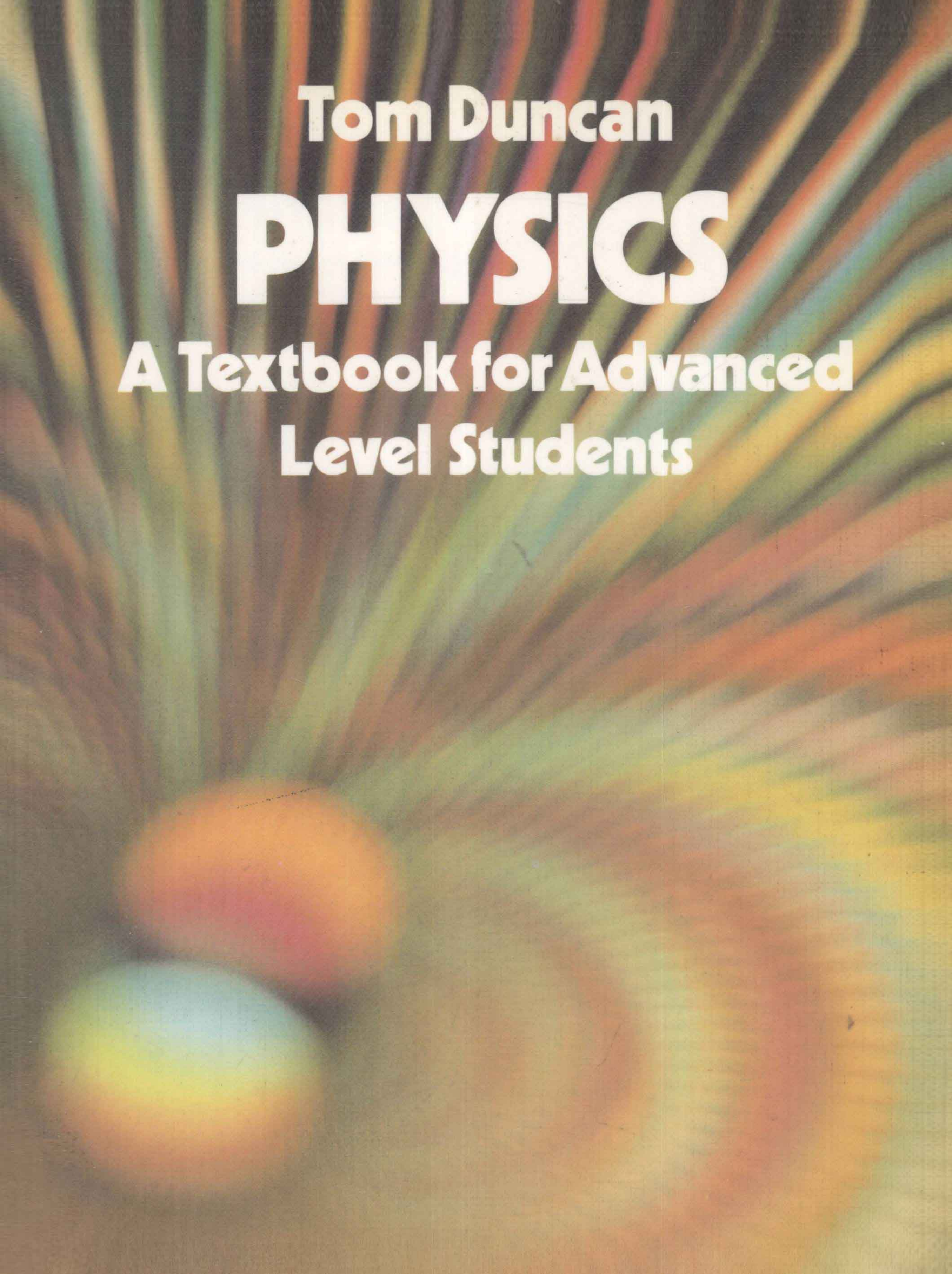
Tom Duncan

PHYSICS

A Textbook for Advanced Level Students



PHYSICS

A Textbook for Advanced Level Students

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PHYSICS

A Textbook for Advanced Level Students

By the same author

Electronics and Nuclear Physics
Practical Modern Physics
Exploring Physics 1–5
Advanced Physics: Materials and Mechanics
Advanced Physics: Fields, Waves and Atoms
Physics for Today and Tomorrow
Adventures with Electronics
Adventures with Physics
Adventures with Microelectronics
Adventures with Digital Electronics
Success in Electronics

Preface

This book is a combined volume of the second editions of the two sixth form textbooks, *Advanced Physics: Materials and Mechanics* and *Advanced Physics: Fields, Waves and Atoms*. These are substantially the same as the original editions but have been updated throughout to be in line with recent syllabus revisions, particularly the chapter on electronics.

Once again I would like to express my gratitude to those who helped with the original manuscripts, Dr J. W. Warren, Mr J. Dawber, Professor J. Stringer, Dr J. C. Gibbings, Mr B. Baker, Drs B. L. N. and H. M. Kennett. Dr M. G. Pellatt of B.D.H. Chemicals Ltd. very kindly gave advice on the new section dealing with liquid crystals. I am also grateful to those students and teachers, particularly Mr J. V. Thornton of St. Mary's College, Crosby, who have written to me with suggestions for improving the earlier books.

Thanks are also due to my wife not only for typing the revision but for continuing to suffer uncomplainingly the inevitable domestic inconvenience that writing or revising a textbook involves.

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T.D.

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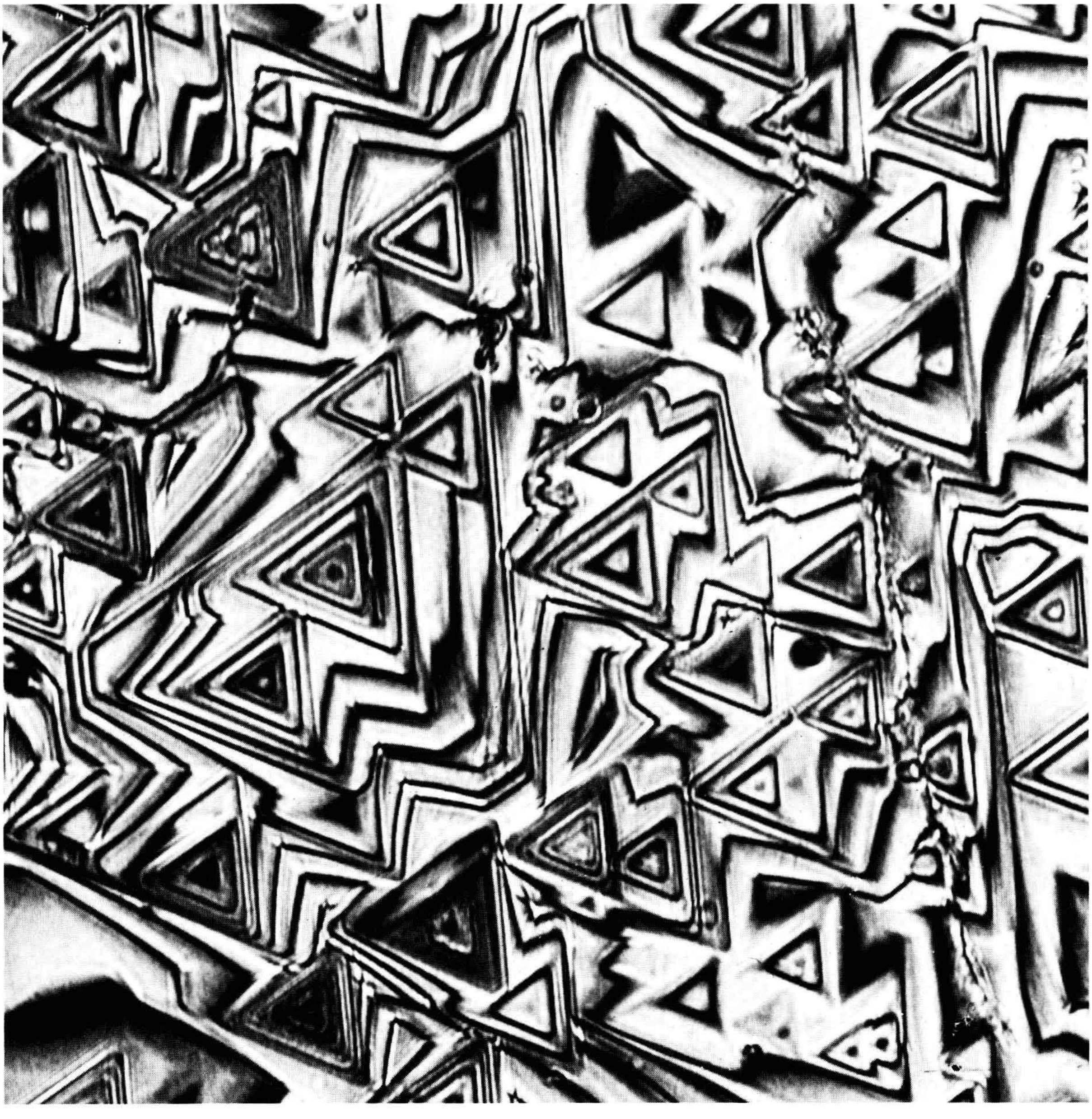
Part 1 MATERIALS



1 Structure of materials

Materials science

Advances in technology depend increasingly on the development of better materials. This is especially true of those industries engaged in aircraft production, space projects, telecommunications, computer manufacture and nuclear power engineering. Structural materials are required to be stronger, stiffer and lighter than existing ones. In some cases they may have to withstand high temperatures or exposure to intense radioactivity. Materials with very precise electrical, magnetic, thermal, optical or chemical properties are also demanded.

A great deal has been known for many years about materials that are useful in everyday life and industry. For example, the metallurgist has long appreciated that alloys can be made by adding one metal to another or that heating, cooling or hammering metals changes their mechanical behaviour. Materials *technology* is a long-established subject. The comparatively new subject of materials *science* is concerned with the study of materials as a whole and not just with their physical, chemical or engineering properties. As well as asking *how* materials behave, the materials scientist also wants to know *why* they behave as they do. Why is steel strong, glass brittle and rubber extensible? To begin to find answers to such questions has required the drawing together of ideas from physics, chemistry, metallurgy and other disciplines.

The deeper understanding of materials which we now have has come from realizing that the properties of matter in bulk depend largely on the way the atoms are arranged when they are close together. Progress has been possible because of the invention of instruments for 'seeing' finer and finer details. The electron microscope, which uses beams of electrons instead of beams of light as in the optical microscope, reveals structure just above the atomic level. The field ion microscope and X-ray apparatus allow investigation at that level.

The scanning electron microscope, Fig. 1.1*a*, is a development from the electron microscope and 'scans' a surface with electrons in the way that a television screen is scanned. It gives higher magnifications and much greater depth of focus than optical microscopes using reflected light. It is useful for examining the surfaces of semiconductors, the hairlike fibres and 'whiskers' that are so important in the manufacture of the new generation of composite materials, man-made fibres, and corroded and fractured surfaces. A view of the end of a torn wire (× 75) is shown in Fig. 1.1*b* and of lead–tin telluride crystals (× 30) in Fig. 1.1*c*.

Materials science is a rapidly advancing subject with exciting prospects for the future. Its importance lies in the help it can give with the selection of materials for particular applications, with the design of new materials and with the improvement of existing ones. The strength of even a tea cup has been improved by research into ceramics, as Fig. 1.1*d* shows.

Atoms, molecules and Brownian motion

The modern atomic theory was proposed in 1803 by John Dalton, an English schoolmaster. He thought of atoms as tiny, indivisible particles, all the atoms of a given element being exactly alike and different from those of other elements in behaviour and mass. By making simple assumptions he explained the gravimetric (i.e. by weight) laws of chemical combination but failed to account satisfactorily for the volume relationships which exist between combining gases. This required the introduction in 1811 by the Italian scientist, Amedeo Avogadro, of the molecule as the smallest particle of an element or compound capable of existing independently and consisting of two or more atoms, not necessarily identical. Thus, whilst we could only have atoms of elements, molecules of both elements and compounds were possible.

(a) (b) (c) (d)

Fig. 1.1

At the end of the nineteenth century some scientists felt that evidence, more direct than that provided by the chemist, was needed to justify the basic assumption that atoms and molecules exist. In 1827 the Scottish botanist, Robert Brown, discovered that fine pollen grains suspended in water were in a state of constant movement, describing small, irregular paths but never stopping. The effect, which has been observed with many kinds of small particles suspended in both liquids and gases, is called *Brownian motion*. It is now considered to be due to the unequal bombardment of the suspended particles by the molecules of the surrounding medium.

Very small particles are essential. If the particle is fairly large, the impacts, occurring on every side and irregularly, will cancel out and there will be no average resultant force on the particle. Only if the particle is small will it suffer impacts with a few hundred molecules at any instant and the chances of these cancelling out are proportionately less. It is then likely that for a short time most of the impacts will be in one direction; shortly afterwards the direction will have changed. The phenomenon can be observed in smoke in a small glass cell which is illuminated strongly from one side and viewed from above with a low-power microscope, Fig. 1.2. How would the random motion be affected by (*i*) cooling the air to a low temperature, (*ii*) using smaller smoke particles?

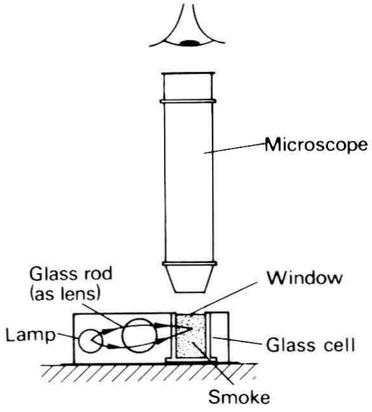


Fig. 1.2

The effect, on its own, does not offer conclusive proof for molecules but it clearly reveals that on the microscopic scale there is great activity in matter which macroscopically (on a large scale) appears to be at rest. The theory of the motion was worked out by Einstein and is found to correspond closely with observation. His basic assumption was that the suspended particles have the same mean kinetic energy as the molecules of the fluid and so behave just like very large molecules. Their motions should therefore be similar to those of the fluid molecules.

The Avogadro constant: mole

Atomic and *molecular masses* (previously called *atomic* and *molecular weights*) give the masses of atoms and molecules compared with the mass of another kind of atom. Originally the hydrogen atom was taken as the standard with atomic mass 1 since it has the smallest mass. In 1960 it was agreed internationally, for various reasons, to base atomic and molecular masses on the atom of carbon (more precisely, on the carbon 12 isotope ${}^{12}_{6}C$). On the carbon scale the atomic mass of carbon 12 is taken as exactly 12 making that of hydrogen 1.008 and of oxygen 16.00. Nowadays atomic masses are found very accurately using a *mass spectrometer*.

It follows from the definition of atomic mass that any number of atoms of carbon will have, near enough, 12 times the mass of the *same* number of atoms of hydrogen. Therefore any mass of hydrogen, say 1 g, will contain the same number of atoms as 12 g of carbon. In general, the atomic mass of any element expressed in grams, contains the same number of atoms as 12 g of carbon. This number is thus, by definition, a constant. It is called the *Avogadro constant* and is denoted by L. Its accepted experimental value is 6.02×10^{23}.

The number of molecules in the molecular mass in grams of a substance is also (because of the way molecular masses are defined) the same for all substances and equal to the Avogadro constant. There are, therefore, 6.02×10^{23} molecules in 2 g of hydrogen (molecular mass 2) and in 18 g of water (molecular mass 18). In fact, the Avogadro constant is useful when dealing with other particles besides atoms and molecules and a quantity which contains 6.02×10^{23} particles is called, especially by chemists, a *mole*. We can thus have a mole of atoms, a mole of molecules, a mole of ions, a mole of electrons, etc.—all contain 6.02×10^{23} particles. We must always have a mole of some kind of particle and so

$$L = 6.02 \times 10^{23} \text{ particles per mole}$$

It should be noted that the mole (abbreviation mol) is based on the gram and not the kilogram, which makes it an anomaly in the SI system of units. Sometimes, however, it is expressed in terms of the number of particles per kilogram-mole and its value then is 6.02×10^{26}.

The Avogadro constant has been measured in various ways. In an early method alpha particles emitted by a radioactive source were counted by allowing those within a small known angle to strike a fluorescent screen. Each particle produced one scintillation on the screen and if it is assumed that one particle is emitted by each radioactive atom an approximate value for L can be obtained (see question 6, p. 18). Other methods give more reliable results—one involves X-ray crystallography.

Size of a molecule

(*a*) *Monolayer experiments.* An experimental determination of the size of a molecule was made by Lord Rayleigh in 1899. He used the fact that certain organic substances, such as olive oil, spread out over a clean water surface to form very thin films.

A simple procedure for performing the experiment is to obtain a drop of olive oil by dipping the end of a loop of thin wire, mounted on a card, into olive oil, quickly withdrawing it and then estimating the diameter of the drop by holding it against a $\frac{1}{2}$ mm scale and viewing the drop and scale through a lens, Fig. 1.3*a*. If the drop is then transferred to the centre of a waxed tray overbrimming with water, the surface of which has been previously cleaned by drawing two waxed booms across it and then lightly dusted with lycopodium powder, Fig. 1.3*b*, it spreads out into a circular film pushing the powder before it. Assuming the drop is spherical, the thickness of the film can be calculated if its diameter is measured. It is found to be about 2×10^{-9} metre, i.e. 2 nanometres (2 nm).

Oil-film experiments do not necessarily prove that matter is particulate but from them we can infer that if molecules exist and if the film is one molecule thick, i.e. a monolayer, then in the case of olive oil one dimension of its molecule is 2 nm.

(*b*) *Predictions from kinetic theory of gases.* Information about the molecular world can sometimes be obtained from observations of the behaviour of matter in bulk, i.e. from macroscopic observations. Thus with the help of the kinetic theory of gases, expressions can be derived relating such properties as rate of diffusion with the size of the gas molecules involved.

(*c*) *Using the Avogadro constant.* Consider copper which has atomic mass 64 and density 9.0 g cm^{-3}. One mole of copper atoms, therefore, has mass 64 g and volume 64/9 cm^3; it contains 6.0×10^{23} atoms. The volume available to each atom is $64/(9 \times 6 \times 10^{23})$ cm^3 and the radius r of a sphere having this volume is given by

$$\frac{4}{3}\pi r^3 = \frac{64}{9 \times 6 \times 10^{23}}$$

$$\therefore \quad r = 0.14 \times 10^{-7}\ \text{cm}$$

$$= 0.14 \times 10^{-9}\ \text{m}$$

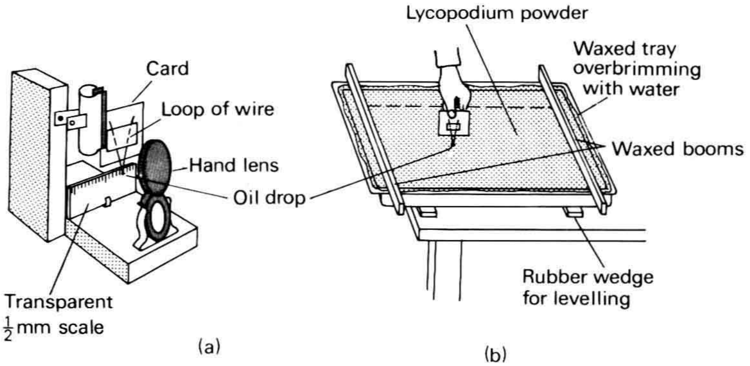
$$= 0.14\ \text{nm}$$

If copper atoms are spherical, would their radius be larger or smaller than this even if they were packed tightly? Why? A more accurate way of calculating the size of a copper atom is indicated in questions 10 to 13 on p. 18.

A word of caution is necessary regarding atomic dimensions. Nowadays atoms and molecules are no longer pictured as having hard, definite surfaces like a ball and there is, therefore, little point in trying to give their diameters too exact values; most are within the range 0.1 to 0.5 nm. Also, although we shall usually treat atoms and molecules as spheres, it is necessary on occasion to consider them as having other shapes.

Periodic table

With the passage of time the early nineteenth-century picture of an indivisible atom came to be doubted in the light of fresh information. During the 1860s chemical knowledge increased sufficiently for it to be clear that there were elements with similar chemical properties. Moreover, atomic masses were being established with greater certainty and attempts were made to relate properties and atomic masses.



(a)

(b)

Fig. 1.3

Table 1.1

Group 1	*Group 2*	*Group 3*	*Group 4*	*Group 5*	*Group 6*	*Group 7*	*Group 0*
1 Hydrogen							2 Helium
3 Lithium	4 Beryllium	5 Boron	6 Carbon	7 Nitrogen	8 Oxygen	9 Fluorine	10 Neon
11 Sodium	12 Magnesium	13 Aluminium	14 Silicon	15 Phosphorus	16 Sulphur	17 Chlorine	18 Argon

It was found that if the elements were arranged in order of increasing atomic masses then, at certain repeating intervals, elements occurred with similar chemical properties. Sometimes it was necessary to place an element of larger atomic mass before one of slightly smaller atomic mass to preserve the pattern. The first eighteen elements of this arrangement, called the Periodic table, are shown in Table 1.1. The third and eleventh (3 + 8) elements are the alkali metals lithium and sodium; the ninth and seventeenth (9 + 8) are the halogens fluorine and chlorine—here the repeating interval is eight. The serial number of an element in the table is called its *atomic number.*

The Periodic table suggests that the atoms of the elements may not be simple entities but are somehow related. There must be similarities between the atoms of similar elements and it would seem that the similarity might be due to the way they are built up.

We now believe that atoms are composed of three types of particles—protons, neutrons and electrons. (Many other subatomic particles, such as positrons, mesons and antiprotons, are known but most are short-lived and are not primary components.) Protons and neutrons are packed together into a very small nucleus which is surrounded by a cloud of electrons, the diameter of the atom as a whole being at least 10 000 times greater than that of the nucleus. The comparative masses and charges of the three basic particles are given in Table 1.2. The nucleus is positively charged and the electron cloud negatively charged but the number of protons equals the number of electrons so that the atom is electrically neutral.

The number of protons in the nucleus of an atom has been found to be the same as its atomic number which therefore means that each element in the Periodic table has one more proton and one more electron in its atom than the previous element. Hydrogen, the first element, has one proton and one electron. Helium, the second element, has two protons and two electrons. Lithium, with atomic number three, has three protons and three electrons. Neutrons are present in all nuclei except that of hydrogen.

Table 1.2

Particle	*Mass*	*Charge*
electron	1	$-e$
proton	1836	$+e$
neutron	1839	0

e = electronic charge

Interatomic bonds

Materials consist of atoms held together by the attractive forces they exert on each other. These forces are electrical in nature and create interatomic bonds of various types. The type formed in any case depends on the outer electrons in the electron clouds of the atoms involved.

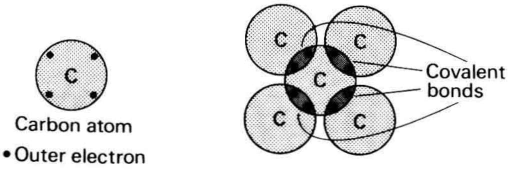
(*a*) *Ionic bond.* This is formed between the atoms of elements at opposite sides of the Periodic table, for example between sodium (Group 1) and chlorine (Group 7) when they are brought together to form common salt. A sodium atom has a loosely held outer electron which is readily accepted by a chlorine atom. The sodium atom becomes a positive ion, i.e. an atom deficient of an electron, and the chlorine atom becomes a negative ion, i.e. an atom with a surplus electron. The two ions are then bonded by the electrostatic attraction between their unlike charges.

A sodium ion attracts *all* neighbouring chloride ions in other pairs of bonded ions and vice versa. Each ion becomes surrounded by ions of opposite sign and the resulting structure depends among other things on the relative sizes of the two kinds of ion.

The ionic bond is strong. Ionic compounds are usually solid at room temperature and have high melting points. They are good electrical insulators in the solid state since the electrons are nearly all firmly bound to particular ions and few are available for conduction.

(*b*) *Covalent bond.* In ionic bonding electron *transfer* occurs from one atom to another. In covalent bonding electron *sharing* occurs between two or more atoms. Thus the atoms of carbon can form covalent bonds with other carbon atoms. Each carbon atom has four outer electrons, Fig. 1.4*a*, and all can be shared with four other carbon atoms to make four bonds, Fig. 1.4*b*, each consisting of two interlocking electron clouds.

Covalent bonds are also strong and many covalent compounds have similar mechanical properties to ionic compounds. However, unlike the latter, they do not conduct electricity when molten.



(a)

(b)

Fig. 1.4

(*c*) *Metallic bond.* Metal atoms have one or two outer electrons that are in general loosely held and are readily lost. In a metal we picture many free electrons drifting around randomly, not attached to any particular atom as they are in covalent bonding. All atoms share *all* the free electrons. The atoms thus exist as positive ions in a 'sea' of free electrons, Fig. 1.5; the strong attraction between the ions and electrons constitutes the metallic bond.

The nature of the metallic bond has a profound influence on the various properties of metals, as we shall see later.

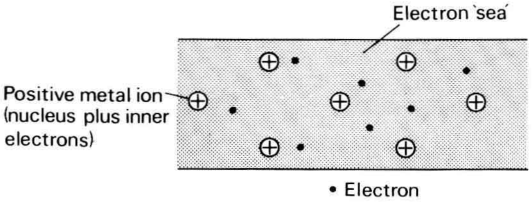


Fig. 1.5

(*d*) *Van der Waals bond.* Van der Waals forces are very weak and are present in all atoms and molecules. They arise because, although the centres of negative and positive charges in an atom coincide over a period of time, they do not coincide at any instant—for reasons too advanced to be considered here. There is a little more of the electron cloud on one side of the nucleus than the other. A weak electric 'dipole' is produced giving rise to an attractive force between opposite ends of such dipoles in neighbouring atoms.

The condensation and solidification at low temperatures of oxygen, hydrogen and other gases is caused by van der Waals forces binding their molecules together. (In the molecules of such gases the atoms are held together covalently.) Van der Waals forces are also important when considering polymers (p. 16).

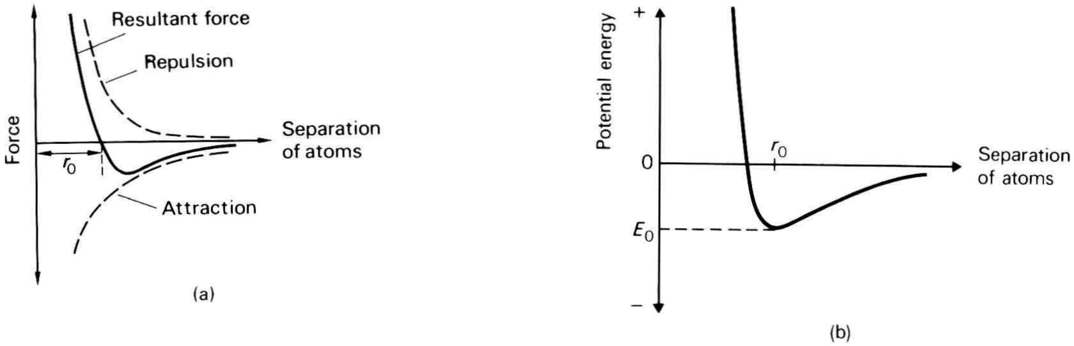
Two further points: first, sometimes more than one of the previous four types of bonding is involved in a given case; second, information about the strength of interatomic bonds in solids is obtained from heat of sublimation measurements in which solid is converted directly to vapour and all atoms separated from their neighbours (see question 8, p. 18) and for liquids latent heat of vaporization measurements provide the information (see page 72).

States of matter

The existence of three states or phases of matter is due to a struggle between interatomic (intermolecular) forces and the motion which atoms (molecules) have because of their internal energy (see p. 65).

(*a*) *Solids.* In the four types of interatomic force just discussed only attractions were considered but there must also be interatomic repulsions, otherwise matter would collapse. Evidence, both theoretical and experimental, suggests that at distances greater than one atomic diameter the attractive force exceeds the repulsive one, whilst for small distances, i.e. less than one atomic diameter, the reverse is true. In Fig. 1.6*a* the dotted graphs show how the *short-range attractive* force and the *very short-range repulsive* force between two atoms vary with the separation of the atoms; the total or resultant force is shown by the continuous graph.

It can be seen that for one value of the separation r_0, the resultant interatomic force is zero. This is the situation that normally exists in a solid, but if the atoms come closer together—for example, when the solid is compressed—they repel each other; they attract when they are pulled farther apart. We have only considered two atoms whereas in a solid each atom has interactions with many of its close neighbours. The conclusion



(a)

(b)

Fig. 1.6

about the existence of an equilibrium separation, however, will still hold good.

In an ionic bond the short-range attractive part of the interatomic force arises from the attraction between positive and negative ions which pulls them together until their electron clouds start to overlap, thus creating a very short-range repulsive force. The attractive and repulsive forces in the other types of bond also arise from the electric charges in atoms.

Now consider the motion of the atoms, the other contestant. In a solid the atoms vibrate about their equilibrium positions alternately attracting and repelling one another but the interatomic forces have the upper hand. The atoms are more or less locked in position and so solids have shape and appreciable stiffness.

The corresponding potential energy–separation curve for two atoms (or molecules) is shown in Fig. 1.6*b*. At the equilibrium separation r_0 when the resultant force is zero, the p.e. must have its minimum value E_0. This is so because any attempt to change the separation involves overcoming an opposing force—an attractive one if the separation increases and a repulsive one if it decreases. E_0 is called the *bonding energy;* it is the energy needed to pull the atoms apart so that their p.e. increases to zero. They are then quite free from one another's influence.

Bonding energy will be considered later in connection with latent heat (p. 72) and surface energy (p. 204).

(*b*) *Liquids.* As the temperature is increased the atoms have larger amplitudes of vibration and eventually they are able partly to overcome the interatomic forces of their immediate neighbours. For short spells they are within range of the forces exerted by other atoms not quite so near. There is less order and the solid melts. The atoms and molecules of a liquid are not much farther apart than in a solid but they have greater speeds, due to the increased temperature, and move randomly in the liquid while continuing to vibrate. The difference between solids and liquids involves a difference of structure rather than a difference of distance between atoms and molecules.

Although the forces between the molecules in a liquid do not enable it to have a definite shape, they must still exist otherwise the liquid would not hold together or exhibit surface tension (i.e. behave as if it had a skin on its surface) and viscosity nor would it have latent heat of vaporization.

(*c*) *Gases.* In a gas or vapour the atoms and molecules move randomly with high speeds through all the space available and are now comparatively far apart. On average their spacing at s.t.p. is about 10 molecular diameters and their mean free path (i.e. the distance travelled between collisions) is roughly 300 molecular diameters. Molecular interaction only occurs for those brief spells when molecules collide and large repulsive forces operate between them.

Conditions in gases and solids are, by comparison, simpler than those in liquids and in general their behaviour is better understood.

Types of solids

There are two main types of solids.

(*a*) *Crystalline.* Most solids, including all metals and many minerals, are crystalline. In substances such as sugar the crystal form is evident but less so in the case of metals, although large crystals of zinc are often visible on a freshly galvanized iron surface.

The crystalline structure of a metal can be revealed by polishing the surface, treating it with an etching chemical, sometimes a dilute acid, and then viewing it under an optical microscope. The metal is seen to