Learning Objectives

• To recognize molecules that are likely to have multiple covalent bonds.

In many molecules, the octet rule would not be satisfied if each pair of bonded atoms shares two electrons. Consider carbon dioxide (CO$_2$). If each oxygen atom shares one electron with the carbon atom, we get the following:

\[
\text{only }6 \text{ e}^-\]

\[
\text{O} \quad \text{C} \quad \text{O}
\]

\[
\text{only }7 \text{ e}^-
\]

This does not give the carbon atom a complete octet; you will find only six electrons in its valence shell. In addition, each oxygen atom has only seven electrons in its valence shell. Finally, no atom makes the number of bonds it typically forms. This arrangement of shared electrons is far from satisfactory.

Sometimes more than one pair of electrons must be shared between two atoms for both atoms to have an octet. In carbon dioxide, a second electron from each oxygen atom is also shared with the central carbon atom, and the carbon atom shares one more electron with each oxygen atom:

\[
\text{O} \quad \text{C} \quad \text{O}
\]

\[
8 \text{ e}^-
\]

In this arrangement, the carbon atom shares four electrons (two pairs) with the oxygen atom on the left and four electrons with the oxygen atom on the right. There are now eight electrons around each atom. Two pairs of electrons shared between two atoms make a double bond between the atoms, which is represented by a double dash:

\[
\text{O} \rightarrow \text{C} \rightarrow \text{O}
\]

Some molecules contain triple bonds, covalent bonds in which three pairs of electrons are shared by two atoms. A simple compound that has a triple bond is acetylene (C$_2$H$_2$), whose Lewis diagram is as follows:
Example 

Draw the Lewis diagram for each molecule.

a. \(\ce{N2}\)
b. \(\ce{CH2O}\) (The carbon atom is the central atom.)

**SOLUTION**

a. The bond between the two nitrogen atoms is a triple bond. The Lewis diagram for \(\ce{N2}\) is as follows:

\[
\text{N} \equiv \text{N}
\]

b. In \(\ce{CH2O}\), the central atom is surrounded by two different types of atoms. The Lewis diagram that fills each atom’s valence electron shell is as follows:

\[
\begin{align*}
\cdot & \cdot \\
\cdot & \cdot \\
\text{H} & \text{H} \\
\end{align*}
\text{ or }
\begin{align*}
\cdot & \cdot \\
\cdot & \cdot \\
\text{H} & \text{H} \\
\end{align*}
\]

Exercise 

Draw the Lewis diagram for each molecule.

a. \(\ce{O2}\)
b. \(\ce{C2H4}\)

formaldehyde and formalin

One application of \(\ce{CH2O}\), also called formaldehyde, is the preservation of biological specimens. Aqueous solutions of \(\ce{CH2O}\) are called formalin and have a sharp, characteristic (pungent) odor.

**Key Takeaway**

- Some molecules must have multiple covalent bonds between atoms to satisfy the octet rule.

**Exercises**

1. What is one clue that a molecule has a multiple bond?

2. Each molecule contains multiple bonds. Draw the Lewis diagram for each. The first element is the central atom.
   a. \(\ce{CS2}\)
b. \( \text{C}_2\text{F}_4 \)

c. \( \text{COCl}_2 \)

3. Each molecule contains double bonds. Draw the Lewis diagram for each. Assume that the first element is the central atom, unless otherwise noted.
   a. \( \text{N}_2 \)
   b. \( \text{HCN} \) (The carbon atom is the central atom.)
   c. \( \text{POCl} \) (The phosphorus atom is the central atom.)

4. Explain why hydrogen atoms do not form double bonds.

5. Why is it incorrect to draw a double bond in the Lewis diagram for \( \text{MgO} \)?

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**Answers**

1. If single bonds between all atoms do not give all atoms (except hydrogen) an octet, multiple covalent bonds may be present.

   a. \[ \begin{array}{ccc}
      & \text{S} & \\
      \text{S} & \text{C} & \text{S} \\
   \end{array} \]

   b. \[ \begin{array}{ccc}
      & \text{F} & \text{C} & \text{C} & \text{F} \\
      \text{F} & \text{C} & \text{C} & \text{F} \\
      \text{F} & \text{F} & \text{F} & \text{F} \\
   \end{array} \]

   c. \[ \begin{array}{ccc}
      & \text{O} & \\
      \text{C} & \text{O} & \text{C} \\
      \text{Cl} & \text{Cl} & \text{Cl} \\
   \end{array} \]

2. Hydrogen can accept only one more electron; multiple bonds require more than one electron pair to be shared.