SECTION 3

Electron Configurations

Objectives

- List the total number of electrons needed to fully occupy each main energy level.
- State the Aufbau principle, the Pauli exclusion principle, and Hund's rule.
- Describe the electron configurations for the atoms of any element using orbital notation, electron-configuration notation, and, when appropriate, noble-gas notation.

The quantum model of the atom improves on the Bohr model because it describes the arrangements of electrons in atoms other than hydrogen. *The arrangement of electrons in an atom is known as the atom's* **electron configuration.** Because atoms of different elements have different numbers of electrons, a unique electron configuration exists for the atoms of each element. Like all systems in nature, electrons in atoms tend to assume arrangements that have the lowest possible energies. The lowest-energy arrangement of the electrons for each element is called the element's *ground-state electron configuration.* A few simple rules, combined with the quantum number relationships discussed in Section 2, allow us to determine these ground-state electron configurations.

Rules Governing Electron Configurations

To build up electron configurations for the ground state of any particular atom, first the energy levels of the orbitals are determined. Then electrons are added to the orbitals one by one according to three basic rules. (Remember that real atoms are not built up by adding protons and electrons one at a time.)

The first rule shows the order in which electrons occupy orbitals. According to the **Aufbau principle,** *an electron occupies the lowest-energy orbital that can receive it.* **Figure 16** shows the atomic orbitals in order of increasing energy. The orbital with the lowest energy is the 1*s* orbital. In a ground-state hydrogen atom, the electron is in this orbital. The 2*s* orbital is the next highest in energy, then the 2*p* orbitals. Beginning with the third main energy level, $n = 3$, the energies of the sublevels in different main energy levels begin to overlap.

Note in the figure, for example, that the 4*s* sublevel is lower in energy than the 3*d* sublevel. Therefore, the 4*s* orbital is filled before any electrons enter the 3*d* orbitals. (Less energy is required for two electrons to pair up in the 4*s* orbital than for those two electrons to

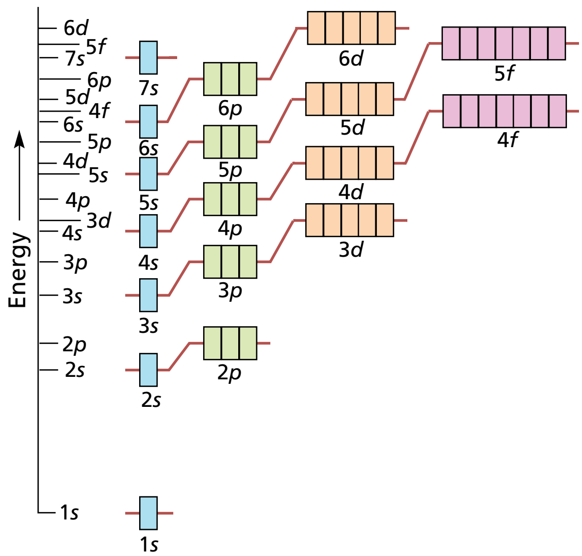

FIGURE 16 The order of increasing energy for atomic sublevels is shown on the vertical axis. Each individual box represents an orbital.