Exercise 8.9.1 – Explanation of Bonding Patterns: Explain the following bonding patterns.

a. C - 4 bonds and no lone pairs, and no formal charge.

Only the highest energy electrons participate in bonding. Carbon is in group 4A, so it has four valence electrons per atom. It is as if one electron is promoted from the 2s orbital to the 2p orbital. The 2s and the three 2p orbitals blend to form four equivalent sp³ hybrid orbitals.

\[
2s \uparrow \quad 2p \uparrow \uparrow \uparrow \rightarrow 2s \uparrow \quad 2p \uparrow \uparrow \uparrow \rightarrow \text{sp}^3 \uparrow \uparrow \uparrow \uparrow
\]

Covalent bonds form to in order to pair unpaired electrons. The four unpaired electrons lead to four covalent bonds. Because there are no pairs of electrons, carbon atoms have no lone pairs when they form four bonds.

\[
4H^+ + \cdot C\cdot \rightarrow H\cdot C\cdot H \quad \text{or} \quad H-C-H
\]

b. N - 4 bonds, no lone pairs, and a +1 formal charge

Nitrogen is in group 5A, so it has five valence electrons per atom. If a nitrogen atom loses one electron from the pair, it will have four unpaired electrons. The 2s and the three 2p orbitals blend to form four equivalent sp³ hybrid orbitals.

\[
2s \uparrow \quad 2p \uparrow \uparrow \uparrow \rightarrow 2s \uparrow \quad 2p \uparrow \uparrow \uparrow \rightarrow \text{sp}^3 \uparrow \uparrow \uparrow \uparrow
\]

Covalent bonds form to in order to pair unpaired electrons. The four unpaired electrons lead to four bonds. There are no pairs of electrons remaining, so there are no lone pairs. The lost electron leads to the +1 formal charge.

\[
4H^+ + \cdot N\cdot \rightarrow \left[H\cdot N\cdot H\right]^+ \quad \text{or} \quad \left[H-N-H\right]^+
\]
c. O - 1 bond, 3 lone pairs, and a −1 formal charge.

Only the highest energy electrons participate in bonding. Oxygen is in group 6A, so it has six valence electrons per atom. If it gains one electron, it will have a total of seven. The 2s and the three 2p orbitals blend to form four equivalent sp³ hybrid orbitals.

\[
\begin{array}{c}
\text{2s} \\
\uparrow \\
\begin{array}{c}
\text{2p} \\
\uparrow \\
\uparrow \\
\uparrow \\
\end{array}
\end{array}
\rightarrow
\begin{array}{c}
\text{2s} \\
\uparrow \\
\begin{array}{c}
\text{2p} \\
\uparrow \\
\uparrow \\
\uparrow \\
\end{array}
\end{array}
\rightarrow
\begin{array}{c}
\text{sp}^3 \\
\uparrow \\
\uparrow \\
\uparrow \\
\uparrow \\
\end{array}
\]

Covalent bonds form to in order to pair unpaired electrons. The one unpaired electron leads to one bond, and the three pairs of electrons give oxygen atoms with an extra electron three lone pairs. The gained electron leads to the −1 formal charge.

\[
\text{H}^+ + \text{O}^{2-} \rightarrow \left[\text{H}^+\text{O}^{2-}\right] \quad \text{or} \quad \left[\text{H}^+\text{O}^{2-}\right]
\]

d. S - 4 bonds, one lone pair, and no formal charge.

Only the highest energy electrons participate in bonding. Sulfur is in group 6A, so it has six valence electrons per atom. It is as if one electron is promoted from the 3p orbital to an empty 3d orbital. The 3s, the three 3p orbitals, and one 3d orbital blend to form 5 equivalent sp³d hybrid orbitals.

\[
\begin{array}{c}
\text{3d} \\
\uparrow \\
\uparrow \\
\uparrow \\
\uparrow \\
\text{3p} \\
\uparrow \\
\uparrow \\
\uparrow \\
\text{3s} \\
\uparrow
\end{array}
\rightarrow
\begin{array}{c}
\text{3d} \\
\uparrow \\
\uparrow \\
\uparrow \\
\uparrow \\
\text{3p} \\
\uparrow \\
\uparrow \\
\uparrow \\
\text{3s} \\
\uparrow
\end{array}
\rightarrow
\begin{array}{c}
\text{sp}^3\text{d} \\
\uparrow \\
\uparrow \\
\uparrow \\
\uparrow \\
\uparrow \\
\end{array}
\]

Covalent bonds form to in order to pair unpaired electrons. The four unpaired electrons lead to four bonds, and the one pair of electrons is the lone pair.

e. I - 5 bonds, one lone pair, and no formal charge.

Only the highest energy electrons participate in bonding. Iodine is in group 7A, so it has seven valence electrons per atom. It is as if two electrons are promoted from two 5p orbitals to two empty 5d orbitals. The 5s, the three 5p orbitals, and two 5d orbitals blend to form 6 equivalent sp³d² hybrid orbitals.

\[
\begin{array}{c}
\text{5d} \\
\uparrow \\
\uparrow \\
\uparrow \\
\uparrow \\
\text{5p} \\
\uparrow \\
\uparrow \\
\uparrow \\
\text{5s} \\
\uparrow
\end{array}
\rightarrow
\begin{array}{c}
\text{5d} \\
\uparrow \\
\uparrow \\
\uparrow \\
\uparrow \\
\text{5p} \\
\uparrow \\
\uparrow \\
\uparrow \\
\text{5s} \\
\uparrow
\end{array}
\rightarrow
\begin{array}{c}
\text{sp}^3\text{d}^2 \\
\uparrow \\
\uparrow \\
\uparrow \\
\uparrow \\
\uparrow \\
\uparrow \\
\end{array}
\]

Covalent bonds form to in order to pair unpaired electrons. The five unpaired electrons lead to five bonds, and the one pair of electrons is the lone pair.
Exercise 8.9.2 – Drawing Lewis Structures: Draw a reasonable Lewis structure for each of the following. If the structure has resonance, draw all the reasonable resonance structures and the resonance hybrid.

a. CH$_3$Br
   
   ![Lewis structure for CH$_3$Br]

b. ClF$_3$
   
   ![Lewis structure for ClF$_3$]

c. CH$_2$O
   
   ![Lewis structure for CH$_2$O]

d. CN$^-$
   
   ![Lewis structure for CN$^-$]

e. CF$_3$CHCl$_2$
   
   ![Lewis structure for CF$_3$CHCl$_2$]

f. C$_2$H$_6$O
   
   ![Lewis structure for C$_2$H$_6$O]

or

   ![Alternative Lewis structure for C$_2$H$_6$O]

g. C$_2$H$_4$F$_2$
   
   ![Lewis structure for C$_2$H$_4$F$_2$]

or

   ![Alternative Lewis structure for C$_2$H$_4$F$_2$]

h. C$_3$H$_7$OH
   
   ![Lewis structure for C$_3$H$_7$OH]

or

   ![Alternative Lewis structure for C$_3$H$_7$OH]
i. \( \text{HCO}_2^- \)

\[
\begin{array}{c}
\text{H} \quad \text{C} \quad \text{O} \\
\text{H} \quad \text{C} \quad \text{O}
\end{array}
\]

\[
\begin{array}{c}
\text{H} \quad \text{C} \quad \text{O} \\
\text{H} \quad \text{C} \quad \text{O}
\end{array}
\]

\[
\begin{array}{c}
\text{H} \quad \text{C} \quad \text{O} \\
\text{H} \quad \text{C} \quad \text{O}
\end{array}
\]

j. \( \text{CH}_2\text{CHF} \)

\[
\begin{array}{c}
\text{H} \\
\text{C} \\
\text{C} \\
\text{C} \\
\text{H}
\end{array}
\]

k. \( \text{HCO}_2\text{H} \)

\[
\begin{array}{c}
\text{H} \\
\text{C} \\
\text{O} \\
\text{H}
\end{array}
\]

l. \( \text{NH}_2\text{COCH}_3 \)

\[
\begin{array}{c}
\text{H} \\
\text{N} \\
\text{C} \\
\text{C} \\
\text{H}
\end{array}
\]

\[
\begin{array}{c}
\text{H} \\
\text{N} \\
\text{C} \\
\text{C} \\
\text{H}
\end{array}
\]

\[
\begin{array}{c}
\text{N} \\
\text{C} \\
\text{C} \\
\text{C} \\
\text{H}
\end{array}
\]

m. \( \text{SO}_4^{2-} \)

\[
\begin{array}{c}
\text{S} \\
\text{O} \\
\text{O} \\
\text{O}
\end{array}
\]

\[
\begin{array}{c}
\text{S} \\
\text{O} \\
\text{O} \\
\text{O}
\end{array}
\]

\[
\begin{array}{c}
\text{S} \\
\text{O} \\
\text{O} \\
\text{O}
\end{array}
\]

\[
\begin{array}{c}
\text{S} \\
\text{O} \\
\text{O} \\
\text{O}
\end{array}
\]

\[
\begin{array}{c}
\text{S} \\
\text{O} \\
\text{O} \\
\text{O}
\end{array}
\]
Exercise 8,9.3 – Predicting Molecular Polarity: Decide whether the molecules represented by the following formulas are polar or nonpolar. (You may need to draw Lewis structures and geometric sketches in order to do this.)

a. $\text{BCl}_3$

The 1.12 difference in electronegativity between B (2.04) and Cl (3.16) tells us that the bonds are polar covalent bonds, but there is a symmetrical distribution of these polar bonds. $\text{BCl}_3$ molecules are nonpolar.

b. $\text{C}_3\text{H}_8$

All hydrocarbons are composed of nonpolar molecules.

c. $\text{CS}_2$

The 0.03 difference in electronegativity between C (2.55) and S (2.58) tells us that the bonds are nonpolar covalent bonds. If there are no polar bonds, the molecules are nonpolar.

d. $\text{NH}_3$

The 0.84 difference in electronegativity between N (3.04) and H (2.20) tells us that the bonds are polar covalent bonds. The asymmetrical distribution of these polar bonds leads to polar $\text{NH}_3$ molecules. All molecular compounds with the N-H bond are polar.
e. CBr₄

![Diagram of CBr₄]

The 0.44 difference in electronegativity between C (2.55) and Br (2.96) tells us that the bonds are polar covalent bonds, but there is a symmetrical distribution of these polar bonds. CBr₄ molecules are nonpolar.

f. CH₃CO₂H

This is a carboxylic acid (acetic acid), and all carboxylic acids are composed of polar molecules.

g. Br₂O

![Diagram of Br₂O]

The 0.48 difference in electronegativity between O (3.44) and Br (2.96) tells us that the bonds are polar covalent bonds. The asymmetrical distribution of these polar bonds leads to polar Br₂O molecules.

h. IF₃

![Diagram of IF₃]

The 1.32 difference in electronegativity between I (2.66) and F (3.98) tells us that the bonds are polar covalent bonds. The asymmetrical distribution of these polar bonds leads to polar IF₃ molecules.
Exercise 8.9.4 – Predicting Type of Attraction:  What is the primary type of attraction that holds the particles of each of these substances in the liquid or solid form.

a. Ir metallic bonds
b. AlF₃ ionic bonds
c. C₂H₆ London forces
d. HCN dipole-dipole attractions
e. C(dia) covalent bonds
f. CH₃NH₂ hydrogen bonds
g. NH₄OH ionic bonds
h. BrF₃ dipole-dipole attractions
i. CH₂O dipole-dipole attractions
j. C₃H₇OH hydrogen bonds

Exercise 8.9.5 - Predicting Types and Strengths of Attractions Between Particles: For each of the following substances, write the name for the type of particle that forms its basic structure and the name of the primary type of attraction that can form between these particles. From each pair of substances, choose the one that you expect would have stronger interparticle attractions.

a. ethanol, C₂H₅OH (in alcoholic beverages)
    Polar molecules and Hydrogen bonds (and London forces)
    or  dimethyl ether, CH₃OCH₃ (refrigerant and solvent for propellant sprays)
    Polar molecules and dipole-dipole attractions (and London forces)
    For molecules of about the same size, hydrogen bonds are stronger than dipole-dipole attractions, so we expect C₂H₅OH to have the stronger attractions.

b. manganese, Mn (improves corrosion resistance and hardness when alloyed with other metals)
    Cations in a sea of electrons and metallic bonds
    or  carbon tetrachloride, CCl₄ (used for metal degreasing and to add chlorine atoms to organic compounds)
    Nonpolar molecules and London forces
    Metallic bonds are generally stronger than London forces, so we expect Mn to have the stronger attractions.
c. propane, $C_3H_8$ (fuel and aerosol propellant)
   or n-heptane, $C_7H_{16}$ (standard for octane rating – 100% n-hexane is zero octane)

   **Both nonpolar molecules and London forces**
   The larger the molecules are, the stronger the London forces between them, so $C_7H_{16}$ has stronger attractions.

d. acetone, $CH_3COCH_3$ (solvent for paints and varnishes)

   **Polar molecules and dipole-dipole attractions (and London forces)**
   or lithium nitrate, $LiNO_3$ (used to make ceramics and in rocket propellants)

   **Cations and anions held together by ionic bonds**
   Ionic bonds are generally stronger than dipole-dipole attractions, so $LiNO_3$ has the stronger attractions.

e. antimony tribromide, $SbBr_3$ (a molecular compound used as a mordant, which fixes dyes to textiles)

   ![Polar Molecules and dipole-dipole attractions (and London forces)]

   or antimony trichloride, $SbCl_3$ (a molecular compound used for fireproofing textiles)

   ![Polar Molecules and dipole-dipole attractions (and London forces)]

   London forces play a part in the attractions between polar molecules as well as nonpolar molecules. The larger the molecules are, the stronger the London forces between them are. Therefore, $SbBr_3$ has stronger attractions.