杜 军 主编

通用专业英语

（热能动力工程、建筑环境与设备工程、轮机工程）

HEUP 哈尔滨工程大学出版社

通用专业英语(热能动力工程、建筑环境与设备工程、轮机工程)

General Specialized English for Thermal Power Engineering, Building Environment and Equipment Engineering and Marine Engineering

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内 容 简 介

本书的编写充分考虑到建筑环境与设备工程、轮机工程与热能动力工程专业的相互渗透性和互补性。全书共分6章:第1章介绍了学科基础知识,包括工程热力学、流体力学和传热学基础;第2章介绍了储能系统基础知识及应用;第3章介绍了太阳能基本知识及其热利用;第4章介绍了燃气轮机基本知识及燃烧器设计;第5章介绍了燃烧基本知识及垃圾焚烧应用;第6章介绍了供热通风及空调系统相关知识。将专业英语词汇融入到国际前沿的学科研究领域,这样既能使读者学习专业英语,又能拓展他们的专业知识,还能使其了解有关领域的最新研究进展。

本书可作为高等院校热能动力工程、建筑环境与设备工程、轮机工程的专业英语教材,也可供相关专业人员参考使用。

前　言

　　本书旨在培养学生在专业英语方面的阅读、翻译、表达及写作能力,提高学生的科技英语水平与国际学术交流能力。供理工科大学热能动力工程、建筑环境与设备工程、轮机工程及相关专业本科生或研究生使用,也可供同等英语程度相关领域的科技人员使用。建议课时数为 28 至 32 学时,授课内容可以根据不同专业需求进行选讲。

　　本书充分考虑到热能动力工程、建筑环境与设备工程和轮机工程三个专业的相互渗透性和互补性,将专业英语词汇融入到国际前沿的学科研究领域,既能使学生们充分学习专业英语,又能拓展他们的专业知识,还能使其了解有关领域的最新研究进展。本书共分 6 个章节:第 1 章介绍学科基础知识,包括工程热力学、流体力学和传热学基础;第 2 章介绍储能系统基本知识及应用;第 3 章介绍太阳能基本知识及其热利用;第 4 章介绍燃气轮机基本知识及燃烧器设计;第 5 章介绍燃烧基本知识及垃圾焚烧应用;第 6 章介绍供热通风及空调系统相关知识。书中每章内容后附有专业术语词汇表、难点注释及习题,具有选材新颖、内容广泛、体系分明等特点。

　　本书编写融入了所有编者的辛勤劳动,由杜军任主编,杨兴林任副主编,共同拟定编写提纲。杜军编写了第 1 章、第 5 章和第 6 章,杨兴林编写了第 2 章,赵忠超编写了第 4 章,黄波编写了第 3 章,刘宏、刘勇和姜波共同编写了专业词汇表、难点注释、练习题及书后答案,刘宏和姜波编辑了公式和表格,并对全书进行了统稿。

　　最后,谨向全书编写有关人员、出版人员表示感谢。由于编者水平有限,书中缺点和不妥之处在所难免,敬请专家和读者批评指正。

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编者

2015 年 1 月

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CONTENTS

Chapter 1 Introduction to Thermal Sciences

This chapter is an abbreviated version of standard thermodynamics, fluid mechanics and heat transfer texts, covering topics that engineering students are most likely to need in their professional lives. The thermal sciences portion of this chapter is based on *Fundamentals of Thermal Fluid Sciences* by Yunus Cengel and Robert Turner, the thermodynamics portion is based on the text *Thermodynamics: An Engineering Approach* by Y. A. Çengel and M. A. Boles, the fluid mechanics portion is based on *Fluid Mechanics: Fundamentals and Applications* by Y. A. Çengel and J. M. Cimbala and the heat transfer portion is based on *Heat Transfer: A Practical Approach* by Y. A. Çengel. Most contents are practically independent of each other and can be covered in any order. The book is well suited for curriculums that have a common introductory course.

It is recognized that all topics of thermodynamics, fluid mechanics and heat transfer cannot be covered adequately in a typical three-semester-hour course. Therefore, sacrifices must be made from depth if not from the breadth. Selecting the right topics and finding the proper level of depth and breadth are no small challenge for the instructors, and the book is intended to serve as the ground for such selection. Students in a combined thermal-fluids course can gain a basic understanding of energy and energy interactions, various mechanisms of heat transfer and fundamentals of fluid flow. Such a course can also instill in students the confidence and the background to do further reading of their own and to be able to communicate effectively with specialists in thermal-fluid sciences.

The word *thermal* stems from the Greek word *therme*, which means *heat*. Therefore, thermal sciences can loosely be defined as the sciences that deal with heat. The recognition of different forms of energy and its transformations has forced this definition to be broadened. The physical sciences that deal with energy and the transfer, transport and conversion of energy are usually referred to as thermal-fluid sciences or just thermal sciences today. Traditionally, the thermal-fluid sciences are studied under the subcategories of thermodynamics, heat transfer and fluid mechanics. In this book we present the basic principles of these sciences and apply them to situations that the engineers are likely to encounter in their practices.

The design and analysis of most thermal systems such as power plants, automotive engines and refrigerators involve all categories of thermal-fluid sciences as well as other sciences. For example, designing the radiator of a car involves the determination of the amount of energy transfer from a knowledge of the properties of the coolant using *thermodynamics*, the determination of the size and shape of the inner tubes and the outer fins using *heat transfer*, and the determination of the size and type of the water pump using *fluid mechanics*. Of course the determination of the materials and the thickness of the tubes require the use of material science as well as strength of materials. The reason for studying different sciences separately is simply to facilitate learning without being overwhelmed. Once the basic principles are mastered, they can then be synthesized by sol-

ving comprehensive real-world practical problems. But first we will present an overview of thermal-fluid sciences.

1. 1　Fundamentals of Engineering Thermodynamics

Thermodynamics can be defined as the science of *energy*. Although everybody has a feeling of what energy is, it is difficult to give a precise definition for it. Energy can be viewed as the ability to cause changes.

The name *thermodynamics* stems from the Greek words *therme* (heat) and *dynamis* (power), which is most descriptive of the early efforts to convert heat into power. Today the same name is broadly interpreted to include all aspects of energy and energy transformations, including power production, refrigeration and relationships among the properties of matter.

One of the most fundamental laws of nature is the conservation of energy principle. It simply states that during an interaction, energy can change from one form to another but the total amount of energy remains constant. That is, energy cannot be created or destroyed. A rock falling off a cliff, for example, picks up speed as a result of its potential energy being converted to kinetic energy. The conservation of energy principle also forms the backbone of the diet industry: a person who has a greater energy input (food and drinks) than energy output (exercise and metabolism with environmental conditions) will gain weight (store energy in the form of tissue and fat), and a person who has a smaller energy input than output will lose weight. The change in the energy content of a body or any other system is equal to the difference between the energy input and the output, and the energy balance is expressed as $E_{in} - E_{out} = \Delta E$.

The first law of thermodynamics is simply an expression of the conservation of energy principle, and it asserts that *energy* is a thermodynamic property. The second law of thermodynamics asserts that energy has *quality* as well as *quantity*, and actual processes occur in the direction of decreasing quality of energy. For example, a cup of hot coffee left on a table eventually cools to room temperature, but a cup of cool coffee in the same room never gets hot by itself. The high-temperature energy of the coffee is degraded (transformed into a less useful form at a lower temperature) once it is transferred to the surrounding air.

Although the principles of thermodynamics have been in existence since the creation of the universe, thermodynamics did not emerge as a science until the construction of the first successful atmospheric steam engines in England by Thomas Savery in 1697 and Thomas Newcomen in 1712. These engines were very slow and inefficient, but they opened the way for the development of a new science. The first and second laws of thermodynamics emerged simultaneously in the 1850s, primarily out of the works of William Rankine, Rudolph Clausius and Lord Kelvin (formerly William Thomson). The term *thermodynamics* was first used in a publication by Lord Kelvin in 1849. The first thermodynamic textbook was written in 1859 by William Rankine, a professor at the University of Glasgow.

It is well known that a substance consists of a large number of particles called *molecules*. The

properties of the substance naturally depend on the behavior of these particles. For example, the pressure of a gas in a container is the result of momentum transfer between the molecules and the walls of the container. But it does not need to know the behavior of the gas particles to determine the pressure in the container. It would be sufficient to attach a pressure gage to the container. This macroscopic approach to the study of thermodynamics that does not require a knowledge of the behavior of individual particles is called classical thermodynamics. It provides a direct and easy way to the solution of engineering problems. A more elaborate approach, based on the average behavior of large groups of individual particles, is called statistical thermodynamics. This microscopic approach is rather involved and is used in this text only in the supporting role.

1.1.1   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. For example, the commonly used mercury-in-glass thermometer is based on the expansion of mercury with temperature. Temperature is also measured by using several other temperature-dependent properties.

Two bodies reaching thermal equilibrium after being brought into contact in an isolated enclosure. 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 which is at a different temperature, heat is transferred from the body at higher temperature to the lower one until both bodies attain the same temperature (Fig. 1.1). At that point, the heat transfer stops, the two bodies are called to have reached thermal equilibrium. The equality of temperature is the only requirement for thermal equilibrium.
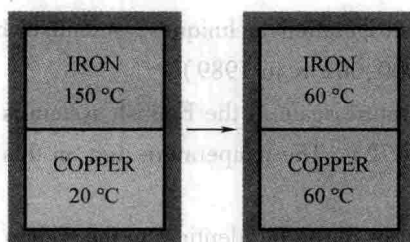

Figure 1.1   The example of the zeroth law of thermodynamics

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.

Temperature Scales

Temperature scales enable us to use a common basis for temperature measurements, and several have been introduced throughout history. All temperature scales are based on some easily reproducible states such as the freezing and boiling points 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 vapor at 1 atm pressure is said to be at the ice point, and a mixture of liquid water and water vapor (with no air) in equilibrium at 1 atm pressure is said to be at the steam point.

The temperature scales used in the SI and in the English system today are the Celsius scale (formerly called the *centigrade scale*; it was renamed after the Swedish astronomer A. Celsius, 1702—1744, who devised it in 1948) and the Fahrenheit scale (named after the German instrument maker G. Fahrenheit, 1686—1736), respectively. On the Celsius scale, the ice and steam points are assigned the values of 0 °C and 100 °C, respectively. The corresponding values on the Fahrenheit scale are 32 °F and 212 °F. These are often referred to as *two-point scales* since temperature values are assigned at two different points.

In thermodynamics, it is very desirable to have a temperature scale that is independent of the properties of any substances. Such a temperature scale is called the thermodynamic temperature scale, which is developed later in conjunction with the second law of thermodynamics. The thermodynamic temperature scale in the SI is the Kelvin scale, named after Lord Kelvin (1824—1907). The temperature unit on this scale is the kelvin, which is designated by K (the degree symbol was officially dropped from kelvin in 1967). The lowest temperature on the Kelvin scale is 0 K. Using non-conventional refrigeration techniques, scientists have approached absolute zero kelvin (they achieved 0.000,000,002 K in 1989).

The thermodynamic temperature scale in the English system is the Rankine scale, named after William Rankine (1820—1872). The temperature unit on this scale is the rankine, which is designated by R.

A temperature scale that turns out to be identical to the Kelvin scale is the ideal-gas temperature scale. The temperatures on this scale are measured using a constant-volume gas thermometer, which is a rigid vessel filled with a gas basically, usually hydrogen or helium, at low pressure. This thermometer is based on the principle that *at low pressures, the temperature of a gas is proportional to its pressure at constant volume.* That is, the temperature of a gas fixed volume varies *line-*

arly with pressure at sufficiently low pressures. Then the relationship between the temperature and the pressure of the gas in the vessel can be expressed as

$$T = a + bP \tag{1.1}$$

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 this relation by immersing the rigid vessel of the gas thermometer into the medium and measuring the gas pressure when thermal equilibrium is established between the medium and the gas in the vessel whose volume is held constant.

An 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 the steam points) and assigning suitable values to temperatures at 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 Eq. 1.1. Then the unknown temperature T of a medium corresponding to a pressure reading P can be determined from that equation by a simple calculation. The values of the constants will be different for each thermometer, depending on the type and the amount of the gas in the vessel, and the temperature values assigned at the two reference points. If the ice and steam points are assigned the values 0 and 100, respectively, the gas temperature scale will be identical to the Celsius scale. In this case the value of the constant a (which corresponds to an absolute pressure of zero) is determined to be -273.15 °C regardless of the type and the amount of the gas in the vessel of the gas thermometer. That is, on a P-T diagram, all the straight lines passing through the data points in this case will intersect the temperature axis at -273.15 °C when extrapolated, as shown in Fig. 1.2. This is the lowest temperature that can be obtained by a gas thermometer, and thus we can obtain an *absolute gas temperature scale* by assigning a value of zero to the constant a in Eq. 1.1. In that case Eq. 1.1 reduces to $T = bP$, and thus we need to specify the temperature at only *one* point to define an absolute gas temperature scale.
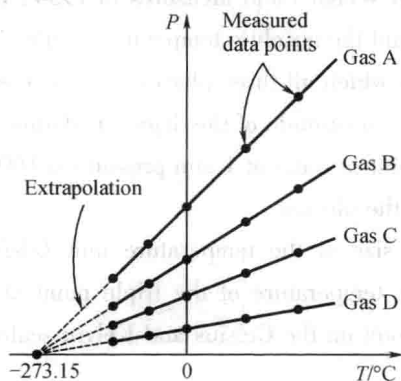
Figure 1.2 P versus T plots of the experimental data obtained from a constant-volume gas thermometer using four different gases at different pressures

It should be noted that the absolute gas temperature scale is not a thermodynamic temperature scale, since it cannot be used at very low temperatures (due to condensation) and at very high

temperatures (due to dissociation and ionization). However, absolute gas temperature is identical to the thermodynamic temperature in the temperature range in which the gas thermometer can be used, and thus we can view the thermodynamic temperature scale at this point as an absolute gas temperature scale that utilizes an "ideal" or "imaginary" gas that always acts as a low-pressure gas regardless of the temperature. If such a gas thermometer existed, it would read zero kelvin at absolute zero pressure, which corresponds to -273.15 °C on the Celsius scale (Fig. 1.3).

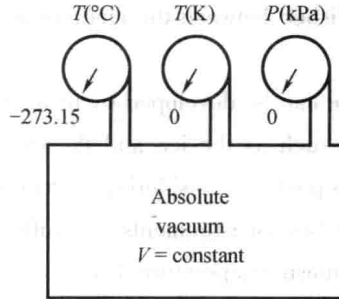**Figure 1.3　A constant-volume gas thermometer would read -273.15 °C at absolute zero pressure**

The Kelvin scale is related to the Celsius scale by

$$T(\mathrm{K}) = T(\mathrm{°C}) + 273.15 \tag{1.2}$$

The Rankine scale is related to the Fahrenheit scale by

$$T(\mathrm{R}) = T(\mathrm{°F}) + 459.67 \tag{1.3}$$

It is common practice to round the constant in Eq. 1.2 to 273 and that in Eq. 1.3 to 460.

The temperature scales in the two unit systems are related by

$$T(\mathrm{R}) = 1.8 \, T(\mathrm{K}) \tag{1.4}$$

$$T(\mathrm{°F}) = 1.8 \, T(\mathrm{°C}) + 32 \tag{1.5}$$

A comparison of various temperature scales is given in Fig. 1.4.

At the Tenth Conference on Weights and Measures in 1954, the Celsius scale was redefined in terms of a single fixed point and the absolute temperature scale. The selected single point is the *triple point* of water (the state at which all three phases of water coexist in equilibrium), which is assigned the value 0.01 °C. The magnitude of the degree is defined from the absolute temperature scale. As before, the boiling point of water at 1 atm pressure is 100.00 °C. Thus the new Celsius scale is essentially the same as the old one.

On the Kelvin scale, the size of the temperature unit *kelvin* is defined as " the fraction 1/273.16 of the thermodynamic temperature of the triple point of water, which is assigned the value of 273.16 K". The ice point on the Celsius and Kelvin scales are 0 °C and 273.15 K, respectively.

Note that the magnitudes of each division of 1 K and 1 °C are identical (Fig. 1.5). Therefore, when we are dealing with temperature differences ΔT, the temperature interval on both scales is the same. Raising the temperature of a substance by 10 °C is the same as raising it by 10 K. That is

$$T(\mathrm{K}) = \Delta T(\mathrm{°C}) \tag{1.6}$$

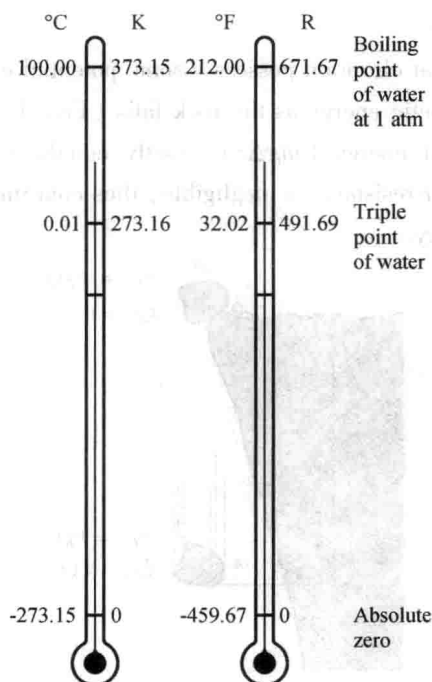$$T(\mathrm{R}) = \Delta T(\mathrm{°F}) \tag{1.7}$$

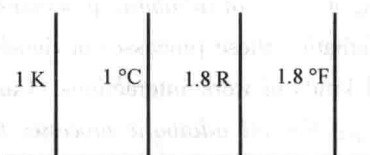Figure 1.4   Comparison of temperature scales

Figure 1.5   Comparison of magnitudes of various temperature units

Some thermodynamic relations involve the temperature T and the question often arises of whether it is in K or °C. If the relation involves temperature differences (such as $a = b\ \Delta T$), it makes no difference and either can be used. However, if the relation involves temperatures only instead of temperature differences (such as $a = bT$), then K must be used. When in doubt, it is always safe to use K because there are virtually no situations in which the use of K is incorrect, but there are many thermodynamic relations that will yield an erroneous result if °C is used.

1.1.2   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 elevation possesses some potential energy, and part of this potential energy is converted to kinetic energy as the rock falls (Fig. 1.6). Experimental data shows that the decrease in potential energy ($mg\Delta z$) exactly equals the increase in kinetic energy $[m(V_2^2 - V_1^2)/2]$ when the air resistance is negligible, thus confirming the conservation of energy principle for mechanical energy.
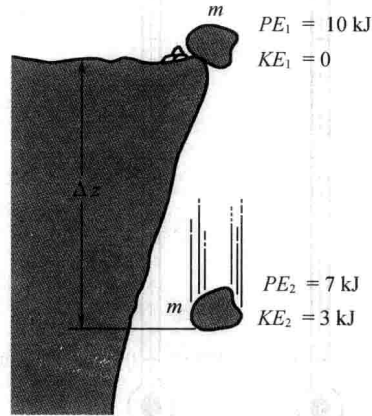
**Figure 1.6　Energy cannot be created or destroyed, it can only change forms**

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 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 an 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 re-

lation 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. 1.7). 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.

$Q_{in} = 5$ kJ
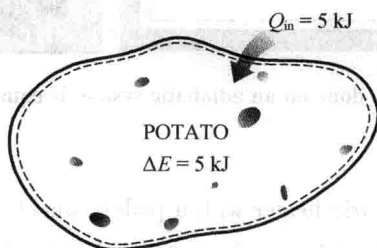
POTATO
$\Delta E = 5$ kJ

Figure 1.7   The increase in the energy of a potato in an oven is equal to the amount of heat transferred to it

As another example, consider the heating of water in a pan on top of a range (Fig. 1.8). 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 transfers to water, which is 12 kJ.

$Q_{out} = 3$ kJ

$\Delta E = Q_{net} = 12$ kJ

$Q_{in} = 15$ kJ
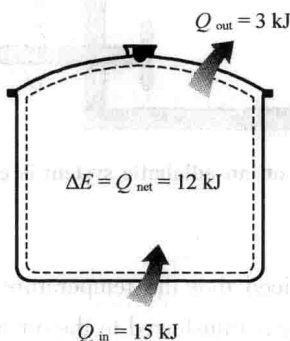
Figure 1.8   In the absence of any work interactions, the energy change of a system is equal to the net heat transfer

Now consider a well-insulated (i.e., adiabatic) room heated by an electric heater as our system (Fig. 1.9). 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 to the increase in energy of the system.

(Adiabatic)

$W_{\text{in}} = 5 \text{ kJ}$

$\Delta E = 5 \text{ kJ}$

Battery
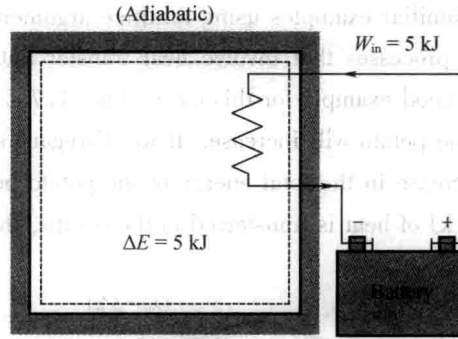
Figure 1.9　The work (electrical) done on an adiabatic system is equal to the increase in the energy of the system

Next, let us replace the electric heater with a paddle wheel (Fig. 1.10). As a result of the stirring process, the energy of the system will increase. Again, since there is no heat interaction between the system and its surroundings ($Q = 0$), the shaft work done on the system must show up as an increase in the energy of the system.

(Adiabatic)

$\Delta E = 8 \text{ kJ}$

$W_{\text{sh, in}} = 8 \text{ kJ}$
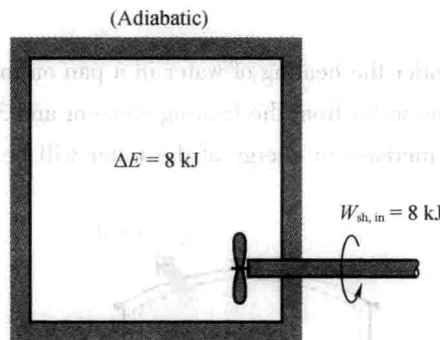
Figure 1.10　The work (shaft) done on an adiabatic system is equal to the increase in the energy of the system

Many of you have probably noticed that the temperature of air rises when it is compressed (Fig. 1.11). This is because energy is transferred to the air in the form of boundary work. In the absence of any heat transfer ($Q = 0$), the entire boundary work will be stored in the air as part of its total energy. The conservation of energy principle requires that the increase in the energy of the system be equal to the boundary work done on the system again.

We can extend these discussions to systems that involve various heat and work interactions simultaneously. For example, if a system gains 12 kJ of heat during a process while 6 kJ of work is done on it, the increase in the energy of the system during that process is 18 kJ (Fig. 1.12). That is, the change in the energy of a system during a process is simply equal to the net energy transfer to (or from) the system.

$W_{b,in} = 10$ kJ

$\Delta E = 10$ kJ
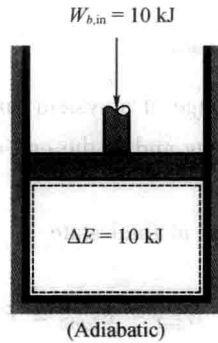
(Adiabatic)

Figure 1. 11 **The work (boundary) done on an adiabatic system is equal to the increase in the energy of the system**

$Q_{out} = 3$ kJ

$\Delta E = (15-3)+6$
$= 18$ kJ

$W_{sh,in} = 6$ kJ

$Q_{in} = 15$ kJ
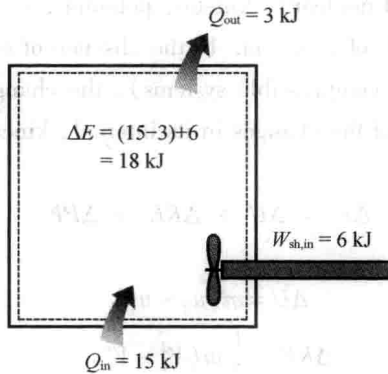
Figure 1. 12 **The energy change of a system during a process is equal to the *net* work and heat transfer between the system and its surroundings**

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 leaving system during that process*. That is

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

or

$$E_{in} - E_{out} = \Delta E_{system} \qquad (1.8)$$

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.