

* 29-5 Bonding in Solids

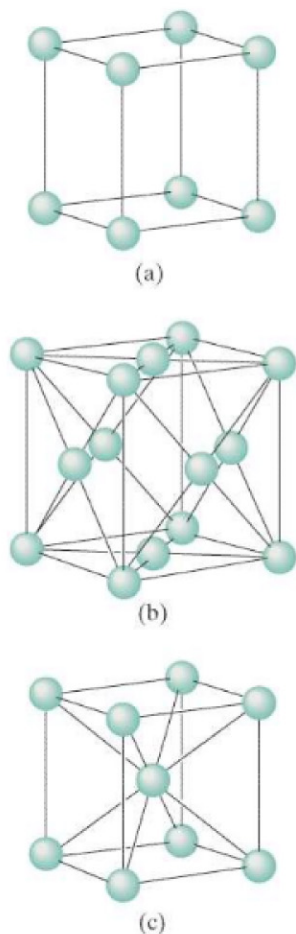
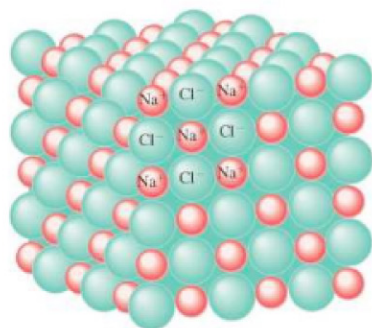


FIGURE 29-19 Arrangement of atoms in (a) a simple cubic crystal, (b) face-centered cubic crystal (note the atom at the center of each face), and (c) body-centered cubic crystal. Each diagram shows the relationship of the bonds. Each of these “cells” is repeated in three dimensions to the edges of the macroscopic crystal.

FIGURE 29-20 Diagram of an NaCl crystal, showing the “packing” of atoms.



Quantum mechanics has been a great tool for understanding the structure of solids. This active field of research today is called **solid-state physics**, or **condensed-matter physics** so as to include liquids as well. The rest of this Chapter is devoted to this subject, and we begin with a brief look at the structure of solids and the bonds that hold them together.

Although some solid materials are *amorphous* in structure (such as glass), in that the atoms and molecules show no long-range order, we are interested here in the large class of *crystalline* substances whose atoms, ions, or molecules are generally accepted to form an orderly array in a geometric arrangement known as a **lattice**. Figure 29-19 shows three of the possible arrangements of atoms in a crystal: simple cubic, face-centered cubic, and body-centered cubic. The NaCl crystal lattice is face-centered cubic, with one Na^+ ion or one Cl^- ion at each lattice point (see Fig. 29-20).

The molecules of a solid are held together in a number of ways. The most common are by *covalent* bonding (as between the carbon atoms of the diamond crystal) and by *ionic* bonding (as in a NaCl crystal). Often the bonds are partially covalent and partially ionic. Our discussion of these bonds earlier in this Chapter for molecules applies equally well here to solids.

Let us look for a moment at the NaCl crystal of Fig. 29-20. Each Na^+ ion feels an attractive Coulomb potential due to each of the six “nearest neighbor” Cl^- ions surrounding it. Note that one Na^+ does not “belong” exclusively to one Cl^- , so we must not think of ionic solids as consisting of individual molecules. Each Na^+ also feels a repulsive Coulomb potential due to other Na^+ ions, although this is weaker since the Na^+ ions are farther away.

A different type of bond, not discussed in Section 29-1, occurs in metals. Metal atoms have relatively loosely held outer electrons. **Metallic bond** theories propose that in a metallic solid, these outer electrons roam rather freely among all the metal atoms which, without their outer electrons, act like positive ions. The electrostatic attraction between the metal ions and this negative electron “gas” is believed to be, at least in part, responsible for holding the solid together. The binding energy of metal bonds is typically 1 to 3 eV, somewhat weaker than ionic or covalent bonds (5 to 10 eV in solids). The “free electrons” are responsible for the high electrical and thermal conductivity of metals. This theory also nicely accounts for the shininess of smooth metal surfaces: the free electrons can vibrate at any frequency, so when light of a range of frequencies falls on a metal, the electrons can vibrate in response and reemit light of those same frequencies. Hence, the reflected light will consist largely of the same frequencies as the incident light. Compare this to nonmetallic materials that have a distinct color—the atomic electrons exist only in certain energy states, and when white light falls on them, the atoms absorb at certain frequencies, and reflect other frequencies which make up the color we see.

Here is a brief comparison of important strong bonds:

- ionic: an electron is stolen from one atom by another
- covalent: electrons are shared by atoms within a single molecule
- metallic: electrons are shared by all atoms in the metal

The atoms or molecules of some materials, such as the noble gases, can form only **weak bonds** with each other. As we saw in Section 29-3, weak bonds have very low binding energies and would not be expected to hold atoms together as a liquid or solid at room temperature. The noble gases condense only at very low temperatures, where the atomic (thermal) kinetic energy is small and the weak attraction can then hold the atoms together.