SECTION 3

OBJECTIVES

- Define *atomic* and *ionic radii, ionization energy, electron affinity,* and *electronegativity.*
- Compare the periodic trends of atomic radii, ionization energy, and electronegativity, and state the reasons for these variations.
- Define *valence electrons,* and state how many are present in atoms of each main-group element.
- Compare the atomic radii, ionization energies, and electronegativities of the *d*-block elements with those of the main-group elements.

Electron Configuration and Periodic Properties

So far, you have learned that the elements are arranged in the periodic table according to their atomic number and that there is a rough correlation between the arrangement of the elements and their electron configurations. In this section, the relationship between the periodic law and electron configurations will be further explored.

Atomic Radii

Ideally, the size of an atom is defined by the edge of its orbital. However, this boundary is fuzzy and varies under different conditions. Therefore, to estimate the size of an atom, the conditions under which the atom exists must be specified. One way to express an atom's radius is to measure the distance between the nuclei of two identical atoms that are chemically bonded together, then divide this distance by two. As illustrated in **Figure 12, atomic radius** *may be defined as one-half the distance between the nuclei of identical atoms that are bonded together.*

Period Trends

Figure 13 gives the atomic radii of the elements and **Figure 14** presents this information graphically. Note that there is a gradual decrease in atomic radii across the second period from lithium, Li, to neon, Ne. *The*

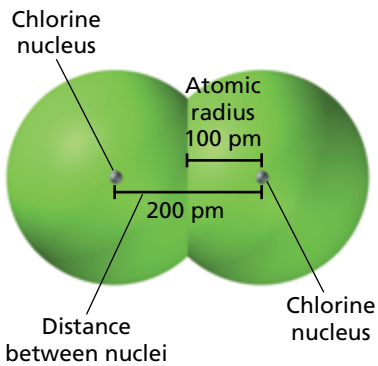

FIGURE 12 One method of determining atomic radius is to measure the distance between the nuclei of two identical atoms that are bonded together in an element or compound, then divide this distance by two. The atomic radius of a chlorine atom, for example, is 100 picometers (pm).