from the potato), the increase in the total energy of the potato becomes equal to the amount of heat transfer. That is, if 5 kJ of heat is transferred to the potato, the energy increase of the potato will also be 5 kJ.

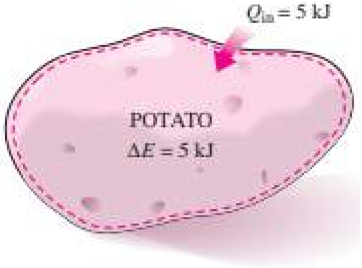


FIGURE 1–9The increase in the energy of a potato in an oven is equal to the amount of heat transferred

As another example, consider the heating of water in a pan on top of a range (Fig. 2–39). If 15 kJ of heat is transferred to the water from the heating element and 3 kJ of it is lost from the water to the surrounding air, the increase in energy of the water will be equal to the net heat transfer to water, which is 12 kJ.

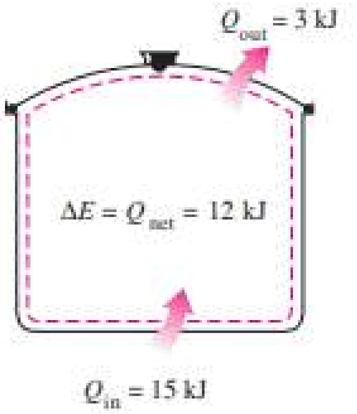


FIGURE:2–39 In the absence of any work interactions, the energy change of a system is equal to the net heat transfer.

Now consider a well-insulated (i.e., adiabatic) room heated by an electric heater as our system (Fig. 2–40). As a result of electrical work done, the energy of the system will increase. Since the system is adiabatic and cannot have any heat transfer to or from the surroundings (Q 0), the conservation of energy principle dictates that the electrical work done on the system must equal the increase in energy of the system.

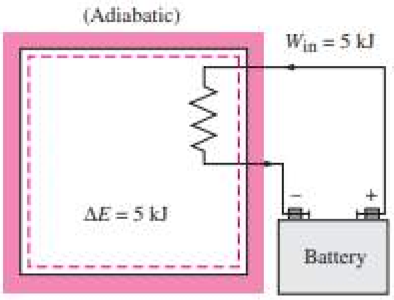


FIGURE 2–40 The work (electrical) done on an adiabatic system is equal to the increase in the energy of the system.

ENERGY BALANCE

In the light of the preceding discussions, 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,

(Total energy entering the system) – (Total energy leaving the system)=(Change in the total energy of the system)

$$E_{in} - E_{out} = \Delta E_{system}$$

This relation is often referred to as the energy balance and is applicable to any kind of system undergoing any kind of process. The successful use of this relation to solve engineering problems depends on understanding the various forms of energy and recognizing the forms of energy transfer.

UNIT II

THE SECOND LAW OF THERMODYNAMICS

INTRODUCTION TO THE SECOND LAW

The *first law of thermodynamics*, or the *conservationof energy principle*, to processes involving closed and open systems. As pointed out repeatedly in those 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.

The first law places no restriction on the direction of a process, but satisfying the first law does notensure 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*.

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 _ 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 modeled accurately as thermal energy reservoirs because of their large thermal
energy storage capabilities or thermal masses (Fig. 6–6). The *atmosphere*, for example, does not warm up as a result of heat losses from residential buildings in winter. Likewise, megajoules of waste energy
dumped in large rivers by power plants do not cause any significant change in water temperature.

A *two-phase system* can be modeled as a reservoir also 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. 6–7). Thermal energy reservoirs are often referred to as **heat reservoirs** since they supply or absorb energy in the form of heat.

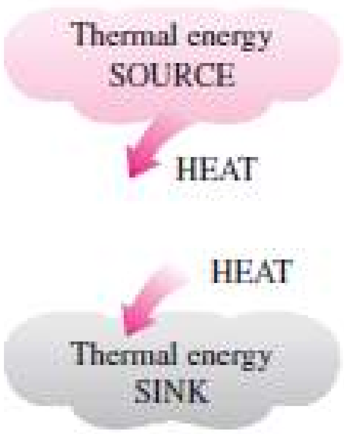


FIGURE 2–1: Source supplies energy in the formof heat, and a sink absorbs it.

Heat transfer from industrial sources to the environment is of major concern to environmentalists as well as to engineers. Irresponsible management of waste energy can significantly increase the temperature of portions of the environment, causing what is called *thermal pollution*. If it is not carefully controlled, thermal pollution can seriously disrupt marine life in lakes and rivers. However, by careful design and management, the waste energy dumped into large bodies of water can be used to improve the qualityof marine life by keeping the local temperature increases within safe and desirable levels.

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 mechanicalwork done by the shaft shown in Fig. 6–8, for example, is first convertedto the internal energy of the water. This energy may then leave the water asheat. We know from experience that any attempt to reverse this process willfail. 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 toheat directly and completely, but converting heat to work requires the use ofsome special devices. These devices are called heat engines.

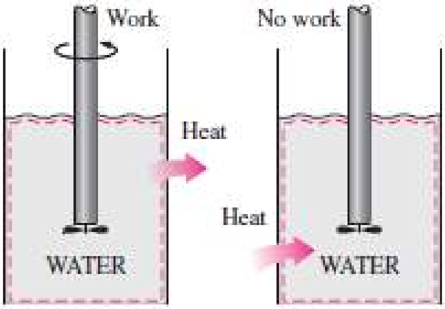


FIG 2-2 : Work can always be converted to heat directly and completely, but the reverse is not true.

Heat engines differ considerably from one another, but all can be characterizedby the following (Fig. 6–9):

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 rotatingshaft).
3. They reject the remaining waste heat to a low-temperature sink (theatmosphere, rivers, etc.).
4. They operate on a cycle.

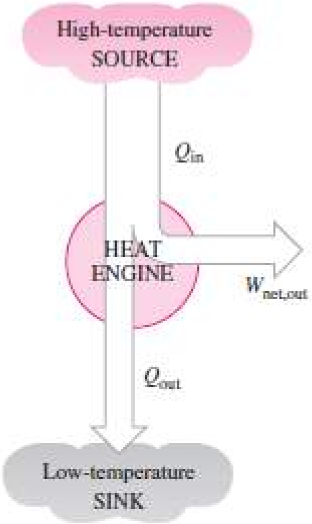


FIGURE 2–3: Part of the heat received by a heat engine is converted to work, while the rest is rejected to a sink.

Heat engines and other cyclic devices usually involve a fluid to and fromwhich 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 producingdevices that do not operate in a thermodynamic cycle. Enginesthat involve internal combustion such as gas turbines and car engines fall intothis category. These devices operate in a mechanical cycle but not in athermodynamic cycle since the working fluid (the combustion gases) doesnot undergo a complete cycle. Instead of being cooled to the initial temperature,the exhaust gases are purged and replaced by fresh air-and-fuel mixtureat the end of the cycle.

The work-producingThe work-producing device that best fits into the definition of a heatengine is the *steam power plant*, which is an external-combustion engine.That is, combustion takes place outside the engine, and the thermal energyreleased during this process is transferred to the steam as heat. Theschematic of a basic steam power plant is shown in Fig. 6–10. This is arather simplified diagram, and the discussion of actual steam power plantsis given in later chapters. The various quantities shown on this figure areas follows:

Qin = amount of heat supplied to steam in boiler from a high-temperaturesource (furnace)
Qout= amount of heat rejected from steam in condenser to a lowtemperature sink (the atmosphere, a river, tc.)
Wout= amount of work delivered by steam as it expands in turbine
Win =amount of work required to compress water to boiler pressure

Notice that the directions of the heat and work interactions are indicatedby the subscripts *in* and *out*. Therefore, all four of the described quantitiesare always *positive*.

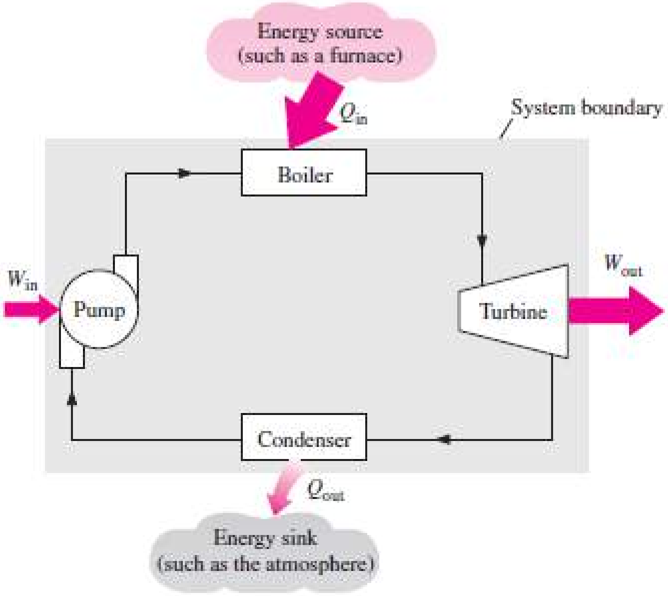


FIGURE 2–4: Schematic of a steam power plant

The net work output of this power plant is simply the difference between the total work output of the plant and the total work input (Fig. 6–11):

$$W\text{net,out} = W\text{out} - W\text{inkJ}$$

The net work 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 they should be treated as open systems. These components, together with the connecting pipes, however, always contain the

same fluid (not countingthe steam that may leak out, of course). No mass enters or leaves this combination system, which is indicated by the shaded area on Fig. 6–10; thus, it can be analyzed as a closed system. Recall that for a closed system undergoing a cycle, the change in internal energy _*U* is zero, and therefore the net work
output of the system is also equal to the net heat transfer to the system:

$$W_{net,out}= Q_{in}-Q_{out}\,\text{kJ}$$

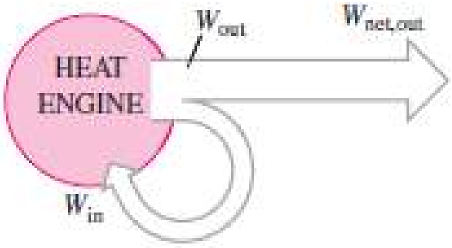


FIGURE 6–11: A portion of the work output of a heat engine is consumed internally to maintain continuous operation.

Thermal Efficiency:

*Q*out represents the magnitude of the energy wasted in order to complete the cycle. But *Q*out is never zero; thus, the net work 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 net work output is a measure of the performance of a heat engine and is called the **thermal efficiency** hth (Fig. 6–12). For heat engines, the desired output is the net work 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

$$\text{Thermal efficiency}=\frac{\text{Net work output}}{\text{Total heat input}}$$

$$\eta_{th} = \frac{W_{net,out}}{Q_{in}}$$

It can also be expressed as

$$\eta_{th} = 1 - \frac{Q_{out}}{Q_{in}}$$

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

QH = magnitude of heat transfer between the cyclic device and the hightemperaturemedium at temperature *TH*
QL = magnitude of heat transfer between the cyclicdevice and the lowtemperaturemedium at temperature *TL*
Notice that both *QL* and *QH* are defined as *magnitudes* and therefore arepositive quantities. The direction of *QH* and *QL* is easily determined byinspection. Then the net work output and thermal efficiency relations forany heat engine (shown inFig. 6–13) can also be expressed as

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

$$\eta_{\text{th}} = \frac{W_{\text{net,out}}}{Q_H}$$

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

The Second Law of Thermodynamics:

Kelvin–Planck Statement

We have demonstrated earlier with reference to the heat engine shown inFig. 6–15 that, even under ideal conditions, a heat engine must reject someheat 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 limitationon the thermal efficiency of heat engines forms the basis for theKelvin–Planck statement of the second law of thermodynamics, which isexpressed as follows:

That is, a heat engine must exchange heat with a low-temperature sink as wellas a high-temperature source to keep operating. The Kelvin–Planck statementcan also be expressed as *no heat engine can have a thermal efficiency of100 percent* (Fig. 2–8), or as *for a power plant to operate, the working fluidmust exchange heat with the environment as well as the furnace*.

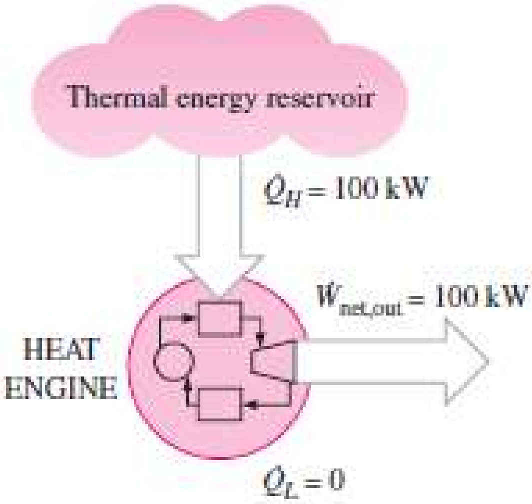


FIGURE 2–8: A heat engine that violates theKelvin–Planck statement of thesecond law.

Note that the impossibility of having a 100 percent efficient heat engine isnot due to friction or other dissipative effects. It is a limitation that appliesto both the idealized and the actual heat engines. Later in this chapter, wedevelop a relation for the maximum thermal efficiency of a heat engine. Wealso demonstrate that this maximum value depends on the reservoir temperatures only.

REFRIGERATORS AND HEAT PUMPS

We all know from experience that heat is transferred in the direction ofdecreasing temperature, that is, from high-temperature mediums to lowtemperatureones. This heat transfer process occurs in nature without requiringany devices. The reverse process, however, cannot occur by itself. The transfer of heat from a low-temperature medium to a high-temperature onerequires special devices called **refrigerators.**

Refrigerators, like heat engines, are cyclic devices. The working fluidused in the refrigeration cycle is called a **refrigerant.** The most frequentlyused refrigeration cycle is the *vapor-compression refrigeration cycle*, whichinvolves four main components: a compressor, acondenser, an expansionvalve, and an evaporator, as shown in Fig. 2–9.

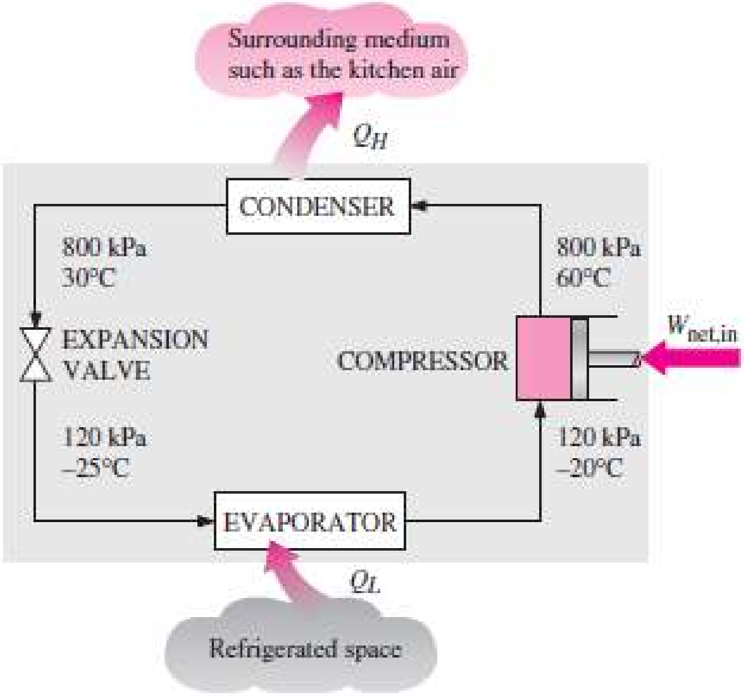


FIGURE 2–9 Basic components of a refrigerationsystem and typical operatingconditions.

The refrigerant enters the compressor as a vapor and is compressed to thecondenser pressure. It leaves the compressor at a relatively high temperatureand cools down and condenses as it flows through the coils of the condenserby rejecting heat to the surrounding medium. It then enters a capillary tubewhere its pressure and temperature drop drastically due to the throttling effect.The low-temperature refrigerant then enters the evaporator, where it evaporatesby absorbing heat from the refrigerated space. The cycle is completed asthe refrigerant leaves the evaporator and reenters the compressor.

Coefficient of Performance

The *efficiency* of a refrigerator is expressed in terms of the **coefficient ofperformance** (COP), denoted by COPR. The objective of a refrigerator is toremove heat (*QL*) from the refrigerated space. To accomplish this objective,it requires a work input of *W*net,in. Then the COP of a refrigerator can beexpressed as

$$COP_R = \frac{\text{Desired output}}{\text{Required input}}$$

$$COP_R = \frac{\text{Desired output}}{\text{Required input}} = \frac{Q_L}{W_{\text{net,in}}}$$

This relation can also be expressed in rate form by replacing *QL* by *Q L* and *W*net,in by *W*net,in.

$$W_{\text{net,in}} = Q_H - Q_L \qquad (\text{kJ})$$

Then the COP relation becomes

Notice that the value of COPR 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.

Heat Pumps

Another device that transfers heat from a low-temperature medium to a high-temperature one is the **heat pump,** shown schematically in Fig. 6–21. Refrigerators and heat pumps operate on the same cycle but differ in theirobjectives. 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 inwinter, and supplying this heat to the high-temperature medium such as a house 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 pumpsince 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** COPHP, defined as

$$\text{COP}_{\text{HP}} = \frac{\text{Desired output}}{\text{Required input}} = \frac{Q_H}{W_{\text{net,in}}}$$

which can also be expressed as

$$\text{COP}_{\text{HP}} = \frac{Q_H}{Q_H - Q_L} = \frac{1}{1 - Q_L/Q_H}$$

A comparison of aboveEqs. reveals that

$$\text{COP}_{\text{HP}} = \text{COP}_{\text{R}} + 1$$

ENTROPY:

The second law of thermodynamics often leads to expressionsthat involve inequalities. An irreversible (i.e., actual) heat engine, for example, is less efficient than a reversible one operating between the same two thermalenergy reservoirs. Likewise, an irreversible refrigerator or a heat pump has a lower coefficient of performance (COP) than a reversible one operating between the same temperature limits. Another important inequality that has major consequences in thermodynamics is the **Clausius inequality.** It was first stated by the German physicist R. J. E.

$$\oint \frac{\delta Q}{T} \leq 0$$

Clausius (1822–1888), one of the founders of thermodynamics, and is expressed asThat is, *the cyclic integral of* d*Q/T is always less than or equal to zero*. This inequality is valid for all cycles, reversible or irreversible. The symbol _ (integral symbol with a circle in the middle) is used to indicate that the integration is to be performed over the entire cycle. Any heat transfer to or from a system can be considered to consist of differential amounts of heat transfer. Then the cyclic integral of d*Q*/*T* can be viewed as the sum of all these differential amounts of heat transfer divided by the temperature at the boundary.

To demonstrate the validity of the Clausius inequality, consider a systemconnected to a thermal energy reservoir at a constant thermodynamic (i.e.,absolute) temperature of *TR* through a *reversible* cyclic device (Fig. 7–1).The cyclic device receives heat d*QR*from the reservoir and supplies heat d*Q*to the system whose temperature at that part of the boundary is *T* (a variable) while producing work d*W*rev. The system produces work d*W*sys as aresult of this heat transfer. Applying the energy balance to the combinedsystem identified by dashed lines yields

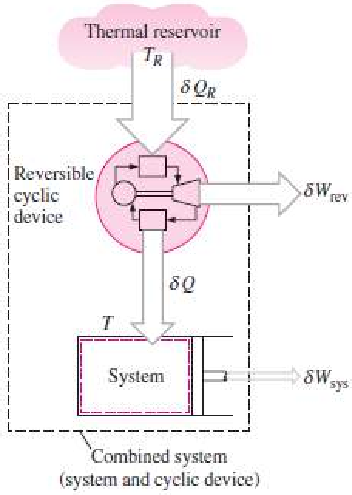


FIGURE 7–1: The system considered in thedevelopment of the Clausiusinequality.

THE MAXWELL RELATIONS

The equations that relate the partial derivatives of properties *P*, *v*, *T*, and *s*of a simple compressible system to each other are called the *Maxwell relations*.They are obtained from the four Gibbs equations by exploiting theexactness of the differentials of thermodynamic properties.

Two of the Gibbs relations expressed as

$$du = T\,ds - P\,dv$$

$$dh = T\,ds + v\,dP$$

The other two Gibbs relations are based on two new combination properties—the **Helmholtz function** *a* and the **Gibbs function** *g*, defined as

$$a = u - Ts$$

$$g = h - Ts$$

Differentiating, we get

$$da = du - T\,ds - s\,dT$$

$$dg = dh - T\,ds - s\,dT$$

Simplifying the above relations by above Eqs. Weobtain the other two Gibbs relations for simple compressible systems:

$$da = -s\,dT - P\,dv$$

$$dg = -s\,dT + v\,dP$$

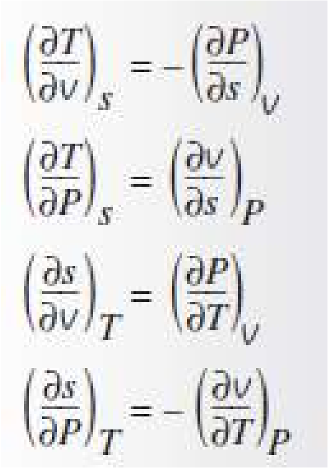
A careful examination of the four Gibbs relations reveals that they are of theform

$$dz = M\,dx + N\,dy$$

With

$$\left(\frac{\partial M}{\partial y}\right)_x = \left(\frac{\partial N}{\partial x}\right)_y$$

since*u*, *h*, *a*, and *g* are properties and thus have exact differentials. ApplyingEq. 12–5 to each of them, we obtain



$$\left(\frac{\partial T}{\partial v}\right)_s = -\left(\frac{\partial P}{\partial s}\right)_v$$

$$\left(\frac{\partial T}{\partial P}\right)_s = \left(\frac{\partial v}{\partial s}\right)_P$$

$$\left(\frac{\partial s}{\partial v}\right)_T = \left(\frac{\partial P}{\partial T}\right)_v$$

$$\left(\frac{\partial s}{\partial P}\right)_T = -\left(\frac{\partial v}{\partial T}\right)_P$$

These are called the **Maxwell relations** They are extremelyvaluable in thermodynamics because they provide a means of determiningthe change in entropy, which cannot be measured directly, by simply measuringthe changes in properties *P*, *v*, and *T*. Note that the Maxwell relationsgiven above are limited to simple compressible systems. However, othersimilar relations can be written just as easily for nonsimple systems such asthose involving electrical, magnetic, and other effects.

UNIT III

PURE SUBSTANCE

A substance that has a fixed chemical composition throughout is called apure 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 orcompound, however. A mixture of various chemical elements or compoundsalso 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 tobe a pure substance because it has a uniform chemical composition(Fig. 3–1). However, 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, formingtwo chemically dissimilar regions.

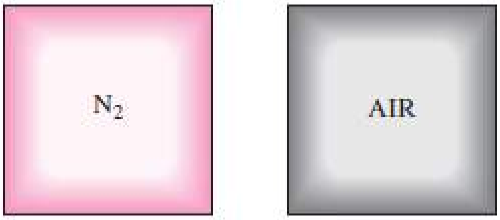
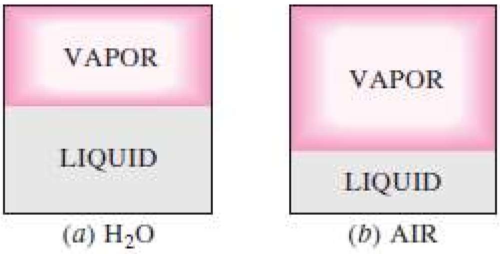


FIGURE 3–1: Nitrogen and gaseous air are puresubstances.

A mixture of two or more phases of a pure substance is still a pure substanceas long as the chemical composition of all phases is the same(Fig. 3–2). A mixture of ice and liquid water, for example, is a pure substancebecause both phases have the same chemical composition. A mixture of liquid air and gaseous air, however, is not a pure substance since thecomposition of liquid air is different from the composition of gaseous air,and thus the mixture is no longer chemically homogeneous. This is dueto different components in air condensing at different temperatures at a specified pressure.



(a) H_2O (b) AIR

FIGURE 3–2A mixture of liquid and gaseous wateris a pure substance, but a mixture of liquid and gaseous air is not.

PHASES OF A PURE SUBSTANCE

We all know from experience that substances exist in different phases. Atroom temperature and pressure, copper is a solid, mercury is a liquid, andnitrogen is a gas. Under different conditions, each may appear in a differentphase. Even though there are three principal phases—solid, liquid, andgas—a substance may have several phases within a principal phase, eachwith a different molecular structure. Carbon, for example, may exist asgraphite or diamond in the solid phase. Helium has two liquid phases; ironhas three solid phases. Ice may exist at seven different phases at high pressures.A phase is identified as having a distinct molecular arrangement thatis homogeneous throughout and separated from the others by easily identifiableboundary surfaces.

The two phases of H2O in iced water represent agood example of this.When studying phases or phase changes in thermodynamics, one does notneed to be concerned with the molecular structure and behavior of differentphases. However, it is very helpful to have some understanding of the molecularphenomena involved in each phase, and a brief discussion of phasetransformations follows.

Intermolecular bonds are strongest in solids and weakest in gases. Onereason is that molecules in solids are closely packed together, whereas ingases they are separated by relatively large distances.

The molecules in a **solid** are arranged in a three-dimensional pattern (lattice)that is repeated throughout (Fig. 3–3). Because of the small distancesbetween molecules in a solid, the attractive forces of molecules on eachother are large and keep the molecules at fixed positions Notethat the attractive forces between molecules turn to repulsive forces as thedistance between the molecules approaches zero, thus preventing the moleculesfrom piling up on top of each other. Even though the molecules in asolid cannot move relative to each other, they continually oscillate abouttheir equilibrium positions. The velocity of the molecules during these oscillationsdepends on the temperature. At sufficiently high temperatures, thevelocity (and thus the momentum) of the molecules may reach a pointwhere the intermolecular forces are partially overcome and groups of moleculesbreak away (Fig. 3–5). This is the beginning of the melting process.

The molecular spacing in the **liquid** phase is not much different from thatof the solid phase, except the molecules are no longer at fixed positions relativeto each other and they can rotate and translate freely. In a liquid, theintermolecular forces are weaker relative to solids, but still relatively strongcompared with gases. The distances between molecules generally experiencea slight increase as a solid turns liquid, with water being a notableexception.

In the gas phase, the molecules are far apart from each other, and a molecularorder is nonexistent. Gas molecules move about at random, continuallycolliding with each other and the walls of the container they are in. Particularlyat low densities, the intermolecular forces are very small, and collisionsare the only mode of interaction between the molecules. Molecules inthe gas phase are at a considerably higher energy level than they are in theliquid or solid phases. Therefore, the gas must release a large amount of itsenergy before it can condense or freeze.

PHASE-CHANGE PROCESSESOF PURE SUBSTANCES:

There are many practical situations where two phases of a pure substanceco exist 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 homeowners 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.

PROPERTY DIAGRAMS FOR PHASE-CHANGEPROCESSES:

The variations of properties during phase-change processes are best studied and understood with the help of property diagrams. Next, we develop and discuss the *T-v*, *P-v*, and *P-T* diagrams for pure substances.

The *T-v* Diagram

The phase-change process of water at 1 atm pressure was described in detail in the last section and plotted on a *T-v* diagram in Fig. 3–11. Now we repeat this process at different pressures to develop the *T-v* diagram. Let us add weights on top of the piston until the pressure inside the cylinder reaches 1 MPa. 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 theprocess path at 1 atm pressure, as shown in Fig. 3–16, but there are some notice able differences. First, water starts boiling at a much higher tempera-ture (179.9°C) 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 shorterAs the pressure is increased further, this saturation line continues toshrink, as shown in Fig. 3–16, and it becomes a point when the pressurereaches 22.06 MPa for the case of water. This point is called the **criticalpoint,** and it is defined as *the point at which the saturated liquid and saturatedvapor states are identical.*

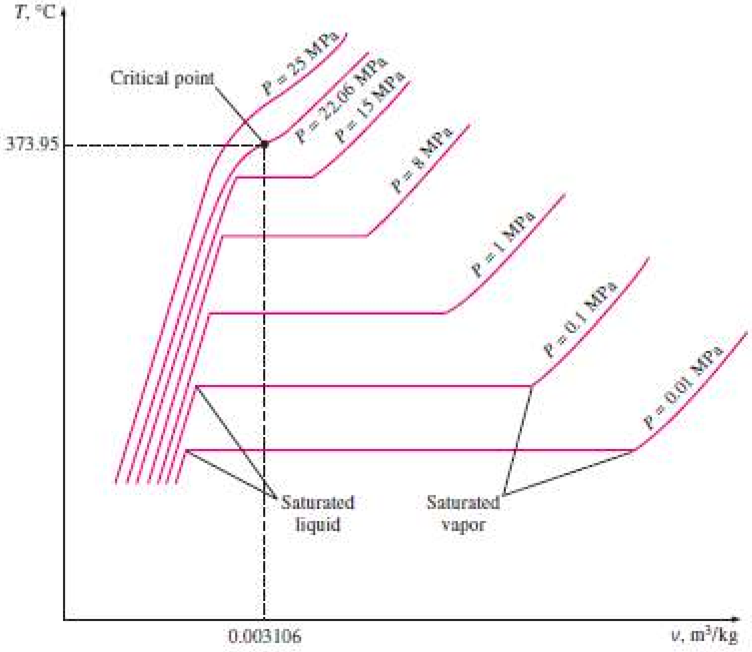


FIGURE 3–16: *T*-*v* diagram of constant-pressure phase-change processes of a pure substance at various pressures (numerical values arefor water).

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 *critical specific volume v*cr. The critical-point properties ofwater are *P*cr _ 22.06 MPa, *T*cr _ 373.95°C, and *v*cr _ 0.003106 m3/kg.For helium, they are 0.23 MPa, _267.85°C, and 0.01444 m3/kg. The criticalproperties for various substances are given in Table A–1 in the appendix.

At pressures above the critical pressure, there is not a distinct phasechange process (Fig. 3–17). Instead, the specific volume of the substance continually increases, and at all times there is only one phase present. Eventually, it resembles a vapor, but we can never tell when the change

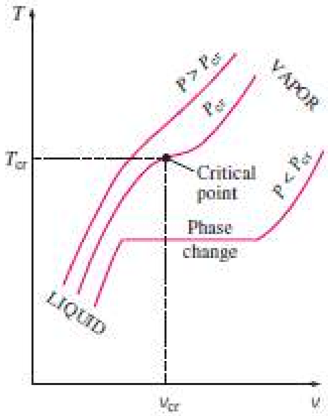


FIGURE 3–17: At supercritical pressures (*P* _ *P*cr),there is no distinct phase-change(boiling) process.

The saturated liquid states in Fig. 3–16 can be connected by a line calledthe **saturated liquid line,** and saturated vapor states in the same figure canbe connected by another line, called the **saturated vapor line.** These twolines meet at the critical point, forming a dome as shown in Fig. 3–18. Allthe compressed liquid states are located in the region to the left of the saturatedliquid line, called the **compressed liquid region.** All the superheatedvapor states are located to the right of the saturated vapor line, called the**superheated vapor region.** In these two regions, the substance exists in asingle phase, a liquid or a vapor. All the states that involve both phases inequilibrium are located under the dome, called the **saturated liquid–vapormixture region,** or the **wet region.**

The *P-v* Diagram:

The general shape of the *P-v* diagram of a pure substance is very much like the *T-v* diagram, but the *T* _ constant lines on this diagram have a downward trend, as shown in Fig. 3–19.Consider again a piston–cylinder device that contains liquid water at 1MPa and 150°C. Water at this state exists as a compressed liquid. Now theweights on top of the piston are removed one by one so that the pressureinside the cylinder decreases gradually (Fig. 3–20). The water is allowed toexchange heat with the surroundings so its temperature remains constant. Asthe pressure decreases, the volume of the water increases slightly. When thepressure reaches the saturation-pressure value at the specified temperature(0.4762 MPa), the water starts to boil.

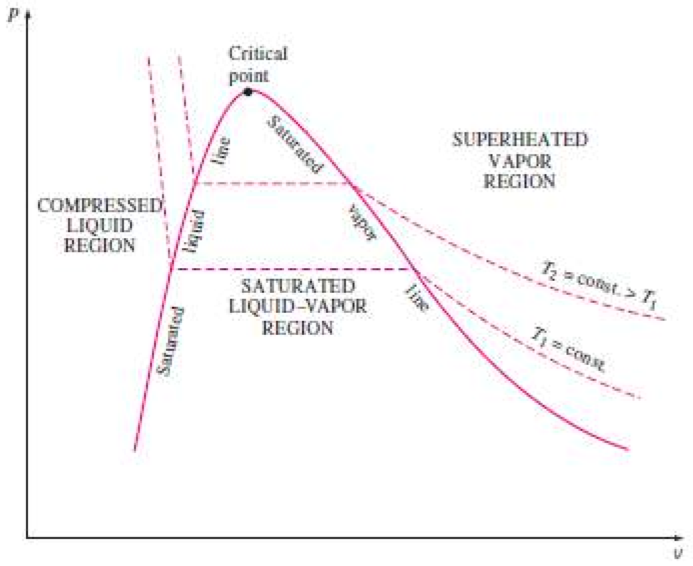


FIGURE 3–18: *T*-*v* diagram of a pure substance.

During this vaporization process,both the temperature and the pressure remain constant, but the specific volumeincreases. Once the last drop of liquid is vaporized, further reduction inpressure results in a further increase in specific volume. Notice that duringthe phase-change process, we did not remove any weights. Doing so wouldcause the pressure and therefore the temperature to drop [since *T*sat _*f* (*P*sat)], and the process would no longer be isothermal.When the process is repeated for other temperatures, similar paths areobtained for the phase-change processes. Connecting the saturated liquidand the saturated vapor states by a curve, we obtain the *P*-*v* diagram of apure substance, as shown in Fig. 3–19.

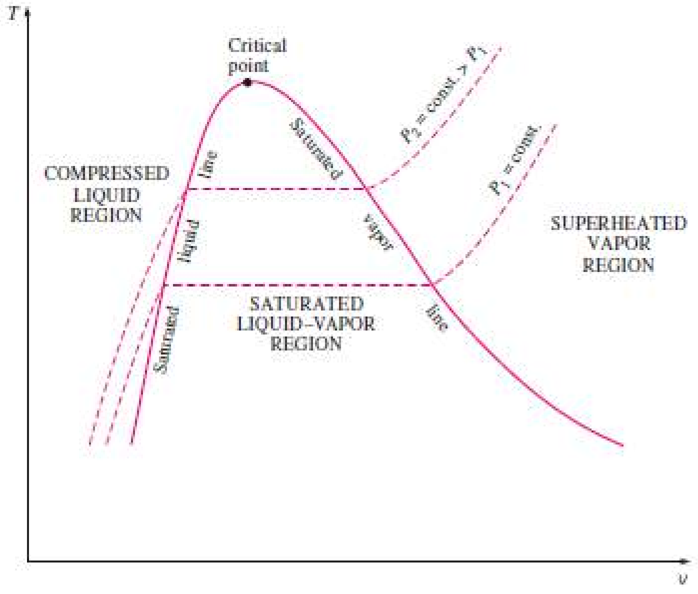


FIGURE 3–19: *P*-*v* diagram of a pure substance

The *P*-*T* Diagram:

Figure 3–25 shows the *P*-*T* diagram of a pure substance. This diagram isoften called the **phase diagram** since all three phases are separated fromeach other by three lines. The sublimation line separates the solid and vaporregions, the vaporization line separates the liquid and vapor regions, and themelting (or fusion) line separates the solid and liquid regions. These threelines meet at the triple point, where all three phases coexist in equilibrium.The vaporization line ends at the critical point because no distinction can bemade between liquid and vapor phases above the critical point. Substancesthat expand and contract on freezing differ only in the melting line on the*P*-*T* diagram.

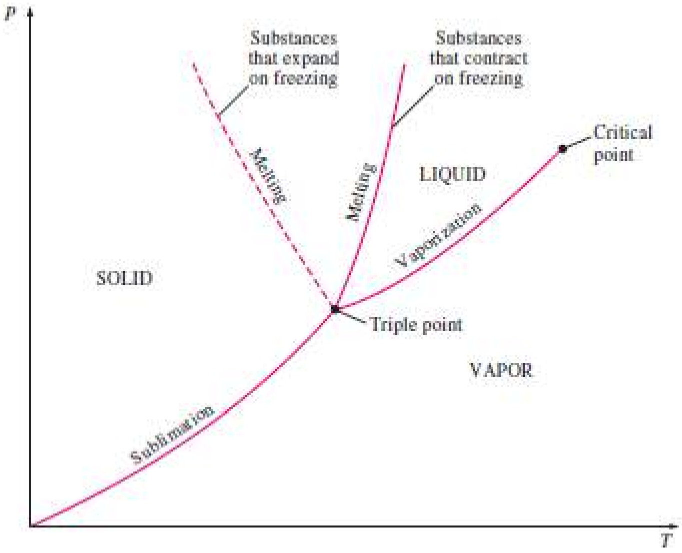


FIGURE 3–25: *P*-*T* diagram of pure substances.

THE *P-V-T* SURFACE

The state of a simple compressible substance is fixed by any two independent,intensive properties. Once the two appropriate properties are fixed, allthe other properties become dependent properties. Remembering that anyequation with two independent variables in the form z _ $z(x, y)$ represents asurface in space, we can represent the *P*-*v*-*T* behavior of a substance as asurface in space, as shown in Figs. 3–26 and 3–27. Here T and v may beviewed as the independent variables (the base) and P as the dependent variable(the height)

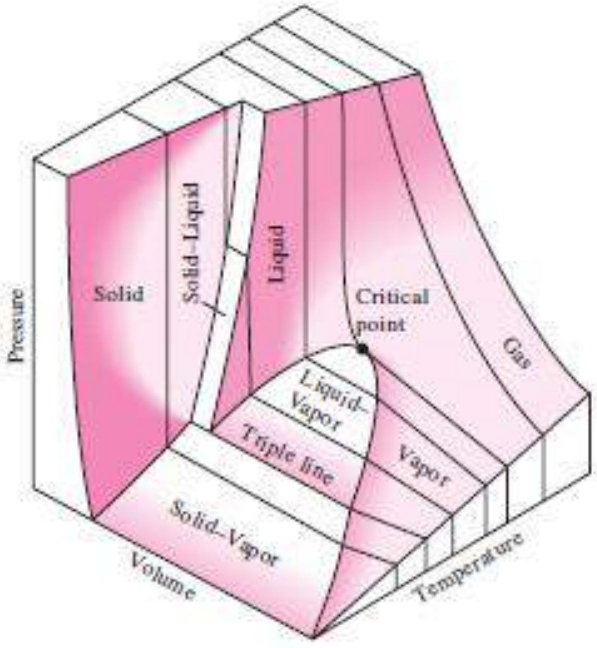


FIGURE 3–26: *P-v-T* surface of a substance that*contracts* on freezing

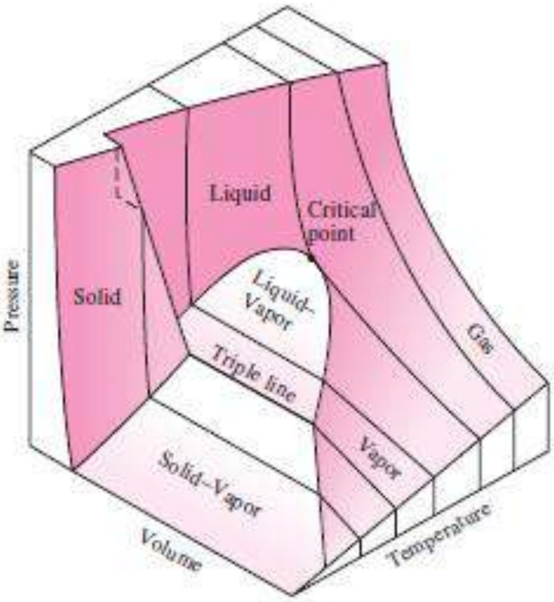


FIGURE 3–27*P-v-T* surface of a substance that*expands* on freezing (like water).

All the points on the surface represent equilibrium states. All states alongthe path of a quasi-equilibrium process lie on the *P-v-T* surface since such aprocess must pass through equilibrium states. The single-phase regionsappear as curved surfaces on the *P-v-T* surface, and the two-phase regionsas surfaces perpendicular to the *P-T* plane. This is expected since the projectionsof two-phase regions on the *P-T* plane are lines.

All the two-dimensional diagrams we have discussed so far are merely projectionsof this three-dimensional surface onto the appropriate planes. A *P-v*diagram is just a projection of the *P-v-T* surface on the P-v plane, and a T-$v$diagram is nothing more than the bird’s-eye view of this surface. The P-v-Tsurfaces present a great deal of information at once, but in a thermodynamicanalysis it is more convenient to work with two-dimensional diagrams, suchas the P-v and T-v diagrams.

OTHER EQUATIONS OF STATE:

The ideal-gas equation of state is very simple, but its range of applicabilityis limited. It is desirable to have equations of state that represent the P-v-Tbehavior of substances accurately over a larger region with no limitations.Such equations are naturally more complicated. Several equations have beenproposed for this purpose (Fig. 3–57), but we shall discuss only three: the *van der Waals* equation because it is one of the earliest, the *Beattie-Bridgeman*equation of state because it is one of the best known and is reasonablyaccurate, and the *Benedict-Webb-Rubin* equation because it is one of themore recent and is very accurate.

Van der Waals Equation of State

The van der Waals equation of state was proposed in 1873, and it has twoconstants that are determined from the behavior of a substance at the criticalpoint. It is given by

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

Van der Waals intended to improve the ideal-gas equation of state byincluding two of the effects not considered in the ideal-gas model: the *intermolecularattraction forces* and the *volume occupied by the molecules themselves*.The term a/v2 accounts for the intermolecular forces, and b accountsfor the volume occupied by the gas molecules. In a room at atmosphericpressure and temperature, the volume actually occupied by molecules isonly about one-thousandth of the volume of the room. As the pressureincreases, the volume occupied by the molecules becomes an increasinglysignificant part of the total volume. Van der Waals proposed to correct thisby replacing v in the ideal-gas relation with the quantity v _ b, where $b$represents the volume occupied by the gas molecules per unit mass.

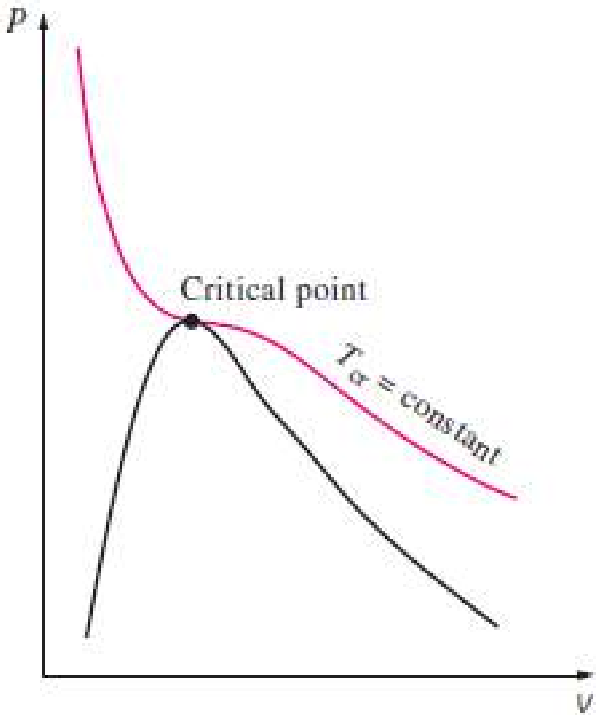
The determination of the two constants appearing in this equation is basedon the observation that the critical isotherm on a P-v diagram has a horizontalinflection point at the critical point (Fig. 3–58). Thus, the first and thesecond derivatives of P with respect to v at the critical point must be zero.That is,

$$\left(\frac{\partial P}{\partial v}\right)_{T=T_{cr}=\text{const}} = 0 \quad \text{and} \quad \left(\frac{\partial^2 P}{\partial v^2}\right)_{T=T_{cr}=\text{const}} = 0$$

By performing the differentiations and eliminating *v*cr, the constants *a* and *b*are determined to be

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

The constants *a* and*b* can be determined for any substance from the criticalpointdata alone (Table A–1).The accuracy of the van der Waals equation of state is often inadequate,but it can be improved by using values of *a* and*b* that are based on theactual behavior of the gas over a wider range instead of a single point.Despite its limitations, the van der Waals equation of state has a historicalvalue in that it was one of the first attempts to model the behavior of realgases. The van der Waals equation of state can also be expressed on a unitmolebasis by replacing the *v* in Eq. 3–22 by and the *R* in Eqs. 3–22 and3–23 by *Ru*.



Beattie-Bridgeman Equation of State:

The Beattie-Bridgeman equation, proposed in 1928, is an equation of statebased on five experimentally determined constants. It is expressed as

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

$$A = A_0\left(1 - \frac{a}{\overline{v}}\right) \quad \text{and} \quad B = B_0\left(1 - \frac{b}{\overline{v}}\right)$$

The constants appearing in the above equation are given in Table 3–4 forvarious substances. The Beattie-Bridgeman equation is known to be reasonablyaccurate for densities up to about 0.8rcr, where rcr is the density of thesubstance at the critical point.

Benedict-Webb-Rubin Equation of State:

Benedict, Webb, and Rubin extended the Beattie-Bridgeman equation in1940 by raising the number of constants to eight. It is expressed as

$$P = \frac{R_uT}{\overline{v}} + \left(B_0R_uT - A_0 - \frac{C_0}{T^2}\right)\frac{1}{\overline{v}^2} + \frac{bR_uT - a}{\overline{v}^3} + \frac{a\alpha}{\overline{v}^6} + \frac{c}{\overline{v}^3T^2}\left(1 + \frac{\gamma}{\overline{v}^2}\right)e^{-\gamma/\overline{v}^2}$$

The values of the constants appearing in this equation are given inTable 3–4. This equation can handle substances at densities up to about2.5rcr. In 1962, Strobridge further extended this equation by raising thenumber of constants to 16

VIRIAL EQUATION OF STATE

The equation of state of a substance can also be expressed in a series formas

$$P = \frac{RT}{v} + \frac{a(T)}{v^2} + \frac{b(T)}{v^3} + \frac{c(T)}{v^4} + \frac{d(T)}{v^5} + \ldots$$

GASLAWS

Ideal and Real Gases

Pure Substance: A pure substance is one that has a homogeneous and invariable chemical composition. It may exist in more than one phase but chemical composition is the same in all phases.

Some times the mixture of gases, such as air is considered a pure substance as long as there is no change of phase. Further our emphasis will be on simple compressible substances

Early experiments on the **variables of state** (such as T, P, V, and n) showed that only two of these variables of state need to be known to know the state of a sample of matter.

Extensive variables: depend on the amount of substance present. Examples include the volume, energy, enthalpy, and heat capacity.

Intensive variables: do not depend on the amount of substance present. Examples include the temperature and pressure.

Equations of State

An equation of state is an equation which relates the variables of state (T, P, V, and n). It's particularly useful when you want to know the effect of a change in one of the variables of state

Solids and Liquids: If the pressure on a solid or liquid is increased, the volume does not change much. If the temperature is increased, the volume doesn't change much either. Therefore, an appropriate equation of state describing such systems would be: $V(T,P) = \text{constant}$.

Gases: In contrast, changing the pressure or temperature of a gas will have an easily observable effect on the volume of that gas. For an **ideal gas** (no intermolecular interactions and no molecular volume) n appropriate equation of state would be: $V(T,P,n) = (nRT)/P$.

There are many equations of state describing **real gases**. These equations take in consideration molecular volume and interactions. The most well-known such equations is probably the Van der Waals equation.

Ideal and real gases

An ideal gas is one which follows the ideal gas equation of state, namely

$$PV = (m/M)\,(MR)\,T = n\,R_u\,T$$

The universal gas constant has a value of 8.314 J/mol K or kJ/kmol K and is related to the specific gas constant by the relation $R_u = (R/M)$

The ideal gas equation of state can be derived from the kinetic theory of gases where the following assumptions are made:

The molecules are independent of each other. In other words, there are no attractive forces between the molecules. The molecules do not occupy any volume. That is the volume occupied by the molecules is quite negligible compared to the volume available for motion of the molecules. The internal energy of an ideal gas is a function of temperature only and is independent of pressure and volume. That is,

$$u = u(T)$$

$$(\partial u/\partial P)_T = 0, \quad (\partial u/\partial v)_T = 0$$

Enthalpy and specific heat

$$h = u + Pv$$

For an ideal gas u = u(T) only and PV = mRT and hence h = h(T) only. The specific heat at constant volume is defined as the amount of energy transferred as heat at constant volume, per unit mass of a system to raise its temperature by one degree. That is,

$$Cv = (dq/dT)v$$

The specific heat at constant pressure is defined as the energy transferred as heat at constant pressure, per unit mass of a substance to raise its temperature by one degree. That is Cp = (dq/dT)P

Real gases:

The ideal gas law is only an approximation to the actual behavior of gases.At high densities, that is at high pressures and low temperatures, the behavior of actual or real gases deviate from that predicted by the ideal gas law. In general, at sufficiently low pressures or at low densities all gases behave like ideal gases.

Compressibility Factor:

The deviation from ideal behavior of a gas is expressed in terms of the compressibility factor Z, which is defined as the ratio of the actual volume to the volume predicted by the ideal gas law.

Z = Actual volume/volume predicted by ideal gas law = v/RT/P = Pv/RT

For an ideal gas Pv = RT and hence Z = 1 at all temperatures and pressures.

The experimental P-v-T data is used to prepare the compressibility chart.

Reduced pressure, PR = P/Pc,

Reduced temperature, TR = T/Tc

Reduced volume, vR = v/vc

Where Pc, Tc and vc denote the critical pressure, temperature and volume respectively. These equations state that the reduced property for a given state is the value of this property in this state divided by the value of this same property by at the critical point.The striking fact is that when such Z versus Pr diagrams are prepared for a number of different substances, all of them very nearly coincide, especially when the substances have simple, essentially spherical molecules.

We need to know only critical temperature and critical pressure to use this basic generalized chart. In general it can be noted that idealized gas behavior for very low pressures as compared to critical) regardless of temperature. Furthermore, at high temperatures (greater than twice Tc), the ideal-gas model can be assumed to good accuracy to pressures as high as 4-5 times Pc.

UNIT- IV

MIXTURE OF PERFECT GASES

The properties of a gas mixture obviously depend on the properties of the individual gases (called components or constituents) as well as on the amount of each gas in the mixture.

COMPOSITION OF A GAS MIXTURE MASS AND MOLE FRACTIONS To determine the properties of a mixture, we need to know the composition of the mixture as well as the properties of the individual components. There are two ways to describe the composition of a mixture: either by specifying the number of moles of each component, called molar analysis, or by specifying the mass of each component, called gravimetric analysis.

MIXTURE OF IDEAL GASES

Basic assumption is that the gases in the mixture do not interact with each other. Consider a mixture with components *l = 1,2,3*... with masses m1, m2, m3 ...m*i* andwith number of moles.
The total mixture occupies a volume V, has a total pressure P and temperature T (which is also the temperature of each of the component species)

The total mass

$$m = \sum m_i$$

Total number of mole N

$$N = \sum N_i$$

Mass fraction of species i

$$\phi_i = \frac{m_i}{m}$$

Mole fraction of species i

$$Y_i = \frac{N_i}{N}$$

The mass and number of moles of species i are related by

$$m_i = N_i M_i$$

N_i is the number of moles of species i and M_i is the molar mass of species *i*
Also to be noted

$$\sum \phi_i = 1 \quad \text{and} \quad \sum Y_i = 1$$

We can also define a molar mass of the mixture as

$$M = \frac{m}{N} = \frac{\sum N_i M_i}{N}$$

or,

$$M = \frac{\sum N_i}{N} M_i$$

or,

$$M = \sum Y_i M_i$$

Dalton 's Law of partial pressure

Total pressure of an ideal gas mixture is equal to the sum of the partial pressures of the constituent components, That is

$$P = \sum P_i$$

P is the total pressure of the mixture
P_i is the partial pressure of species *i*
= pressure of the species if it existed alone in the given **temperature T** and **volume V**

$$P_i = \frac{N_i \hat{R} T}{V}$$

$\hat{R}$ is the universal gas constant = 8.314 kJ/k mol K

Dalton 's Law

$$P = \sum P_i = \frac{\sum N_i \hat{R} T}{V} = \frac{RT}{V} \sum N_i$$

or

$$P = \frac{N \hat{R} T}{V}$$

The pressure function of species *i*

$$\frac{P_i}{P} = \frac{N_i \hat{R} T}{V} \frac{V}{N \hat{R} T} = \frac{N_i}{N} = Y_i$$

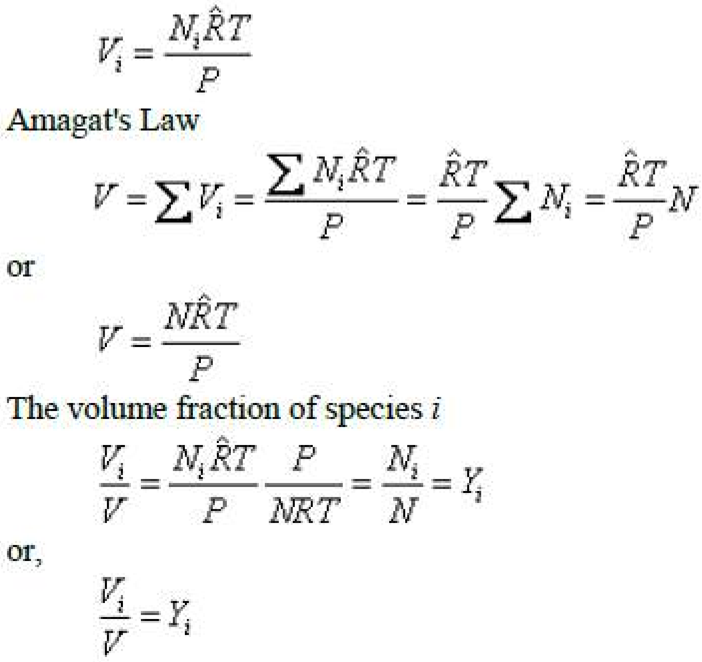
Therefore,
Pressure fraction = Mole fraction

Amagat's Law:

Volume of an ideal gas mixture is equal to the sum of the partial volumes

$$V = \sum V_i$$

V = total volume of the mixture
V_i = partial volume of the species *i*
= volume of the species if it existed alone in the given temperature T and pressure P
For an ideal gas



$$V_i = \frac{N_i \hat{R} T}{P}$$

Amagat's Law

$$V = \sum V_i = \frac{\sum N_i \hat{R} T}{P} = \frac{\hat{R} T}{P} \sum N_i = \frac{\hat{R} T}{P} N$$

or

$$V = \frac{N \hat{R} T}{P}$$

The volume fraction of species *i*

$$\frac{V_i}{V} = \frac{N_i \hat{R} T}{P} \frac{P}{NRT} = \frac{N_i}{N} = Y_i$$

or,

$$\frac{V_i}{V} = Y_i$$

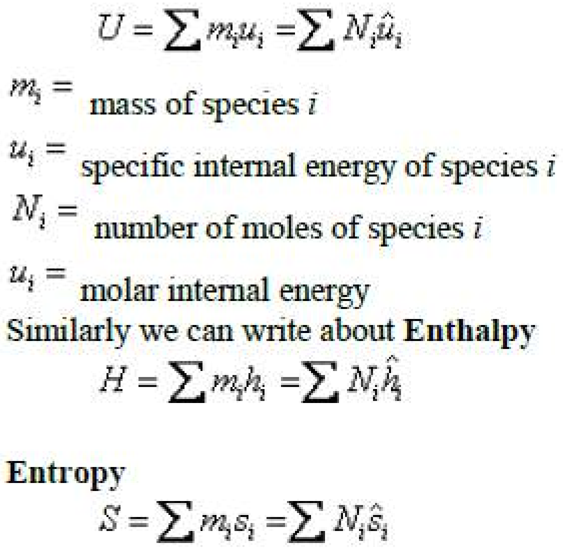
Volume fraction = Mole fraction

Mass based analysis is known as gravimetric analysis

Mole based analysis is known as molar analysis

Therdynamic Properties of Mixtures

Internal Energy



$$U = \sum m_i u_i = \sum N_i \hat{u}_i$$

$m_i =$ mass of species *i*

$u_i =$ specific internal energy of species *i*

$N_i =$ number of moles of species *i*

$u_i =$ molar internal energy

Similarly we can write about **Enthalpy**

$$H = \sum m_i h_i = \sum N_i \hat{h}_i$$

Entropy

$$S = \sum m_i s_i = \sum N_i \hat{s}_i$$

Entropy

$$S = \sum m_i s_i = \sum N_i \hat{s}_i$$

Specific internal energy of the mixture

$$u = \frac{U}{M} = \frac{\sum m_i u_i}{M} = \frac{\sum m_i}{M} u_i$$

or

$$u = \sum \phi_i u_i$$

$\phi_i =$ mass fraction of species *i*

$u_i =$ specific internal energy of species *i*

Molar internal energy of mixture

$$\hat{u}_i = \frac{U}{N} = \frac{\sum N_i \hat{u}_i}{N} = \frac{\sum N_i}{N} \hat{u}_i = \sum Y_i \hat{u}_i$$

$Y_i =$ mole fraction of species *i*

$u_i =$ molar internal energy of species *i*

Similarly we can write for specific enthalpy and molar enthalpy

$$h = \sum \phi_i h_i \quad \text{and} \quad \hat{h} \sum Y_i \hat{h}_i$$

We can also write for specific entropy and molar entropy

$$s = \sum \phi_i s_i \quad \text{and} \quad \hat{s} \sum Y_i \hat{s}_i$$

Change in u, h and s

$$\Delta u = \sum \phi_i \Delta u_i = \sum (\phi_i u_i)_2 - \sum (\phi_i u_i)_1$$
$$= \sum \phi_i (u_{i2} - u_{i1})$$

Assume ϕ_i remains unchanged

$$\Delta u = \sum \phi_i c_{v_i} \Delta T$$

For ideal gas mixture

$$\Delta u = \Delta T \sum \phi_i c_{v_i}$$

as ΔT is the same for all species as a result of thermodynamics equilibrium

or

$$\Delta u = c_v \Delta T$$

$$\Delta u = c_v \Delta T$$

Where $C_v = \sum \phi_i c_{v_i} =$ definition of mixture c_v

Similarly it can be shown that

$$\Delta \hat{u} = \hat{c}_v \Delta T$$

Where

$$\hat{C}_v = \sum Y_i \hat{c}_{v_i} =$$ mixture $\hat{c}_v$ (molar basis)

Let us also recall that $\sum \phi_i = 1$ and $\sum Y_i = 1$

We can also write similar relations for mixture enthalpy

$$\Delta h = c_p \Delta T$$

Where

$$c_p = \sum \phi_i c_{p_i}$$ definition of mixture c_p

and

$$\Delta \hat{h} = \hat{c}_p \Delta T$$

Where

$$\hat{c}_p = \sum Y_i \, c_{p_i}$$ definition of mixture c_p (molar basis)

Therefore, $\Delta u = c_v \Delta T$ and $\Delta h = c_p \Delta T$. The equations are similar to the equations for a single (ideal gas) species.

Properties of Atmospheric Air:

Dry air is a mechanical mixture of the following gases: Oxygen, nitrogen, carbon dioxide, hydrogen, argon, neon, krypton, helium, ozone, and xenon. Dry air is considered to consist of 21% oxygen and 79% nitrogen **by volume**. It consists of 23% oxygen, and 77% nitrogen **by mass**. **Completely dry air does not exist in nature**. Water vapour in varying amount is diffused through it. If **Pa** and **Pw**are the **partial pressures** of **dry air** and **water vapour**respectively, then by Dalton's law of **partial pressure**

$$Pa + Pw = P$$

Where P is the atmospheric pressure

Mole – fraction of dry air,

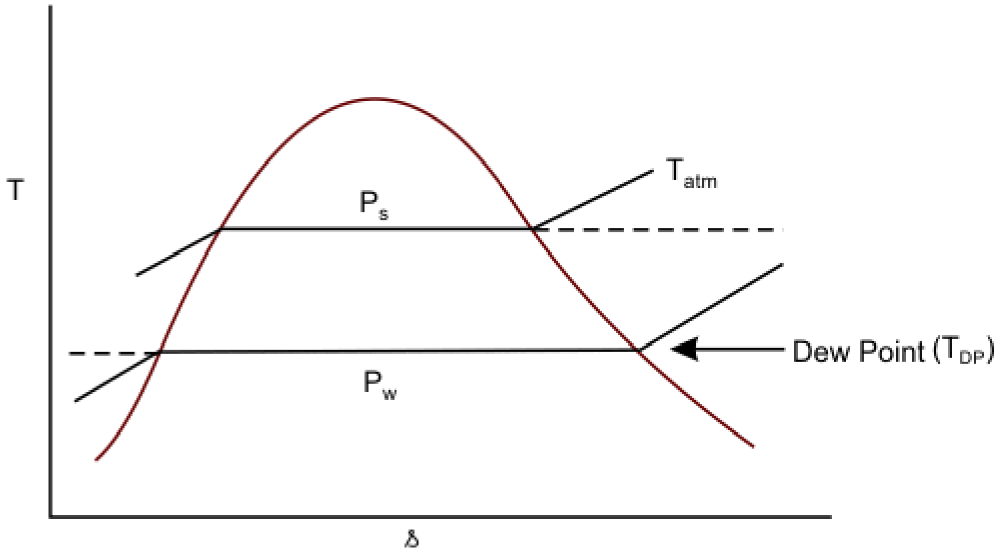
$$Y_a = \frac{P_a}{P} = P_a$$

P_m is considered to be 1 atm

Mole – fraction of water vapour,

$$Y_w = \frac{P_w}{P} = P_w$$

Since P_w is very small, the saturation temperature of water vapour at P_w is less than the atmospheric temperature . So the water vapour in air exists in the superheated state, and the air is said to be unsaturated. If the air- water vapour mixture which is initially not saturated, is cooled at constant pressure, the partial pressure of water vapour in the mixture remains constant till it is equal to the saturation pressure of water. Further cooling result in condensation of water vapour. The temperature at which the vapour condenses when the air-water vapour mixture is cooled at constant pressure, is called Dew Point



Relative Humidity (RH)

Relative humidity is defined as the ratio of partial pressure of water vapour, P_w , in a mixture to the saturation pressure,of P_s pure water at the temperature of the mixture

$$RH = \frac{P_w}{P_s} = \phi$$

If water is injected into unsaturated air in a container, water will evaporate, which will increase the moisture content of the air. and P_w will increase. This will continue till the air becomes saturated at that temperature and there will be no more evaporation of water. For saturated air,

$$P_w V = m_w R_{H_2O} T = N_W \bar{R} T$$

$$P_s V = m_s R_{H_2O} T = N_s \bar{R} T$$

relative humidity is 100%. Assuming water vapour as an ideal gas,

$$\boxed{RH = \frac{P_w}{P_s} = \frac{m_w}{m_s}} = \phi$$

RH= mass of water vapour in a given volume of air at temperature T / mass of water vapour when the same volume of air is saturated at temperature T

$$= \frac{N_w}{N_s} = \frac{Y_w}{Y_s}$$

Specific Humidity or Humidity Ratio

SH is defined as the **mass of water vapour**per unit **mass of dry air** in a mixture of air and water vapour.

If m_a = mass of dry air, m_w = mass of water vapour

$$SH=\frac{m_w}{m_a}=w$$

Also we know that

$$P_w=m_wR_wT \quad \text{and} \quad P_aV=m_aR_aT$$

$$SH=\frac{m_w}{m_a}=\frac{P_w}{P_a}\cdot\frac{R_a}{R_w}=w$$

or,

$$SH=\frac{P_w}{P_a}\cdot\frac{8.3143/28.96}{8.3143/18}=0.622\frac{P_w}{P_a}=w$$

or

$$\boxed{SH=0.622\frac{P_w}{(P-P_w)}=w}$$

Where P is the **atmospheric pressure**

Relative humidity,

$$RH=\phi=\frac{P_w}{P_s}$$

$$w=0.622\frac{P_w}{P_s}\cdot\frac{P_s}{P_a}=0.622\phi\frac{P_s}{P_a}$$

or

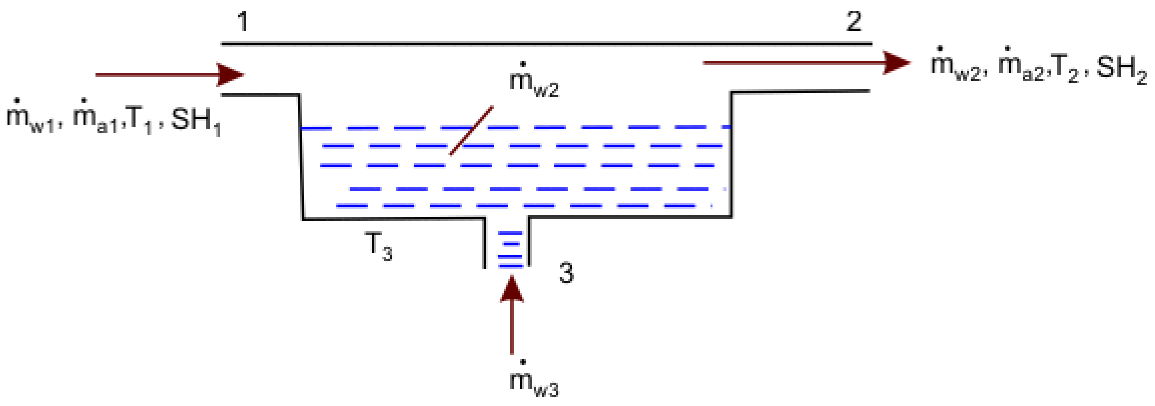
$$w=0.622\phi\frac{P_s}{P-P_w}=0.622\phi\frac{P_s}{P-\phi P_s}$$

Or

If a mixture of air and superheated (or unsaturated) water vapour is cooled at constant pressure, the partial pressure of each constituent remains constant until the water vapour reaches its saturated state. Further cooling causes condensation. The temperature at which water vapour starts condensing is called the dew point temperature T_{DP} of the mixture. it is equal to the saturation temperature at the partial pressure, P_w of the water vapour in mixture.

ADIABATIC SATURATION

Specific humidity or the relative humidity of an air – water vapour mixture can be measured in principle with the help of a device called the adiabatic saturator



The air – water vapour mixture flows steadily into the device. The SH or w of the incoming mixture has to be determined.

The air – water vapour mixture leaves the adiabatic saturator as saturated mixture. Let the device be insulated so that there is no energy loss.

Since the unsaturated air – water vapour mixture is sweeping over a layer of liquid water, some water evaporates. The energy needed for the evaporation comes from the air mixture. Hence, the air – water vapour mixture leaves the adiabatic saturator at a temperature lower than that of the entering air. As the air leaving the adiabatic saturator is in equilibrium with the liquid water, the temperature of the liquid water is equal to the temperature of the saturated air – water vapour mixture.

Mass balance

$$\dot{m}_{a1} = \dot{m}_{a2}$$

for air Mass

$$\dot{m}_{w1} + \dot{m}_{w3} = \dot{m}_{w2}$$

balance for

water

Energy balance

$$\dot{m}_{a1}h_{a1} + \dot{m}_{w1}h_{w1} + \dot{m}_{w3}h_{w3} = \dot{m}_{a2}h_{a2} + \dot{m}_{w2}h_{w2}$$

Dividing (35.17) by $\dot{m}_{a1} \left(= \dot{m}_{a2}\right)$

$$\frac{\dot{m}_{w1}}{\dot{m}_{a1}} + \frac{\dot{m}_{w3}}{\dot{m}_{a1}} = \frac{\dot{m}_{w2}}{\dot{m}_{a2}}$$

$$SH_1 + \frac{\dot{m}_{w3}}{\dot{m}_{a1}} = SH_2$$

$$\dot{m}_{w3} = \dot{m}_{a1}\left(SH_2 - SH_1\right)$$

$$SH_1 = \frac{\left(h_{a2} - h_{a1}\right) + SH_2\left(h_{w2} - h_{w3}\right)}{\left(h_{w1} - h_{w3}\right)} = w_1$$

The quantity

$$\left(h_{a2} - h_{a1}\right)$$
$$= c_p\left(T_2 - T_1\right)$$

The adiabatic saturation temperature depends only on the conditions of the entering fluid.

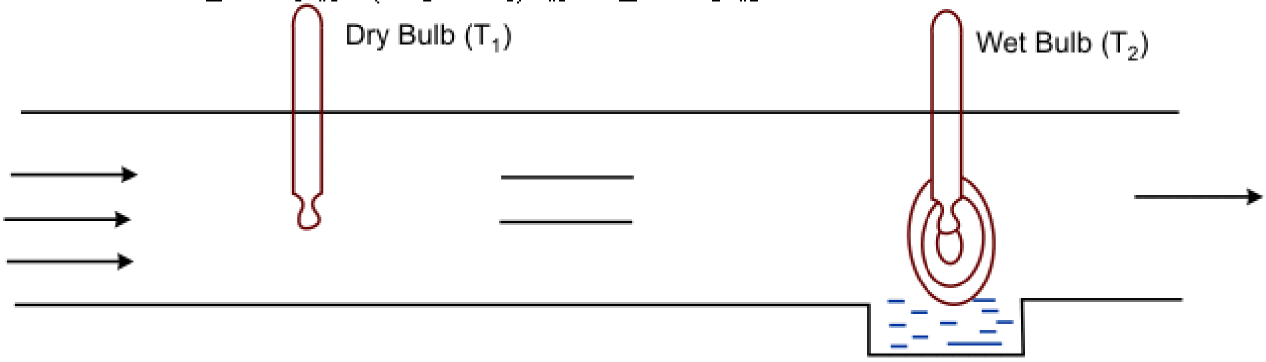
$$SH = w_1 = \frac{c_p\left(T_2 - T_1\right) + w_2 h_{fg2}}{h_{g1} - h_{f2}}$$

Finally

PSYCHROMETER:

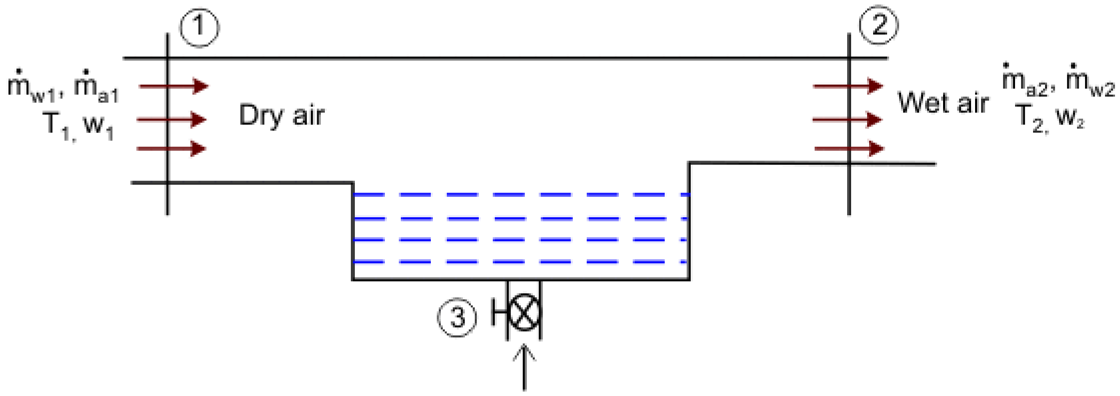
Psychrometer is an instrument to measure the wet-bulb and dry-bulb temperature of an air-water vapour mixture. This instrument uses the **principle of adiabatic saturation**. The specific humidity and relative humidity of air-water vapour mixture can be determined with knowledge of dry-bulb and wet-bulb temperatures figure 36.1

$$h_{a1} + SH_1 h_{w1} + \left(SH_2 - SH_1\right) h_{w3} = h_{a2} + SH_2 h_{w2}$$



- The dry-bulb temperature is the temperature of the incoming mixture.

- The wet-bulb temperature is the temperature of the saturated air-water vapour mixture.
- For air-water vapour mixtures, the wet- bulb temperature is found to be approximately (AST).
- The enthalpy of air -water vapour mixture is expressedas



We can write

$$h_{a1} + w_1 hw_1 + (w_2 - w_1)\, hw_3 = h_{a2} + w_2 hw_2$$

$$hw_3 \langle\langle hw_1, hw_2$$

In theaboveequation, isvapourenthalpy, is also enthalpy of vapour.Thetermis

meant for enthalpy of liquidwater

Also $(w_2 - w_1)\langle w_1, w_2$

For the above reasons, we can rewrite (36.2) as invoking (36.1), the above equation

$$h_1 + w_1 hw_1 = ha_2 + w_2 hw_2$$

becomesthe conserved property in an adiabatic saturation process. (Remember that state 2

$$h_1^* = h_2^*$$

h^* need no

$$h^* = h_a + whw$$

Where $ha = c_p T$ and

$$hw = h_g(T)$$

$$(\sim 10^0 C),$$

If ΔT is small

$$h^* = c_p T + wh_g$$

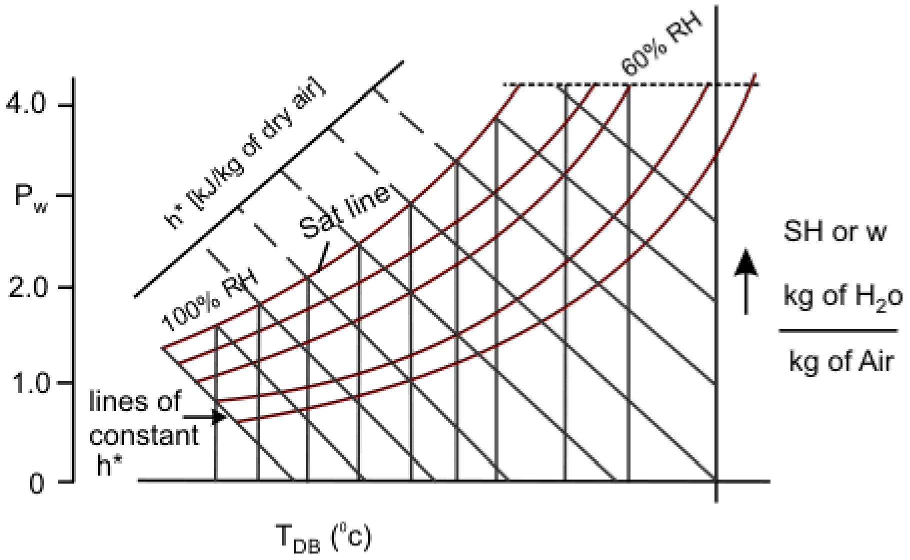
$h_g =$ constant and we get

$$h_1^* = h_2^*$$

A long an adiabatic miniaturization process will remain constant or

$$c_p(T_2 - T_1) + (w_2 - w_1)h_g = 0$$

PSYCHROMETRIC CHART



Abscissa is the dry bulb temperature. The right hand side ordinate provides humidity ratio. The equation shows a direct relationship between w and P_w

$$w = 0.622\frac{Pw}{P - Pw}$$

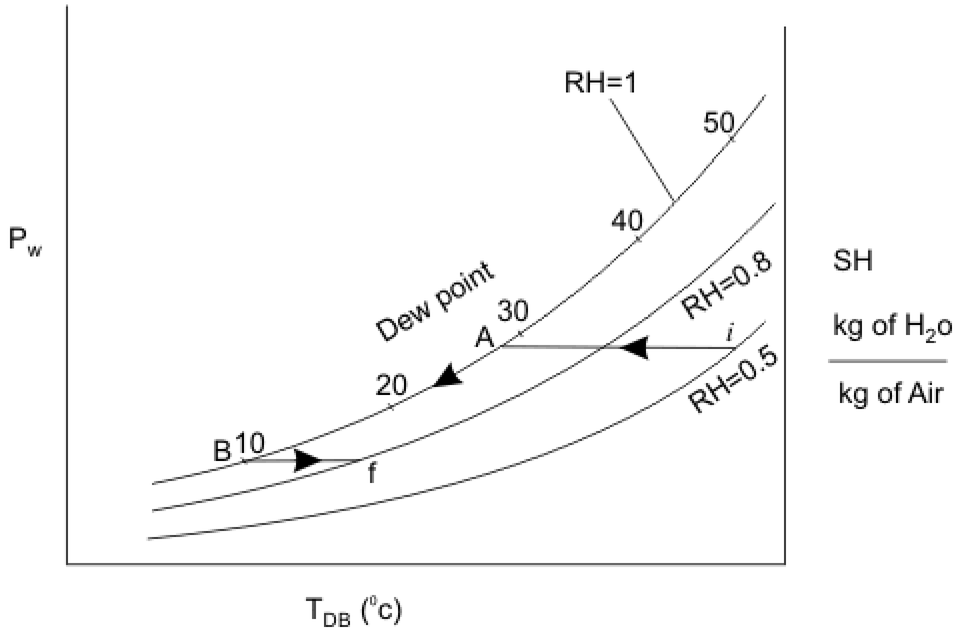
According, the vapour pressure can also be shown as the ordinate . The curves of constant relative humidity are also drawn on the Psychometric Chart. On figure 36.3, the curves are labeled as Φ = 100%, 60% etc. Psychometric Charts also gives values of the mixture enthalpy per unit mass of dry air in the mixture. The constant wet bulb temperature T_{WB} lines run from the upper left to lower right of the chart. The relationship between the wet bulb temperature and other chart quantities are provided by using Eqn

$$w_1 = \frac{c_p\left(T_{WB} - T\right)_1 + w_2 h_{fg2}}{h_{g1} - h_{f2}}$$

The lines of wet bulb temperature are approximately the lines of constant mixture enthalpy per unit mass of dry air.

Dehumidification:

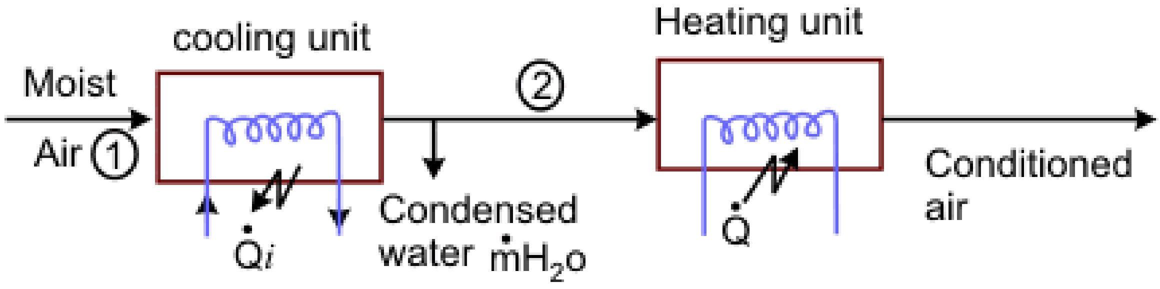
When a moist air steam is cooled at constant mixture pressure to a temperature below its dew point temperature, some condensation of water vapour would occur. Refer to fugure 36.4 for understanding the process.



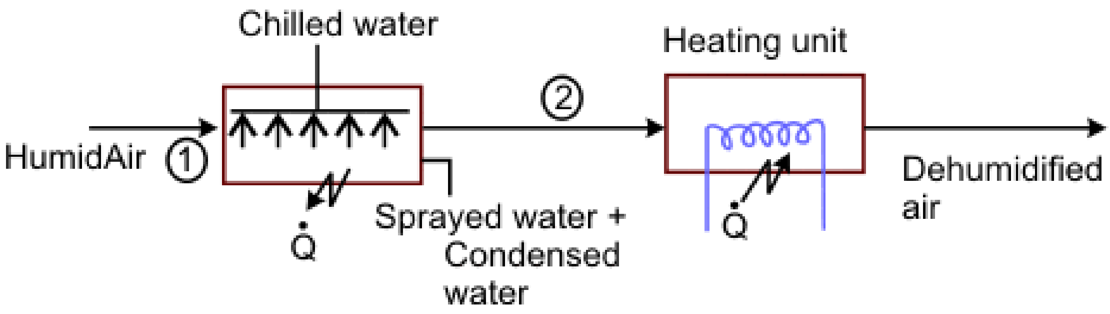
Dehumidification of air-water vapour mixture can be achieved by cooling the mixture below its dew point temperature (path i-A-B) allowing some water to condense, and then heating the mixture (path B-f) to the desired temperature.

Dehumidification and Cooling

For cooling the mixture, the mixture can be made to pass over the cooling coils through which a cold refrigerant is circulatedthe mixture) is sprayed into the air to be dehumidified. Then the air leaves with less humidity at the temperature of the chilled water. Next the air is heated to the desired temperature

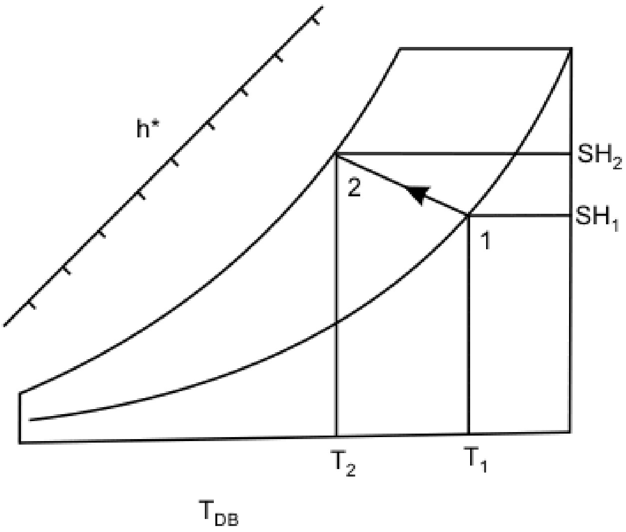


the mixture) is sprayed into the air to be dehumidified. Then the air leaves with less humidity at the temperature of the chilled water. Next the air is heated to the desired temperature

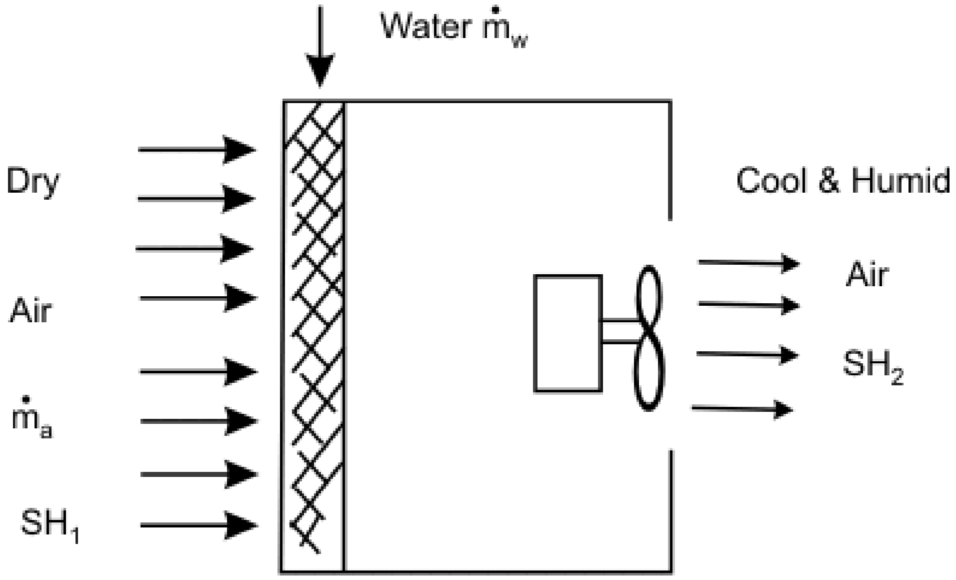


These two cooling and heating processes constitute an air conditioning plant. Humidification with Cooling

The process is same as the Adiabatic saturation except that the air may leaves unsaturated



Extensively used in desert coolers which is used for cooling homes in not & dry climates. An unsaturated air-water vapour mixture is made to flow through porous pads soaked in water (figure 36.8).


Water $\dot{m}_w$
Dry
Air
$\dot{m}_a$
SH_1
Cool & Humid
Air
SH_2

UNIT-V

AIR CONDITIONING SYSTEM

Otto cycle The Otto cycle is the ideal cycle for spark-ignition reciprocating engines. Itis named after Nikolaus A. Otto, who built a successful four-stroke enginein 1876 in Germany using the cycle proposed by Frenchman Beau deRochas in 1862. In most spark-ignition engines, the piston executes fourcomplete strokes (two mechanical cycles) within the cylinder, and thecrankshaft completes two revolutions for each thermodynamic cycle. Theseengines are called four-stroke internal combustion engines. A schematic ofeach stroke as well as a *P*-*v* diagram for an actual four-stroke spark-ignitionengine is given in Fig. 9–13(*a*).

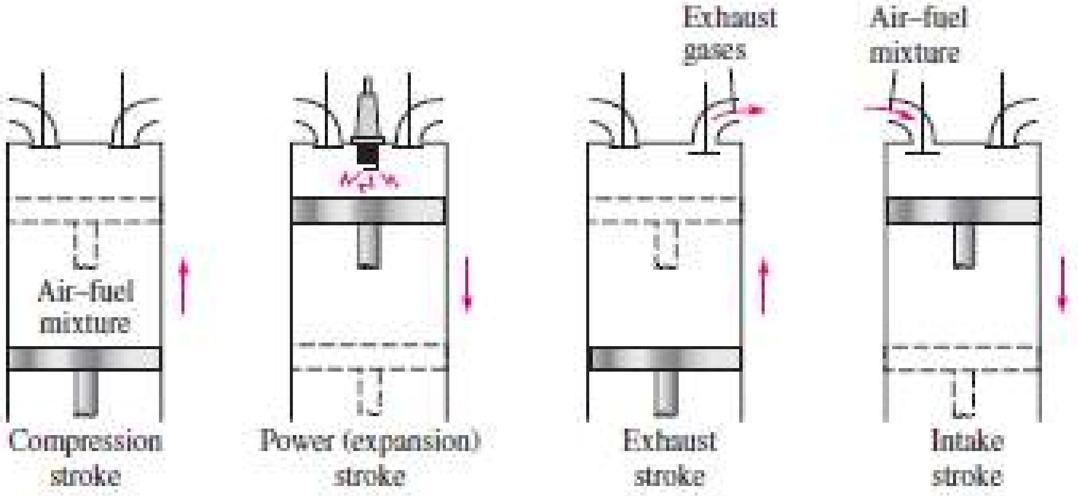


Fig: 5.1. (*a*) Actual four-stroke spark-ignition engine

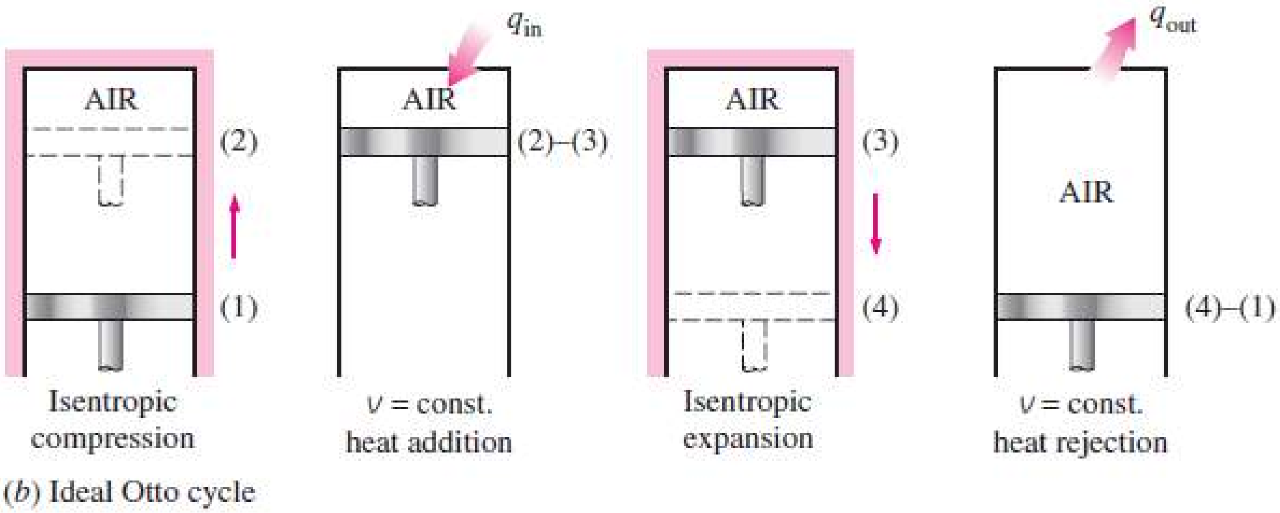


Fig: 5.2 (*b*) Ideal Otto cycle.

Initially, both the intake and the exhaust valves are closed, and the piston isat its lowest position (BDC). During the *compression stroke,* the piston movesupward, compressing the air–fuel mixture. Shortly before the piston reaches its highest position (TDC), the spark plug fires and the mixture ignite,increasing the pressure and temperature of the system. The high-pressuregases force the piston down, which in turn forces the crankshaft to rotate,producing a useful work output during the *expansion* or *power stroke.* At theend of this stroke, the piston is at its lowest position (the completion of thefirst mechanical cycle), and the cylinder is filled with combustion products.Now the piston moves upward one more time, purging the exhaust gasesthrough the exhaust valve (the *exhaust stroke*), and down a second time,drawing in fresh air–fuel mixture through the intake valve (the *intakestroke*). Notice that the pressure in the cylinder is slightly above the atmosphericvalue during the exhaust stroke and slightly below during the intakestroke.

In **two-stroke engines,** all four functions described above are executed injust two strokes: the power stroke and the compression stroke. In theseengines, the crankcase is sealed, and the outward motion of the piston isused to slightly pressurize the air–fuel mixture in the crankcase, as shown inFig. 9–14. Also, the intake and exhaust valves are replaced by openings inthe lower portion of the cylinder wall. During the latter part of the powerstroke, the piston uncovers first the exhaust port, allowing the exhaust gasesto be partially expelled, and then the intake port, allowing the fresh air–fuelmixture to rush in and drive most of the remaining exhaust gases out of thecylinder. This mixture is then compressed as the piston moves upward duringthe compression stroke and is subsequently ignited by a spark plug.

Advances in several technologies—such as direct fuel injection, stratifiedcharge combustion, and electronic controls—brought about a renewed interestin two-stroke engines that can offer high performance and fuel economywhile satisfying the stringent emission requirements. For a given weight anddisplacement, a well-designed two-stroke engine can provide significantlymore power than its four-stroke counterpart because two-stroke engines producepower on every engine revolution instead of every other one. In the newtwo-stroke engines, the highly atomized fuel spray that is injected into thecombustion chamber toward the end of the compression stroke burns muchmore completely. The fuel is sprayed after the exhaust valve is closed, whichprevents unburned fuel from being ejected into the atmosphere. With stratifiedcombustion, the flame that is initiated by igniting a small amount of therich fuel–air mixture near the spark plug propagates through the combustionchamber filled with a much leaner mixture, and this results in much cleanercombustion. Also, the advances in electronics have made it possible to ensurethe optimum operation under varying engine load and speed conditions.S

Major car companies have research programs underway on two-strokeengines which are expected to make a comeback in the future.The thermodynamic analysis of the actual four-stroke or two-stroke cyclesdescribed is not a simple task. However, the analysis can be simplified significantlyif the air-standard assumptions are utilized. The resulting cycle,which closely resembles the actual operating conditions, is the ideal **Ottocycle.** It consists of four internally reversible processes:

1-2 Isentropic compression
2-3 Constant-volume heat addition
3-4 Isentropic expansion
4-1 Constant-volume heat rejection

The execution of the Otto cycle in a piston–cylinder device together witha *P*-*v* diagram is illustrated in Fig. 5–1. The *T*-*s* diagram of the Otto cycleis given in Fig. 5–2.

The Otto cycle is executed in a closed system, and disregarding thechanges in kinetic and potential energies, the energy balance for any of theprocesses is expressed, on a unit-mass basis, s

$$(q_{\text{in}} - q_{\text{out}}) + (w_{\text{in}} - w_{\text{out}}) = \Delta u \qquad (\text{kJ/kg}) \tag{9–5}$$

No work is involved during the two heat transfer processes since both take place at constant volume. Therefore, heat transfer to and from the working fluid can be expressed as

$$q_{\text{in}} = u_3 - u_2 = c_v(T_3 - T_2) \tag{9–6a}$$

and

$$q_{\text{out}} = u_4 - u_1 = c_v(T_4 - T_1) \tag{9–6b}$$

Then the thermal efficiency of the ideal Otto cycle under the cold air standard assumptions becomes

$$\eta_{\text{th,Otto}} = \frac{w_{\text{net}}}{q_{\text{in}}} = 1 - \frac{q_{\text{out}}}{q_{\text{in}}} = 1 - \frac{T_4 - T_1}{T_3 - T_2} = 1 - \frac{T_1(T_4/T_1 - 1)}{T_2(T_3/T_2 - 1)}$$

Processes 1-2 and 3-4 are isentropic, and $v_2 = v_3$ and $v_4 = v_1$. Thus,

$$\frac{T_1}{T_2} = \left(\frac{v_2}{v_1}\right)^{k-1} = \left(\frac{v_3}{v_4}\right)^{k-1} = \frac{T_4}{T_3} \tag{9–7}$$

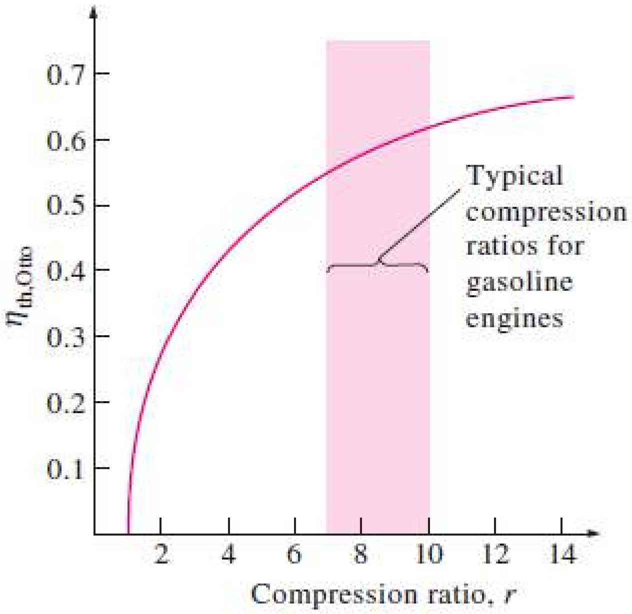
Substituting these equations into the thermal efficiency relation and simplifyinggive

$$\eta_{\text{th,Otto}} = 1 - \frac{1}{r^{k-1}}$$

$$r = \frac{V_{\max}}{V_{\min}} = \frac{V_1}{V_2} = \frac{v_1}{v_2}$$

is the **compression ratio** and k is the specific heat ratio *cp* /*cv*.

Equation 9–8 shows that under the cold-air-standard assumptions, thethermal efficiency of an ideal Otto cycle depends on the compression ratioof the engine and the specific heat ratio of the working fluid. The thermalefficiency of the ideal Otto cycle increases with both the compression ratioand the specific heat ratio. This is also true for actual spark-ignition internalcombustion engines. A plot of thermal efficiency versus the compressionratio is given in Fig. 9–17 for k _ 1.4, which is the specific heat ratio valueof air at room temperature. For a given compression ratio, the thermal efficiencyof an actual spark-ignition engine is less than that of an ideal Ottocycle because of the irreversibility, such as friction, and other factors suchas incomplete combustion.



The thermal efficiency curve is rathersteep at low compression ratios but flattens out starting with a compressionratio value of about 8. Therefore, the increase in thermal efficiency with the compression ratio is not as pronounced at high compression ratios. Also,when high compression ratios are used, the temperature of the air–fuel mixturerises above the autoignition temperature of the fuel (the temperature atwhich the fuel ignites without the help of a spark) during the combustionprocess, causing an early and rapid burn of the fuel at some point or pointsahead of the flame front, followed by almost instantaneous inflammation ofthe end gas. This premature ignition of the fuel, called **autoignition,** producesan audible noise, which is called **engine knock.** Autoignition inspark-ignition engines cannot be tolerated because it hurts performance andcan cause engine damage. The requirement that autoignition not be allowedplaces an upper limit on the compression ratios that can be used in spark ignitioninternal combustion engines.

Improvement of the thermal efficiency of gasoline engines by utilizinghigher compression ratios (up to about 12) without facing the auto ignitionproblem has been made possible by using gasoline blends that have goodantiknock characteristics, such as gasoline mixed with tetraethyl lead.

Tetraethyl lead had been added to gasoline since the 1920s because it is aninexpensive method of raising the *octane rating,* which is a measure of theengine knock resistance of a fuel. Leaded gasoline, however, has a veryundesirable side effect: it forms compounds during the combustion processthat are hazardous to health and pollute the environment. In an effort tocombat air pollution, the government adopted a policy in the mid-1970s thatresulted in the eventual phase-out of leaded gasoline. Unable to use lead, therefiners developed other techniques to improve the antiknock characteristicsof gasoline. Most cars made since 1975 have been designed to use unleadedgasoline, and the compression ratios had to be lowered to avoid engineknock. The ready availability of high octane fuels made it possible to raisethe compression ratios again in recent years. Also, owing to the improvementsin other areas (reduction in overall automobile weight, improvedaerodynamic design, etc.), today’s cars have better fuel economy and consequentlyget more miles per gallon of fuel. This is an example of how engineeringdecisions involve compromises, and efficiency is only one of theconsiderations in final design.

The second parameter affecting the thermal efficiency of an ideal Ottocycle is the specific heat ratio *k*. For a given compression ratio, an idealOtto cycle using a monatomic gas (such as argon or helium, *k* _ 1.667) asthe working fluid will have the highest thermal efficiency. The specific heatratio *k*, and thus the thermal efficiency of the ideal Otto cycle, decreases asthe molecules of the working fluid get larger (Fig. 9–2). At room temperatureit is 1.4 for air, 1.3 for carbon dioxide, and 1.2 for ethane. The workingfluid in actual engines contains larger molecules such as carbon dioxide,and the specific heat ratio decreases with temperature, which is one of thereasons that the actual cycles have lower thermal efficiencies than the idealOtto cycle. The thermal efficiencies of actual spark-ignition engines rangefrom about 25 to 30 percent.

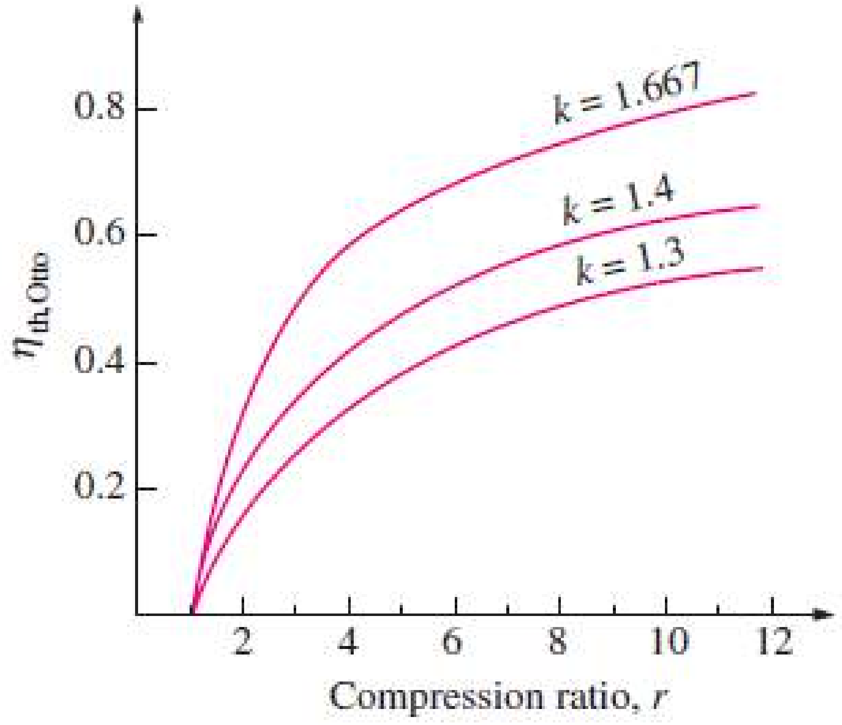


FIGURE 5–3: The thermal efficiency of the Ottocycle increases with the specific heat ratio*k* of the working fluid.

DIESEL CYCLE: THE IDEAL CYCLEFOR COMPRESSION-IGNITION ENGINES:

The Diesel cycle is the ideal cycle for CI reciprocating engines. The CIengine, first proposed by Rudolph Diesel in the 1890s, is very similar to theSI engine discussed in the last section, differing mainly in the method ofinitiating combustion. In spark-ignition engines (also known as *gasolineengines*), the air–fuel mixture is compressed to a temperature that is belowthe autoignition temperature of the fuel, and the combustion process is initiatedby firing a spark plug. In CI engines (also known as *diesel engines*),the air is compressed to a temperature that is above the autoignition temperatureof the fuel, and combustion starts on contact as the fuel is injected intothis hot air. Therefore, the spark plug and carburetor are replaced by a fuelinjector in diesel engines (Fig. 9–4).

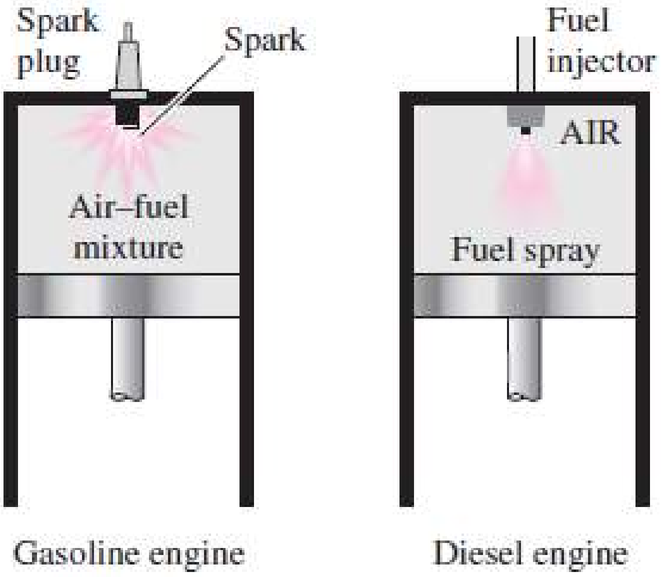
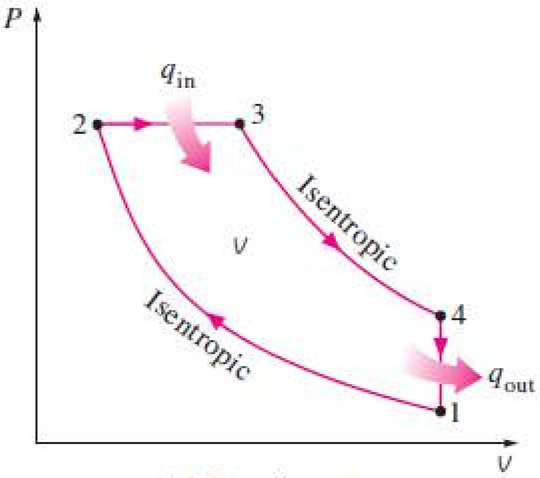


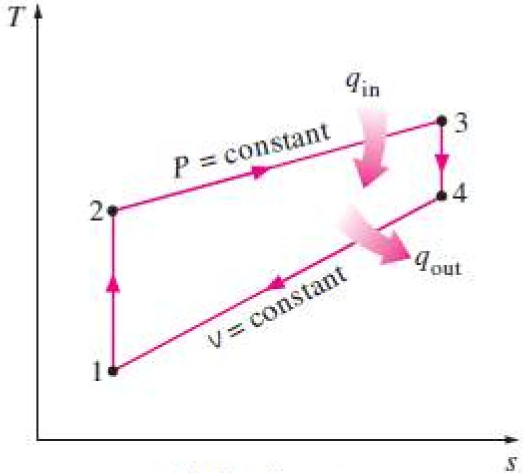
FIGURE 5–4: In diesel engines, the spark plug isreplaced by a fuel injector, and onlyair is compressed during thecompression process.

In gasoline engines, a mixture of air and fuel is compressed during thecompression stroke, and the compression ratios are limited by the onset ofautoignition or engine knock. In diesel engines, only air is compressed duringthe compression stroke, eliminating the possibility of auto ignition. Therefore, diesel engines can be designed to operate at much higher compressionratios, typically between 12 and 24. Not having to deal with theproblem of auto ignition has another benefit: many of the stringent requirementsplaced on the gasoline can now be removed, and fuels that are lessrefined (thus less expensive) can be used in diesel engines.

The fuel injection process in diesel engines starts when the pistonapproaches TDC and continues during the first part of the power stroke.Therefore, the combustion process in these engines takes place over alonger interval. Because of this longer duration, the combustion process inthe ideal Diesel cycle is approximated as a constant-pressure heat-additionprocess. In fact, this is the only process where the Otto and the Dieselcycles differ. The remaining three processes are the same for both idealcycles. That is, process 1-2 is isentropic compression, 3-4 is isentropicexpansion, and 4-1 is constant-volume heat rejection. The similaritybetween the two cycles is also apparent from the *P*-*v* and *T*-*s* diagrams ofthe Diesel cycle, shown in Fig. 9–4.



(*a*) *P*- v diagram



(*b*) *T*-*s* diagram

FIGURE 5–5*T-s* and *P-v* diagrams for the idealDiesel cycle.

Noting that the Diesel cycle is executed in a piston–cylinder device,which forms a closed system, the amount of heat transferred to the workingfluid at constant pressure and rejected from it at constant volume can beexpressed as

$$q_{in} - w_{b,out} = u_3 - u_2 \rightarrow q_{in} = P_2(v_3 - v_2) + (u_3 - u_2)$$
$$= h_3 - h_2 = c_p(T_3 - T_2) \qquad \textbf{(9–10}\textbf{\textit{a}}\textbf{)}$$

and

$$-q_{out} = u_1 - u_4 \rightarrow q_{out} = u_4 - u_1 = c_v(T_4 - T_1) \qquad \textbf{(9–10}\textbf{\textit{b}}\textbf{)}$$

Then the thermal efficiency of the ideal Diesel cycle under the cold-air-standard assumptions becomes

$$\eta_{th,Diesel} = \frac{w_{net}}{q_{in}} = 1 - \frac{q_{out}}{q_{in}} = 1 - \frac{T_4 - T_1}{k(T_3 - T_2)} = 1 - \frac{T_1(T_4/T_1 - 1)}{kT_2(T_3/T_2 - 1)}$$

We now define a new quantity, the **cutoff ratio** *rc*, as the ratio of the cylindervolumes after and before the combustion process:

$$r_c = \frac{V_3}{V_2} = \frac{v_3}{v_2}$$

Utilizing this definition and the isentropic ideal-gas relations for processes1-2 and 3-4, we see that the thermal efficiency relation reduces to

$$\eta_{\text{th,Diesel}} = 1 - \frac{1}{r^{k-1}}\left[\frac{r_c^k - 1}{k(r_c - 1)}\right]$$

where*r* is the compression ratio defined by Eq. 9–9. Looking at Eq. 9–12carefully, one would notice that under the cold-air-standard assumptions, theefficiency of a Diesel cycle differs from the efficiency of an Otto cycle bythe quantity in the brackets. This quantity is always greater than 1. Therefore,

$$\eta_{\text{th,Otto}} > \eta_{\text{th,Diesel}}$$

when both cycles operate on the same compression ratio. Also, as the cutoffratio decreases, the efficiency of the Diesel cycle increases (Fig. 9–22). For thelimiting case of *rc*_ 1, the quantity in the brackets becomes unity (can youprove it?), and the efficiencies of the Otto and Diesel cycles become identical.Remember, though, that diesel engines operate at much higher compressionratios and thus are usually more efficient than the spark-ignition (gasoline)engines. The diesel engines also burn the fuel more completely since theyusually operate at lower revolutions per minute and the air–fuel mass ratio ismuch higher than spark-ignition engines. Thermal efficiencies of large dieselengines range from about 35 to 40 percent.

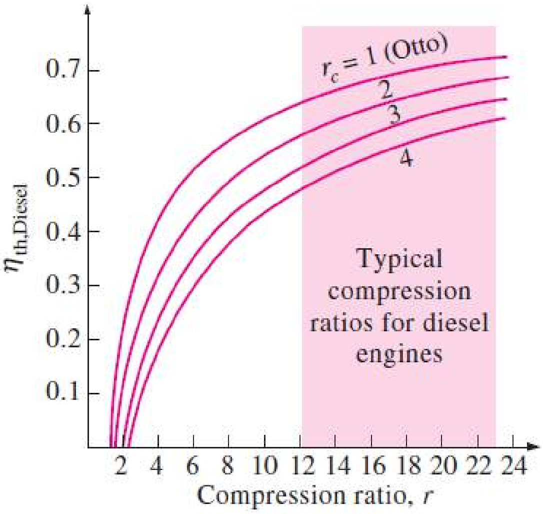


FIGURE 5–6 Thermal efficiency of the ideal Dieselcycle as a function of compression andcutoff ratios (*k* _ 1.4).

The higher efficiency and lower fuel costs of diesel engines make themttractive in applications requiring relatively large amounts of power, suchas in locomotive engines, emergency power generation units, large ships,and heavy trucks. As an example of how large a diesel engine can be, a 12-cylinder diesel engine built in 1964 by the Fiat Corporation of Italy had anormal power output of 25,200 hp (18.8 MW) at 122 rpm, a cylinder boreof 90 cm, and a stroke of 91 cm.Approximating the combustion process in internal combustion engines as aconstant-volume or a constant-pressure heat-addition process is overly simplisticand not quite realistic. Probably a better (but slightly more complex)approach would be to model the combustion process in both gasoline anddiesel engines as a combination of two heat-transfer processes, one at constantvolume and the other at constant pressure. The ideal cycle based on this conceptis called the **dual cycle,** and a *P-v* diagram for it is given in Fig. 9–23.The relative amounts of heat transferred during each process can be adjusted toapproximate the actual cycle more closely. Note that both the Otto and theDiesel cycles can be obtained as special cases of the dual cycle.

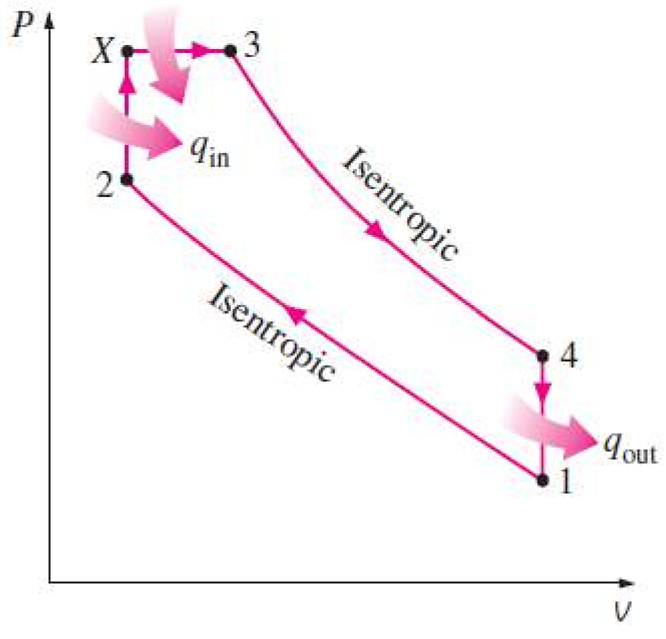


FIGURE 5–7*P-v* diagram of an ideal dual cycle.

BRAYTON CYCLE: THE IDEAL CYCLEFOR GAS-TURBINE ENGINES:

The Brayton cycle was first proposed by George Brayton for use in the reciprocatingoil-burning engine that he developed around 1870. Today, it is usedfor gas turbines only where both the compression and expansion processestake place in rotating machinery. Gas turbines usually operate on an *opencycle,* as shown in Fig. 9–29. Fresh air at ambient conditions is drawn intothe compressor, where its temperature and pressure are raised. The highpressureair proceeds into the combustion chamber, where the fuel is burnedat constant pressure. The resulting high-temperature gases then enter the turbine,where they expand to the atmospheric pressure while producingpower. The exhaust gases leaving the turbine are thrown out (not recirculated),causing the cycle to be classified as an open cycle.

The open gas-turbine cycle described above can be modeled as a *closedcycle,* as shown in Fig. 9–30, by utilizing the air-standard assumptions. Herethe compression and expansion processes remain the same, but the combustionprocess is replaced by a constant-pressure heat-addition process froman external source, and the exhaust process is replaced by a constantpressureheat-rejection process to the ambient air. The ideal cycle that theworking fluid undergoes in this closed loop is the **Brayton cycle,** which ismade up of four internally reversible processes:

1-2 Isentropic compression (in a compressor)
2-3 Constant-pressure heat addition
3-4 Isentropic expansion (in a turbine)
4-1 Constant-pressure heat rejection

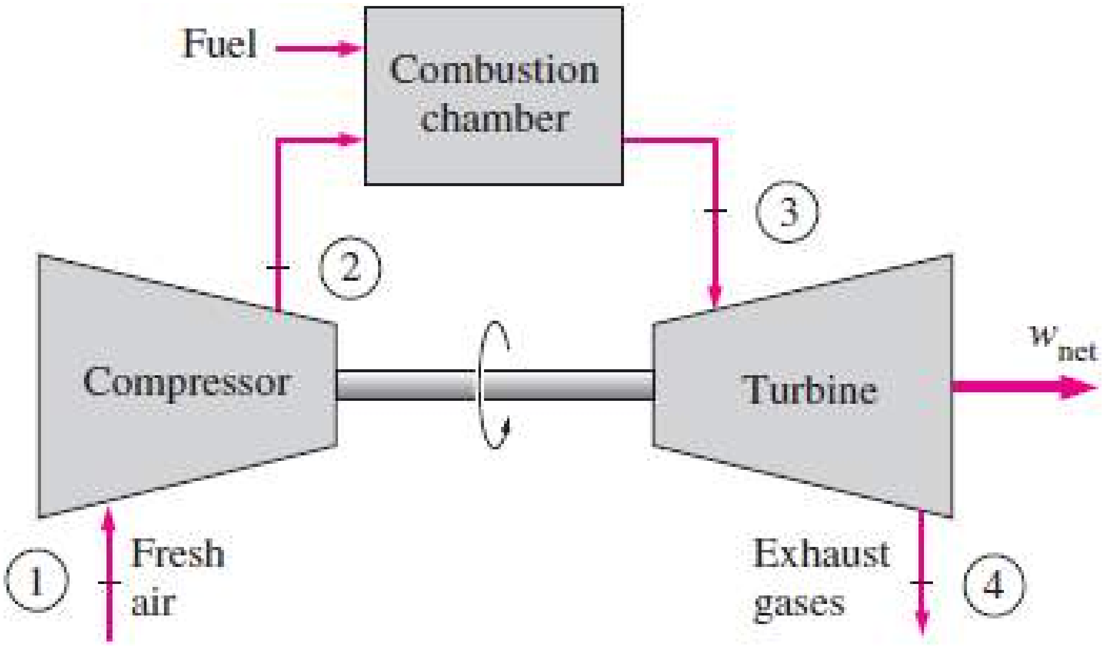


FIGURE 5–8: An open-cycle gas-turbine engine.

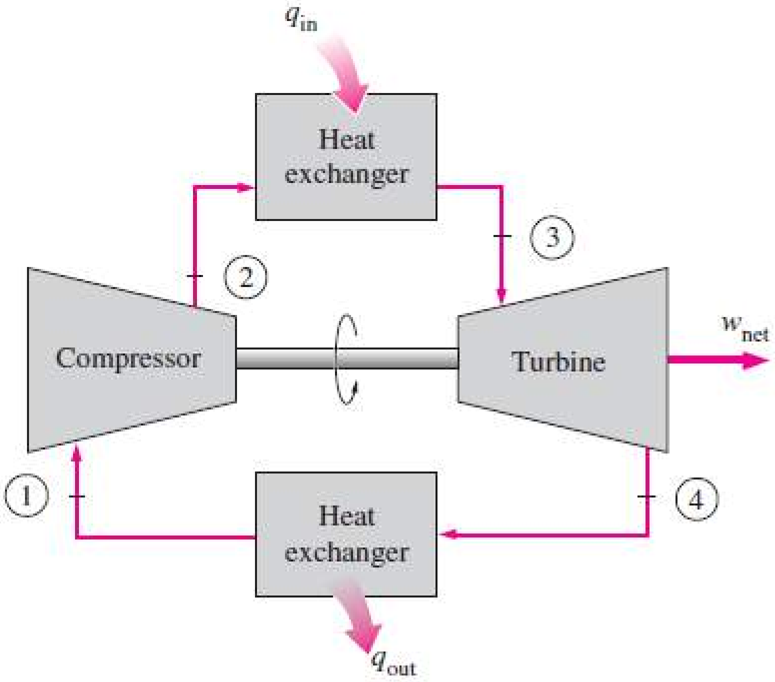
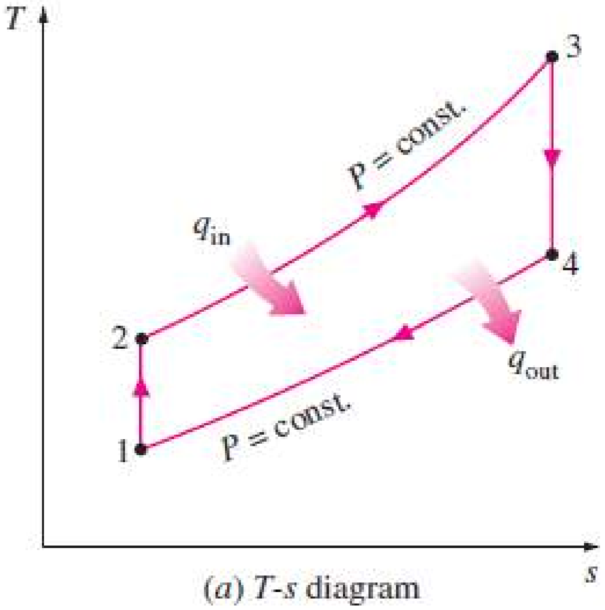


FIGURE 5–9: A closed-cycle gas-turbine engine.

The *T-s* and *P-v* diagrams of an ideal Brayton cycle are shown in Fig. 9–31.Notice that all four processes of the Brayton cycle are executed in steadyflowdevices; thus, they should be analyzed as steady-flow processes. Whenthe changes in kinetic and potential energies are neglected, the energy balancefor a steady-flow process can be expressed, on a unit–mass basis, as



(*a*) *T-s* diagram

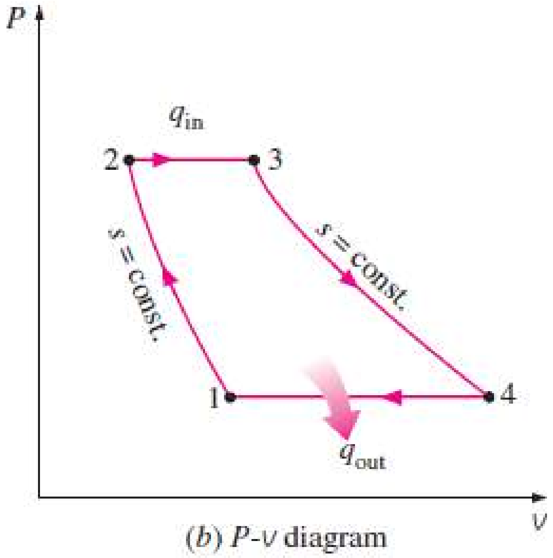


FIGURE 5–10: *T-s* and *P-v* diagrams for the idealBrayton cycle.

$$(q_{in} - q_{out}) + (w_{in} - w_{out}) = h_{exit} - h_{inlet} \tag{9–15}$$

Therefore, heat transfers to and from the working fluid are

$$q_{in} = h_3 - h_2 = c_p(T_3 - T_2) \tag{9–16a}$$

and

$$q_{out} = h_4 - h_1 = c_p(T_4 - T_1) \tag{9–16b}$$

Then the thermal efficiency of the ideal Brayton cycle under the cold-air-standard assumptions becomes

$$\eta_{th,Brayton} = \frac{w_{net}}{q_{in}} = 1 - \frac{q_{out}}{q_{in}} = 1 - \frac{c_p(T_4 - T_1)}{c_p(T_3 - T_2)} = 1 - \frac{T_1(T_4/T_1 - 1)}{T_2(T_3/T_2 - 1)}$$

Processes 1-2 and 3-4 are isentropic, and $P_2 = P_3$ and $P_4 = P_1$. Thus,

$$\frac{T_2}{T_1} = \left(\frac{P_2}{P_1}\right)^{(k-1)/k} = \left(\frac{P_3}{P_4}\right)^{(k-1)/k} = \frac{T_3}{T_4}$$

Substituting these equations into the thermal efficiency relation and simplifying give

$$\eta_{th,Brayton} = 1 - \frac{1}{r_p^{(k-1)/k}} \tag{9–17}$$

where

$$r_p = \frac{P_2}{P_1}$$

is the **pressure ratio** and *k* is the specific heat ratio. Equation 9–17 showsthat under the cold-air-standard assumptions, the thermal efficiency of anideal Brayton cycle depends on the pressure ratio of the gas turbine and thespecific heat ratio of the working fluid. The thermal efficiency increases withboth of these parameters, which is also the case for actual gas turbines.A plot of thermal efficiency versus the pressure ratio is given in Fig. 9–32 for*k* _ 1.4, which is the specific-heat-ratio value of air at room temperature.

The highest temperature in the cycle occurs at the end of the combustionprocess (state 3), and it is limited by the maximum temperature that the turbineblades can withstand. This also limits the pressure ratios that can beused in the cycle. For a fixed turbine inlet temperature *T*3, the net work outputper cycle increases with the pressure ratio, reaches a maximum, andthen starts to decrease, as shown in Fig. 9–33. Therefore, there should be acompromise between the pressure ratio (thus the thermal efficiency) and thenet work output. With less work output per cycle, a larger mass flow rate(thus a larger system) is needed to maintain the same power output, whichmay not be economical. In most common designs, the pressure ratio of gasturbines ranges from about 11 to 16.

The air in gas turbines performs two important functions: It supplies thenecessary oxidant for the combustion of the fuel, and it serves as a coolantto keep the temperature of various components within safe limits. The secondfunction is accomplished by drawing in more air than is needed for thecomplete combustion of the fuel. In gas turbines, an air–fuel mass ratio of50 or above is not uncommon. Therefore, in a cycle analysis, treating thecombustion gases as air does not cause any appreciable error. Also, the massflow rate through the turbine is greater than that through the compressor, thedifference being equal to the mass flow rate of the fuel. Thus, assuming aconstant mass flow rate throughout the cycle yields conservative results foropen-loop gas-turbine engines.

The two major application areas of gas-turbine engines are *aircraft propulsion*and *electric power generation.* When it is used for aircraft propulsion,the gas turbine produces just enough power to drive the compressor and asmall generator to power the auxiliary equipment. The high-velocity exhaustgases are responsible for producing the necessary thrust to propel the aircraft.

Gas turbines are also used as stationary power plants to generate electricityas stand-alone units or in conjunction with steam power plants on thehigh-temperature side. In these plants, the exhaust gases of the gas turbineserve as the heat source for the steam. The gas-turbine cycle can also be executedas a closed cycle for use in nuclear power plants. This time the workingfluid is not limited to air, and a gas with more desirable characteristics(such as helium) can be used.

The majority of the Western world’s naval fleets already use gas-turbineengines for propulsion and electric power generation. The General ElectricLM2500 gas turbines used to power ships have a simple-cycle thermal efficiencyof 37 percent. The General Electric WR-21 gas turbines equipped withintercooling and regeneration have a thermal efficiency of 43 percent and

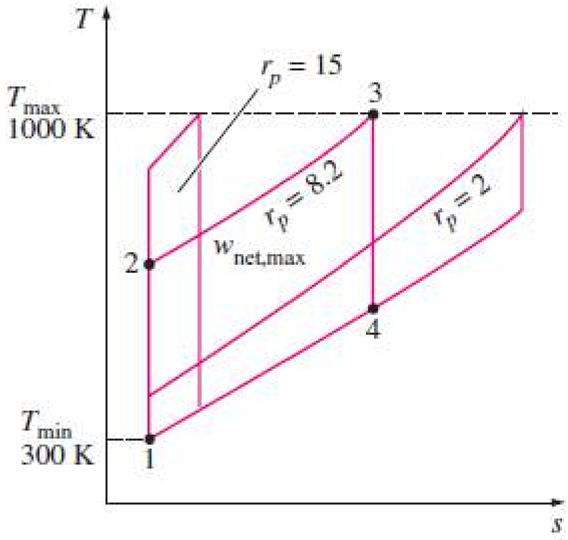


FIGURE 5–11: For fixed values of *T*min and *T*max,the net work of the Brayton cyclefirst increases with the pressureratio, then reaches a maximum at*rp*_ (*T*max/*T*min)*k*/[2(*k* _ 1)], andfinally decreases.

Produce 21.6 MW (29,040 hp). The regeneration also reduces the exhaust temperaturefrom 600°C (1100°F) to 350°C (650°F). Air is compressed to 3 atmbefore it enters the intercooler. Compared to steam-turbine and dieselpropulsionsystems, the gas turbine offers greater power for a given size andweight, high reliability, long life, and more convenient operation. The enginestart-up time has been reduced from 4 h required for a typical steampropulsionsystem to less than 2 min for a gas turbine. Many modern marinepropulsion systems use gas turbines together with diesel engines because of thehigh fuel consumption of simple-cycle gas-turbine engines. In combined dieseland gas-turbine systems, diesel is used to provide for efficient low-power andcruise operation, and gas turbine is used when high speeds are needed.

In gas-turbine power plants, the ratio of the compressor work to the turbinework, called the **back work ratio,** is very high (Fig. 9–34). Usuallymore than one-half of the turbine work output is used to drive the compressor.The situation is even worse when the isentropic efficiencies of the compressorand the turbine are low. This is quite in contrast to steam powerplants, where the back work ratio is only a few percent. This is not surprising,however, since a liquid is compressed in steam power plants instead ofa gas, and the steady-flow work is proportional to the specific volume of theworking fluid.

A power plant with a high back work ratio requires a larger turbine toprovide the additional power requirements of the compressor. Therefore, theturbines used in gas-turbine power plants are larger than those used in steampower plants of the same net power output

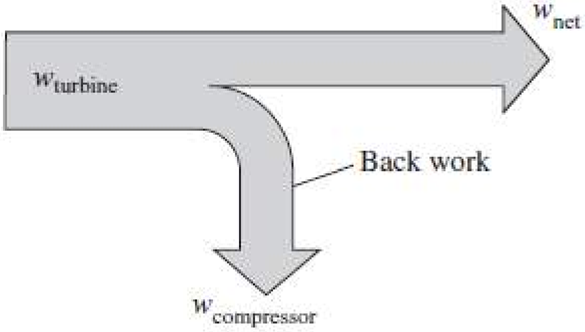


FIGURE 5–12: The fraction of the turbine work used to drive the compressor is called the back work ratio.

THE BRAYTON CYCLE WITH REGENERATION:

In gas-turbine engines, the temperature of the exhaust gas leaving the turbineis often considerably higher than the temperature of the air leaving thecompressor. Therefore, the high-pressure air leaving the compressor can beheated by transferring heat to it from the hot exhaust gases in a counter-flowheat exchanger, which is also known as a *regenerator* or a *recuperator*.A sketch of the gas-turbine engine utilizing a regenerator and the *T-s*diagram of the new cycle are shown in Figs. 5–12.

The thermal efficiency of the Brayton cycle increases as a result of regenerationsince the portion of energy of the exhaust gases that is normally rejectedto the surroundings is now used to preheat the air entering the combustionchamber. This, in turn, decreases the heat input (thus fuel) requirements forthe same net work output. Note, however, that the use of a regenerator is recommendedonly when the turbine exhaust temperature is higher than the compressorexit temperature. Otherwise, heat will flow in the reverse direction (*to*the exhaust gases), decreasing the efficiency. This situation is encountered ingas-turbine engines operating at very high pressure ratios.

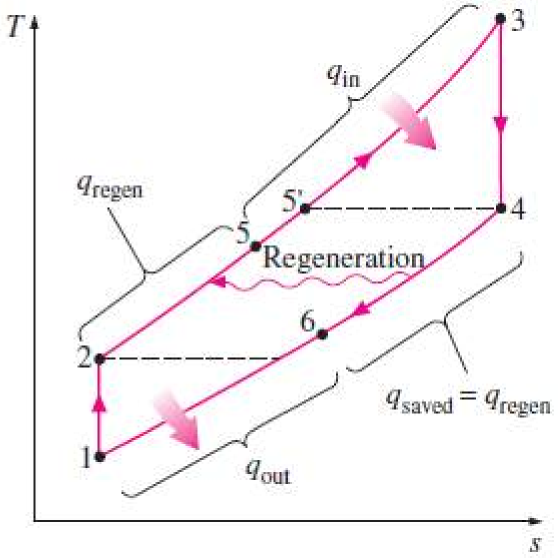


FIGURE 5–13:*T*-*s* diagram of a Brayton cycle withregeneration.

The highest temperature occurring within the regenerator is *T*4, the temperatureof the exhaust gases leaving the turbine and entering the regenerator.Under no conditions can the air be preheated in the regenerator to atemperature above this value. Air normally leaves the regenerator at a lowertemperature, *T*5. In the limiting (ideal) case, the air exits the regenerator atthe inlet temperature of the exhaust gases *T*4. Assuming the regenerator tobe well insulated and any changes in kinetic and potential energies to benegligible, the actual and maximum heat transfers from the exhaust gases tothe air can be expressed as

$$q_{\text{regen,act}} = h_5 - h_2 \qquad \text{(9–21)}$$

and

$$q_{\text{regen,max}} = h_{5'} - h_2 = h_4 - h_2 \qquad \text{(9–22)}$$

The extent to which a regenerator approaches an ideal regenerator is called the **effectiveness** ϵ and is defined as

$$\epsilon = \frac{q_{\text{regen,act}}}{q_{\text{regen,max}}} = \frac{h_5 - h_2}{h_4 - h_2} \qquad \text{(9–23)}$$

When the cold-air-standard assumptions are utilized, it reduces to

$$\epsilon \cong \frac{T_5 - T_2}{T_4 - T_2}$$

A regenerator with a higher effectiveness obviously saves a greateramount of fuel since it preheats the air to a higher temperature prior to combustion.However, achieving a higher effectiveness requires the use of alarger regenerator, which carries a higher price tag and causes a larger pressuredrop. Therefore, the use of a regenerator with a very high effectivenesscannot be justified economically unless the savings from the fuel costsexceed the additional expenses involved. The effectiveness of most regeneratorsused in practice is below 0.85.Under the cold-air-standard assumptions, the thermal efficiency of anideal Brayton cycle with regeneration is

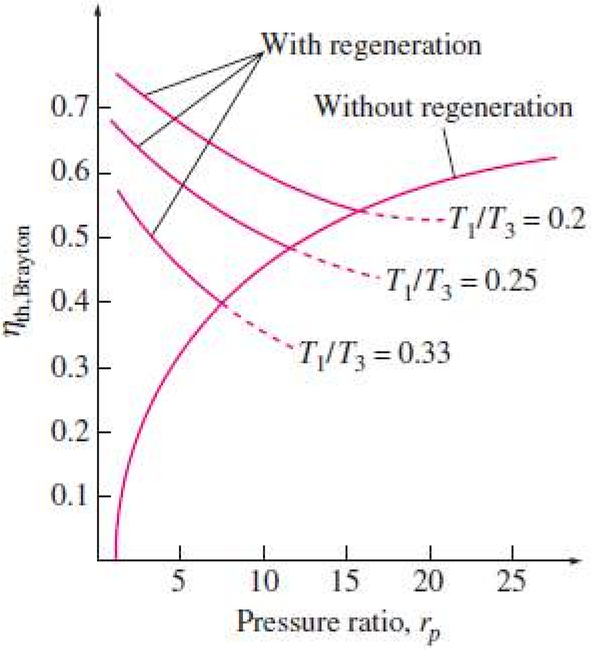


FIGURE 5–14: Thermal efficiency of the idealBrayton cycle with and withoutregeneration.

Therefore, the thermal efficiency of an ideal Brayton cycle with regenerationdepends on the ratio of the minimum to maximum temperatures as wellas the pressure ratio. The thermal efficiency is plotted in Fig. 9–40 for variouspressure ratios and minimum-to-maximum temperature ratios. This figureshows that regeneration is most effective at lower pressure ratios andlow minimum-to-maximum temperature ratios.

THE BRAYTON CYCLE WITHINTERCOOLING, REHEATING,AND REGENERATION:

The net work of a gas-turbine cycle is the difference between the turbinework output and the compressor work input, and it can be increased byeither decreasing the compressor work or increasing the turbine work, orboth. It was shown in Chap. 7 that the work required to compress a gasbetween two specified pressures can be decreased by carrying out the compressionprocess in stages and cooling the gas in between (Fig. 9–14)—thatis, using *multistage compression with intercooling.* As the number of stagesis increased, the compression process becomes nearly isothermal at thecompressor inlet temperature, and the compression work decreases.

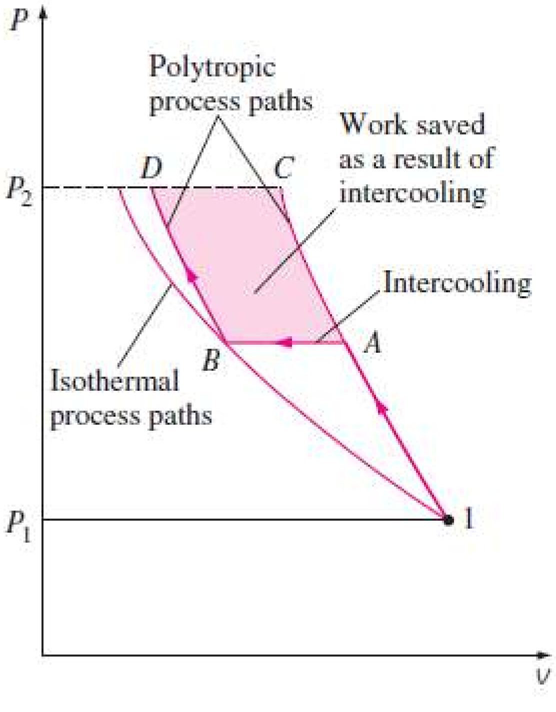
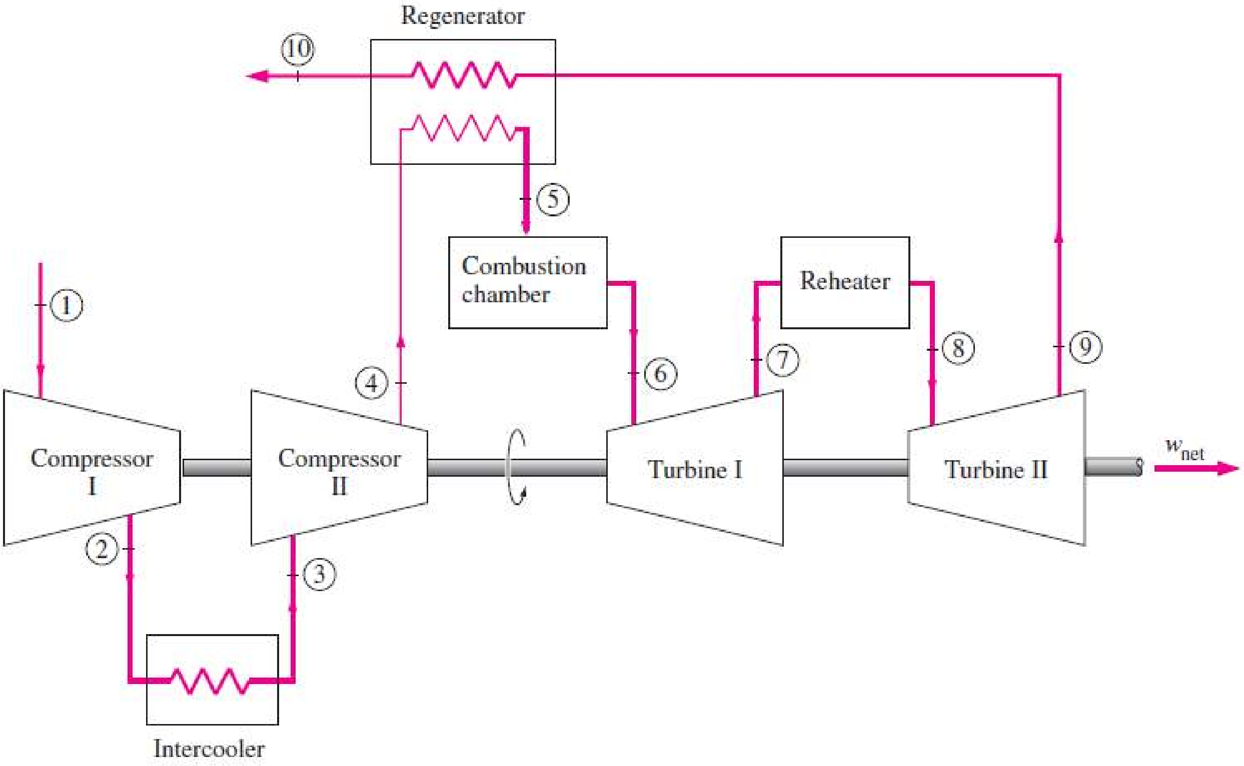


FIGURE 5–15: Comparison of work inputs to asingle-stage compressor (1*AC*) and atwo-stage compressor withintercooling (1*ABD*).

Combustion in gas turbines typically occurs at four times the amount ofair needed for complete combustion to avoid excessive temperatures. Therefore,the exhaust gases are rich in oxygen, and reheating can be accomplishedby simply spraying additional fuel into the exhaust gases between two expansion states.The working fluid leaves the compressor at a lower temperature, and the turbine at a higher temperature, when intercooling and reheating are utilized.

This makes regeneration more attractive since a greater potential forregeneration exists. Also, the gases leaving the compressor can be heated toa higher temperature before they enter the combustion chamber because ofthe higher temperature of the turbine exhaust.

A schematic of the physical arrangement and the *T*-*s* diagram of an idealtwo-stage gas-turbine cycle with intercooling, reheating, and regeneration areshown in Figs. 9–43 and 9–44. The gas enters the first stage of the compressorat state 1, is compressed isentropically to an intermediate pressure *P*2, iscooled at constant pressure to state 3 (*T*3 _ *T*1), and is compressed in the second stageisentropically to the final pressure *P*4. At state 4 the gas enters theregenerator, where it is heated to *T*5 at constant pressure. In an ideal regenerator,the gas leaves the regenerator at the temperature of the turbine exhaust,that is, *T*5 _ *T*9. The primary heat addition (or combustion) process takesplace between states 5 and 6. The gas enters the first stage of the turbine atstate 6 and expands isentropically to state 7, where it enters the reheater. It isreheated at constant pressure to state 8 (*T*8 _ *T*6), where it enters the secondstage of the turbine. The gas exits the turbine at state 9 and enters the regenerator,where it is cooled to state 10 at constant pressure. The cycle is completedby cooling the gas to the initial state (or purging the exhaust gases).



That the work input to a two-stage compressor isminimized when equal pressure ratios are maintained across each stage. Itcan be shown that this procedure also maximizes the turbine work output.Thus, for best performance we have

$$\frac{P_2}{P_1} = \frac{P_4}{P_3} \quad \text{and} \quad \frac{P_6}{P_7} = \frac{P_8}{P_9}$$

In the analysis of the actual gas-turbine cycles, the irreversibilities that arepresent within the compressor, the turbine, and the regenerator as well as thepressure drops in the heat exchangers should be taken into consideration.

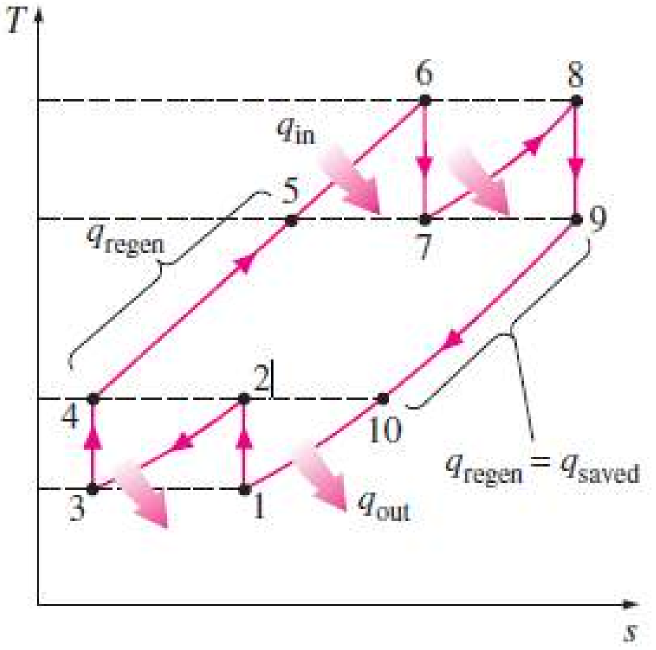


FIGURE 5-16: *T-s* diagram of an ideal gas-turbinecycle with intercooling, reheating, andregeneration.

If the number of compression and expansion stages is increased, the idealgas-turbine cycle with intercooling, reheating, and regeneration approachesthe Ericsson cycle, as illustrated in Fig. 9–45, and the thermal efficiencyapproaches the theoretical limit (the Carnot efficiency). However, the contributionof each additional stage to the thermal efficiency is less and less, andthe use of more than two or three stages cannot be justified economically.

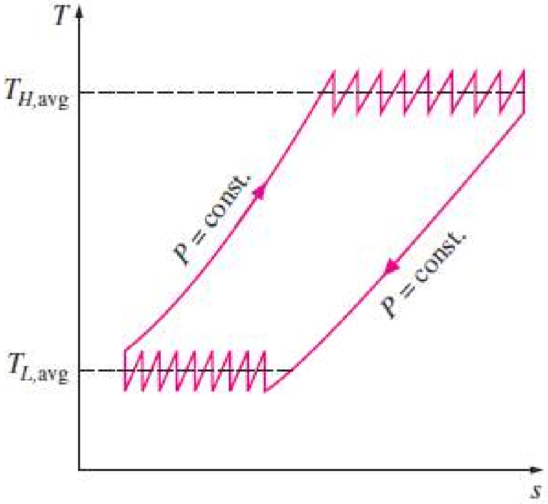


FIGURE 5–17: As the number of compression andexpansion stages increases, the gas turbinecycle with intercooling,reheating, and regenerationapproaches the Ericsson cycle.

THE CARNOT VAPOR CYCLE

Steady-flow *Carnot cycle* executed within the saturation domeof a pure substance, as shown in Fig. 5-18*a*. The fluid is heated reversiblyand isothermally in a boiler (process 1-2), expanded isentropically in a turbine(process 2-3), condensed reversibly and isothermally in a condenser (process 3-4), and compressed isentropically by a compressor to the initialstate (process 4-1).

Several impracticalities are associated with this cycle:

1. Isothermal heat transfer to or from a two-phase system is not difficultto achieve in practice since maintaining a constant pressure in thedevice automatically fixes the temperature at the saturation value. Therefore,processes 1-2 and 3-4 can be approached closely in actual boilers and condensers.Limiting the heat transfer processes to two-phase systems, however,severely limits the maximum temperature that can be used in the cycle(it has to remain under the critical-point value, which is 374°C for water).Limiting the maximum temperature in the cycle also limits the thermal efficiency.Any attempt to raise the maximum temperature in the cycle involvesheat transfer to the working fluid in a single phase, which is not easy toaccomplish isothermally.

2. The isentropic expansion process (process 2-3) can be approximatedclosely by a well-designed turbine. However, the quality of the steam decreasesduring this process, as shown on the *T*-*s* diagram in Fig. 10–1*a*. Thus theturbine has to handle steam with low quality, that is, steam with a highmoisture content. The impingement of liquid droplets on the turbine bladescauses erosion and is a major source of wear. Thus steam with qualities lessthan about 90 percent cannot be tolerated in the operation of power plants.

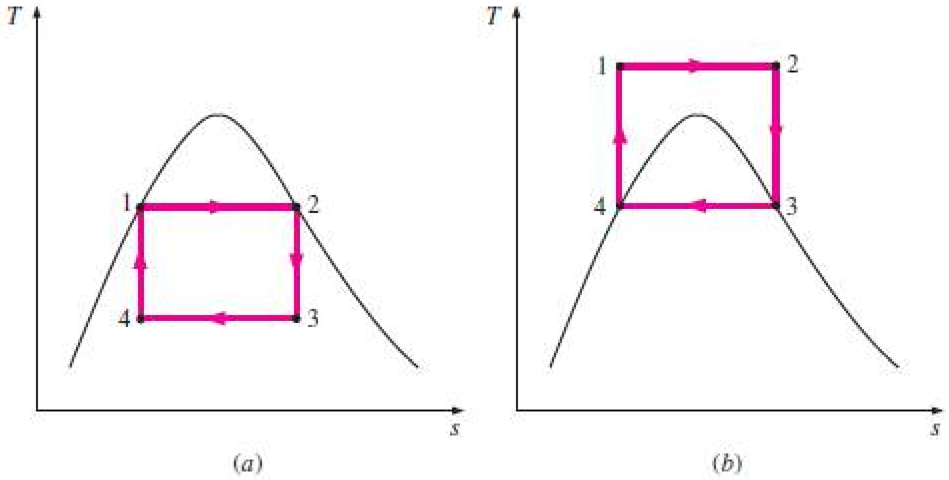


FIGURE 5–18: *T-s* diagram of two Carnot vaporcycles.

3. The isentropic compression process (process 4-1) involves the compressionof a liquid–vapor mixture to a saturated liquid. There are two difficultiesassociated with this process. First, it is not easy to control the condensationprocess so precisely as to end up with the desired quality at state second, itis not practical to design a compressor that handles two phases.

RANKINE CYCLE: THE IDEAL CYCLEFOR VAPOR POWER CYCLES

Many of the impracticalities associated with the Carnot cycle can be eliminatedby superheating the steam in the boiler and condensing it completelyin the condenser, as shown schematically on a *T-s* diagram in Fig. 10–2. Thecycle that results is the **Rankine cycle,** which is the ideal cycle for vaporpower plants. The ideal Rankine cycle does not involve any internal irreversibilities
and consists of the following four processes:
1-2 Isentropic compression in a pump
2-3 Constant pressure heat addition in a boiler
3-4 Isentropic expansion in a turbine
4-1 Constant pressure heat rejection in a condenser.

Water enters the *pump* at state 1 as saturated liquid and is compressedisentropically to the operating pressure of the boiler. The water temperatureincreases somewhat during this isentropic compression process due to aslight decrease in the specific volume of water. The vertical distance between states 1 and 2 on the *T-s* diagram is greatly exaggerated for clarity.

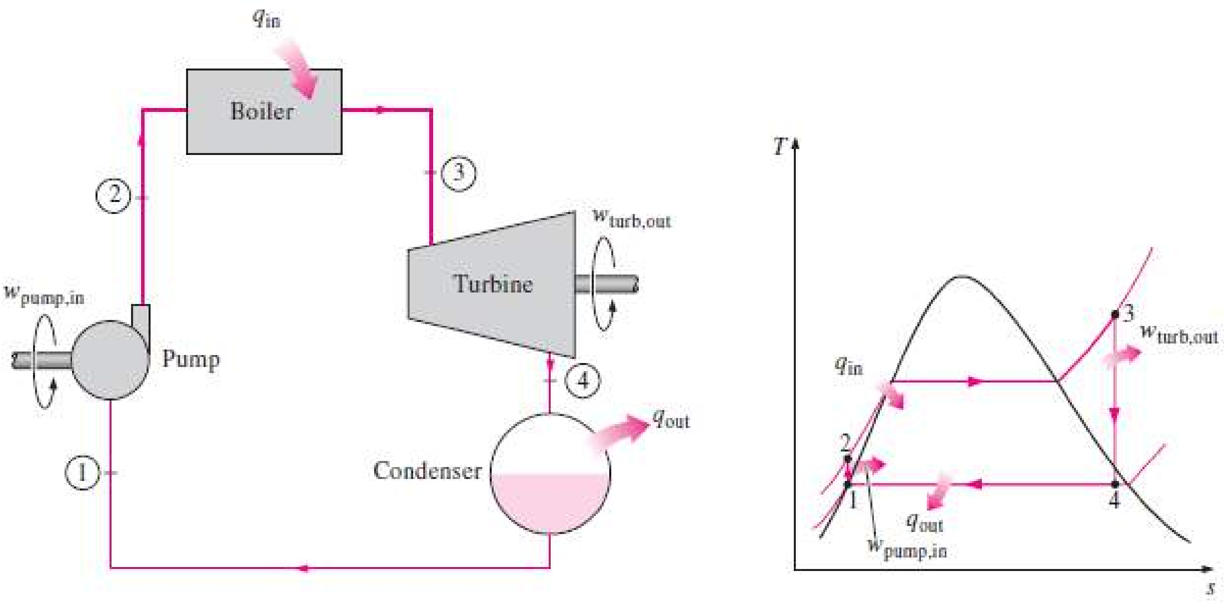


FIGURE 5–19: The simple ideal Rankine cycle.

Energy Analysis of the Ideal Rankine Cycle

All four components associated with the Rankine cycle (the pump, boiler,turbine, and condenser) are steady-flow devices, and thus all four processesthat make up the Rankine cycle can be analyzed as steady-flow processes.The kinetic and potential energy changes of the steam are usually small relativeto the work and heat transfer terms and are therefore usuallyneglected. Then the *steady-flow energy equation* per unit mass of steamreduces to

$$(q_{in} - q_{out}) + (w_{in} - w_{out}) = h_e - h_i \qquad \text{(kJ/kg)}$$

The boiler and the condenser do not involve any work, and the pump andthe turbine are assumed to be isentropic. Then the conservation of energyrelation for each device can be expressed as follows:

Pump (q = 0): $w$pump,in = h2 _ h1

The thermal efficiency of the Rankine cycle is determined from

$$\eta_{th} = \frac{w_{net}}{q_{in}} = 1 - \frac{q_{out}}{q_{in}}$$

Where

$w$net=$q$in-$q$out=$w$turb,out-$w$pump,in

The conversion efficiency of power plants in the United States is oftenexpressed in terms of **heat rate,** which is the amount of heat supplied, inBtu’s, to generate 1 kWh of electricity. The smaller the heat rate, the greaterthe efficiency. Considering that 1 kWh _ 3412 Btu and disregarding the losses associated with the conversion of shaft power to electric power, therelation between the heat rate and the thermal efficiency can be expressed as

$$\eta_{th} = \frac{3412\ (\text{Btu/kWh})}{\text{Heat rate}\ (\text{Btu/kWh})}$$

Lowering the Condenser Pressure (*Lowers T*low,avg)

Steam exists as a saturated mixture in the condenser at the saturationtemperature corresponding to the pressure inside the condenser. Therefore,lowering the operating pressure of the condenser automatically lowers thetemperature of the steam, and thus the temperature at which heat is rejected.

The effect of lowering the condenser pressure on the Rankine cycle efficiencyis illustrated on a *T*-*s* diagram in Fig. 5–20. For comparison purposes,the turbine inlet state is maintained the same. The colored area onthis diagram represents the increase in net work output as a result of lowering the condenser pressure from *P*4 to . The heat input requirements alsoincrease (represented by the area under curve 2_-2), but this increase is verysmall. Thus the overall effect of lowering the condenser pressure is anincrease in the thermal efficiency of the cycle.

To take advantage of the increased efficiencies at low pressures, the condensersof steam power plants usually operate well below the atmosphericpressure. This does not present a major problem since the vapor powercycles operate in a closed loop. However, there is a lower limit on the condenserpressure that can be used. It cannot be lower than the saturation pressurecorresponding to the temperature of the cooling medium. Consider, forexample, a condenser that is to be cooled by a nearby river at 15°C. Allowinga temperature difference of 10°C for effective heat transfer, the steamtemperature in the condenser must be above 25°C; thus the condenser pressuremust be above 3.2 kPa, which is the saturation pressure at 25°C.

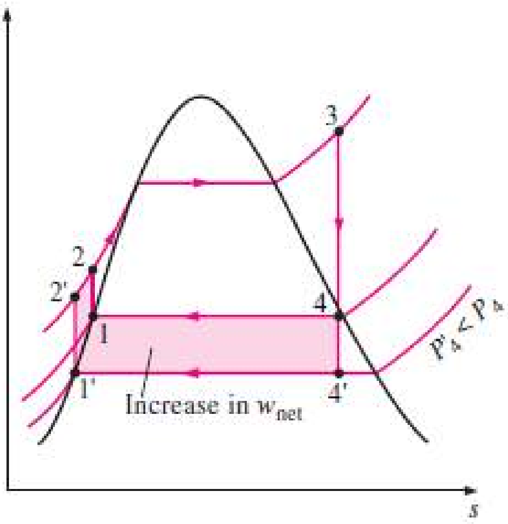


FIGURE 5–20: The effect of lowering the condenserpressure on the ideal Rankine cycle.

Superheating the Steam to High Temperatures(*Increases T*high,avg):

The average temperature at which heat is transferred to steam can be increased without increasing the boiler pressure by superheating the steam to high temperatures. The effect of superheating on the performance of vapor power cycles is illustrated on a *T*-*s* diagram in Fig. 5–21. The colored area on this diagram represents the increase in the net work. The total area under theprocess curve 3-3_ represents the increase in the heat input. Thus both the network and heat input increase as a result of superheating the steam to a highertemperature. The overall effect is an increase in thermal efficiency, however,since the average temperature at which heat is added increases.

Superheating the steam to higher temperatures has another very desirable effect: It decreases the moisture content of the steam at the turbine exit, ascan be seen from the *T*-*s* diagram (the quality at state 4_ is higher than thatat state 4).

The temperature to which steam can be superheated is limited, however, bymetallurgical considerations. Presently the highest steam temperature allowedat the turbine inlet is about 620°C (1150°F). Any increase in this valuedepends on improving the present materials or finding new ones that canwithstand higher temperatures. Ceramics are very promising in this regard.

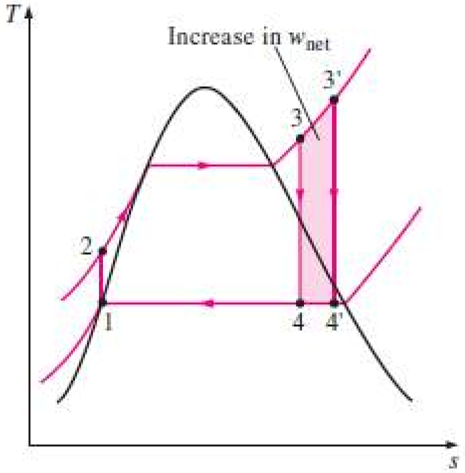


FIGURE 5–21: The effect of superheating the steam tohigher temperatures on the ideal Rankine cycle.

Increasing the Boiler Pressure (*Increases T*high,avg):

Another way of increasing the average temperature during the heat-additionprocess is to increase the operating pressure of the boiler, which automaticallyraises the temperature at which boiling takes place. This, in turn, raisesthe average temperature at which heat is transferred to the steam and thusraises the thermal efficiency of the cycle.

The effect of increasing the boiler pressure on the performance of vaporpower cycles is illustrated on a *T-s* diagram in Fig. 5–22. Notice that for afixed turbine inlet temperature, the cycle shifts to the left and the moisture contentof steam at the turbine exit increases. This undesirable side effect can becorrected, however, by reheating the steam, as discussed in the next section

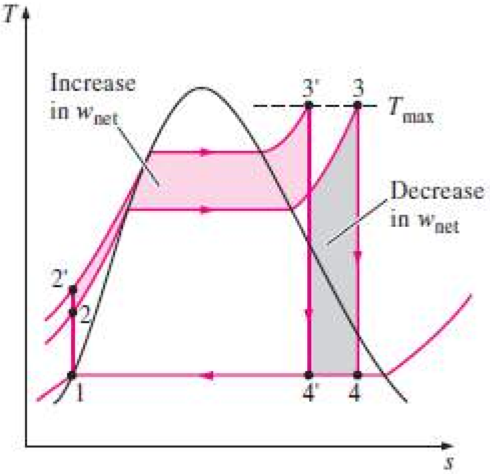


FIGURE 5–22 The effect of increasing the boilerpressure on the ideal Rankine cycle.

Operating pressures of boilers have gradually increased over the yearsfrom about 2.7 MPa (400 psia) in 1922 to over 30 MPa (4500 psia) today,generating enough steam to produce a net power output of 1000 MW or morein a large power plant. Today many modern steam power plants operate atsupercritical pressures (*P* _ 22.06 MPa) and have thermal efficiencies ofabout 40 percent for fossil-fuel plants and 34 percent for nuclear plants.There are over 150 supercritical-pressure steam power plants in operation inthe United States. The lower efficiencies of nuclear power plants are due tothe lower maximum temperatures used in those plants for safety reasons. The *T*-*s* diagram of a supercritical Rankine cycle is shown in Fig. 5–23.

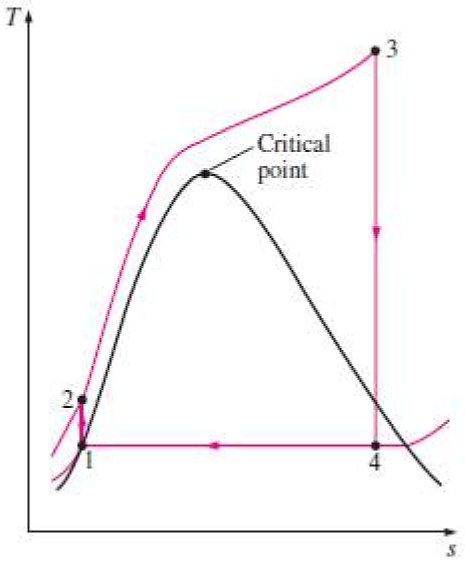


FIGURE 5–23: A supercritical Rankine cycle.

THE IDEAL REGENERATIVE RANKINE CYCLE

A careful examination of the *T*-*s* diagram of the Rankine cycle redrawn inFig.5-24 reveals that heat is transferred to the working fluid duringprocess 2-2_ at a relatively low temperature. This lowers the average heatadditiontemperature and thus the cycle efficiency.

To remedy this shortcoming, we look for ways to raise the temperature ofthe liquid leaving the pump (called the *feedwater*) before it enters the boiler.One such possibility is to transfer heat to the feedwater from the expandingsteam in a counterflow heat exchanger built into the turbine, that is, to use**regeneration.** This solution is also impractical because it is difficult todesign such a heat exchanger and because it would increase the moisturecontent of the steam at the final stages of the turbine.

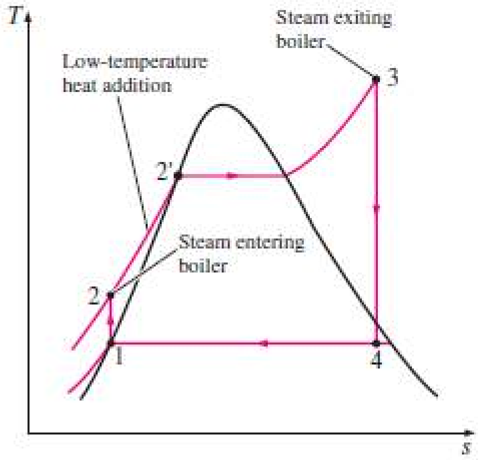


FIGURE 5–24 The first part of the heat-additionprocess in the boiler takes place at relatively low temperatures.

A practical regeneration process in steam power plants is accomplished byextracting, or “bleeding,” steam from the turbine at various points. This steam,which could have produced more work by expanding further in the turbine, isused to heat the feedwater instead. The device where the feedwater is heatedby regeneration is called a **regenerator,** or a **feedwater heater (FWH).**

Closed Feed water Heaters

Another type of feed water heater frequently used in steam power plants isthe **closed feed water heater,** in which heat is transferred from the extractedsteam to the feed water without any mixing taking place. The two streamsnow can be at different pressures, since they do not mix. The schematic of asteam power plant with one closed feed water heater and the *T-s* diagram of the cycle are shown in Fig. 5–25. In an ideal closed feed water heater, thefeed water is heated to the exit temperature of the extracted steam, whichideally leaves the heater as a saturated liquid at the extraction pressure. Inactual power plants, the feed water leaves the heater below the exit temperature *w*pumpture of the extracted steam because a temperature difference of at least a few degrees is required for any effective heat transfer to take place.

The condensed steam is then either pumped to the feedwater line or routedto another heater or to the condenser through a device called a **trap.** A trapallows the liquid to be throttled to a lower pressure region but *traps* thevapor. The enthalpy of steam remains constant during this throttling process.

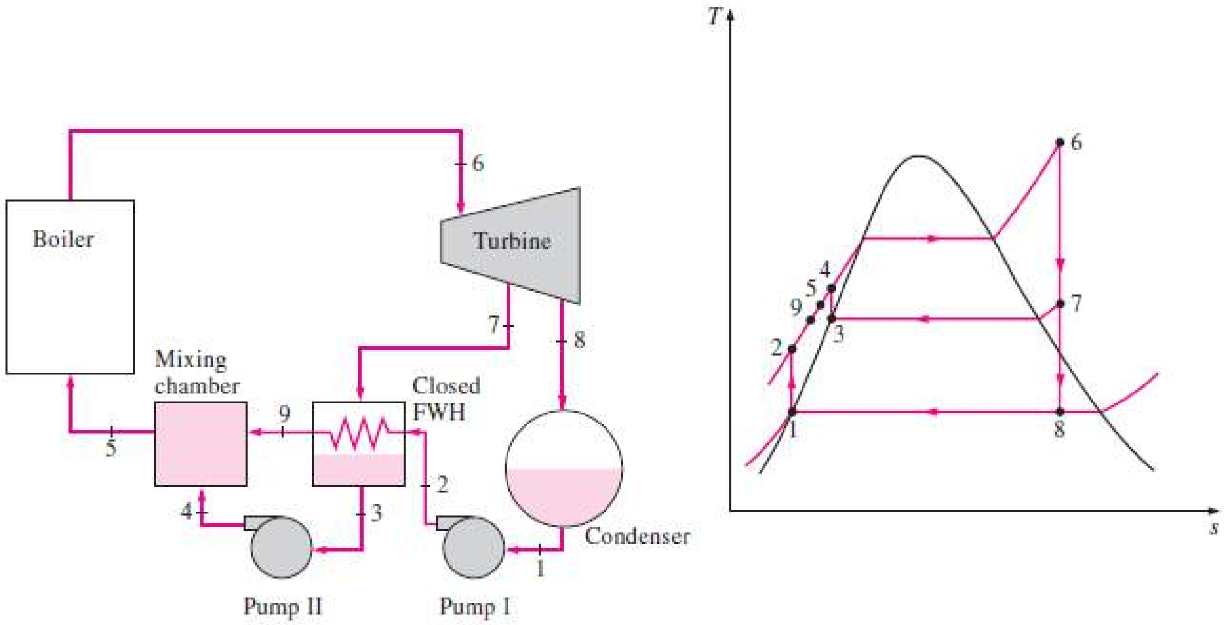


FIGURE 5–25: The ideal regenerative Rankine cycle with a closed feedwater heater.

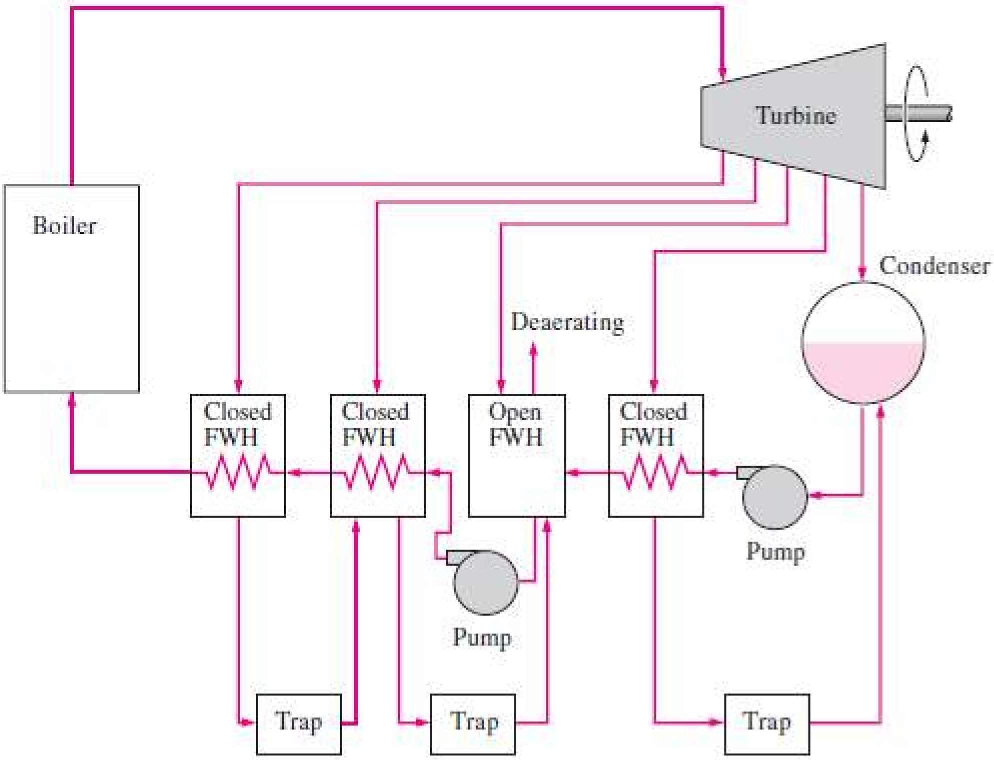


FIGURE 5–26A steam power plant with one open and three closed feed water heaters.

The open and closed feedwater heaters can be compared as follows. Openfeedwater heaters are simple and inexpensive and have good heat transfercharacteristics. They also bring the feedwater to the saturation state. Foreach heater, however, a pump is required to handle the feedwater. The closed feedwater heaters are more complex because of the internal tubingnetwork, and thus they are more expensive. Heat transfer in closed feedwaterheaters is also less effective since the two streams are not allowed to bein direct contact. However, closed feedwater heaters do not require a separatepump for each heater since the extracted steam and the feedwater canbe at different pressures. Most steam power plants use a combination ofopen and closed feedwater heaters, as shown in Fig. 5–26.

REFRIGERATORS AND HEAT PUMPS

We all know from experience that heat flows in the direction of decreasingtemperature, that is, from high-temperature regions 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 heatfrom a low-temperature region to a high-temperature one requires specialdevices called **refrigerators.**

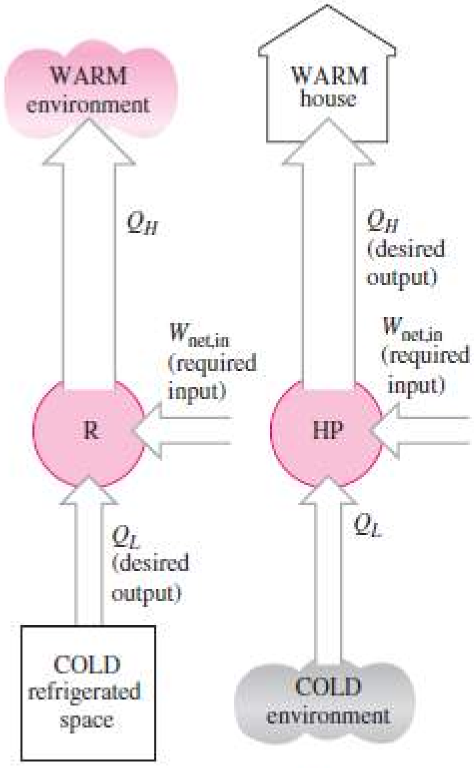
Refrigerators are cyclic devices, and the working fluids used in the refrigerationcycles are called **refrigerants.** A refrigerator is shown schematicallyin Fig. 5–27. Here QL is the magnitude of the heat removed from the refrigeratedspace at temperature TL ,$QH$is the magnitude of the heat rejected tothe warm space at temperature TH , and Wnet,in is the net work input to therefrigerator. QL and QH represent magnitudes andthus are positive quantities.

Another device that transfers heat from a low-temperature medium to ahigh-temperature one is the **heat pump.** Refrigerators and heat pumps areessentially the same devices; they differ in their objectives only. The objectiveof a refrigerator is to maintain the refrigerated space at a low temperatureby removing heat from it. Discharging this heat to a higher-temperaturemedium is merely a necessary part of the operation, not the purpose. Theobjective of a heat pump, however, is to maintain a heated space at a hightemperature. This is accomplished by absorbing heat from a low-temperaturesource, such as well water or cold outside air in winter, and supplying thisheat to a warmer medium such as a house (Fig. 5–27*b*).

The performance of refrigerators and heat pumps is expressed in terms ofthe **coefficient of performance** (COP), defined as

$$\mathrm{COP_R} = \frac{\text{Desired output}}{\text{Required input}} = \frac{\text{Cooling effect}}{\text{Work input}} = \frac{Q_L}{W_{\mathrm{net,in}}}$$

$$\mathrm{COP_{HP}} = \frac{\text{Desired output}}{\text{Required input}} = \frac{\text{Heating effect}}{\text{Work input}} = \frac{Q_H}{W_{\mathrm{net,in}}}$$



(*a*) Refrigerator (*b*) Heat pump

FIGURE 5–27 The objective of a refrigerator is toremove heat (*QL*) from the coldmedium; the objective of a heat pumpis to supply heat (*QH*) to a warmmedium.

These relations can also be expressed in the rate form by replacing thequantities *QL*, *QH*, and *W*net,in by *QH*, and *W*net,in, respectively. Notice thatboth COPR and COPHP can be greater than 1. A comparison of for fixed values of *QL* and *QH*. This relation implies that COPHP _ 1 sinceCOPR is a positive quantity. That is, a heat pump functions, at worst, as aresistance heater, supplying as much energy to the house as it consumes. Inreality, however, part of *QH* is lost to the outside air through piping andother devices, and COPHP may drop below unity when the outside air temperatureis too low. When this happens, the system normally switches to thefuel (natural gas, propane, oil, etc.) or resistance-heating mode.

The *cooling capacity* of a refrigeration system—that is, the rate of heatremoval from the refrigerated space—is often expressed in terms of **tons ofrefrigeration.** The capacity of a refrigeration system that can freeze 1 ton(2000 lbm) of liquid water at 0°C (32°F) into ice at 0°C in 24 h is said to be1 ton. One ton of refrigeration is equivalent to 211 kJ/min or 200 Btu/min.
The cooling load of a typical 200-m2 residence is in the 3-ton (10-kW)range.

THE REVERSED CARNOT CYCLE

The Carnot cycle is a totally reversible cycle thatconsists of two reversible isothermal and two isentropic processes. It has themaximum thermal efficiency for given temperature limits, and it serves as astandard against which actual power cycles can be compared.Since it is a reversible cycle, all four processes that comprise the Carnotcycle can be reversed. Reversing the cycle does also reverse the directionsof any heat and work interactions. The result is a cycle that operates in thecounterclockwise direction on a T-s diagram, which is called the **reversedCarnot cycle.** A refrigerator or heat pump that operates on the reversedCarnot cycle is called a **Carnot refrigerator** or a **Carnot heat pump.**

Consider a reversed Carnot cycle executed within the saturation dome of arefrigerant, as shown in Fig. 5–28. The refrigerant absorbs heat isothermallyfrom a low-temperature source at TL in the amount of QL (process 1-2), iscompressed isentropically to state 3 (temperature rises to TH), rejects heatisothermally to a high-temperature sink at TH in the amount of QH (process3-4), and expands isentropically to state 1 (temperature drops to TL). Therefrigerant changes from a saturated vapor state to a saturated liquid state inthe condenser during process 3-4.

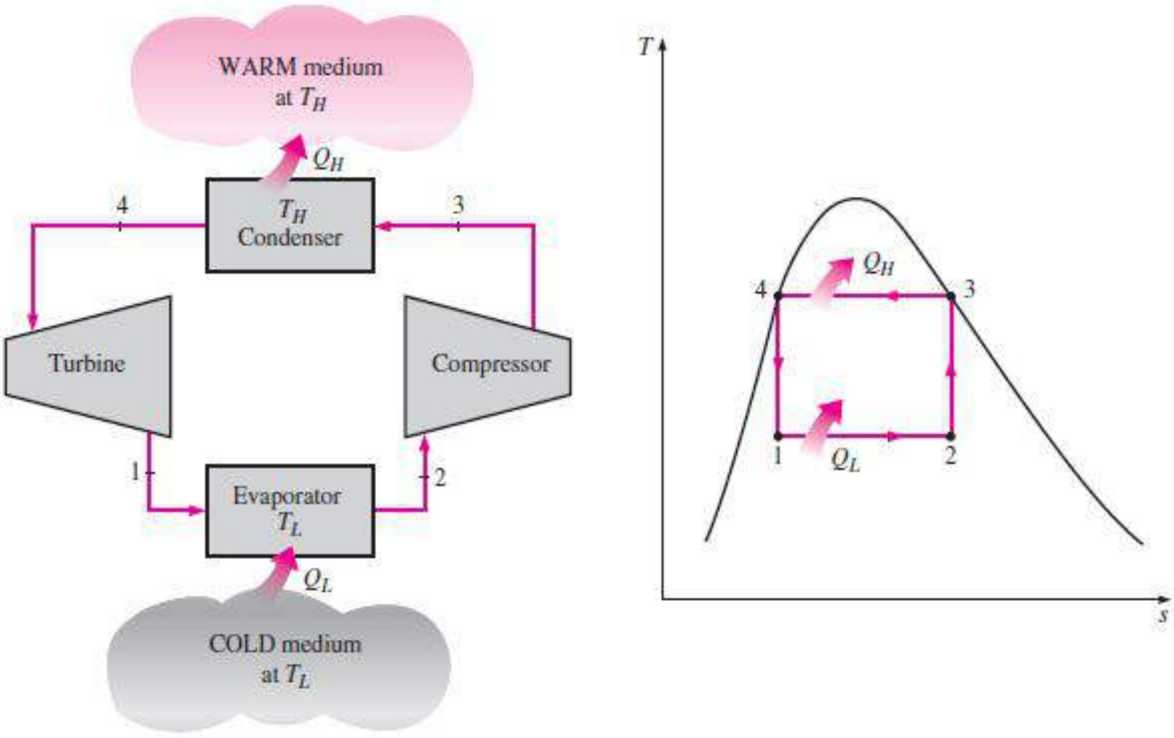


FIGURE 5–28: Schematic of a Carnot refrigerator and T-s diagram of the reversed Carnot cycle.

The reversed Carnot cycle is the *most efficient* refrigeration cycle operatingbetween two specified temperature levels. Therefore, it is natural to look at itfirst as a prospective ideal cycle for refrigerators and heat pumps. If we could,we certainly would adapt it as the ideal cycle. As explained below, however,the reversed Carnot cycle is not a suitable model for refrigeration cycles.

The two isothermal heat transfer processes are not difficult to achieve inpractice since maintaining a constant pressure automatically fixes the temperatureof a two-phase mixture at the saturation value. Therefore, processes1-2 and 3-4 can be approached closely in actual evaporators and condensers.However, processes 2-3 and 4-1 cannot be approximated closely in practice.
This is because process 2-3 involves the compression of a liquid–vapor mixture,which requires a compressor that will handle two phases, and process4-1 involves the expansion of high-moisture-content refrigerant in a turbine

THE IDEAL VAPOR-COMPRESSIONREFRIGERATION CYCLE

Many of the impracticalities associated with the reversed Carnot cycle canbe eliminated by vaporizing the refrigerant completely before it is compressedand by replacing the turbine with a throttling device, such as anexpansion valve or capillary tube. The cycle that results is called the **idealvapor-compression refrigeration cycle,** and it is shown schematically andon a *T-s* diagram in Fig. 11–3. The vapor-compression refrigeration cycle isthe most widely used cycle for refrigerators, air-conditioning systems, andheat pumps. It consists of four processes:

1-2 Isentropic compression in a compressor
2-3 Constant-pressure heat rejection in a condenser
3-4 Throttling in an expansion device
4-1 Constant-pressure heat absorption in an evaporator

In an ideal vapor-compression refrigeration cycle, the refrigerant enters thecompressor at state 1 as saturated vapor and is compressed isentropically tothe condenser pressure. The temperature of the refrigerant increases duringthis isentropic compression process to well above the temperature of the surroundingmedium. The refrigerant then enters the condenser as superheatedvapor at state 2 and leaves as saturated liquid at state 3 as a result of heatrejection to the surroundings. The temperature of the refrigerant at this stateis still above the temperature of the surroundings.

The saturated liquid refrigerant at state 3 is throttled to the evaporatorpressure by passing it through an expansion valve or capillary tube. Thetemperature of the refrigerant drops below the temperature of the refrigeratedspace during this process. The refrigerant enters the evaporator at state4 as a low-quality saturated mixture, and it completely evaporates byabsorbing heat from the refrigerated space. The refrigerant leaves the evaporatoras saturated vapor and reenters the compressor, completing the cycle.

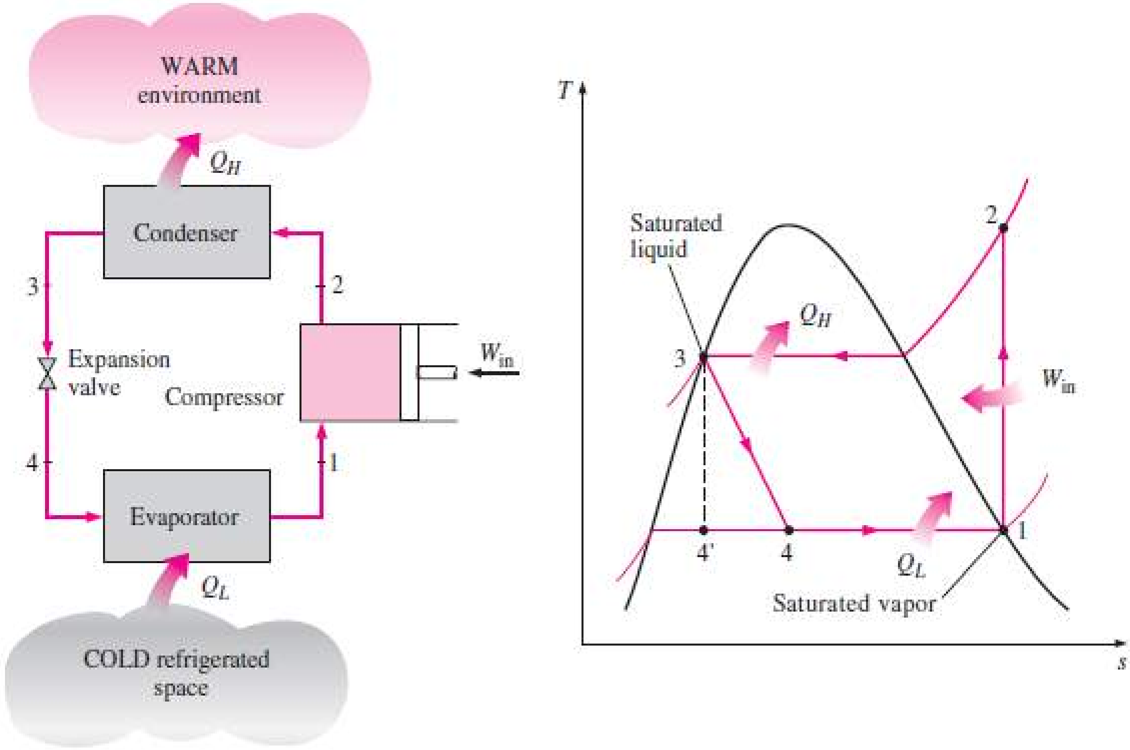


FIGURE 5–29 Schematic and *T*-*s* diagram for the ideal vapor-compression refrigeration cycle.

Liquefaction of Gases

The liquefaction of gases has always been an important area of refrigerationsince many important scientific and engineering processes at cryogenic temperatures(temperatures below about _100°C) depend on liquefied gases.Some examples of such processes are the separation of oxygen and nitrogenfrom air, preparation of liquid propellants for rockets, the study of material properties at low temperatures, and the study of some exciting phenomenasuch as superconductivity.

At temperatures above the critical-point value, a substance exists in thegas phase only. The critical temperatures of helium, hydrogen, and nitrogen(three commonly used liquefied gases) are _268, _240, and _147°C,respectively. Therefore, none of these substances exist in liquid form at atmospheric conditions. Furthermore, low temperatures of this magnitudecannot be obtained by ordinary refrigeration techniques.

Makeup gas is mixed with the uncondensed portion of the gas from theprevious cycle, and the mixture at state 2 is compressed by a multistagecompressor to state 3. The compression process approaches an isothermalprocess due to intercooling. The high-pressure gas is cooled in an aftercoolerby a cooling medium or by a separate external refrigeration system tostate 4. The gas is further cooled in a regenerative counter-flow heatexchanger by the uncondensed portion of gas from the previous cycle tostate 5, and it is throttled to state 6, which is a saturated liquid–vapor mixturestate. The liquid (state 7) is collected as the desired product, and thevapor (state 8) is routed through the regenerator to cool the high-pressuregas approaching the throttling valve. Finally, the gas is mixed with freshmakeup gas, and the cycle is repeated.

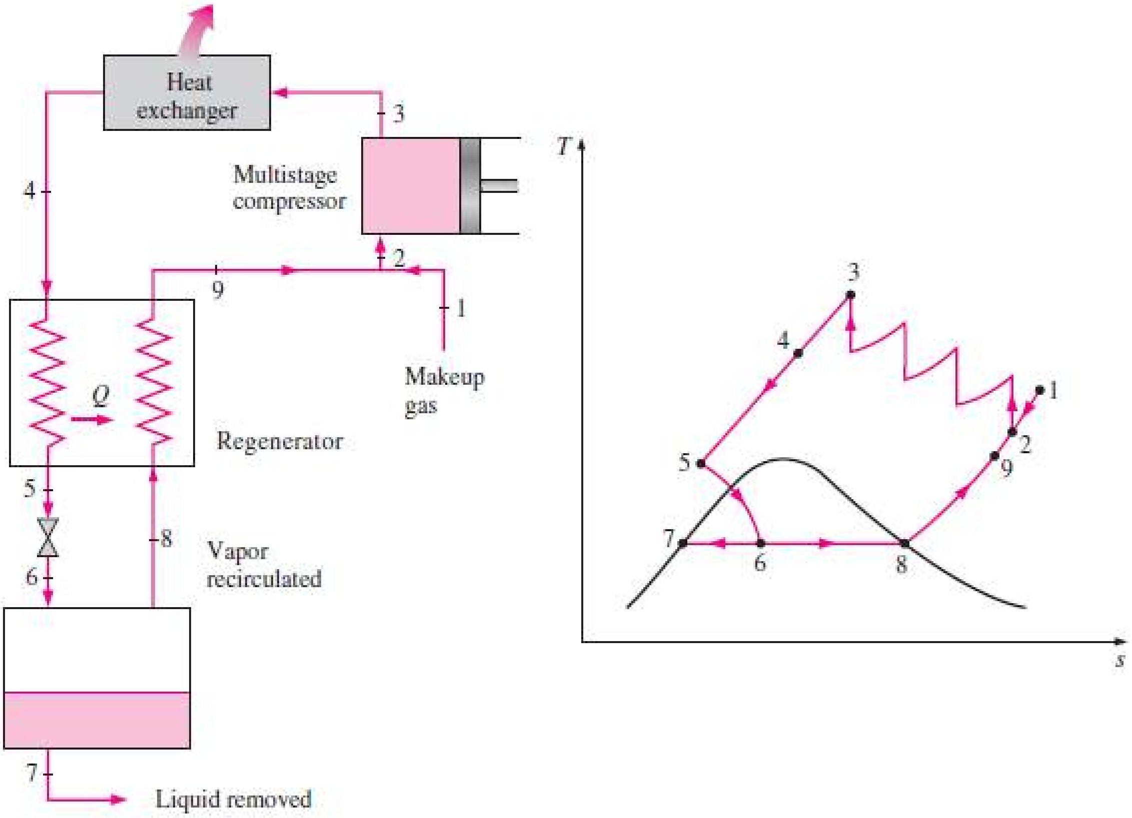


FIGURE 5–30 Linde-Hampson system for liquefying gases.

This and other refrigeration cycles used for the liquefaction of gases canalso be used for the solidification of gases.

11–8 GAS REFRIGERATION CYCLES

The Carnot cycle (the standard of comparison forpower cycles) and the reversed Carnot cycle (the standard of comparisonfor refrigeration cycles) are identical, except that the reversed Carnot cycleoperates in the reverse direction. This suggests that the power cycles discussedin earlier chapters can be used as refrigeration cycles by simplyreversing them. In fact, the vapor-compression refrigeration cycle is essentiallya modified Rankine cycle operating in reverse.

Another example is thereversed Stirling cycle, which is the cycle on which Stirling refrigeratorsoperate. In this section, we discuss the *reversed Brayton cycle,* better knownas the **gas refrigeration cycle.**

Consider the gas refrigeration cycle shown in Fig. 5–31. The surroundingsare at *T*0, and the refrigerated space is to be maintained at *TL*. The gasis compressed during process 1-2. The high-pressure, high-temperature gasat state 2 is then cooled at constant pressure to *T*0 by rejecting heat to thesurroundings. This is followed by an expansion process in a turbine, duringwhich the gas temperature drops to *T*4. (Can we achieve the cooling effectby using a throttling valve instead of a turbine?) Finally, the cool gasabsorbs heat from the refrigerated space until its temperature rises to *T*1.

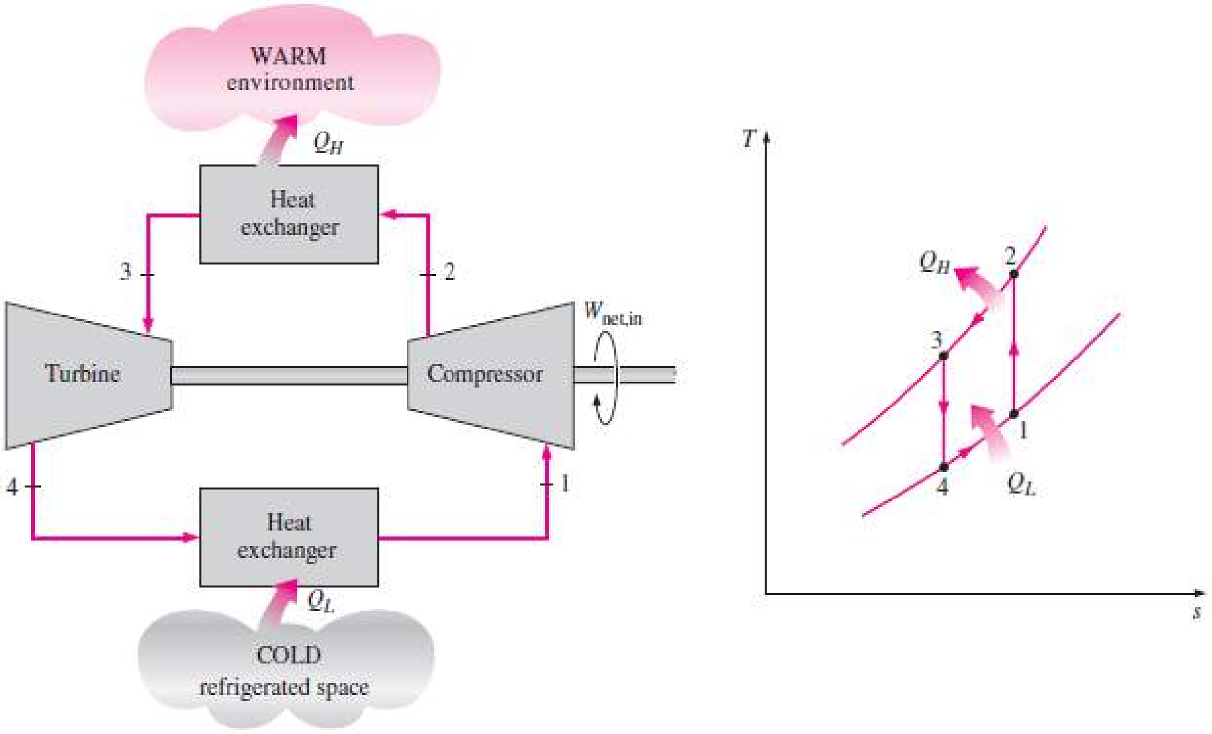
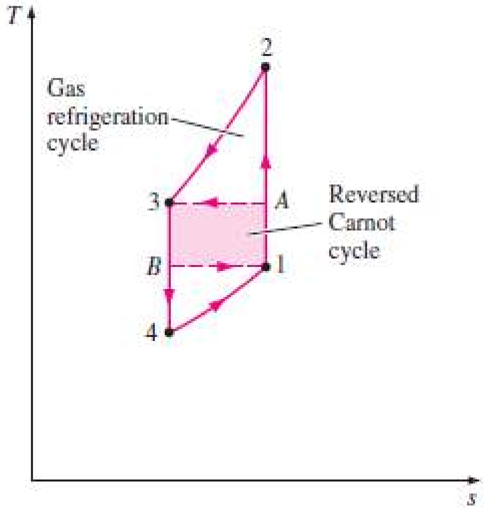


FIGURE 5–31 Simple gas refrigeration cycle.

All the processes described are internally reversible, and the cycle executedis the *ideal* gas refrigeration cycle. In actual gas refrigeration cycles,the compression and expansion processes deviate from the isentropic ones,and *T*3 is higher than *T*0 unless the heat exchanger is infinitely large.On a *T-s* diagram, the area under process curve 4-1 represents the heatremoved from the refrigerated space, and the enclosed area 1-2-3-4-1 representsthe net work input. The ratio of these areas is the COP for the cycle,which may be expressed as

$$\text{COP}_R = \frac{q_L}{w_{\text{net,in}}} = \frac{q_L}{w_{\text{comp,in}} - w_{\text{turb,out}}}$$



The gas refrigeration cycle deviates from the reversed Carnot cyclebecause the heat transfer processes are not isothermal. In fact, the gas temperaturevaries considerably during heat transfer processes. Consequently, thegas refrigeration cycles have lower COPs relative to the vapor-compressionrefrigeration cycles or the reversed Carnot cycle. This is also evident fromthe *T*-*s* diagram in Fig. 5–31. The reversed Carnot cycle consumes a fractionof the net work (rectangular area 1*A*3*B*) but produces a greater amountof refrigeration (triangular area under *B*1).

Despite their relatively low COPs, the gas refrigeration cycles have twodesirable characteristics: They involve simple, lighter components, whichmake them suitable for aircraft cooling, and they can incorporate regeneration,which makes them suitable for liquefaction of gases and cryogenic applications. An open-cycle aircraft cooling system is shown in Fig. 11–18.Atmospheric air is compressed by a compressor, cooled by the surroundingair, and expanded in a turbine. The cool air leaving the turbine is thendirectly routed to the cabin.

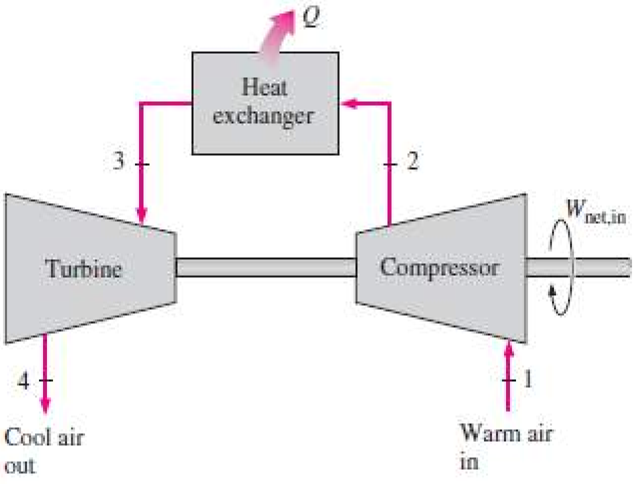


FIGURE 5–32 An open-cycle aircraft cooling system.

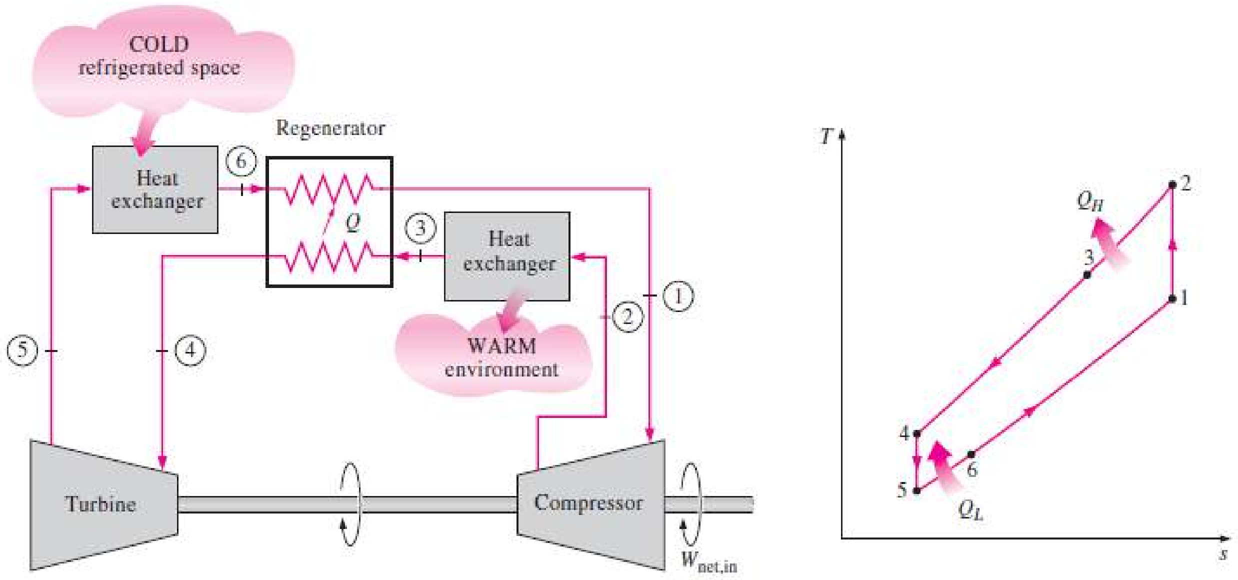


FIGURE 5–33 Gas refrigeration cycle with regeneration.

The regenerative gas cycle is shown in Fig. 5–33. Regenerative coolingis achieved by inserting a counter-flow heat exchanger into the cycle. Withoutregeneration, the lowest turbine inlet temperature is $T0$, the temperatureof the surroundings or any other cooling medium. With regeneration, thehigh-pressure gas is further cooled to $T4$ before expanding in the turbine.Lowering the turbine inlet temperature automatically lowers the turbine exittemperature, which is the minimum temperature in the cycle. Extremely lowtemperatures can be achieved by repeating this process.