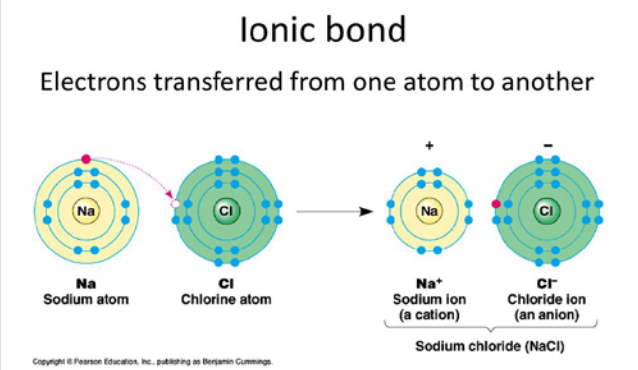
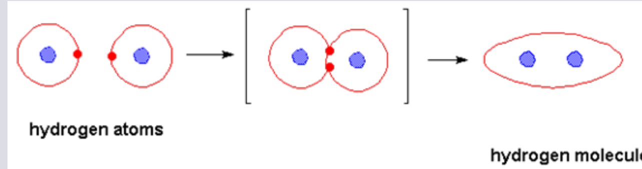
FIGURE 26



A Lewis structure depicting the formation of an ionic bond between sodium and chlorine.

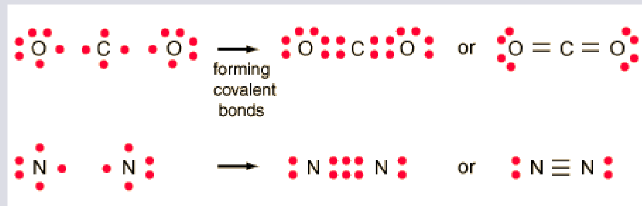
FIGURE 27



A depiction of a single covalent bond.

Image courtesy of Dr. Gary Trammell, University of Illinois at Springfield.

FIGURE 28



A depiction of double and triple covalent bonds.

Lewis structures can be drawn for individual atoms, ions, and atoms in compounds. Complete Lewis structures sometimes show all the electrons, though usually only valence electrons are shown. Lewis also decided that covalent bonds should each have a pair of electrons. The structures shown in Figure 25 are representative Lewis structures.

Notice that the structures for CCl_4, NH_3, O_3 have electron pairs in bonds, but also electron pairs that are not involved in bonding. These are known as non-bonding electron pairs, or simply "lone pairs." The total number of electrons in each neutral structure must add up to the total number of protons in the nuclei of the atoms. In ions, this number can differ by the amount of the value of the ion's charge.

Figure 26 shows how a complete Lewis structure can be used to show the change as a sodium atom gives up its one outer valence electron to the nearly full valence shell of the chlorine atom to form ions of opposite charge that then attract one another in an ionic bond.

Valence Bond Concept

As chemists developed the idea that electrons in atoms were in orbitals, they needed to fit the successful Lewis electron pair bond that explained compound formation and formulas into this picture. This was done by supposing that an electron in an orbital on atom A interacted with an electron in an orbital on atom B with which it formed a bond.

Figure 27 illustrates the idea that the electron orbital shapes, which represent the map of electron density, overlap between the atoms to provide the attractive force in the bond. Such a bond is called a "single" covalent bond.

In some cases, atoms can overlap more than one orbital at the same time to form "double" or even "triple" bonds as shown in Figure 28. Carbon and oxygen share two pairs of electrons in forming two double bonds in carbon dioxide, while nitrogen atoms share three pairs of electrons in forming a triple bond in nitrogen gas.

Hybridization of Atomic Orbitals as an Explanation for Molecular Shapes

Sometimes the interaction of the two atomic orbitals doesn't seem to predict the correct shapes of the molecules. In response to this concern, chemists developed the idea that two (or more) electron orbitals in an atom could combine to form new orbitals that had a different shape. This process is known as **hybridization** and can be used to rationalize the symmetric shapes of many molecules. Figure 29 shows how one s and three p type orbitals can hybridize to form four sp^3 orbitals.

The Formation of Molecular Orbitals

When two or more atoms combine to form a chemical molecule, the bonding electrons take up new physical arrangements between the atoms, pulling the atoms together and forming the bond. Theoretical (mathematical) models can trace out the "electron waves" that constitute these electron positions and