

SOLIDS AND MODERN MATERIALS

The content in this topic is the basis for mastering Learning Objectives 2.1, 2.19, 2.20, and 2.22 through 2.32 as found in the Curriculum Framework.

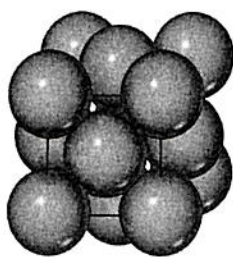
When you finish reviewing this topic, be sure you are able to:

- Construct atomic-level visual representations for the structural features of the major classes of solids: ionic, metallic, covalent-network, and molecular
- Predict characteristic properties and structure of solids based on their chemical formulas
- Understand the forces that bind the atoms and molecules of each type of solid
- Use three-dimensional representations and the interaction of the particles to explain the common macroscopic properties of solids
- Understand how the electron-sea model of a metal with its delocalized electrons explains common metallic properties
- Know what an alloy is and how its structure and properties relate to that of a pure metallic solid
- Distinguish the structures of substitutional alloys and interstitial alloys
- Use chemical formulas to associate and recognize the classification, structure, and bonding of solid substances
- Design a plan to collect or interpret data to classify a solid substance based on its observable properties

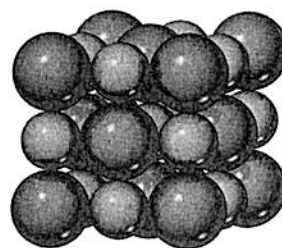
Classifications of Solids

Section 12.1

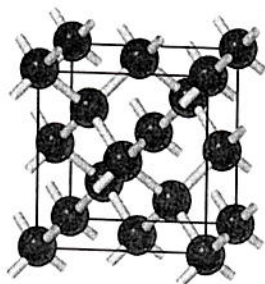
The structures of solids and their physical properties are determined by the types of bonds that hold their atoms in place. Classification of solids is based on bonding type. Often classification can be deduced from the chemical formula. Figure 12.1 illustrates the three-dimensional nature of the four bonding types. Table 12.1 summarizes the bonding and properties of different types of solid materials.

**Metallic solids**

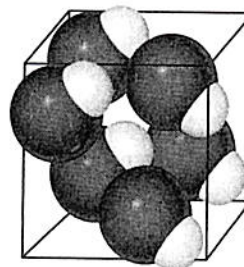
Extended networks of atoms held together by metallic bonding (Cu, Fe)

**Ionic solids**

Extended networks of ions held together by ion-ion interactions (NaCl, MgO)

**Covalent-network solids**

Extended networks of atoms held together by covalent bonds (C, Si)

**Molecular solids**

Discrete molecules held together by intermolecular forces (HBr, H₂O)

Figure 12.1 Classifications of solids according to predominant bonding type.

Table 12.1 Summary of the four bonding types in solids.

Solid	Recognizing a Formula with Examples	Structure and Bonding	Properties
Ionic solids	Compounds of metal cations and nonmetal anions: NaCl, Al ₂ O ₃ , K ₂ SO ₄ , FeS	Three-dimensional crystal lattice of ions locked in place by relatively strong ionic bonds	Brittle, high melting nonconductors (insulators) in pure form, conductors in water solution
Metallic solids	Metal atoms only: Al, Fe, Cr, Ni	Three-dimensional arrays of metal ions surrounded by a uniform sea of delocalized valence electrons	Good conductors of heat and electricity, malleable and ductile; mixtures of metals form alloys
Covalent-network solids	C (graphite and diamond); metalloids: Si, Ge; and compounds of metalloids: SiO ₂ , SiC, BN	An extended network of covalent bonds	Hard, brittle, high melting, poor conductivity or semiconductors
Molecular solids	Compounds of nonmetals only: H ₂ O(s), P ₂ O ₅ , C ₆ H ₁₂ O ₆	Individual covalently bonded molecules held together by weak intermolecular forces	Soft, low melting nonconductors

High thermal conductivity of metals results because the mobile electrons can readily transfer kinetic energy (heat) through the solid.

This model also explains why metals deform. Metals are **malleable** (they can be shaped) and **ductile** (they can be drawn into thin wires). Mobile electrons easily redistribute themselves when reshaping of the metal causes changes in atomic positions.

An **alloy** has the characteristic properties of a pure metal but contains two or more elements. Alloys are useful because they impart desirable properties to the metal. For example, pure gold is too soft to be used as jewelry but small amounts of added silver make it much harder. The structure of an alloy resembles that of the electron-sea model of a pure metal. Figure 12.3 shows the structures of two common alloys, both solid solutions.

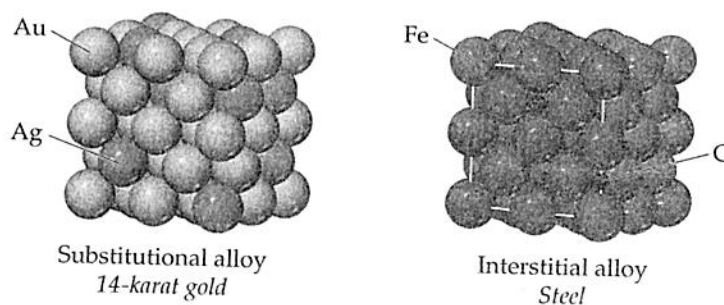


Figure 12.3 The distribution of solute and solvent atoms in a substitutional alloy and an interstitial alloy. Both types of alloys are solid solutions and, therefore, homogenous mixtures.

A **substitutional alloy** is a solid solution where atoms of another element replace some of the metal atoms. For example, silver atoms, which are about the same size as gold atoms, and copper atoms, which are only slightly smaller, readily replace atoms of gold in its crystal structure. The slightly different-sized replacement atoms deform the crystal structure enough to improve strength and hardness because the crystal cannot break or bend easily along nonuniform planes. Also, a substitutional alloy takes on some of the properties of the substituted metal. For example, silver is harder than gold and copper modifies its color, so a gold alloy is harder than pure gold and has a different color.

In an **interstitial alloy**, another element occupies the “holes” between the metal atoms. Typically, an interstitial element is a nonmetal with a much smaller bonding radius than the metal. Its atoms fit into the spaces between the metal atoms and form covalent bonds with the metal. These extra bonds cause the metal lattice to become harder, stronger, and less malleable and ductile. Steel, an interstitial alloy of iron and carbon that forms covalent iron carbide, is a good example.

(Note: Except for alloys, you need NOT know the structural details of solids described in Sections 12.2 and 12.3 of *Chemistry: The Central Science*.)

Describe the atomic-level structures of iron and steel and explain why steel is much harder and stronger than iron. Write your answer in the space provided.

Your Turn 12.2

Ionic Solids

Section 12.3

Ionic solids consist of cations and anions locked together by ionic bonds in a three-dimensional regularly ordered **crystal lattice** structure as shown in Figure 12.4. High melting points and boiling points of ionic solids indicate that these ionic bonds are very strong.

Coulomb's law guides our thinking to determine relative strengths of ionic bonds. Recall that the strength of an ionic bond depends on the size and charge of each ion.

$$F = kQ_1Q_2/d^2$$

Small ions with large charges will produce very strong forces of attraction. Larger ions with smaller charges will be less tightly bonded.

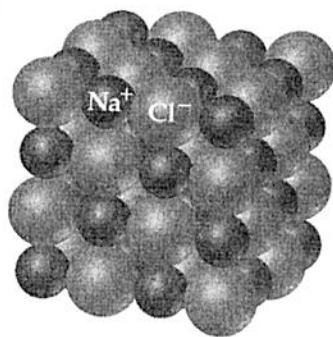


Figure 12.4 The crystal lattice structure of sodium chloride, an ionic solid.

Unlike metallic solids, ionic substances are insulators rather than conductors.

The valence electrons in ionic solids are confined, so they do not carry an electric current in the solid phase. However, when melted, the ions are free to move and conduct electricity. Aqueous solutions of ionic compounds also conduct electricity.

because the dissolving process breaks down the lattice and frees the ions to move throughout the solution. Ionic compounds do not dissolve in nonpolar solvents because the ionic bonds in the solid are much stronger in the solid than the attractive forces of the nonpolar solute.

Ionic solids are brittle because of repulsive forces of ions of like charge. Figure 12.5 shows that when a stress is applied to an ionic solid, a fracture occurs between planes of ions. Movement along a plane shifts the alignment of ions from an attractive cation–anion–cation–anion arrangement to a repulsive cation–cation and anion–anion configuration. This causes the planes to break away from each other resulting in a brittle property.

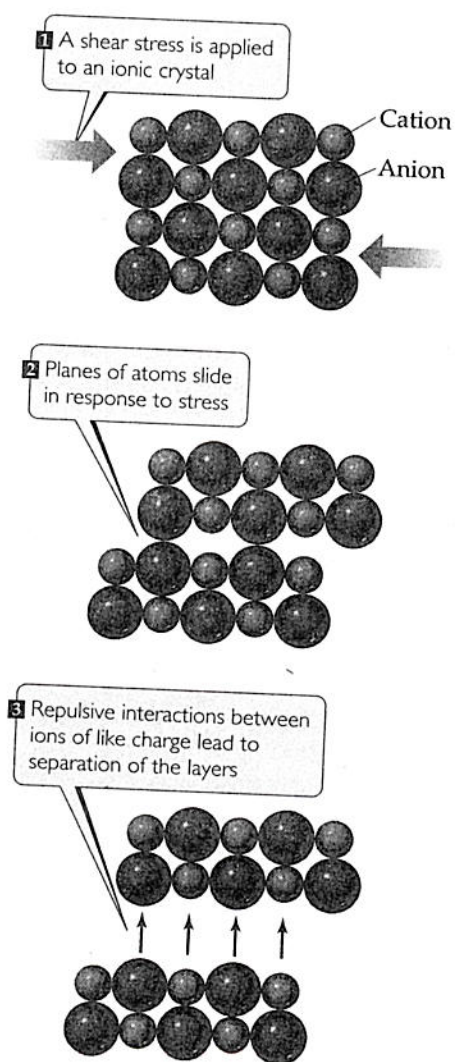


Figure 12.5 Brittleness in ionic crystals.

When ionic solids dissolve in water, the ionic bonds break and the ions attach themselves to water molecules as they dissociate into the aqueous solution (Figure 12.6). Such an ionic compound is called an **electrolyte** because its aqueous solution conducts electricity. On dissolving, the ions are no longer locked into place. The charged particles are free to move about the solution to carry a charge.

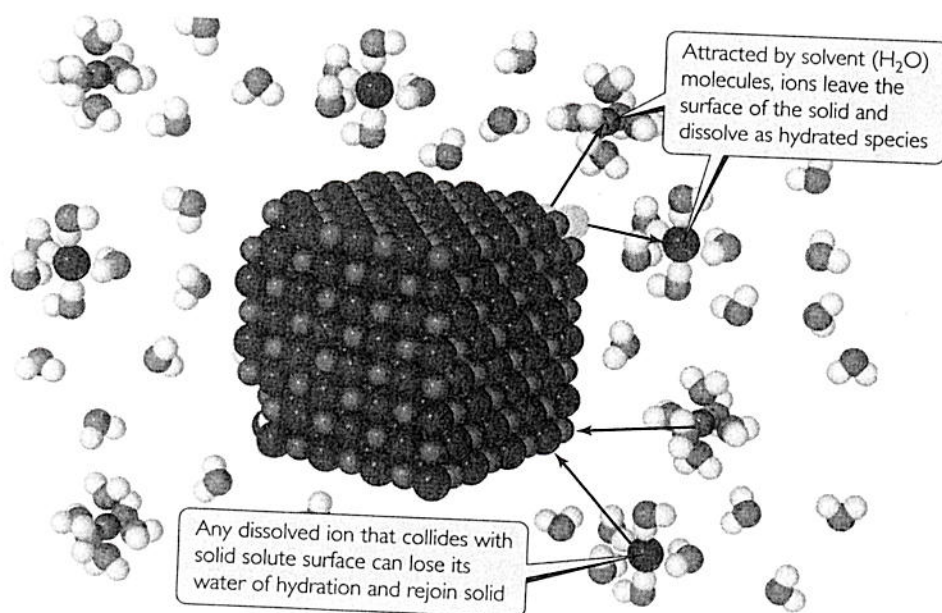


Figure 12.6 Dynamic equilibrium in a saturated solution with excess ionic solute.

Molecular Solids

Section 12.4

Molecular solids usually consist of discrete molecules made up of exclusively nonmetals having covalent bonds. Molecular solids are held together by weak dipole–dipole, London dispersions, and/or hydrogen bonds. Because of these relatively weak intermolecular forces, molecular solids tend to be soft, low-melting substances. Because their electrons are confined to strong covalent bonds, molecular solids are nonconductors. Common examples include sucrose (common table sugar, Figure 12.7), candle wax ($\text{C}_{20}\text{H}_{42}$), and aspirin ($\text{C}_9\text{H}_8\text{O}_4$). Low molar mass molecular substances such as water, carbon dioxide, and argon tend to be liquids or gases at room temperature and form molecular solids only at cold temperatures.

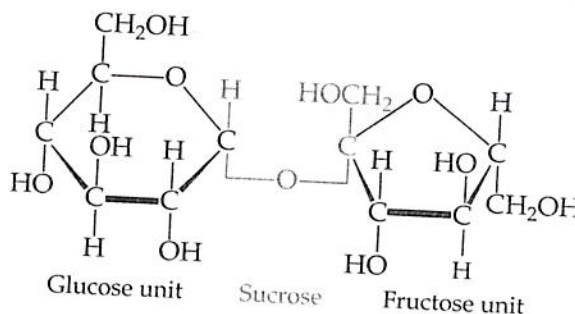


Figure 12.7 Sucrose forms discrete molecules that are held in a solid structure by relatively weak intermolecular forces of attraction.

Your Turn 12.3

Examine the structure of sucrose. Tell what kinds of intermolecular forces likely hold sucrose molecules together. Predict whether sucrose has a relatively high or low melting point compared to table salt (NaCl) and other molecular solids. Justify your answers. Write your answers in the space provided.

Section 12.5

Covalent-Network Solids

Covalent-network solids consist of the two allotropes of carbon, diamond and graphite, as well as metalloids such as silicon and germanium and compounds of metalloids such as quartz (SiO_2), silicon carbide (SiC), and boron nitride (BN). Covalent-network solids are held together by large networks or chains of covalent bonds. Because covalent bonds are much stronger than intermolecular forces, covalent-network solids tend to be much harder and have higher melting points than molecular solids. The structures of diamond and graphite are shown in Figure 12.8.

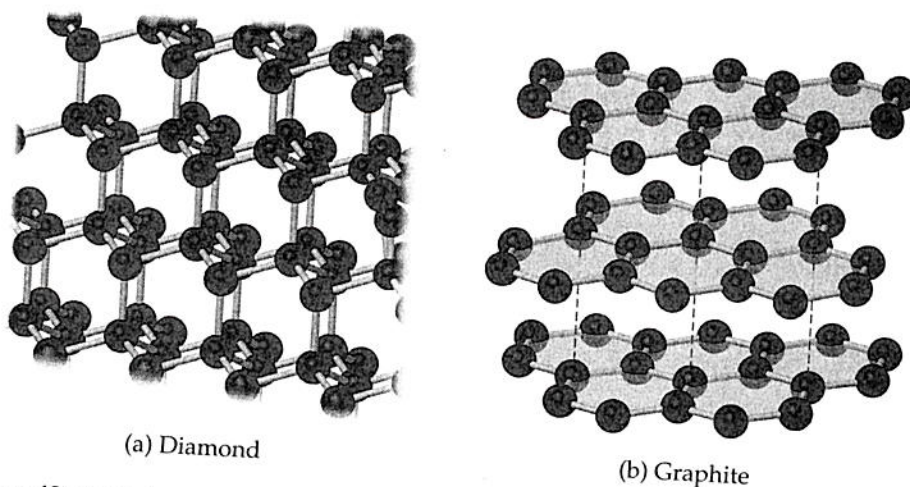


Figure 12.8 The structures of two covalent-network solids: (a) diamond and (b) graphite. Notice the planar structure of graphite.

Doping is the process of adding small amounts of impurities to the crystal lattice of a material to influence its electrical conductivity.

An **n-type semiconductor**, one that can carry a negative charge, is made by replacing a few atoms of pure silicon, each of which contains four valence electrons, with atoms of phosphorus, which contain five valence electrons. The “extra” valence electrons greatly increase the conductivity of silicon.

A **p-type semiconductor** consists of silicon doped with a Group 3A element having three valence electrons. The resulting material has a deficiency of valence electrons called **holes**. A hole can be thought of as having a positive charge. Conductivity is increased because electrons can jump from hole to hole as they move through the material. The junction of an n-type semiconductor and a p-type semiconductor is the basis for many electronic devices such as diodes, transistors, and solar cells.

Multiple Choice Questions

Questions 1–4 refer to the following information.

A solid is a poor conductor of electricity, is very hard, and has a high melting point.

1. The solid is probably
 - A) metallic.
 - B) an alloy.
 - C) ionic.
 - D) molecular.
 - E) covalent network.
2. The solid might be
 - A) quartz.
 - B) tin.
 - C) brass.
 - D) table sugar.
 - E) rock salt.
3. The solid could have the chemical formula:
 - A) CaSO_4
 - B) Pb
 - C) $\text{C}_{20}\text{H}_{42}$
 - D) SiC
 - E) P_2O_5
4. The properties of this solid can be attributed to
 - A) an interlocking pattern of atoms.
 - B) intermolecular hydrogen bonding.
 - C) the electron-sea model.
 - D) strong ion–ion interactions.
 - E) interstitial atoms occupying the spaces between atoms.

Examine the atomic-level structures of graphite and diamond in Figure 12.8 and answer Questions 5–8.

5. The orbital hybridization of graphite and diamond, respectively, are
 - A) sp^3 and sp^3 .
 - B) sp^2 and sp^2 .
 - C) sp^3 and sp^2 .
 - D) sp^2 and sp^3 .
 - E) sp and sp^2 .
6. Which contains delocalized electrons?
 - A) Neither graphite nor diamond
 - B) Graphite only
 - C) Diamond only
 - D) Both graphite and diamond
 - E) There is not enough information.
7. Which conducts electricity?
 - A) Neither graphite nor diamond
 - B) Graphite only
 - C) Diamond only
 - D) Both graphite and diamond
 - E) There is not enough information.
8. Graphite is characterized by
 - A) adjacent layers that can slide past each other easily.
 - B) London dispersions that hold the layers together.
 - C) strong covalent bonds between the carbon atoms.
 - D) a high melting, relatively soft material.
 - E) all of these.

Free Response Questions

1. *Using some common laboratory equipment and materials, devise a plan to deduce the type of bonding in a sample of unknown solid. Specify the tests you would perform and discuss what the results mean.*
2.
 - a. *Draw an atomic representation of a binary ionic solid. Use the interactions of the particles to justify the structure and stability of the solid.*
 - b. *Use your model to explain these macroscopic properties of ionic solids:*
 - i. *Melting point*
 - ii. *Conductivity*
 - iii. *Solubility*
 - c. *Make a two-dimensional drawing of the solid showing at the atomic level why it is brittle.*
3. *Draw a molecular depiction of the structure of ice. Label and name the major forces of attraction between the molecules and tell why these forces operate.*
4.
 - a. *Draw an atomic representation of a metal and describe the relative positions of atoms and the electrons.*
 - b. *Use your depiction to explain on an atomic level the following macroscopic properties of metals:*
 - i. *Electrical conductivity*
 - ii. *Malleability*
 - iii. *High thermal conductivity*
5. *Draw a model of an interstitial alloy showing the relative positions of the atoms and the electrons. Explain the composition of a steel alloy and why it is harder and stronger than pure iron.*