The periodic table is the most important organizing principle in chemistry. If you know the properties of any one element in a group, or column, of the periodic table, you can make a good guess at the properties of every other element in the same group and even of the elements in neighboring groups. Although the periodic table was originally constructed from empirical observations, its scientific underpinnings have long been established and are well understood. To see why it's called the periodic table, look at the graph of atomic radius versus atomic number in Figure 5.1, which shows a periodic rise-and-fall pattern. Beginning on the left with atomic number 1 (hydrogen), the size of the atoms increases to a maximum at atomic number 3 (lithium), then decreases to a minimum, then increases again to a maximum at atomic number 11 (sodium), then decreases, and so on. It turns out that all the maxima occur for atoms of group 1A elements—Li, Na, K, Rb, Cs, and Fr—and that the minima occur for atoms of the group 7A elements—F, Cl, Br, and I.

LIGHT AND THE ELECTROMAGNETIC SPECTRUM

What fundamental property of atoms is responsible for the periodic variations we observe in atomic radii and in so many other characteristics of the elements? This question occupied the thoughts of chemists for more than 50 years after Mendeleev, and it was not until well into the 1920s that the answer was established. To understand how the answer slowly emerged, it's necessary to look first at the nature of visible light and other forms of radiant energy. Historically, studies of the interaction of radiant energy with matter have provided immense insight into atomic structure. Although they appear quite different to our senses, visible light, infrared radiation, microwaves, radio waves, X rays, and other forms of radiant energy are all different kinds of electromagnetic energy. Collectively, they make up the electromagnetic spectrum, shown in Figure 5.2