8 SOLIDS

Solids are classified into two groups; crystalline solids in which the constituent particles are orderly arranged, and amorphous solids in which constituent particles lack any complete regularity. The study of crystals has a long history probably because crystals were easier to study than amorphous solids

The most important development in the continuing study of crystals is X-ray crystallographic analysis. Initially this technique was barely able to treat materials of even simple structure such as common salt (NaCl). During the last eighty years, however, the progress of X-ray crystallographic analysis has been so rapid that proteins with considerably large molecular weight can now be studied with this technique.

8.1 Crystalline and amorphous solids

There are several ways to classify solids, which include a wide range of varieties. However, the simplest classification is to divide solids into two groups: crystalline solids in which the constituent particles are very regularly arrayed and amorphous solids in which little or no regularity is present.

(a) Crystalline solids

In some crystalline solids, the constituent particles are arrayed so regularly that their regularity is evident to the naked eye. Familiar crystals such as sodium chloride, hydrated copper sulfate and quartz are typical examples.

The locations of the constituent particles of crystalline solids (ions, atoms or molecules) are usually represented by a **lattice**, and the location of each particle is a **lattice point**. The smallest repeating unit of a lattice is called a **unit cell**.
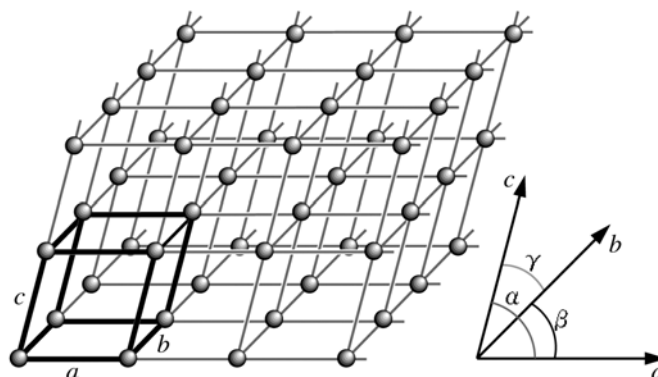

Figure 8.1 The definition of a unit cell.
Unit cell is constituted by lattice points and is identified by bold lines.
The distances between two points along each axis are defined as *a*, *b* and *c*.
The angles made by two axes are defined as α, β and γ.

The simplest unit cell is a cube. The three axes of a cube and of some other unit cells are perpendicular to each other but axes of other unit cells are not perpendicular. The factors that determine the unit cell are the distances between points, and the angles between lattice axes. These are called the **lattice constants** (Fig. 8.1).

In 1848, the French crystallographer Auguste Bravais (1811-1863) classified the crystal lattices based on their symmetries, and found that there are 14 kinds of crystal lattices as indicated in Fig. 8.2. These are called the **Bravais lattices**. These 14 Bravais lattices are classified into seven crystal systems. In this book, only three well known **cubic systems--simple cubic lattice**, **body-centered cubic lattice** and **face-centered cubic lattice**--will be treated.
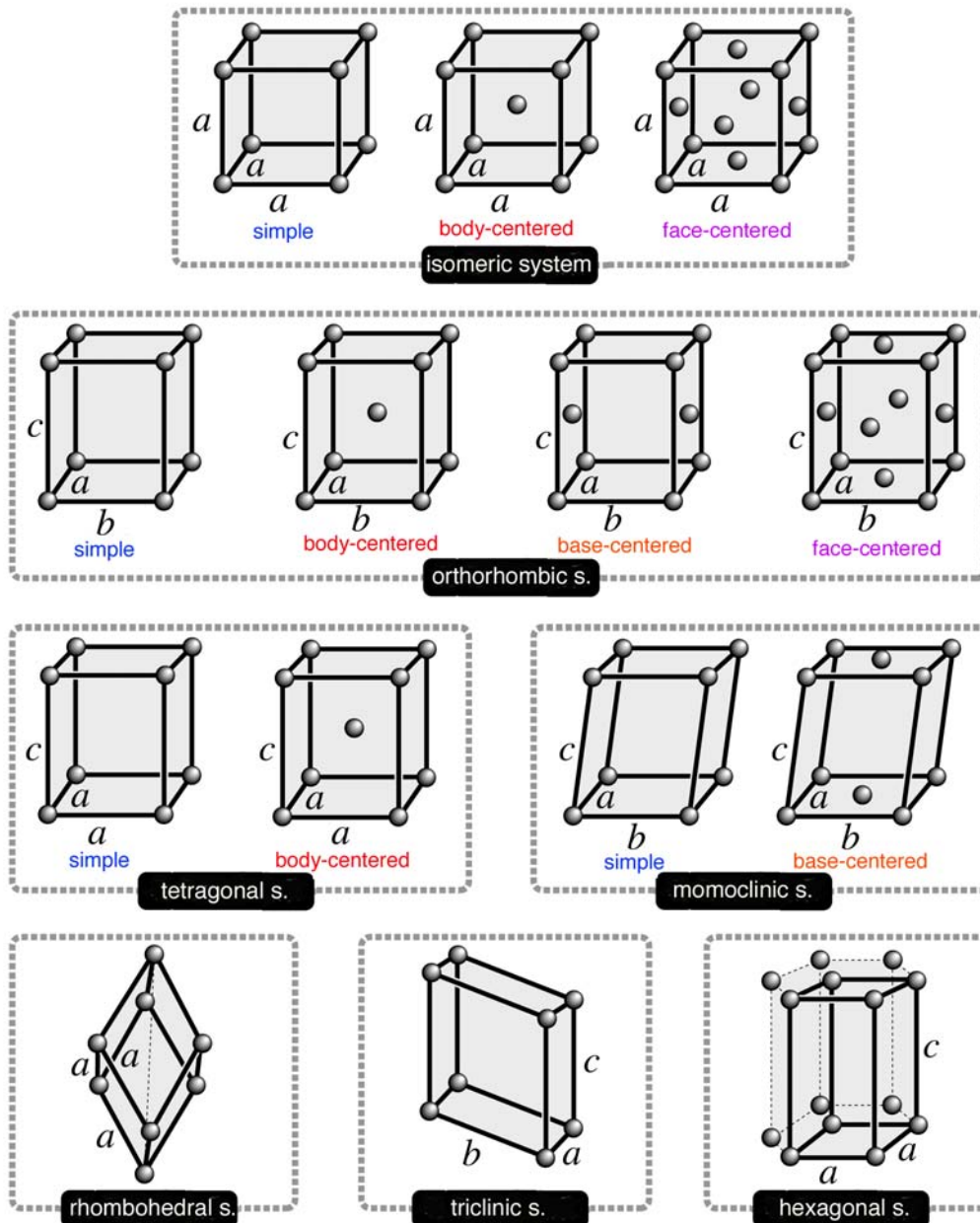
Figure 8.2 Bravais lattices.
Crystals are classified into 14 Bravais lattices and 7 crystalline systems.

The magnitude of a unit cell can be determined by Bragg's condition, which was proposed by the English physicist William Lawrence Bragg (1890-1971) in 1912. In order to obtain the detailed information on the accurate arrangement of particles in crystals, measurements of the intensity of spots in the diffraction pattern is necessary.

(b) Amorphous solids

The array of particles in amorphous solids is partly regular and thus resembles crystalline solids to some extent. This regularity is, however, limited and is not present in the whole solid. Many amorphous solids around you--common glass, rubber and polyethylene are some examples—have partial regularity (Fig. 8.3).

This feature of the structure of amorphous solids may be regarded as intermediate between that of liquids and that of solids. Recently much attention has been focused on such artificial materials as optical fibers and amorphous silicon (Table 8.1).
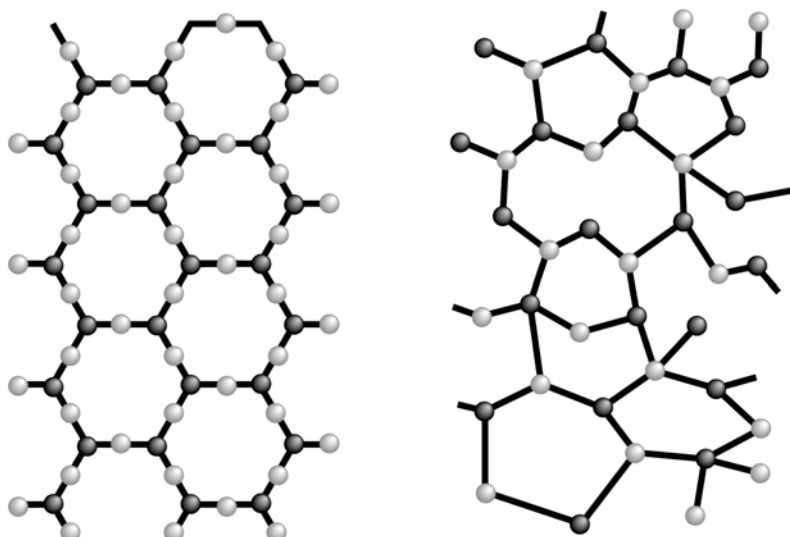
Figure 8.3 Crystalline and amorphous solids
There is a large difference in the regularity of the array of constituent particles.

Some scientists insist that amorphous solids may be regarded as the 4th state of matter.

Table 8.1 Some examples of functional amorphous solids

amorphous materials	uses
quartz glass	optical fiber
chalcogenide glass	selenium membranes for copy machines
amorphous silicon	solar cell
amorphous alloy	iron/cobalt metal (magnetic material)
polymer	polystyrene
amorphous carbon	carbon black (adsorbent)
gel	silica gel (adsorbent)

8.2 The structure of crystalline solids

(a) Closest packing

Many compounds, especially metallic crystals and molecular crystals have a common property that maximizes their densities by packing particles as tight as possible. Such crystal structures are called the **closest packed structure**.

As an example, examine the closest packing of a metallic crystal constituted by spherical atoms. The spheres are packed in layers. The 1st layer will necessarily be as shown in Fig. 8.4(a) to attain the closest packing. Each of the spheres in the 2nd layer occupies the pit formed by three spheres in the 1st layer. This is the most efficient way of using the available space (Fig. 8.4(b)). There are two ways to place the spheres of the 3rd layer. One is to place each sphere directly above the spheres of the 1st layer (Fig. 8.4(c)), and the other is to place each sphere over the pit of the 1st layer which is not used by the spheres of the 2nd layer (Fig. 8.4(d)). For convenience, the 1st way is called the *aba* arrangement, and the resulting structure is the **hexagonal closest packing**. The 2nd way is called the *abc* arrangement, and the resulting structure is the **cubic closest packing**.

Any two closest packing structures have common features: (1) the spheres occupy *ca*. 74% of the space; (2) each sphere is surrounded by twelve spheres; (3) six spheres out of twelve belong to the same layer and three spheres belong to the layer above and the remaining three spheres to the layer below. The number of other spheres which contact the sphere in question is called the

coordination number. For the closest packed structure, the coordination number is twelve, which is the maximum. In these cases, four particles are included in one unit cell.
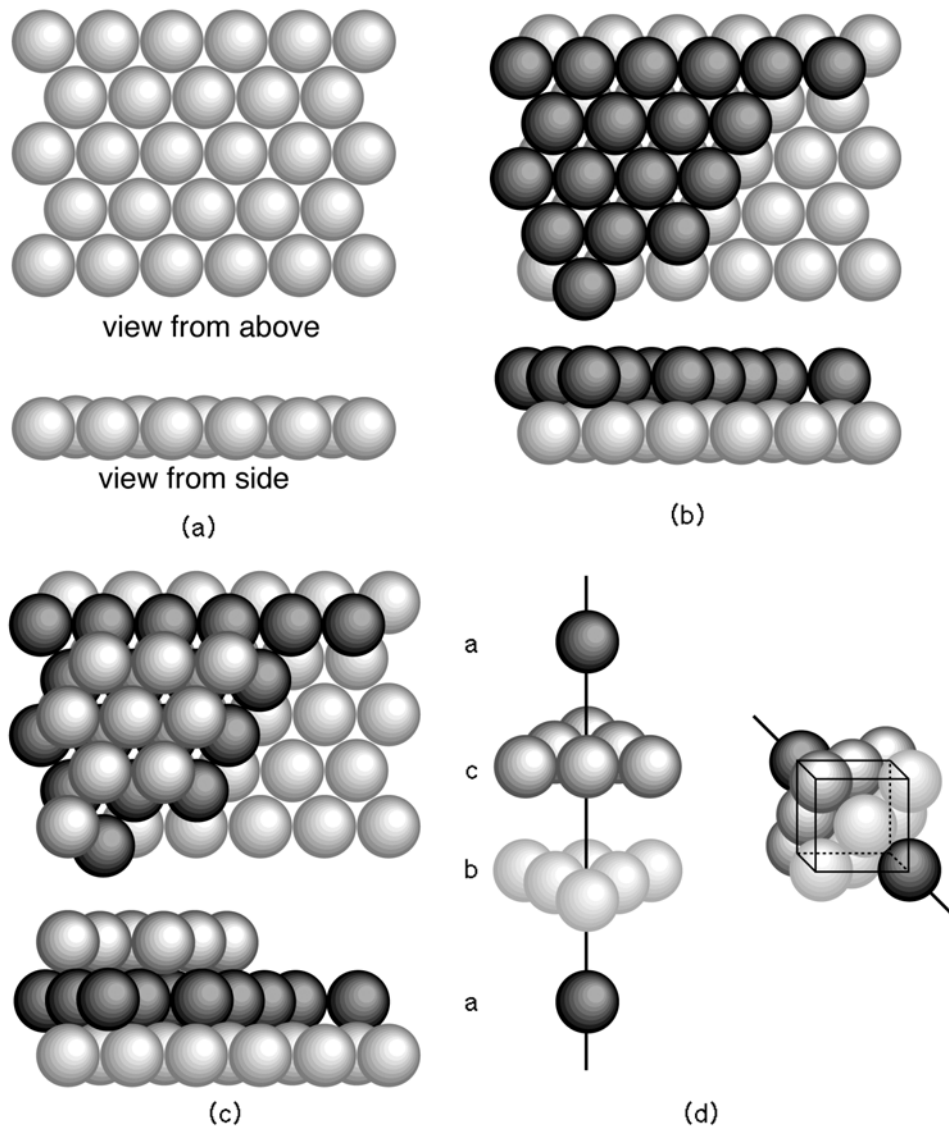

Figure 8.4 Closest packed structures

(a) A typical layer. Each sphere is surrounded by twelve spheres. (b) The 2nd layer is similar to the 1st layer. Each sphere occupies the pit formed by three spheres in the 1st layer. (c) Each of the spheres in the 3rd layer lies directly above the spheres of the 1st layer (the *aba* arrangement). (d) Each sphere of the 3rd layer lies above the pit of the 1st layer which is not used by the spheres of the 2nd layer (the *abc* arrangement).

Silver crystallizes in a cubic closest packed structure. If the crystal is cut as shown in Fig. 8.5, a sphere is located at the center of each face of the cube. Since a sphere (an atom) is located at the center of each face of the cube, this type of lattice is called the **face-centered lattice**.
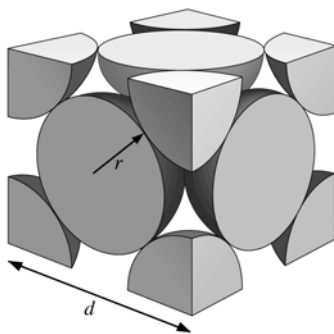
Figure 8.5 face-centered lattice
In such a case, the relation between r, the radius of the sphere, and d,
the length of the edge of the unit cell, can be determined by the Pythagorean theorem.

Sample exercise 8.1 Density of a metal
The atomic radius of a silver atom is 0.144 nm. Knowing that silver crystallizes in a cubic closest packed structure, calculate the density (g cm^{-3}) of silver.
Answer
The packing of silver atoms is shown in Fig. 8.5. You need to determine the volume and the number of silver atoms in the unit cell. Since the length of the diagonal is $4r$, d can be determined by the Pythagorean theorem.

$$d^2 + d^2 = (4r)^2 \qquad \therefore d = r\sqrt{8} = 0.144\sqrt{8} = 0.407 \text{ nm}$$

The number of silver atoms in the unit cell can be obtained from Fig. 8.5. Thus, there are six half-sphere and eight of eighth-sphere. Altogether there are four atoms in the unit cell. The mass m of a silver atom is:

$$m = 107.9 \text{ (g mol}^{-1}\text{)} / 6.022 \times 10^{23} \text{ (atom mol}^{-1}\text{)} = 1.792 \times 10^{-22} \text{ (g atom}^{-1}\text{)}$$

Since the density is given by (mass/volume), the density of silver d_{Ag} is;

$$d_{Ag} = [4(\text{atom}) \times 1.792 \times 10^{-22} \text{ (g atom}^{-1}\text{)}]/(0.407 \times 10^{-7})^3 \text{ (cm}^3\text{)} = 10.63 \text{ (g cm}^{-3}\text{)}$$

The experimental value is 10.5 (g cm^{-3}) (20°C)

(b) Body-centered cubic lattice
Some metals, such as alkaline metals, crystallize in a body-centered cubic lattice in which the center of a sphere is located at the center and at each apex (there are eight) of the unit cell as shown in Fig. 8.6. This packing is called the **body-centered cubic lattice**.
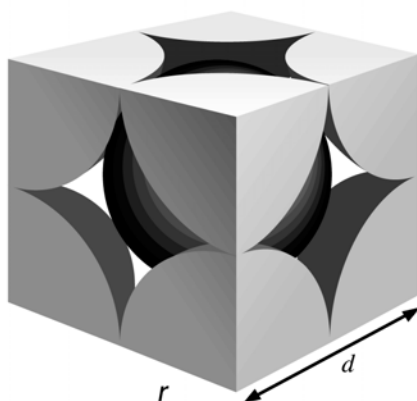

Figure 8.6 Body-centered cubic lattice

Sample exercise 8.2 Packing of the metallic crystals
With the aid of Fig. 8.6, answer the following questions.

(1) Determine the coordination number of the central sphere.
(2) What portion of a sphere is located at each corner of a unit cell?
(3) Determine the coordination number of each sphere at a corner of the cube.

Answer

(1) 8. The central sphere is surrounded by eight spheres, one at each corner.
(2) 8. The sphere at each corner is 1/8 of a sphere.
(3) 8. The sphere at each corner of a unit cell only contacts the eight central spheres around it and does not contact the spheres at the other corners.

Since the central sphere is surrounded by eight other spheres, the coordination number is eight. In other words, body-centered cubic lattice is not closest packed.

(c) X-ray crystallographic analysis

The technique of X-ray crystallographic analysis was first introduced at the beginning of the 20th century, and since then it has been used widely for the determination of the structure of various compounds. This technique perfectly solved several problems that had previously lacked experimental solutions. The initial step was achieved by William Henry Bragg (1862-1942), the father, and William Laurence Bragg (1890-1971), his son, who determined the structure of sodium chloride (common salt) and of diamond.

Until recently, X-ray crystallographic analysis was only carried out by specialists, *i.e.*, crystallographers, no matter what the target molecule was. Indeed, the measurement and the data processing required very high knowledge and experience. Now, however, thanks to the rapid and enormous development in both hard- and software for the technique, good X-ray crystallographic measurements have become possible after only moderate training. Today, even synthetic chemists whose main interests are in syntheses can carry out X-ray crystallographic analysis by themselves. Accordingly, the target molecules of specialists have become more and more complicated, and now even the structure of proteins can be elucidated if their molecular weights of these are not excessively large. As a result, basic knowledge of X-ray crystallographic analysis is required for all chemists in addition to that of NMR (Ch. 13.3).

Diffraction of light takes place within a substance when its particles are arrayed regularly and the wavelength of the incident light is comparable to the distances between particles. The waves of diffracted parallel beams reinforce each other when the waves are in phase, but cancel each other when the waves are out of phase. When crystals are irradiated by monochromatic X-rays, a diffraction pattern is obtained. It depends on the distance between lattice points whether or not two parallel waves reinforce or cancel each other.
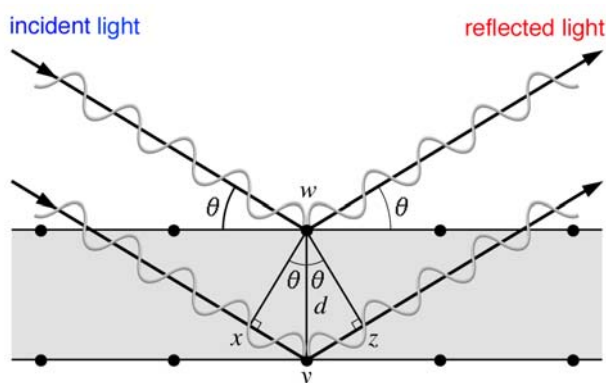
Figure 8.7 Bragg's condition
The diffraction of X-rays by atoms which lie in two layers. When the optical path difference,
$xy + yz = 2d\sin\theta$, is an integral number of wavelengths, the waves reinforce each other.

Suppose the wavelengths of the incident X-ray is λ (Fig. 8.7). When the difference between the

optical paths of X-rays reflected by the atoms in the 1st layer and by the atoms in the 2nd layer is $2d\sin\theta$, the waves reinforce each other to give a diffraction pattern. Thus, the intensity of the diffraction pattern will be maximum when

$$n\lambda = 2d\sin\theta \tag{8.1}$$

This equation is called **Bragg's condition**.

Bragg's condition can be applied for two purposes. If the interatomic distance d is known, the wavelength of X-rays can be determined by measuring the angle of diffraction. Moseley used this method when he determined the wavelength of the characteristic X-rays of various elements. On the contrary, if the wavelength of X-rays λ is known, the interatomic distances can be determined by measuring the angle of diffraction. This is the basic principle of X-ray crystallographic analysis.

Sample exercise 8.3 Bragg's condition
X-rays of wavelengh 0.154 nm were used to analyze the aluminum crystal. A diffraction pattern was obtained at $\theta = 19.3°$. Determine the distance d of two atomic planes, assuming $n = 1$.
Answer
$d = n\lambda/2\sin\theta = (1 \times 0.154)/(2 \times 0.3305) = 0.233$ (nm)

8.3 Various crystals

So far, crystals have been classified by the mode of packing of particles. Crystals can also be classified by the type of particles that constitute the crystal or by the interaction that combines the particles together (Table 8.2).

Table 8.2 Types of metals

Metallic		ionic		molecular		covalent	
Li	38	LiF	246.7	Ar	1.56	C(diamond)	170
Ca	42	NaCl	186.2	Xe	3.02	Si	105
Al	77	AgCl	216	Cl_2	4.88	SiO_2	433
Fe	99	ZnO	964	CO_2	6.03		
W	200			CH_4	1.96		

The values are the energy required to break the crystals apart
into their constituent atoms (or ions, molecules, *etc*) (kcal mol^{-1})

(a) Metallic crystals
The lattice of metallic crystals is constituted by the metallic bond. The valence electrons in metallic atoms are readily removed (because of small ionization energy) to yield cations. When two metallic atoms approach, their outer atomic orbitals overlap to form a molecular orbital. When a 3rd atom approaches the two atoms, interaction between the orbitals takes place and a new molecular orbital is formed. Thus, a great number of molecular orbitals will be formed by many metallic atoms, and the resultant molecular orbitals will spread in three-dimensions. This was already mentioned in Ch. 3.4 (Fig. 3.8).
Since atomic orbitals overlap repeatedly, the electrons in the outer shell of each atom will be affected by many other atoms. Such electrons do not necessarily belong to one particular atom, but move freely in the lattice formed by all these atoms. Thus, these electrons are called **free electrons**.
Useful properties of metals such as malleability, ductility, electric conductivity, and the metallic luster of freshly cut metals are related to the properties of the metallic bonds. For instance, metals can maintain their structure even if some deformation takes place. This is because there exists a strong interaction in various directions between atoms and the free electrons around them (Fig. 8.8).
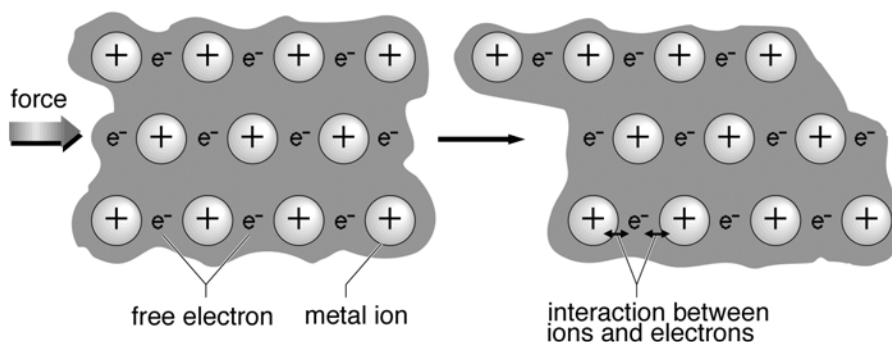
Figure 8.8 Deformation of the structure of metals
Metals will be deformed if a strong force is applied, but they are not readily broken.
This property is due to the strong (electric) interaction between metallic ions and free electrons.

The high thermal conductivity of metals can also be explained by free electrons. When one end of a metal piece is heated, the kinetic energy of electrons around that part is increased. The increased kinetic energy is quickly transferred to other free electrons. Electric conductivity is explained in the same manner. If voltage is applied to the both ends of a metal piece, the cloud of free electrons flow toward the positive charge.

Metallic luster is due to the numerous molecular orbitals of metallic crystals. Since there are so many molecular orbitals, the gap between energy levels is very small. When the surface of the metal is irradiated by light, electrons will absorb its energy and be excited. As a result, the range of wavelengths for such absorbed light is very wide. When the excited electrons discharge the acquired energy and return to their original energy levels, light with a wide range of wavelengths is emitted, which is observed as metallic luster.

(b) Ionic crystals

Ionic crystals such as sodium chloride (NaCl) are formed by the electrostatic attraction between positively charged and negatively charged ions. Ionic crystals generally have high melting points and low electrical conductivity. However, in solution or in a molten state, ionic crystals are dissociated into ions and hence possesses electrical conductivity.

It is assumed that bonds are formed between cations and anions. In crystals of sodium chloride, sodium ions and chloride ions are bonded by ionic bonds. Contrary to covalent bonds, ionic bonds have no directionality, and as a result, a sodium ion interacts with all chloride ions in the crystal although the intensity of interaction varies. Similarly, a chloride ion interacts with all sodium ions in the crystal.

The most stable array of an ionic crystal is the one in which the amount of contact between oppositely charged particles is largest, or in other words, the coordination number is largest. However, the size of cations is different from that of anions, and as a result, there is a tendency for the larger anions to pack in a closest packed structure, and for the smaller cations to locate in the gaps between anions.

In the case of sodium chloride, the chloride anion (ionic radius is 0.181 nm) form a closest packed face-centered cubic lattice with somewhat elongated interatomic distances so that smaller sodium cations (0.098 nm) can be readily accommodated in the space (Fig. 8.9(a)). Each sodium ion is surrounded by six chloride anions (coordination number = 6). Similarly, each chloride ion is surrounded by six sodium ions (coordination number = 6) (Fig. 8.9(b)). Thus, a 6:6 coordination is achieved.
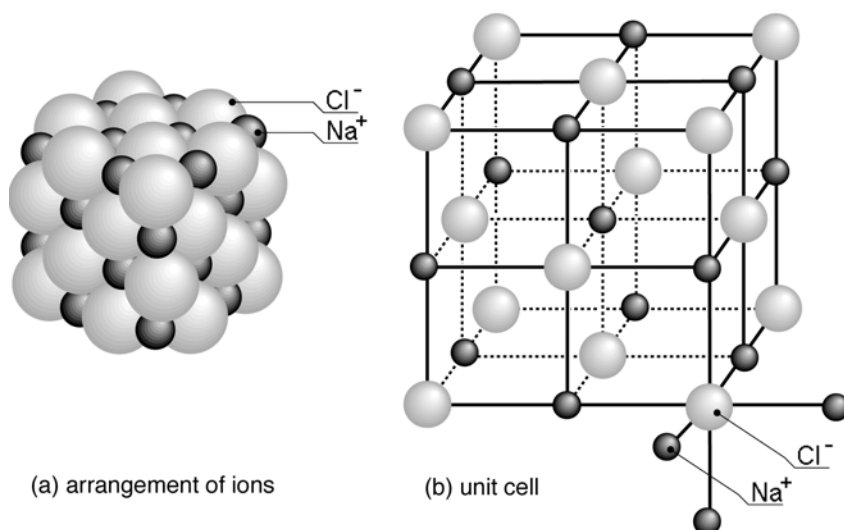
Figure 8.9 Crystal structure of sodium chloride
Each ion is surrounded by six ions with an opposite charge. This structure is not closest packed.

In cesium chloride, a cesium ion is much larger (0.168nm) than a sodium ion, and hence it is surrounded by eight chloride ions to form a 8:8 coordination. Both cesium ions and chloride ions independently form simple cubic lattices, one in which a chloride ion is located at the center of a unit cell made by cesium ions, and one in which a cesium ion is located at the center of a unit cell made by chloride ions. (Fig. 8.10)
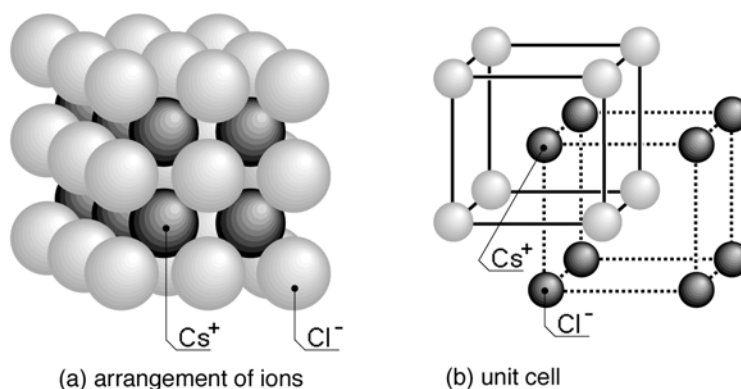

Figure 8.10 Crystal structure of cesium chloride
Each ion is surrounded by eight ions with opposite charges.
This structure is also not closest packed.

It is clear that the crystal structure of salts depends on the ratio of the sizes of cations and anions. When the ratio (the radius of cation)/(the radius of anion) (r_C/r_A) is smaller than that in sodium chloride, the coordination number is smaller than six. In zinc sulfide ZnS, a zinc ion is surrounded by only four sulfide ions. These points are summarized in Table 8.3.

Table 8.3 The ratio of the radius of cation r_C and that of anion r_A, and the coordination number

the ratio of radii r_C/r_A	coordination number	example
0.225-0.414	4	ZnS
0.414-0.732	6	most of halides of alikaline metal elements
>0.732	8	CsCl, CsBr, CsI

Sample exercise 8.4 Packing in ionic crystals

Using the values of ionic radii (nm) given below, propose the packing pattern of lithium fluoride LiF and rubidium bromide RbBr.

Li$^+$ = 0.074, Rb$^+$ = 0.149, F$^-$ = 0.131, Br$^-$ = 0.196

Answer

For LiF, r_C/r_A = 0.074/0.131 = 0.565. This is a six-coordinated crystal as are other halides of alkaline metal elements. For RbBr, r_C/r_A = 0.149/0.196 = 0.760. It is an eight-coordinated crystal and has a structure of the cesium salt type.

(c) Molecular crystals

Crystals in which molecules are bound by an intermolecular force such as the van der Waals force are called **molecular crystals**. Crystals thus far discussed are made of some kind of chemical bond between atoms or ions. However, crystals can be formed, without the aid of bonding, by weak interactions between molecules. Even rare gases crystallize at an extremely low temperature. Argon crystallizes by virtue of extremely weak van der Waals interactions, and its melting point is −189.2°C. Since argon is a monatomic molecule, it can be regarded as a sphere like a metal atom. In fact the structure of solid argon is the closest packed cube.

A diatomic molecule such as iodine cannot be approximated as a sphere. Though regularly arrayed in crystals, the direction of the molecules alternates (Fig. 8.11). However, because of their simple structure, the surfaces of the crystals are regular. This is the reason why crystals of iodine have a luster.
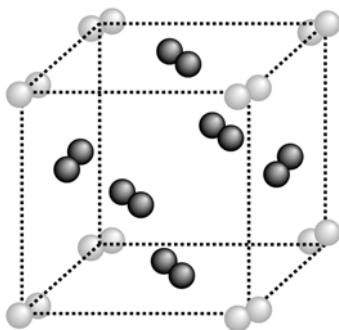

Fig. 8.11 Crystal structure of iodine
The structure is an orthorhombic face-centered lattice.
The molecules at the center of each face are black-colored.

The packing pattern of crystals of organic compounds with more complicated structures are investigated exclusively by means of X-ray crystallographic analysis. Their shapes are in most cases very similar or essentially identical with those of isolated molecules in gaseous phases or in solution.

(d) Covalent crystals

Many crystals have polymer-like or giant molecule-like structures. In such crystals all constituent atoms (not necessarily one type) are repeatedly bonded by covalent bonds to such an extent that the resultant particles are observable to the naked eye. Diamond is a typical example of this type, and its hardness originates from a strong network formed by covalent bonds between sp^3 hybridized carbon atoms (Fig. 8.12). Diamond is stable up to 3500°C, and at this temperature or above it sublimes.

Crystals such as silicon carbide (SiC)$_n$ or boron nitride (BN)$_n$ have much the same structure as diamond. A well-known example is silicon dioxide (quartz; SiO$_2$) (Fig. 8.13). Silicon is tetravalent as is carbon, and binds to four oxygen atoms to form a tetrahedron. Each oxygen atom binds to another silicon atom. The melting point of quartz is 1700°C.
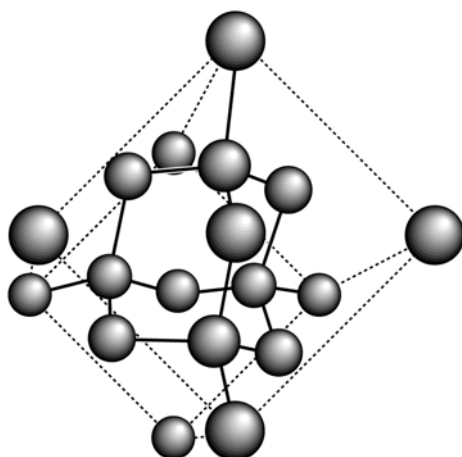
Fig. 8.12 Crystal structure of diamond
The ∠C-C-C angle is the tetrahedral angle, and each
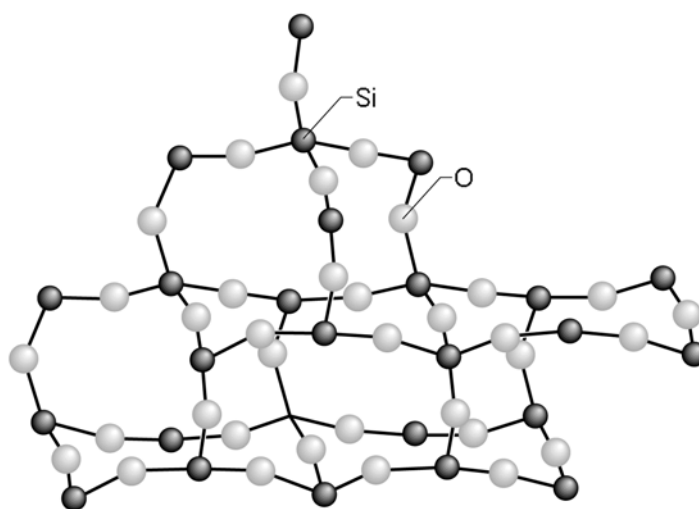carbon atom is surrounded by four other carbon atoms.

Figure 8.13 Crystal structure of silicon dioxide
If oxygen atoms are ignored, silicon atoms form a diamond-like
structure. Oxygen atoms are present between silicon atoms.

Sample question 8.5 Classification of solids
The crystals given below are either metallic, ionic, covalent or molecular crystals. Identify the type of crystal for each.

solid	m.p.(°C)	solubility in water	electroconductivity
A	150	not soluble	no conductivity
B	1450	not soluble	with conductivity
C	2000	not soluble	no conductivity
D	1050	soluble	no conductivity

Answer
A = molecular crystal, B = metallic crystal, C = covalent crystal, D = ionic crystal

Crystals are usually classified as in sample exercise 8.5. In another method, crystals are classified depending on their constituent particles, *i.e.*, atoms, molecules or ions. Crystals made of atoms include metallic crystals, covalent crystals such as diamond, and molecular crystals such as

those of rare gases. Table 8.4 summarizes this type of classification of crystals.

Table 8.4 Classification of solids

unit particles		type of bond	typical properties	examples
atoms	network	covalent bonds with direction	hard high melting points insulator	diamond
	metal	covalent bonds without direction	variable hardness variable melting points conductor	silver iron
	rare gases	intermolecular forces	very low melting points	argon
molecules		dipole-dipole interaction (polar molecules)	soft low melting points insulator	ice dry ice
ions		ionic bonds	hard high melting points insulator	sodium chloride

(e) Liquid crystals

Crystals have constant melting points, in other words, they maintain a constant temperature from the beginning to the end of their melting. On the other hand, melting points of amorphous substances have a wide range, and the temperature varies during the melting process.

There are some solids which turn into a turbid liquid phase at a constant temperature called the **transition temperature** before they finally melt. This liquid phase possesses the characteristic properties of liquids such as fluidity and surface tension. However, in this liquid phase, molecules maintain order to some extent, and the optical properties of this liquid phase are rather close to these of crystals. Such materials are called **liquid crystals**. Molecules that can become liquid crystals have a common structural feature, *i.e.*, they possess planar structural units such as benzene rings. In Fig. 8.14, some representative liquid crystals are shown.
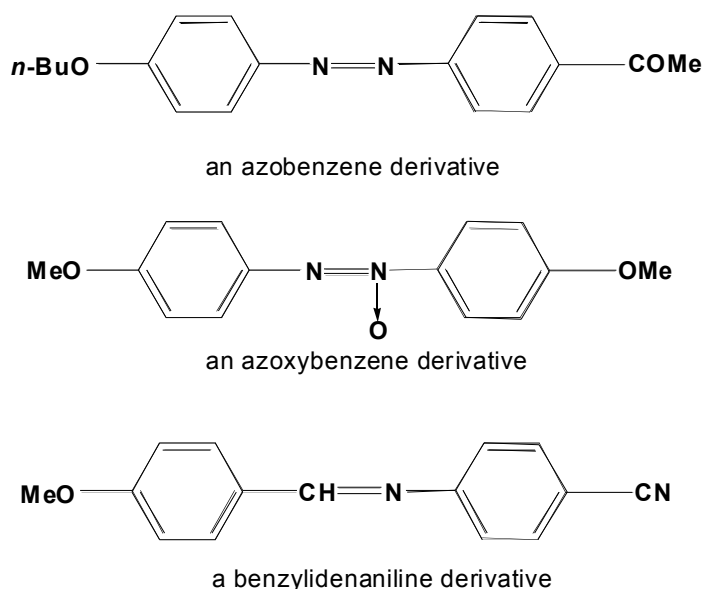
n-BuO—⬡—N═N—⬡—COMe

an azobenzene derivative

MeO—⬡—N═N—⬡—OMe
  |
  O

an azoxybenzene derivative

MeO—⬡—CH═N—⬡—CN

a benzylidenaniline derivative

Figure 8.14 Some examples of liquid crystals
In these liquid crystals, two benzene rings form the planar skeleton.

There are three types of liquid crystals: **smectic**, **nematic**, and **cholesteric**. The structural relation among solid crystals--smectic, nematic and cholesteric--are schematically represented in Fig.

8.15. Liquid crystals are widely used for practical purposes such as displays on watches or TVs.
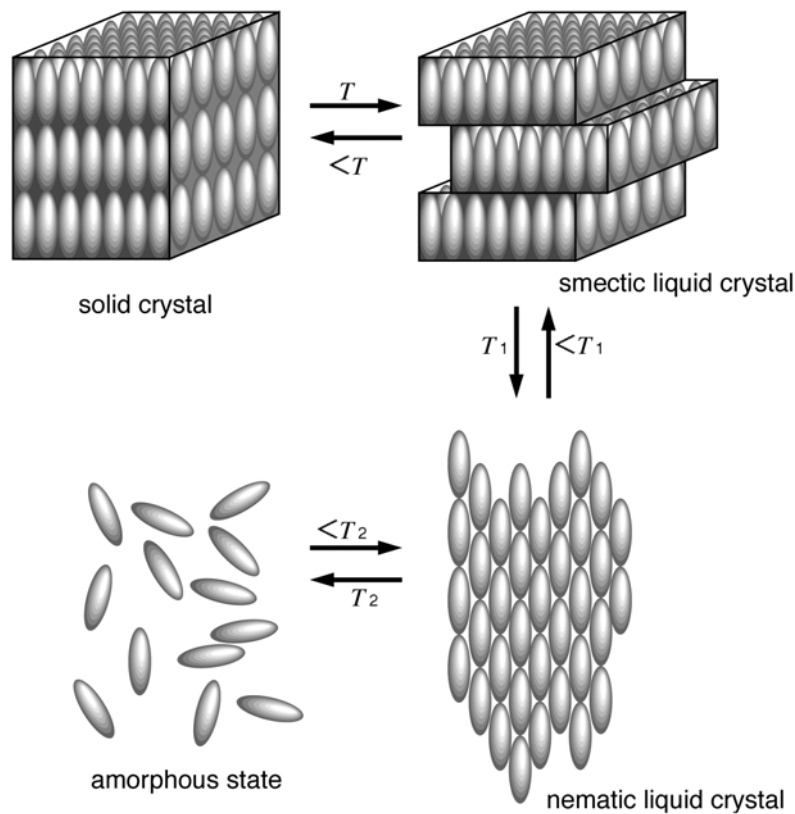

Figure 8.15 The order in liquid crystals
The order is three-dimensional in crystals. It may be said that the order is
two-dimensional in smectic liquid crystals and one-dimensional in nematic
liquid crystals. T is the transition temperature.

Exercises

8.1 Bragg's condition
A crystal was examined with X-rays (λ = 0.1541 nm) to obtain a diffraction pattern at θ = 15.55°. Determine the distance between layers when n = 1.
8.1 Answer
0.2874 nm

8.2 Crystals of sodium chloride type
In ionic crystals of sodium chloride type (coordination number = 6), a cation is surrounded by six anions. Determine the ratio of the radius of the cation to the radius of anion (r_C/r_A) when the two ions just contact.
8.2 Answer
$r_C + r_A = \sqrt{2}\, r_A$. $\therefore$ r_C/r_A = 0.414. If the ratio r_C/r_A is smaller than this value, there is no contact between cations and anions.

8.3 Density of crystals
The radius of a nickel (Ni) atom is 1.24×10^{-10} m, and the crystals are cubic closest packed (face-centered cube). Calculate the density of crystals of nickel.
8.3 Answer
Take the length of one edge of the unit cell of crystals of nickel as d, then $(4r)^2 = 2d^2$ by the

Pythagorean theorem. $\therefore$ $d = \sqrt{2}r$. The volume of a unit cell V is: $V = 22.63\ r^3$. In a unit cell, four atoms are included. Hence, the weight of a unit cell w is:

$w = (4 \times 58.70\ (\text{g mol}^{-1}))/(6.022 \times 10^{23}\ (\text{mol}^{-1})) = 3.900 \times 10^{-22}\ (\text{g})$.

$\therefore$ $d = w/V = 9.04\ \text{g cm}^{-3}$.

The experimental value is $8.90\ \text{g cm}^{-3}$. A good agreement.

8.4 Body-centered cubic lattice

The crystals of titanium have a body-centered cubic lattice structure and its density is $4.50\ \text{g cm}^{-3}$. Calculate the length of the edge of the unit cell l, and the atomic radius r of a titanium atom. In a body-centered cubic lattice, the atoms contact each other along the diagonal of the unit cell.

8.4 Answer

Since two atoms are contained in one unit cell of the body-centered cubic lattice, the following equation is obtained.

$4.50\ (\text{g cm}^{-3}) = (2 \times 47.88\ (\text{g mol}^{-1}))/(6.022 \times 10^{23}\ (\text{mol}^{-1}) \times l^3\ (\text{cm}^3))$

$\therefore$ $l = 3.28 \times 10^{-8}\ \text{cm}$.

By the Pythagorean theorem,

$(4r)^2 = (3.28 \times 10^{-8})^2 + [\sqrt{2}(3.28 \times 10^{-8})]^2$ $\therefore$ $r = 1.42 \times 10^{-8}\ \text{cm}$.

8.5 Classification of crystals

Classify the crystals given below based on the classification shown in Table 8.2.
(a) dry ice (CO_2); (b) graphite; (c) CaF; (d) MnO_2; (e) $C_{10}H_8$ (naphthalene);
(f) P_4; (g) SiO_2; (h) Si; (i) CH_4; (j) Ru; (k) I_2; (l) KBr; (m) H_2O; (n) NaOH;
(o) U; (p) $CaCO_3$; (q) PH_3.

8.5 Answer

(a) molecular; (b) covalent; (c) ionic; (d) ionic; (e) molecular; (f) molecular;
(g) covalent; (h) metallic; (i) molecular; (j) metallic; (k) molecular; (l) ionic;
(m) molecular; (n) ionic; (o) metallic; (p) ionic; (q) molecular.

Coffee Break

Great news on the battlefield

William Henry Bragg (1862-1942), the father, and William Laurence Bragg (1890-1971), the son, received the Nobel Prize for physics in their joint names in 1915 due to their contribution to the establishment of the method of X-ray crystallographic analysis. Though there are now some other examples where both a father and a son have receive the Nobel prize, there is no other case of a father and son being point recipients of the prize.
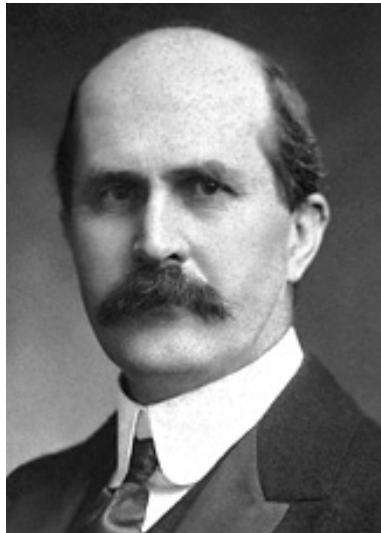
Sir William Henry Bragg   William Lawrence Bragg

When William Laurence received the prize, he was very young (25 years old). This meant that the new field of science, to explore the structure of atoms, was also young. William Laurence received the news at a military base in Belgium where he was engaged in the development of devices that were designed to detect the location of a hostile army with the aid of sound.

At the beginning of the 1st World War, there had not been any system to recruit scientists for military research. As the war went on for several years, scientists were called up for military research. The calling up of scientists became very systematic and extensive during the 2nd World War as exemplified by the Manhattan project.

At the beginning of the 1st World War, many young scientists volunteered for military service and went to the battlefield. Moseley was one of them. He applied to join the British army soon after the war began, and in 1915, was killed by the Turkish army in a tragic venture to attack Gallipoli Peninsula.

In view of the great scientific value of Moseley's law, it is almost certain that he would have received a Novel prize in 1916 or soon after if he had survived the war. It was indeed a great pity and ironical that a most promising candidate for a Nobel Prize was killed by a reckless venture planned by the ambitious Winston Churchill (1874-1965). Churchill did later become a Nobel Laureate.