



普通高等教育“十二五”规划教材

A Concise Course in University Physics

Second Edition Volume 2

大学物理简明教程 (英文版)

第二版

下册

王安安 伏云昌 主编



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伏云昌 王安安 陈劲波 编写

科学出版社

北 京

内 容 简 介

本书是按照《理工科类大学物理课程教学基本要求(2010年版)》修订的,涵盖了所有A类的内容,选择了部分B类的内容,如非惯性系、质心、气体的范德瓦尔斯方程、玻尔兹曼分布、几何光学、固体能带论和激光简介等.为方便上、下两学期的教学安排,本次改版将原三册改编为上、下两册.全书共19章:上册为力学和电磁学;下册为热学、振动与波动、光学和近代物理.本书配有双语课件光盘.

本书可供理工科非物理专业112~128学时双语教学使用,也可供在某一部分内容进行双语教学试点选用,还可供对英文物理感兴趣的广大读者自学或作参考书用.

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第二版前言

根据《理工科类大学物理课程教学基本要求(2010年版)》对教学过程的基本要求第5点“双语教学——在保证教学效果的前提下,有条件的学校可开展物理课程的双语教学,以提高学生查阅外文资料和科技外语交流的能力”,为满足坚持双语教学的师生对英文物理教材的需要,我们对 *A Concise Course in University Physics* (《大学物理简明教程(英文版)》,1998~2000年出版)进行了部分修订. 按2010年版教学基本要求,第二版涵盖了所有A类的内容,保留原有的并新选了部分B类内容,如增加了非惯性系和惯性力、质心和质心运动定理、热力学第零定律、范德瓦尔斯方程及几何光学等内容. 为方便目前工科物理两学期教学内容的安排,将第一版的三册改编为上、下两册,供非物理专业112~128学时双语物理教学之用. 第二版保留了第一版的主要特色,力求系统完整、理论叙述准确、文字简明易懂,以达到教师好用、学生易学的目的. 因篇幅所限,第二版省去了各章总结和三个阅读材料.

全书19章编写分工如下:

上册:力学(第1~5章)由王安安编写;电磁学(第6~9章)主要由樊则宾编写.

下册:热学(第10,11章)由王安安编写;振动与波动部分第12、14章及第15章中的几何光学由伏云昌编写,第15章中的波动光学由陈劲波编写,第13章由王安安和陈劲波编写;近代物理(第16~19章)由伏云昌编写. 王安安和伏云昌负责全书的统稿、修改和定稿工作. 为方便本书读者进行双语多媒体课堂教学,伏云昌编制了配套教学课件.

北京大学陆果教授对第二版全书进行了认真的审定,并提出了宝贵的意见和建议,谨此致以诚挚谢意! 编者也对昆明理工大学校领导多年来对我们工作的支持表示衷心的感谢!

双语教学贵在坚持,我们编写及使用英文物理教材的初衷未改,通过双语教学提高学生的综合素质始终是我们的目标,编者愿与致力于这项事业的同行共同努力,持之以恒,为物理教学改革尽绵薄之力!

因时间仓促、水平有限,书中难免有不当之处,恳请同行与读者提出宝贵意见.

编 者

2011年12月于昆明

第一版前言摘要

我们正处在一个高新技术飞速发展、科技信息量激增、知识更新加快、国际交流日益广泛的时代。我国的进一步改革开放,社会主义市场经济的建立都要求高校毕业生有更强的适应能力,在人才市场上,有效强的外语应用能力、交流型、综合型的毕业生供不应求。在这种形势下,我们的高等教育正向着重视素质教育的方向转变,而素质与能力是密切相关的,素质的培养要以一定的知识和能力为基础,其中包括独立获取知识的能力。毋庸置疑,直接用外语为工具获取知识、进行交流的能力是人才素质的一个重要方面。

然而,由于历史原因、文化背景、经济基础、外语教育模式和各类师资外语水平等诸多因素的影响,我们在应用外语进行教学方面的基础性工作十分薄弱。在普通高校本科生教育中,教材和教学过程基本上只使用中文这一单一语种,实际上已经制约了学生应用外语(主要是英语)获取知识能力的发展。为了改变这种现状,跟上时代的步伐,试用英文教材,使用英语进行教学的改革便应运而生了。

本教材的编写是编者主持的“试用英文物理教材”教改试点工作的继续,也是编者大学物理教学经验的总结。从国外引进的教材,虽有诸多优点,但在系统上与我国的大学物理教学基本要求不完全对应,为了满足师生对英文物理教材的需要,编写一套根据我国工科物理教学基本要求,顺应工科物理教学改革形势,反映编者在多年物理教学实践中总结出来的教学方法与经验,与我们的学生在一、二年级的英文水平相适应的英文“简明物理学教程”的计划就提到日程上来了,这就是我组织编写这套教材的初衷。这套英文物理教材是1996年经国家教委批准列入正式出版计划的。本教材可供普通高等工科院校本科生物理课130~140学时使用。

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全书21章,具体编写分工如下:

第一册:力学(第1~5章),分子运动论和热力学基础(第6、7章)由王安安编写。

第二册:电学及稳恒电流的磁场(第8~10章)由吴光敏编写,磁介质和电磁感应(第11、12章)由樊则宾编写,麦克斯韦方程组(第13章)由王安安编写。

第三册:机械振动(第14章)由伏云昌编写,机械波(第15章)由王安安和陈劲波编写。电磁振荡与电磁波(第16章)由吴光敏编写,波动光学部分(第17章)中干涉与衍射由陈劲波编写,光的偏振由樊则宾编写。近代物理部分(第18~21章)由伏云昌编写。

第一册绝大部分插图由刘富华用计算机绘制,其余插图由李俊昌教授绘制,封面也由李俊昌设计,谨此致以诚挚谢意。

编写大学英文物理教材是一种大胆的尝试,由于编者水平有限,错误疏漏之处在所难免,希望同行和读者批评指正。我们相信这本教材的出版将对物理教学的现代化和物理教学与国际接轨作出有益的贡献。

编 者

1997年5月于昆明

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Part Three Thermodynamics

Chapter 10

The Kinetic Theory of Gases

We come to observations on heat phenomena, such as the temperature of a gas system, the pressure exerted by the gas on the walls of the container, the heat transformation and the work done by a gas system, which we will discuss in detail in this course.

Unlike in mechanics, the object in this part is an intrinsically complex system of a vast number of particles, rather than the simple system of a few particles or objects. There are two distinct methods to study heat phenomena. The first, called the macroscopic method or classical thermodynamics, is established on the basis of macroscopic variables measured as pressure, volume and temperature, deals with the properties of bulk matter, takes no account of the fact that all matter is made up of atoms and molecules.

The second, called microscopic method or kinetic theory, is based on atomic model of matter. The basic assumption of kinetic theory is that the measurable properties of matter like temperature, pressure and volume of a gas system reflect the combined actions of countless numbers atoms and molecules. Kinetic theory attempts to relate the microscopic properties of atoms or molecules which are not directly measurable, such as the mass, velocity, momentum and kinetic energy of a molecule, to the measurable macroscopic parameters of the system by means of the statistic method—investigating the average behavior of the microscopic parameters that characterizes the individual molecule. We shall focus our attention on this method in this chapter.

10.1 The Zeroth Law of Thermodynamics

10.1.1 The thermal equilibrium

Temperature is a central concept in thermodynamics. In order to define it strictly not depending on our subjective sense of touch in which objects feel hot or cold, we must understand the concept about thermal equilibrium first. When we put object *A* and object *B* together in an isolate system without any disturber from environment (refer to section 3-6), and let them in thermal contact so that the thermal energy exchanges from one to another. The experimental facts show that the two objects will eventually reach a stable state at which the thermal energy exchange stops. Then the two objects are defined to be in thermal equilibrium.

Furthermore, if we want to determine whether or not system *A* and system *B* are in thermal equilibrium but without being contact with each other, usually we can employ a third sys-

tem C . Suppose that system A and system C are in thermal equilibrium; while system B and system C are in thermal equilibrium respectively, experiments indicate that system A and system B are in thermal equilibrium too. Even if we let them in thermal contact there would be no thermal energy exchange between them at all.

10.1.2 The zeroth law of thermodynamics, Definition of temperature

The experimental facts mentioned above are summed up in the zeroth law of thermodynamics: If system A and system B are each in thermal equilibrium with a third system T , then they are in thermal equilibrium with each other. It is called thermal equilibrium law also. According to this law, the objects in thermal equilibrium have a common property so that we define this property as temperature. When two objects are in thermal equilibrium, their temperatures are equal and vice versa.

The zeroth law of thermodynamics allows us to make use of a thermometer, say, to employ the third system T mentioned above to measure the temperature of other system, the only thing we need to do is to calibrate it.

10.2 State Parameters, Equilibrium State, and Ideal Gas Law

10.2.1 State parameters of gas

In mechanics, to describe the states of a mechanical system, the position and velocity were needed. In thermodynamics, the states of a thermodynamic system are described by pressure p , volume V , and temperature T . The p , V , T are called state parameters of the system.

Because molecules are constantly in random motion, the container is filled with gases at any time, so, the volume of gas actually is the volume of the container and in SI unit, its unit is m^3 . The pressure p is the force exerted by a gas on per unit area of the wall of the container, and it must surely be related to the steady drumbeat of molecules on the walls of the container. The SI unit of pressure is pascal (Pa), $1 \text{ Pa} = 1 \text{ N/m}^2$. Meanwhile, the following units are often used

$$1 \text{ cmHg} = 10 \text{ mmHg} = 1.333 \times 10^3 \text{ Pa}$$

$$1 \text{ atm} = 76 \text{ cmHg} = 1.013 \times 10^5 \text{ Pa}$$

$$1 \text{ kgf/cm}^2 = 9.8 \times 10^4 \text{ Pa}$$

We have made definition of temperature in the last section, based on thermal equilibrium, or the zeroth law of thermodynamics. The unit of temperature is Kelvin called the thermodynamic scale or absolute temperature, labeled as K, defined as the ratio of the temperature of triple-point of water to 273.15. It is that the temperature of triple-point of water, 273.15 K is defined as a constant temperature in the system of thermodynamic scale. The Celsius temperature scale is also often used, labeled as $^\circ\text{C}$, and it employs a degree of the same magnitude as that of the Kelvin scale but its zero point is shifted by 273.15 degrees. Thus if t denotes the

Celsius temperature, and T denotes the Kelvin temperature (i. e. absolute temperature), we have

$$T/\text{K} = t/^{\circ}\text{C} + 273.15$$

The pressure, volume and temperature of a given system, all can be determined by measurements, so that we call them as macroscopic properties of that system.

10.2.2 Equilibrium state and equilibrium process

When the temperature and pressure are the same at all points in a system, the system is said to be in an equilibrium state. Hence if heat is added at some point to a system in an equilibrium state, we must wait until the process of heat transfer within the system has brought about a new uniform temperature before the system is again in equilibrium state.

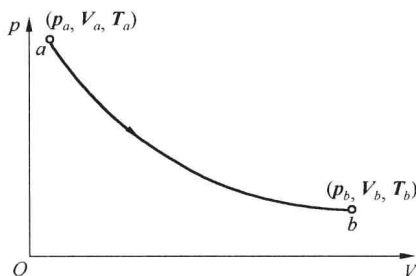


Fig. 10-1 The thermal equilibrium states and process are represented on the p - V diagram

An equilibrium state of a thermodynamic system can be described by its pressure p , volume V and temperature T ; for example, $(p_a, V_a, T_a,)$ describes state a , (p_b, V_b, T_b) describes state b . An equilibrium state can be represented by a dot on the pressure-volume diagram (briefly as p - V diagram) shown in Fig. 10-1.

The operation of changing the system from its initial state to its final state is called a thermodynamic process. During such processes, energy may be transferred into (or out of) the system from the environment. We assume that such transfers are carried out very slowly, so that the system remains approximately in thermodynamic equilibrium at all stages. This process is called thermal equilibrium process. A thermal equilibrium process is represented by a smooth curve on the p - V diagram. In Fig. 10-1, curve ab shows an equilibrium process from state a to state b .

10.2.3 The ideal gas law

Experiments show that there is a simple relationship among the macroscopic parameters that characterize a gas in equilibrium state. If we take one mole samples of various gases (oxygen, hydrogen, nitrogen or any others), confine them in containers of identical volume, and hold them at the same temperature, we find that their measured pressure are nearly—though not exactly—the same. If we repeat the measurements at lower gas densities, we find that these small differences in the measurements tend to disappear. Further experiments indicate that, at low enough densities, all real gases tend to obey the relation

$$\frac{pV}{T} = \text{constant} \quad (10-1)$$

Eq. (10-1) holds when the sum mass of the system remains unchanged. At normal temperature and pressure, real gases obey this law quite well; but if a real gas is compressed to an

excessively high density, then its behavior will deviate from this law. An ideal gas is a gas that obeys Eq. (10-1) exactly. Although there is no such thing in nature as a truly ideal gas, all gasses approach the ideal state at low enough density, that is, under conditions in which molecules are far enough apart. Thus, ideal gas is an ideal model that allows us to gain useful insights into the limiting behavior of real gas.

Suppose p_0 , V_0 and T_0 are the parameters of a gas under the standard condition, Eq. (10-1) can be rewritten as

$$\frac{pV}{T} = \frac{p_0 V_0}{T_0}$$

where $p_0 = 1.013 \times 10^5 \text{ Pa}$, $T_0 = 273.15 \text{ K}$. If M is the sum mass of the system, and μ is its mass per mole, the number of mole is then M/μ . Because the volume of one mole is identical for all gases, i. e., $V_{\text{mol}} = 22.4 \times 10^{-3} \text{ m}^3/\text{mol}$, the volume of a gas of mass M is therefore $V_0 = V_{\text{mol}}(M/\mu)$. Substituting those into Eq. (10-1), we have

$$\frac{pV}{T} = \frac{M}{\mu} \frac{p_0 V_{\text{mol}}}{T_0}$$

Because $p_0 V_{\text{mol}}/T_0$ is a constant for all gases, called mole gas constant, labeled as R , so that, the equation becomes

$$pV = \frac{M}{\mu} RT \quad (10-2)$$

which is called ideal gas law or state equation of ideal gas, and it holds for equilibrium states. The numerical value of R depends on the units in which p , V , M and T are expressed:

(1) In SI system, the unit of pressure is Pa, $1 \text{ Pa} = 1 \text{ N/m}^2$, the unit of volume is m^3 , and the temperature is in K, the value of R is therefore

$$R = \frac{p_0 V_{\text{mol}}}{T_0} = \frac{1.013 \times 10^5 \text{ N/m}^2 \times 22.4 \times 10^{-3} \text{ m}^3/\text{mol}}{273.15 \text{ K}} = 8.31 \text{ J}/(\text{mol} \cdot \text{K})$$

(2) Note that heat has the same unit as energy, sometimes we use cal as the unit of energy, $1 \text{ J} = 0.24 \text{ cal}$, so that

$$R = 8.31 \text{ J}/(\text{mol} \cdot \text{K}) \times 0.24 \text{ cal/J} = 2 \text{ cal}/(\text{mol} \cdot \text{K})$$

(3) If volume is commonly expressed in liters (L), pressure in atmospheres and temperature in K, thus

$$R = \frac{1 \text{ atm} \times 22.4 \text{ L/mol}}{273.15 \text{ K}} = 0.082 \text{ atm} \cdot \text{L}/(\text{mol} \cdot \text{K})$$

Example 10-1 The volume of an oxygen tank is 50 L. As oxygen is withdrawn from the tank, the reading of a pressure gauge drops from $2.17 \times 10^6 \text{ Pa}$ to $7.9 \times 10^5 \text{ Pa}$ and the temperature of the gas remaining in the tank drops from 30°C to 10°C .

(1) How many kilograms of oxygen were withdrawn?

(2) How many kilograms of oxygen were in the tank originally?

(3) What volume would be occupied by the oxygen withdrawn from the tank at a pressure of 1 atm and a temperature of 20°C ?

Solution (1) Let us express volume in cubic meters, and temperature in Kelvins. Thus $50 \text{ L} = 0.05 \text{ m}^3$, $30 \text{ L} = 0.3 \text{ m}^3$, $30^\circ\text{C} = 303 \text{ K}$, and $10^\circ\text{C} = 283 \text{ K}$. According to Eq. (10-2) the

initial number of moles is

$$n_1 = \frac{p_1 V}{RT_1} = \frac{2.17 \times 10^6 \text{ Pa} \times 0.05 \text{ m}^3}{8.31 \text{ J/(mol} \cdot \text{K)} \times 303 \text{ K}} = 43.1 \text{ mol}$$

The original mass is therefore

$$m_1 = 43.1 \text{ mol} \times 32 \text{ g/mol} = 1379 \text{ g} = 1.379 \text{ kg}$$

(2) The number of moles remaining in the tank is

$$n_2 = \frac{p_2 V}{RT_2} = \frac{7.9 \times 10^5 \text{ Pa} \times 0.05 \text{ m}^3}{8.31 \text{ J/(mol} \cdot \text{K)} \times 283 \text{ K}} = 16.8 \text{ mol}$$

And the mass remaining is

$$m_2 = 16.8 \text{ mol} \times 32 \text{ g/mol} = 538 \text{ g} = 0.538 \text{ kg}$$

The mass withdrawn is, therefore

$$\Delta m = m_1 - m_2 = 1.379 \text{ kg} - 0.538 \text{ kg} = 0.841 \text{ kg}$$

(3) The number of moles withdrawn is

$$\Delta n = 43.1 \text{ mol} - 16.8 \text{ mol} = 26.3 \text{ mol}$$

And the volume occupied by the oxygen withdrawn would be

$$V = \frac{nRT}{p} = \frac{26.3 \text{ mol} \times 8.31 \text{ J} \cdot \text{mol/K} \times 293 \text{ K}}{1.01 \times 10^5 \text{ Pa}} = 0.634 \text{ m}^3$$

10.3 Essential Concepts of the Kinetic Theory of Gases

The hypothesis supported by thousands of physical and chemical observations is that all matter is composed of tiny particles called molecules. The following essential concepts of kinetic theory of gases are given on the basis of the experimental observations.

10.3.1 All matters consist of a very large number of molecules, molecules are separated

Experimental evidence for assuming that the gas contains many molecules is related to the determination of Avogadro's number:

A mole of any pure substance contains a definite number of identical molecules. The number of molecules in a mole is called Avogadro's constant, denoted by N_0 ,

$$N_0 = 6.022 \times 10^{23} \text{ mol}^{-1}$$

The enormously large value of the Avogadro's constant suggests how tiny and how numerous atoms must be in a mole of air. Yet, if these molecules were spread uniformly over the surface of the earth, there would be about 120 000 of them in every square centimeter.

The fact that all matter is compressible indicates that there are separations between molecules. For instance, air can be pumped into the tire of the bicycle or car; the gas enclosed in a cylinder can be compressed and its volume can be decreased to a small part, say, one seventh of the original volume, etc. The fact that a drop of red ink diffuses into the water until uniformly mixture is also an evidence that the molecules are separated.

In other parts of physics and chemistry it is important to consider the structure of the molecule, but it is not necessary at this point. The smallest molecules are of the order of 10^{-10} m in size; the largest is at least 10 000 times of this order.

10.3.2 Molecules are constantly in random motion

Experimental evidence for the random motion of molecules is based on the observation of what we call Brownian motion. In 1827, the English botanist Robert Brown used a microscope to observe the motion of pollen grains suspended in water. The pollen appeared to dance about in an erratic fashion. At first, Brown thought that the pollen was alive and that its motion was some sort of dance! Subsequent studies of liquid suspensions of various inanimate particles convinced observers that the liquid itself was responsible for the erratic motion. Fig. 10-2 shows an example of such Brownian motion.



Fig. 10-2 The trace of the Brownian motion of a pollen grain

Eventually, the irregular motion of the suspended particles was explained in detail by assuming that the liquid was composed of molecules in random motion. The suspended particles are continuously bombarded on all sides by the molecules of the fluid. The numbers of molecules striking opposite sides of the particle in any short time interval, being determined by chance, will not be exactly equal. Because of these fluctuations, a randomly directed unbalanced force will act on the suspended particle, accounting for its “Brownian dance”.

It is evident from experiments that the “random motion” is related to the temperature: as temperature increases, molecule motion becomes more vigorous. So, the random motion of molecules is also called heat motion.

10.3.3 There is interaction between molecules

One essential characteristic of a molecule is the force that exists between it and a neighboring molecule. There is, of course, a force of gravitational attraction between every pair of molecules, but it turns out that this is negligible in comparison with the forces to be considered now. The forces that hold the molecules of a liquid (or solid) together are chiefly of electrical origin and do not follow a simple inverse-square law.

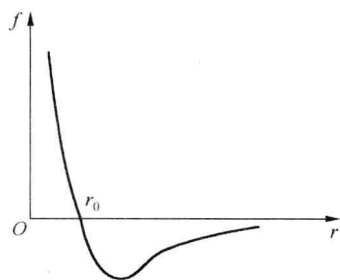


Fig. 10-3 The interaction force between molecules varies with the distance between them

The interaction force is related to the separation between molecules. When the separation of the molecules is large, as in a gas, the force is extremely small and attractive. The attractive force increases as the gas is compressed and its molecules are brought closer together. But since a tremendous pressure is needed to compress a liquid (i. e., to force its molecules closer together than their normal spacing in the liquid state), we conclude that at separations only slightly less than the dimension of a molecule the force is repulsive and relatively large. Thus the force must vary with separation in somewhat the fashion as shown in Fig. 10-3. At large separations

the force is small and attractive. As the molecules are brought closer together, the force of attraction becomes larger, passes through a maximum, and then decreases to zero at a separation r_0 . When the distance between the molecules is less than r_0 , the force becomes repulsive. The order of the distance r_0 is about 10^{-10} m, and when the distance between molecules is larger than 10^{-9} m, the interaction force becomes so small that it can actually be neglected, so, the force between molecules is short-range force. Note that in Fig. 10-3, the attractive force is represented as negative while the repulsive force as positive.

10.4 The Representation of Pressure for Ideal Gas

The first kinetic theory problem here is to find the connection between the macroscopic parameter of a gas—pressure and the microscopic parameter, the speed of the molecules of the gas. We shall use the statistical method by considering the average motion of the vast number of molecules in the system.

10.4.1 The microscopic model of an ideal gas

In order to derive the pressure equation, we must introduce a microscopic model of an ideal gas, including the following assumptions:

- (1) The size of molecules may be thought of as so small that it can be considered negligible in comparison with their average distance.
- (2) Molecules exert no force to each others except for the instantaneous impulsive force during the collisions with the wall of the container holding the gas and the collisions with each others.
- (3) Molecules are in constantly random motion, they collide with one another and also with the walls of the container frequently, and these collisions are perfectly elastic.
- (4) The motion of an individual molecule obeys Newton's laws of motion.

We can make a conclusion about this microscopic model that the molecule of an ideal gas is like infinitesimal elastic particle having no interaction with the others.

It is necessary to introduce a statistical assumption for the great number of molecules in random motion: Under equilibrium state, molecules have equal possibilities (opportunities) to move in all different possible directions. There is no any preference of direction of motion to the others, that is, the numbers of molecules moving in all different directions are equal; otherwise, the molecules would be concentrated at a certain part of the container, this is opposite to the assumption for the system in equilibrium state. When a gas is in equilibrium state, its density is all the same everywhere in the container. The statistical meaning of this assumption is that, as a consequence of the random motion, the average values of the squares of the molecule's velocity components are equal, that is

$$\overline{v_x^2} = \overline{v_y^2} = \overline{v_z^2} \quad (10-3)$$

10.4.2 Derivation of pressure equation of ideal gas

Now, we are ready to discuss the pressure of a gas against the walls of its container, which is due to the impacts of the molecules on the walls during the collision. Suppose there