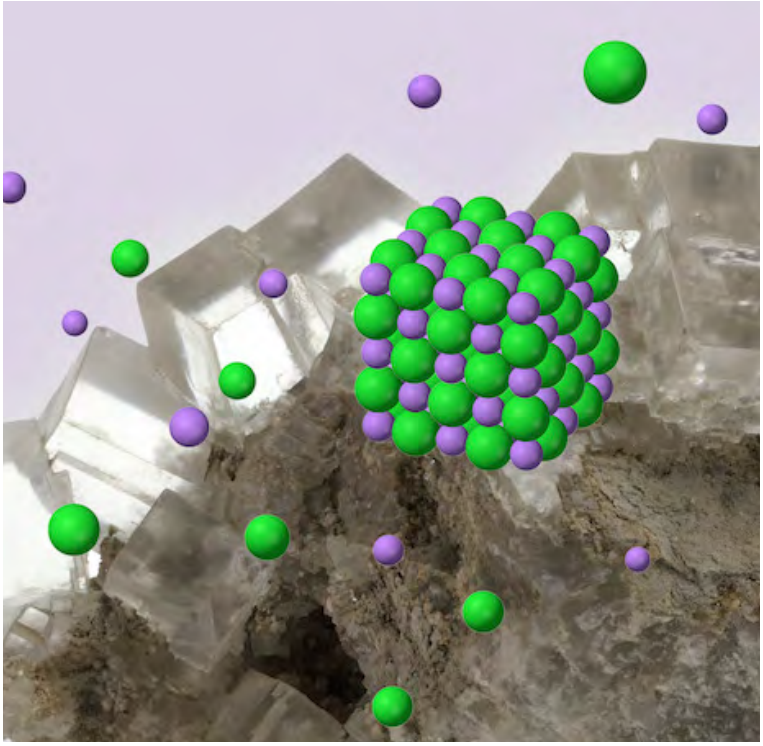


Chapter 6

How Atoms Bond

THE MAIN IDEA

Atoms bond by exchanging or sharing electrons

6.1 Electron-Dot Structures
6.2 Ion Formation
6.3 Ionic Bonds
6.4 Metallic Bonds
6.5 Covalent Bonds
6.6 Molecular Shape
6.7 Polar Covalent Bonds
6.8 Molecular Polarity

Λ These large cubic salt crystals grew many years ago from an evaporating sea in what is now Colorado.

Salt crystals, as shown in the opening photograph, are cubic because that's the way sodium and chlorine join together to form sodium chloride. Smash these crystals with a hammer and you'll merely get smaller cubes! Why are metals opaque to light, and why do they conduct both electricity and heat so well? Again, the answer has to do with how their atoms are bonded. You know carbon dioxide to be a gas at room temperature and water to be a liquid. Would it surprise you that, molecule for molecule, carbon dioxide is over twice as heavy as water? As we explore in this chapter, carbon dioxide molecules are non-sticky, which allows them to float away from each other into a gaseous phase. Conversely, the relatively light water molecules are sticky, which keeps them together in a liquid phase. The reason has to do with, you guessed it, how the atoms within these molecules are bonded. So, if we are to understand why materials behave as they do, it's important that we have some understanding of how it is that their atoms bond.

6.1 Electron-Dot Structures

An atomic model is needed to help us understand how atoms bond. We begin this chapter with a brief overview of the shell model presented in Section 4.8. You may recall how electrons are arranged around an atomic nucleus. Rather than moving in neat orbits like planets around the Sun, electrons are wavelike entities that hover in various volumes of space called *shells*.

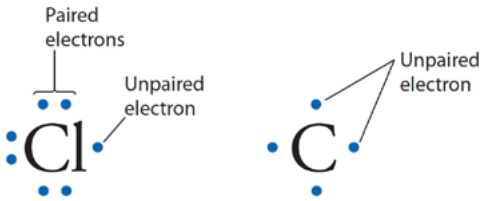
As was shown in Figure 4.32, there are seven shells available to the electrons in an atom, and the electrons fill these shells in order from innermost to outermost. Furthermore, the maximum number of electrons allowed in the first shell is 2, and for the second and third shells it is 8. The fourth and fifth shells can each hold 18 electrons, and the sixth and seventh shells can each hold 32 electrons.* These numbers match the number of elements in each period (horizontal row) of the periodic table. **Figure 6.1** shows how this model applies to the first three elements of group 18.

* As a point of reference for advanced physics students reading this text, these are shells of orbitals grouped by similar energy levels rather than by principal quantum number. They are the "argonian" shells developed by Linus Pauling in the 1930s to explain chemical bonding and the organization of the periodic table. This is an old atomic model, but it works well for a simple description of chemical bonding.

Electrons in the outermost occupied shell of any atom play a significant role in that atom's chemical properties, including its ability to form chemical bonds. To indicate their importance, these outermost electrons are called **valence electrons** (from the Latin valentia, "strength"), and the shell they occupy is called the **valence shell**. Valence electrons can be conveniently represented as a series of dots surrounding an atomic symbol. This notation is called an **electron-dot structure** or, sometimes, a *Lewis dot symbol* in honor of the American chemist Gilbert N. Lewis, who first proposed the concepts of shells and valence electrons (**Figure 6.2**).

The electron-dot structures shown in Figure 6.3 help us to understand ionic and covalent bonds. Electron-dot structures, however, are not so useful in describing metallic bonds. The reason is because within a metallic bond, the bonding electrons readily flow from one atom to the next, as is discussed further in Section 6.4. This is why the metallic groups 3–12 are not included in **Figure 6.3**.

When you look at the electron-dot structure of an atom, you immediately know two important things about that element. You know how many valence electrons it has and how many of these electrons are *paired*. Chlorine, for example, has three pairs of electrons and one unpaired electron, and carbon has four unpaired electrons:



Paired valence electrons are relatively stable with little need to form bonds with other atoms. For this reason, electron pairs in an atom's electron-dot structure are called **nonbonding pairs**. (Do not take this term literally, however, because under the right conditions even "nonbonding" pairs can form a chemical bond. This is explored further in Chapter 10.)

READING CHECK

Electron-dot structures are needed to help us understand what kinds of chemical bonds?

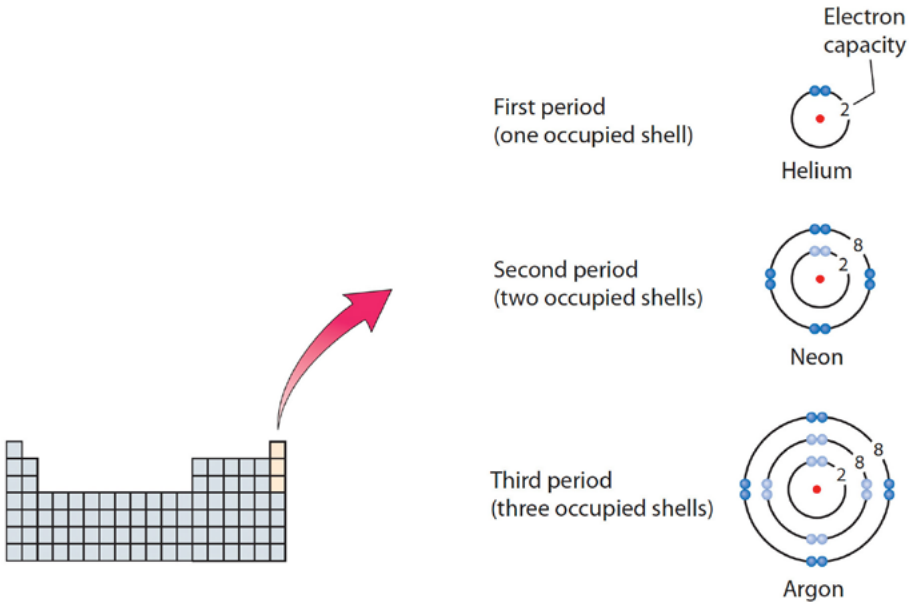


Figure 6.1 >

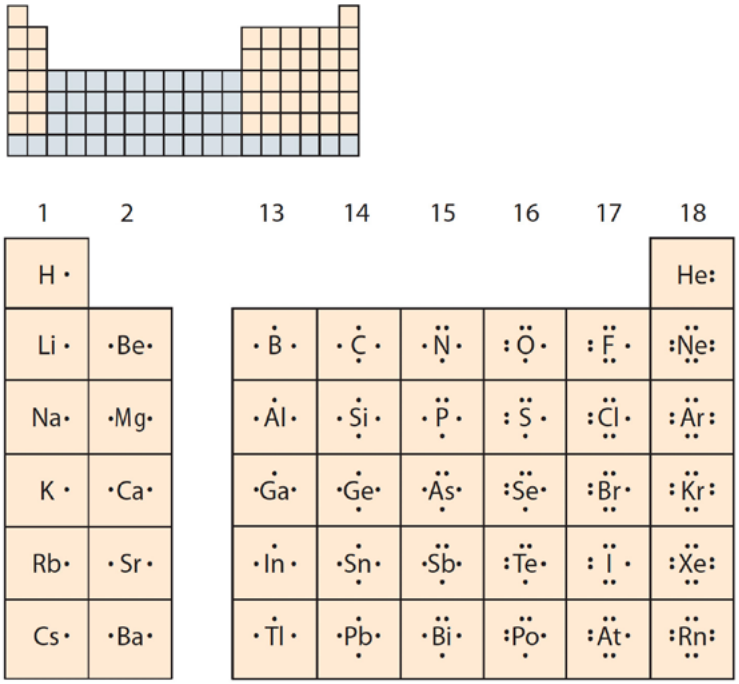
Occupied shells in the group 18 elements helium through argon. Each of these elements has a filled outermost shell. Note that the number of electrons in each shell (2, 8, 8, and so on) corresponds to the number of elements in the periods of the periodic table. For more detail, please review Sections 4.8 and 4.9.

Valence electrons that are *unpaired*, by contrast, have a strong tendency to participate in chemical bonding. By doing so, they become paired with an electron from another atom. The ionic and covalent bonds discussed in this chapter all result from either a transfer or a sharing of unpaired valence electrons.

So why do electrons like to pair up? As was discussed in Chapter 4, electrons have a property called *spin*. Spin can have two states, and these states are analogous to a ball's spinning either clockwise or counterclockwise, as was shown in Figure 4.28. Because an electron has an electric charge, the spinning generates a tiny magnetic field. Two electrons spinning in opposite directions have oppositely aligned magnetic fields, which allows them to come together as a pair.

Λ Figure 6.2

Gilbert Newton Lewis (1875–1946) revolutionized chemistry with his theory of chemical bonding, which he published in 1916. He worked most of his life in the chemistry department of the University of California, Berkeley, where he was not only a productive researcher but also an exceptional teacher. Among his teaching innovations was the idea of providing students with problem sets (homework) as a follow-up to lectures and readings.



< Figure 6.3

The valence electrons of an atom are shown in its electron-dot structure. Note that the first three periods here parallel Figure 4.33. Also note that for larger atoms, not all the electrons in the valence shell are valence electrons. Krypton, Kr, for example, has 18 electrons in its valence shell, but only 8 of these are classified as valence electrons. You can learn the reason for this detail, which involves sub-orbitals, in a follow-up course on advanced chemistry.

CONCEPT **CHECK**

Where are valence electrons located, and why are they important?

CHECK YOUR ANSWER

Valence electrons are located in the outermost occupied shell of an atom. They are important because they play a leading role in determining the chemical properties of the atom.