

PHYSICS

STUDY UNIT 2: BONDING IN SOLIDS

Learning Outcomes

By the end of this unit, you will be able to:

- Describe the packing structures of common solids
- Explain the difference between bonding in a solid and in a molecule
- Determine the equilibrium separation distance given crystal properties
- Determine the dissociation energy of a salt given crystal properties

Beginning in this section, we study crystalline solids, which consist of atoms arranged in an extended regular pattern called a **lattice**. Solids that do not or are unable to form crystals are classified as **amorphous solids**. Although amorphous solids (like glass) have a variety of interesting technological applications, the focus of this chapter will be on crystalline solids.

Atoms arrange themselves in a lattice to form a crystal because of a net attractive force between their constituent electrons and atomic nuclei. The crystals formed by the bonding of atoms belong to one of three categories, classified by their bonding: ionic, covalent, and metallic. Molecules can also bond together to form crystals; these bonds, not discussed here, are classified as molecular. Early in the twentieth century, the atomic model of a solid was speculative. We now have direct evidence of atoms in solids.

PHYSICS

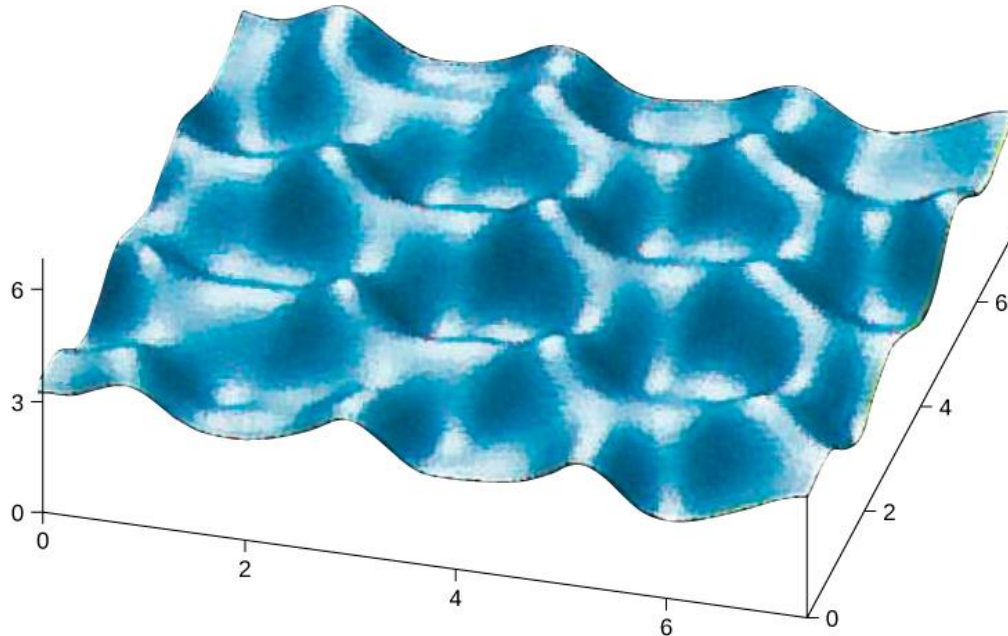


Figure 2.1 : An image made with a scanning tunneling microscope of the surface of graphite. The peaks represent the atoms, which are arranged in hexagons. The scale is in angstroms.

Ionic Bonding in Solids

Many solids form by ionic bonding. A prototypical example is the sodium chloride crystal, as we discussed earlier. Electrons transfer from sodium atoms to adjacent chlorine atoms, since the valence electrons in sodium are loosely bound and chlorine has a large electron affinity. The positively charged sodium ions and negatively charged chlorine (chloride) ions organize into an extended regular array of atoms.

PHYSICS

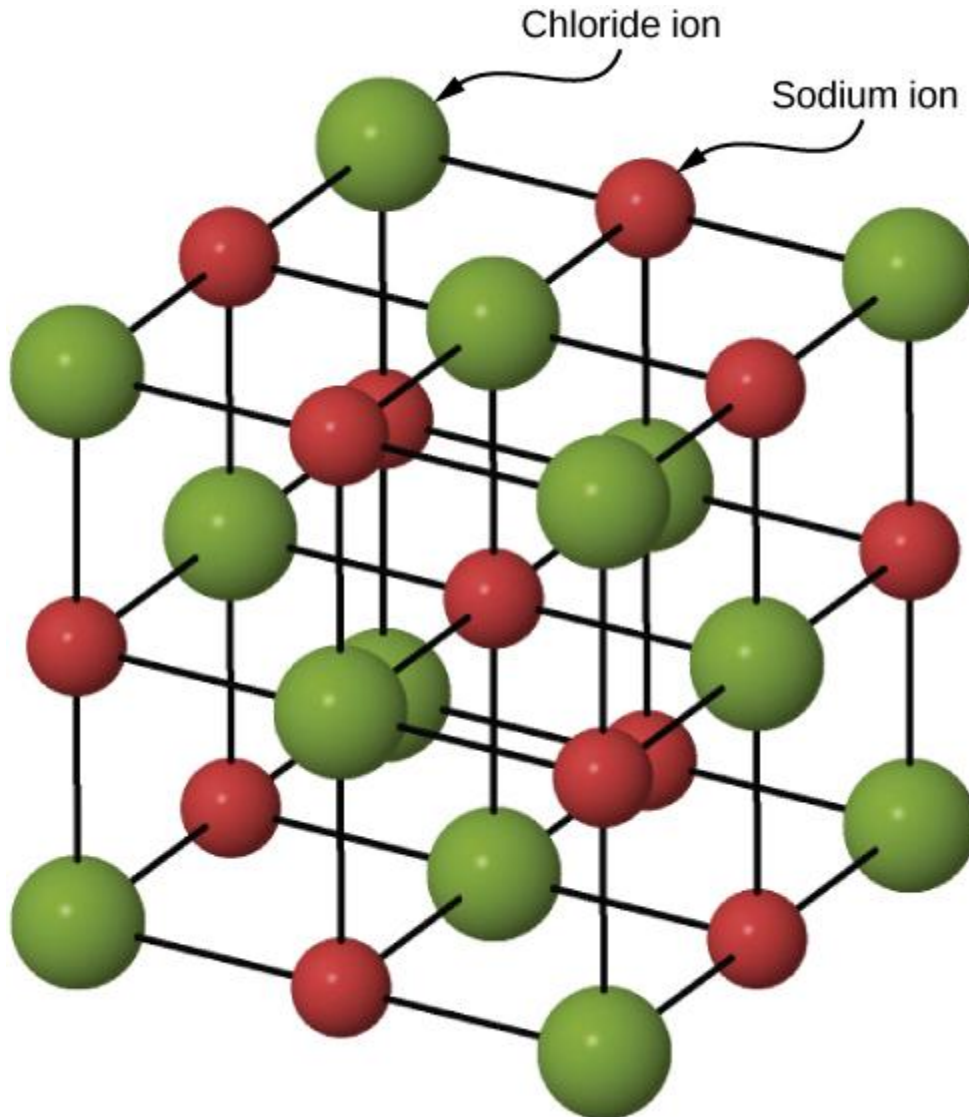


Figure 2.2: Structure of the sodium chloride crystal. The sodium and chloride ions are arranged in a face-centered cubic (FCC) structure.

The charge distributions of the sodium and chloride ions are spherically symmetric, and the chloride ion is about two times the diameter of the sodium ion. The lowest energy arrangement of these ions is called the **face-centered cubic (FCC)** structure. In this structure, each ion is closest to six ions of the other species. The unit cell is a cube—an atom occupies the center and

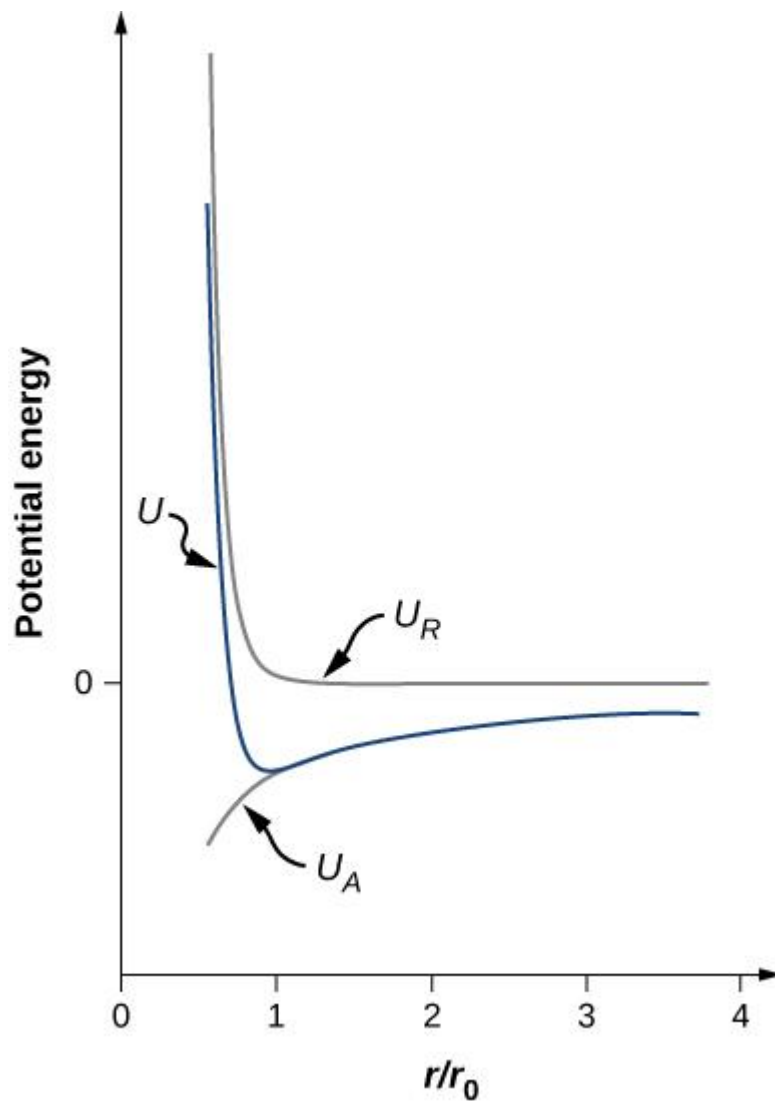
PHY: 1203 INTRODUCTORY SOLID STATE

PHYSICS

corners of each “face” of the cube. The attractive potential energy of the Na^+Na^+ ion due to the fields of these six Cl^-Cl^- ions is written

$$U_1 = -6e^{24}\pi\epsilon_0 r$$

where the minus sign designates an attractive potential (and we identify $k=1/4\pi\epsilon_0 k$).



PHY: 1203 INTRODUCTORY SOLID STATE PHYSICS

Figure 2.3: The potential energy of a sodium ion in a NaCl crystal for $n=9$. The equilibrium bond length occurs when the energy is a minimum.

As long as $n > 1$, the curve for U has the same general shape: U approaches infinity as $r \rightarrow 0$ and U approaches zero as $r \rightarrow \infty$.

The energy per ion pair needed to separate the crystal into ions is the **dissociation energy** of the solid. The dissociation energy can also be used to describe the total energy needed to break a mole of a solid into its constituent ions, often expressed in kJ/mole. The dissociation energy can be determined experimentally using the latent heat of vaporization.

Because the ions in crystals are so tightly bound, ionic crystals have the following general characteristics:

1. They are fairly hard and stable.
2. They vaporize at relatively high temperatures (1000 to 2000 K).
3. They are transparent to visible radiation, because photons in the visible portion of the spectrum are not energetic enough to excite an electron from its ground state to an excited state.
4. They are poor electrical conductors, because they contain effectively no free electrons.
5. They are usually soluble in water, because the water molecule has a large dipole moment whose electric field is strong enough to break the electrostatic bonds between the ions.

ITQ

If the dissociation energy were larger, would that make it easier or more difficult to break the solid apart?

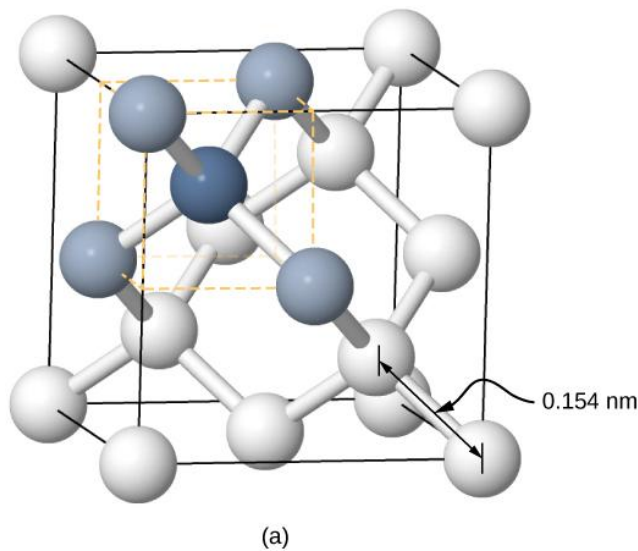
ITA

PHY: 1203 INTRODUCTORY SOLID STATE PHYSICS

more difficult

Covalent Bonding in Solids

Crystals can also be formed by covalent bonding. For example, covalent bonds are responsible for holding carbon atoms together in diamond crystals. The electron configuration of the carbon atom is $1s^2 2s^2 2p^2$ - a He core plus four valence electrons. This electron configuration is four electrons short of a full shell, so by sharing these four electrons with other carbon atoms in a covalent bond, the shells of all carbon atoms are filled. Diamond has a more complicated structure than most ionic crystals. Each carbon atom is the center of a regular tetrahedron, and the angle between the bonds is 109.5° . This angle is a direct consequence of the directionality of the **p** orbitals of carbon atoms.



(b)

Figure 2.4 : Structure of the diamond crystal. (a) The single carbon atom represented by the dark blue sphere is covalently bonded to the four carbon atoms represented by the light blue spheres. (b) Gem-quality diamonds can be cleaved along smooth planes, which gives a large number of angles that cause total internal reflection of incident light, and thus gives diamonds their prized brilliance.

PHY: 1203 INTRODUCTORY SOLID STATE

PHYSICS

Covalently bonded crystals are not as uniform as ionic crystals but are reasonably hard, difficult to melt, and are insoluble in water. For example, diamond has an extremely high melting temperature (4000 K) and is transparent to visible light. In comparison, covalently bonded tin (also known as alpha-tin, which is nonmetallic) is relatively soft, melts at 600 K, and reflects visible light. Two other important examples of covalently bonded crystals are silicon and germanium. Both of these solids are used extensively in the manufacture of diodes, transistors, and integrated circuits. We will return to these materials later in our discussion of semiconductors.

Metallic Bonding in Solids

As the name implies, **metallic bonding** is responsible for the formation of metallic crystals. The valence electrons are essentially free of the atoms and are able to move relatively easily throughout the metallic crystal. Bonding is due to the attractive forces between the positive ions and the conduction electrons. Metallic bonds are weaker than ionic or covalent bonds, with dissociation energies in the range $1\text{--}3\text{eV}$.

References

- Samuel J. Ling (Truman State University), Jeff Sanny (Loyola Marymount University), and Bill Moebs with many contributing authors. This work is licensed by OpenStax University Physics under a [Creative Commons Attribution License \(by 4.0\)](https://creativecommons.org/licenses/by/4.0/).