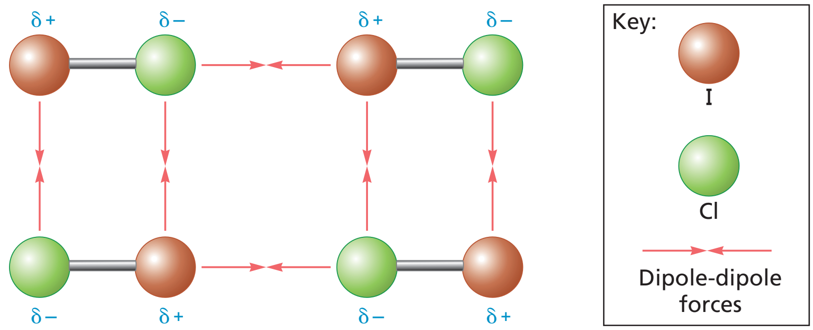

FIGURE 25 Ball-and-stick models illustrate the dipole-dipole forces between molecules of iodine chloride, ICl. In each molecule, the highly electronegative chlorine atom has a partial negative charge, leaving each iodine atom with a partial positive charge. Consequently, the negative and positive ends of neighboring molecules attract each other.

The polarity of diatomic molecules such as ICl is determined by just one bond. For molecules containing more than two atoms, molecular polarity depends on both the polarity and the orientation of each bond. A molecule of water, for example, has two hydrogen-oxygen bonds in which the more-electronegative oxygen atom is the negative pole of each bond. Because the molecule is bent, the polarities of these two bonds combine to make the molecule highly polar, as shown in **Figure 26.** An ammonia molecule is also highly polar because the dipoles of the three nitrogen-hydrogen bonds are additive, combining to create a net molecular dipole. In some molecules, individual bond dipoles cancel one another, causing the resulting molecular polarity to be zero. Carbon dioxide and carbon tetrachloride are molecules of this type.

A polar molecule can *induce* a dipole in a nonpolar molecule by temporarily attracting its electrons. The result is a short-range intermolecular force that is somewhat weaker than the dipole-dipole force. The force of an induced dipole accounts for the solubility of nonpolar O_2 in water. The positive pole of a water molecule attracts the outer electrons

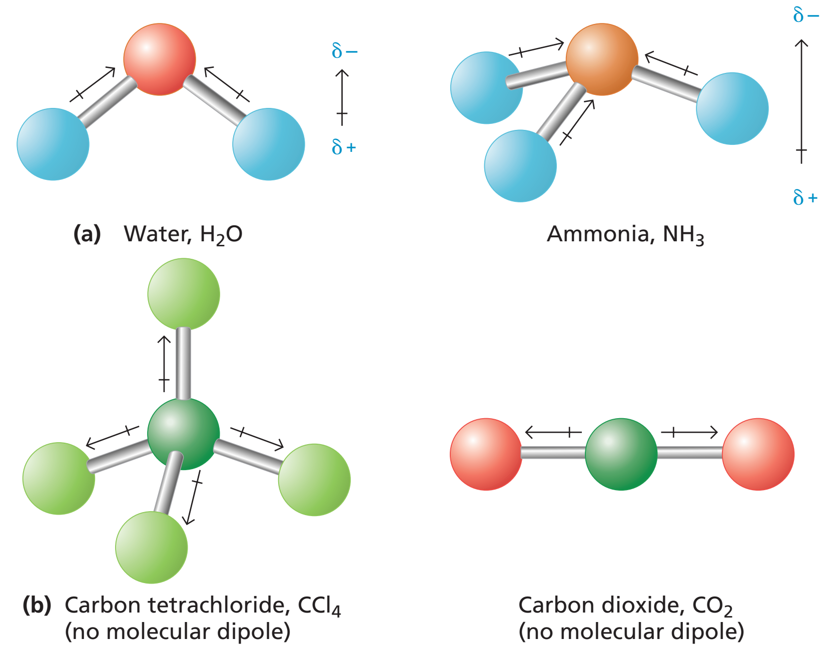

FIGURE 26 (a) The bond polarities in a water or an ammonia molecule are additive, causing the molecule as a whole to be polar. (b) In molecules of carbon tetrachloride and carbon dioxide, the bond polarities extend equally and symmetrically in different directions, canceling each other's effect and causing each molecule as a whole to be nonpolar.