When using Localized Electron Model to describe bonding in a molecule, follow these three steps,
1. Draw the Lewis structure
2. Determine the arrangement of electron pairs using the VSEPR model.
3. Specify the hybrid orbitals needed to accommodate the electron pairs.

Use Localized Electron Model to describe bonding in the following molecules.
1. CO
2. CO₂
3. BF₄⁻
4. XeF₂
5. C₃H₄
6. I₃
7. N₂
Drawing VESPR Structures

Determine the geometry of each of the following structures or ions.

1. CO₂

2. SO₄²⁻

3. BrF₃

4. XeO₄

5. ICl₂⁺
### Predicted Molecular Structure (VSEPR)

**Total Sets of Electronic Regions**

<table>
<thead>
<tr>
<th>Bond Angles</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
</tr>
</thead>
<tbody>
<tr>
<td>180°</td>
<td>Linear</td>
<td>Trigonal Planar</td>
<td>Tetrahedral</td>
<td>Trigonal Bipyramidal</td>
<td>Octahedral</td>
</tr>
<tr>
<td>120°</td>
<td>Bent</td>
<td>Trigonal Pyramidal</td>
<td>See-saw</td>
<td>Square Pyramidal</td>
<td></td>
</tr>
<tr>
<td>109.5°</td>
<td>Linear</td>
<td>Bent</td>
<td>T-shape</td>
<td>Square Planar</td>
<td></td>
</tr>
<tr>
<td>120°/90°</td>
<td>Linear</td>
<td>Linear</td>
<td>T-shape</td>
<td></td>
<td></td>
</tr>
<tr>
<td>90°</td>
<td>Linear</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Unshared Pairs of Electrons (aka Lone Pairs):**

- 0 pairs: $AB_2$
- 1 pair: $AB_3$
- 2 pairs: $AB_4$
- 3 pairs: $AB_5$
- 4 pairs: $AB_6$
- 5 pairs: $AB_7$
- 6 pairs: $AB_8$

### Arrangement of Valence Pairs

(Bonding & Lone)

<table>
<thead>
<tr>
<th>Total Regions</th>
<th>Electronic Structure</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Linear</td>
</tr>
<tr>
<td>3</td>
<td>Trigonal Planar</td>
</tr>
<tr>
<td>4</td>
<td>Tetrahedral</td>
</tr>
<tr>
<td>5</td>
<td>Trigonal Bipyramidal</td>
</tr>
<tr>
<td>6</td>
<td>Octahedral</td>
</tr>
</tbody>
</table>

*An electronic region is the shared electrons between 2 atoms. One region may include one, two or three pairs of electrons.*
Ch. 9
Covalent Bonding

9.1 Hybridization

Problems in Bonding
Example: methane, CH₄

1. If kept separate, the bonding between valence orbitals should be varied:
   - one type for 1s from H and 2s from C
   - one type for 1s from H and 2p from C
   - but all the bonds are equal (bond energies!)

2. 2p orbitals are perpendicular but C-H bonds are at 109.5° angles
Problems in Bonding
Example: methane, CH₄

Solutions for Bonding
Example: methane, CH₄

- carbon must accept a new set of orbitals used for bonding other than its 2s and 2p
- all 4 must be the same
- must form tetrahedral shape (109.5°)

- hybridization - mixing of normal orbitals to form special bonding orbitals
sp\(^3\) hybridization of CH\(_4\)

- four new sp\(^3\) orbitals are created from one 2s and three 2p orbitals
- identical in shape:
  - one large lobe
  - one small lobe
- tetrahedral arrangement (109.5°)
- *When a set of equal tetrahedrally-arranged orbitals is needed, the atom adopts a set of four sp\(^3\) orbitals*
sp\(^3\) hybridization of CH\(_4\)

- three new sp\(^3\) orbitals are created from one 2s and two 2p orbitals
- identical in shape:
  - one large lobe
  - one small lobe
- trigonal planar arrangement (120°)
- When a set of equal trigonal planar-arranged orbitals is needed, the atom adopts a set of three sp\(^2\) orbitals
**sp² hybridization of C₂H₄**

Orbitals in an isolated carbon atom → Carbon orbitals in ethylene

---

**sp² hybridization of C₂H₄**

Diagram showing the hybridization process and orbital overlap.
sp\(^2\) hybridization of C\(_2\)H\(_4\)

- What happens to the other 2p orbital left out of the hybridization in each C?
- They share an electron pair
- Occupies the space above and below the sharing between sp\(^2\) orbitals
- Forms pi (\(\pi\)) bond

\[\text{sp}^2 \text{ hybridization of } \text{C}_2\text{H}_4\]

- The sp\(^2\) orbitals are used to share electrons too
- Shared in the area centered on a direct line between 2 atoms
- Form a sigma (\(\sigma\)) bond

\[
\begin{align*}
\text{Sideways overlap} & \quad \text{-- a } \pi \text{ bond} \\
\text{Head-on overlap} & \quad \text{-- a } \sigma \text{ bond}
\end{align*}
\]
sp² hybridization of C₂H₄

Each double bond consists of:

- one pi
- one sigma

Each single bond consists of:

- one sigma
sp hybridization of CO$_2$

- Two new sp orbitals are created from one 2s and one 2p orbitals.
- Identical in shape:
  - One large lobe
  - One small lobe
- Linear arrangement (180°)
- When a set of equal linearly-arranged orbitals is needed, the atom adopts a set of two sp orbitals:
  $$\text{O} \equiv \text{C} \equiv \text{O}$$
What happens to the other two 2p orbitals left out of the hybridization in each C?

- they share electron pairs with the one extra 2p orbital from sp$^2$ hybridized O
sp hybridization of CO₂

- sigma bond (1 pair of electrons)
- pi bond (1 pair of electrons)

pi bond
(1 pair of electrons)

sp³d hybridization of PCl₅

- five new sp³d orbitals are created from one 3s, three 3p, and one 3d orbitals
- trigonal bipyramidal arrangement
- When a set of equal trigonal bipyramidal -arranged orbitals is needed, the atom adopts a set of five sp³d orbitals
sp\(^3\)d\(^2\) hybridization of SF\(_6\):

- six new sp\(^3\)d\(^2\) orbitals are created from one 3s, three 3p, and two 3d orbitals
- octahedral arrangement
- When a set of equal octahedrally arranged orbitals is needed, the atom adopts a set of six sp\(^3\)d\(^2\) orbitals
Example 1  Describe the bonding in ammonia using the LE model

steric #: 4
PG:tetrahedral
MS: trigonal pyramidal

three sigma bonds, no pi bonds

Example 2  Describe the bonding in $N_2$ using the LE model

one sigma bond
two pi bonds
Example 3
Describe the bonding in $I_3^-$ using the LE model

- steric #: 5
- PG: trigonal bipyramidal
- MS: linear
- requires 5: sp$^3$d for central I
  - 3 hold lone pairs
  - 2 make sigma bonds with outside I's sp$^3$
- requires 4: sp$^3$ for outside I's
  - 3 hold lone pairs
  - 1 makes sigma bond with central I's sp$^3$d

Example 4
Describe the bonding in XeF$_4$ using the LE model

- steric #: 6
- PG: octahedral
- MS: square planar
- requires 6: sp$^3$d$^2$ for central Xe
  - 2 hold lone pairs
  - 4 make sigma bonds with outside F's sp$^3$
- requires 4: sp$^3$ for outside F's
  - 3 hold lone pairs
  - 1 makes sigma bond with central Xe's sp$^3$d$^2$
Example 9.1 A  \( sp^3 \) Hybridization

Describe the bonding in the water molecule using the localized electron model.

Strategy

We must establish the VSEPR structure. This is done by drawing the Lewis structure and determining the number of effective electron pairs around the central atom.

Solution

The Lewis structure of \( \text{H}_2\text{O} \) is (see Sections 8.10 and 8.13)

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

Four effective electron pairs around the oxygen atom indicate a tetrahedral basis leading to a V-shaped structure. Therefore the oxygen is \( sp^3 \) hybridized. Two \( sp^3 \) orbitals are occupied with 1s electrons from hydrogen. The other two \( sp^3 \) orbitals are occupied by lone pairs.

In molecules or ions with a trigonal planar configuration (3 effective electron pairs around an atom), \( sp^2 \) hybridization occurs. With this hybridization, an \( s \) orbital and two \( p \) orbitals are used. That leaves an unhybridized (unchanged) \( p \) orbital perpendicular to the \( sp^2 \) plane.

Look at Figure 9.12 in your textbook.

- Bonds formed from the overlap of orbitals in the plane between two atoms are called sigma (\( \sigma \)) bonds.
- Bonds formed by the overlap of unhybridized \( p \) orbitals (above and below the center plane) are called pi (\( \pi \)) bonds.
- A single bond is a \( \sigma \) bond.
- A double bond consists of one \( \sigma \) and one \( \pi \) bond.
- A triple bond consists of one \( \sigma \) bond and two \( \pi \) bonds.

Example 9.1 B  Sigma and Pi Bonds

How many \( \sigma \) bonds are there in the commercial insecticide, "Sevin," shown below? How many \( \pi \) bonds?

Solution

There are 27 \( \sigma \) bonds (single bonds and one of the bonds in a double bond). There are 6 \( \pi \) bonds (the other bond in a double bond).
Your textbook goes over a variety of different hybridization schemes, all of which center on the idea that the hybridization of the orbitals of an atom depends on the total number of effective electron pairs around it. In order to determine the hybridization of an atom, it is essential that you can figure out its VSEPR structure.

The following table summarizes effective electron pairs around an atom with hybridization. Remember that, for purposes of the VSEPR model, double and triple bonds count as only one effective electron pair.

<table>
<thead>
<tr>
<th>Effective Electron Pairs Around an Atom</th>
<th>Arrangement</th>
<th>Hybridization</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>linear</td>
<td>sp</td>
</tr>
<tr>
<td>3</td>
<td>trigonal planar</td>
<td>sp²</td>
</tr>
<tr>
<td>4</td>
<td>tetrahedral</td>
<td>sp³</td>
</tr>
<tr>
<td>5</td>
<td>trigonal bipyramid</td>
<td>sp³d</td>
</tr>
<tr>
<td>6</td>
<td>octahedral</td>
<td>sp³d²</td>
</tr>
</tbody>
</table>

Keep in mind that a ligand can have a hybridization different from that of the central atom. Each atom in a molecule must be considered separately based on the Lewis and VSEPR structures of that molecule.

**Example 9.1 C  Practice with Hybrid Orbitals**

Give the hybridization, and predict the geometry of each of the central atoms in the following molecules or ions.

a. IF₅⁺

b. OSF₄ (sulfur is the central atom)

c. SiF₆²⁻

d. HCCCH (work with both carbons)

**Solution**

a. Lewis structure: \[ :F—I—F: \]

All three atoms have 4 effective electron pairs. They are all \( sp^3 \) hybridized. The VSEPR structure has a tetrahedral basis. Because the central atom (iodine) has two bonding pairs, it will take on a V-shape, but the bond angle will be smaller than 109.5°.

b. Lewis structure:

Sulfur has 5 effective electron pairs. The VSEPR structure is a trigonal bipyramid. **Sulfur is \( sp^3d \) hybridized.** Each of the fluorines has 4 effective electron pairs.

c. Lewis structure:

Silicon has 6 effective electron pairs. The VSEPR structure is octahedral. **Silicon is \( sp^3d^2 \) hybridized.** Each of the fluorines has 4 effective electron pairs.
d. Lewis structure: 

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

Each carbon has 2 effective electron pairs. The VSEPR structure is linear. Each carbon is sp hybridized. Hydrogen atoms bond using 1s orbitals. The orbitals are unhybridized.

**Example 9.1 D  Summing it All up**

Answer the following questions regarding aspartame (NutraSweet\textsuperscript{TM}).

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

a. How many \(\sigma\) bonds are in the molecule? 

b. How many \(\pi\) bonds? 

c. What is the hybridization on carbon \(a\)? Carbon \(b\)? 

d. What is the hybridization on nitrogen \(a\)? Oxygen \(a\)? 

e. What is the C\(_b\)-O\(_b\)-H\(_a\) bond angle?

**Solution**

You must remember to complete the octets in this "shorthand" Lewis structure. Lone pairs are often "assumed," so always be on the lookout.

a. 39 \(\sigma\) bonds 

b. 6 \(\pi\) bonds 

c. \(sp^3, sp^2\) (3 effective electron pairs) 

d. \(sp^3\) (remember the "assumed" lone pair to complete the octet), \(sp^3\) (3 effective electron pairs, including the two to complete the octet) 

e. O\(_b\) is \(sp^3\) hybridized (complete the octet!); therefore, the angle is based on a tetrahedron, with two lone pairs compressing the C\(_b\)-O\(_b\)-H\(_a\) bond angle to 104.5°.
BOND LENGTH

(Chemistry 6th ed. page 372 / 7th ed. page 352)

As the number of bonds between two atoms increases, the bond grows shorter and stronger.

EXAMPLE: Arrange the following molecules in order of decreasing C—C bond length: C₂H₄, C₂H₂, C₂H₆.

In order of decreasing C—C bond length: C₂H₆, C₂H₄, C₂H₂.

First, draw the Lewis structure for each molecule. The C—C triple bond is the shortest, followed by the double bond, and the single bond.

Molecular Shapes and Bond Angles

<table>
<thead>
<tr>
<th>No. of Electron Pairs</th>
<th>Electron Pair Geometry (Bond Angle)</th>
<th>No. of Bonding Electron Pairs</th>
<th>No. of Lone Electron Pairs</th>
<th>Molecular Geometry</th>
<th>Formula</th>
<th>2D Structure</th>
<th>Hybrid Orbital</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Linear (180°)</td>
<td>2</td>
<td>0</td>
<td>Linear</td>
<td>BeH₂</td>
<td><img src="image" alt="BeH2 structure" /></td>
<td>sp</td>
</tr>
<tr>
<td>3</td>
<td>Trigonal planar (120°)</td>
<td>3</td>
<td>0</td>
<td>Trigonal planar</td>
<td>CO₃²⁻</td>
<td><img src="image" alt="CO32- structure" /></td>
<td>sp²</td>
</tr>
<tr>
<td>3</td>
<td>Trigonal planar (120°)</td>
<td>3</td>
<td>1</td>
<td>Bent</td>
<td>NO₃⁻</td>
<td><img src="image" alt="NO3- structure" /></td>
<td>sp²</td>
</tr>
<tr>
<td>4</td>
<td>Tetrahedral (109.5°)</td>
<td>4</td>
<td>0</td>
<td>Tetrahedral</td>
<td>CH₄</td>
<td><img src="image" alt="CH4 structure" /></td>
<td>sp³</td>
</tr>
<tr>
<td>No. of Electron Pairs</td>
<td>Electron Pair Geometry (Bond Angle)</td>
<td>No. of Bonding Electron Pairs</td>
<td>No. of Lone Electron Pairs</td>
<td>Molecular Geometry</td>
<td>Formula</td>
<td>2D Structure</td>
<td>Hybrid Orbital</td>
</tr>
<tr>
<td>----------------------</td>
<td>------------------------------------</td>
<td>-----------------------------</td>
<td>---------------------------</td>
<td>-------------------</td>
<td>---------</td>
<td>-------------</td>
<td>---------------</td>
</tr>
<tr>
<td>4</td>
<td>Tetrahedral (&gt;109.5°)</td>
<td>3</td>
<td>1</td>
<td>Trigonal pyramidal</td>
<td>NH₃</td>
<td><img src="image" alt="NH₃ Structure" /></td>
<td>sp³</td>
</tr>
<tr>
<td>4</td>
<td>Tetrahedral (&gt;&gt;109.5°)</td>
<td>2</td>
<td>2</td>
<td>Bent</td>
<td>H₂O</td>
<td><img src="image" alt="H₂O Structure" /></td>
<td>sp³</td>
</tr>
<tr>
<td>5</td>
<td>Trigonal bipyramidal (90°, 120°)</td>
<td>5</td>
<td>0</td>
<td>Trigonal bipyramidal</td>
<td>PCl₅</td>
<td><img src="image" alt="PCl₅ Structure" /></td>
<td>sp³d</td>
</tr>
<tr>
<td>5</td>
<td>Tetrahedral (90° and &gt;120°)</td>
<td>4</td>
<td>1</td>
<td>Seesaw unsymmetrical tetrahedron</td>
<td>SF₄</td>
<td><img src="image" alt="SF₄ Structure" /></td>
<td>sp³d</td>
</tr>
<tr>
<td>No. of Electron Pairs</td>
<td>Electron Pair Geometry (Bond Angle)</td>
<td>No. of Bonding Electron Pairs</td>
<td>No. of Lone Electron Pairs</td>
<td>Molecular Geometry</td>
<td>Formula</td>
<td>2D Structure</td>
<td>Hybrid Orbital</td>
</tr>
<tr>
<td>----------------------</td>
<td>-------------------------------------</td>
<td>-------------------------------</td>
<td>----------------------------</td>
<td>-------------------</td>
<td>---------</td>
<td>--------------</td>
<td>---------------</td>
</tr>
<tr>
<td>5</td>
<td>Tetrahedral (90° and 180°)</td>
<td>3</td>
<td>2</td>
<td>T-shaped</td>
<td>BrF₃</td>
<td><img src="image1.png" alt="2D Structure" /></td>
<td>sp³d</td>
</tr>
<tr>
<td>5</td>
<td>Tetrahedral (180°)</td>
<td>2</td>
<td>3</td>
<td>Linear</td>
<td>ICl⁻</td>
<td><img src="image2.png" alt="2D Structure" /></td>
<td>sp³d</td>
</tr>
<tr>
<td>6</td>
<td>Octahedral (90°)</td>
<td>6</td>
<td>6</td>
<td>Octahedral</td>
<td>SF₆</td>
<td><img src="image3.png" alt="2D Structure" /></td>
<td>sp³d²</td>
</tr>
<tr>
<td>6</td>
<td>Octahedral (90°)</td>
<td>5</td>
<td>1</td>
<td>Square pyramidal</td>
<td>BrF₃</td>
<td><img src="image4.png" alt="2D Structure" /></td>
<td>sp³d²</td>
</tr>
</tbody>
</table>
EXAMPLE: Predict the molecular structure of the carbon dioxide molecule. Is this molecule expected to have a dipole moment?

First we must draw the Lewis structure for the CO₂ molecule.

\[ \overset{\cdot}{\overset{\cdot}{O}} = C = \overset{\cdot}{\overset{\cdot}{O}} \]

In this structure for CO₂ there are two effective pairs around the central atom (each double bond is counted as one effective pair). There are two atoms attached to the central atom.

According to the table, a linear arrangement is required.

Each C-O bond is polar, but, since the molecule is linear, the dipoles cancel out and the molecule is nonpolar.

EXAMPLE: Draw the Lewis structure for PF₅ and identify the molecular geometry.

PF₅ has 40 valence electrons. Connecting 5 F atoms to P, the central atom, uses 10 electrons in five single bonds. The 30 remaining electrons can be used to satisfy the octet rule for each F atom. There are 5 electron pairs around P and 5 atoms attached to P, resulting in a trigonal bipyramidal shape.

EXAMPLE: Draw the Lewis structure for CH₃OH.

This molecule has more than one central atom, the carbon and the oxygen. Hydrogen cannot be in the middle since it can form
only one bond. Draw the correct Lewis structure and then using the rules for VSEPR, assign the molecular geometry around each central atom. The carbon atom has four electron pairs around it and four atoms attached resulting in tetrahedral geometry. Oxygen also has four electron pairs around it, but only has two atoms attached. This leads to bent geometry around oxygen.

**EXAMPLE:** Arrange the following molecules in order of increasing bond angles: \( \text{H}_2\text{S}, \text{CCl}_4, \text{NF}_3 \).

First, draw the correct Lewis structure for each molecule and then assign an approximate bond angle for each.

Each of the four molecules has four electron pairs around the central atom, resulting in a tetrahedral arrangement of the electron pairs about the central atom and approximately a 109.5° bond angle. \( \text{H}_2\text{S} \) has two lone pairs on the central atom; \( \text{NF}_3 \) has one lone pair on the central atom. These lone pairs repel more than the bonded electron pairs and will cause the bond angle to be smaller than the expected 109.5°. The more lone pairs, the smaller the bond angle.

In order of increasing bond angles, the answer is \( \text{H}_2\text{S} < \text{NF}_3 < \text{CCl}_4 \).

**Exercises**

**Section 8.1**

1. Indicate whether the bonds between the following would be primarily covalent, polar covalent, or ionic:
   a. O–H  
   b. Ca–Cl  
   c. H–Cl  
   d. Br–Br

2. Calculate the energy of interaction for KCl if the internuclear distance is 0.314 nm.

3. Calculate the energy of interaction between Ag⁺ and Br⁻ if the internuclear distance of AgBr is 0.120 nm (in kJ/mole).

**Section 8.2**

4. Using a periodic table, order the following from lowest to highest electronegativity:
   a. Fr, Mg, Rb  
   b. B, Al, C, N  
   c. P, As, Ga, O  
   d. Cl, S, P

5. Using the periodic chart of elements, place the following in order from the lowest to the highest electronegativity:
   F, Nb, N, Si, Rb, Ca, Pt

6. Using Figure 8.3 in your textbook, calculate the difference in electronegativity (\( \Delta \)) for each of the following bonds:
   a. Cl–Cl  
   b. K–Br  
   c. Fe–O  
   d. H–O  
   e. S–H
7. Place the following in order of increasing polarity:
   NaBr, I₂, H₂O, MnO₂, CN⁻

8. Which of the following molecules contain polar covalent bonds? List in order of increasing bond polarity. (Use Fig. 8.3 in your text).
   O₃, P₄, NO, CO₂, CH₄, H₂S

9. How will the charge be distributed on each of the following molecules: HF, NO, CO, and HCl?

10. Why is it that BeF₂ is ionic, and BeCl₂ is covalent?

Section 8.3

11. Determine the orientation of the dipole of the following, if any.
   a. AlCl₃ (planar with aluminum atom at the center)
   b. CH₃F (tetrahedral with carbon at the center)
   c. N₂O (linear with N–N–O structure)
   d. AgCl₂ (planar molecule, silver atom at center, Ag–Cl bonds 90° apart)

12. Which of the molecules in problem 11 contain one or more polar bonds?

13. Which of the following molecules would you expect to have a dipole moment of zero? Describe the dipole orientation of the other two molecules.
   a. KI
   b. CF₄ (tetrahedral structure)
   c. H₂Se (bent structure)

Section 8.4

14. Determine the most stable ion for each of the following atoms, and indicate which element they would be isoelectronic with if they lost or gained electrons:
   a. O
c. I
e. Na
   b. Be
d. Te

15. List four ions that are isoelectronic with argon and have charges from −2 to +2. Arrange these in order of increasing ionic radius.

16. List four ions that are isoelectronic to Kr. Arrange these in order of increasing radius.

17. Determine the formula of the binary compound formed from the following sets of atoms.
   a. Ca and O
   b. K and Cl
   c. Rb and S
   d. Ba and P

18. Predict formulas for the following binary ionic compounds.
   a. Mg and N
   b. Na and F
   c. Ca and S
   d. Sr and Te
19. Using shorthand notation, list the core electron configurations for the ions in the compounds in problem 18.

20. Place the following in an order of increasing ionic size. Use the shorthand notation to list the core electron configuration for each of the ions.
   a. Ba\(^{2-}\), Te\(^{2-}\), Cs\(^{+}\), I\(^{-}\)
   b. Cs\(^{+}\), S\(^{2-}\), O\(^{2-}\), K\(^{+}\)

21. Place the following in an order of increasing ionic size. Use shorthand notation to list the core electron configuration for each of the ions.
   a. Cl\(^{-}\), F\(^{-}\), Sr\(^{2+}\), Ca\(^{2+}\)
   b. Na\(^{+}\), Mg\(^{2+}\), Li\(^{+}\), Be\(^{2+}\)

Section 8.8

22. Using the bond energy values listed in Table 8.4 of your text, calculate the \(\Delta H\) for the following reactions:
   a. \(2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g)\)
   b. \(2\text{C}_2\text{H}_4(g) + 7\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 6\text{H}_2\text{O}(g)\)
   c. \(\text{HCN}(g) + 2\text{H}_2(g) \rightarrow \text{CH}_3\text{NH}_2(g)\)

23. Use bond energy values from Table 8.4 in your textbook to calculate \(\Delta H\) for the following reactions:
   a. \(\text{H}-\text{C}=\text{C}-\text{H}(g) + \text{H}_2(g) \rightarrow \text{CH}_3=\text{CH}_2(g)\)
   b. \(2\text{CH}_4(g) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 4\text{H}_2\text{O}(g)\)
   c. \(\text{N}_2(g) + 2\text{H}_2(g) \rightarrow \text{N}_2\text{H}_4(g) (\text{NH}_2-\text{NH}_2)\)

24. Compare the values obtained in parts "a" and "b" of Problem 23 to \(\Delta H\) values calculated from \(\Delta H°\) data in Appendix 4 in your textbook.

25. Calculate the enthalpy of reaction \(\Delta H° \) for the following reaction. Use the enthalpies of formation found in Appendix 4 of your textbook.
   \(\text{H}_2(g) + \text{C}_2\text{H}_4(g) \rightarrow \text{C}_2\text{H}_6(g)\)

Section 8.10

26. Draw Lewis dot structures for the following atoms, ions, or molecules:
   a. \(\text{Sr}\)
   b. \(\text{Br}^{-}\)
   c. \(\text{ICN}\)
   d. \(\text{Ga}\)
   e. \(\text{GaCl}_4^{-}\)
   f. \(\text{P}^{3-}\)
   g. \(\text{NH}_2^{-}\)
   h. \(\text{CSe}_2\)

27. Draw Lewis structures for the following:
   a. \(\text{H}^{+}\)
   b. \(\text{C}\)
   c. \(\text{P}\)
   d. \(\text{P}^{5+}\)
   e. \(\text{Cl}^{-}\)

28. Draw Lewis structures for the following:
   a. \(\text{AsF}_3\)
   b. \(\text{O}_3\)
   c. \(\text{H}_2\text{O}^{+}\)
   d. \(\text{BH}_4^{-}\)
   e. \(\text{NH}_4^{+}\)
   f. \(\text{O}_2\)
Section 8.11

29. Draw Lewis dot structures for the following:
   a. $\text{BCl}_3$  
   b. $\text{AsF}_3$  
   c. $\text{BrO}_3^-$  
   d. $\text{S}_2\text{F}_{10}$ (contains a $\text{S-S}$ bond)

30. Draw Lewis dot structures for the following:
   a. $\text{SbCl}_3$  
   b. $\text{AlF}_6^{3-}$  
   c. $\text{PCl}_3$

Section 8.12

31. Assign formal charges to each of the labeled atoms, a-e.

32. Draw the remaining resonance forms for $\text{N}_2\text{O}_4$.

33. How many reasonable resonance structures are there for carbon monoxide, CO. What are the formal charges?

34. Hydrazine, $\text{N}_2\text{H}_4$, is used as a propellant on the Space Shuttle. Draw all reasonable structures for $\text{N}_2\text{H}_4$, and assign formal charges.

35. Draw a Lewis structure and any resonance forms of benzene, $\text{C}_6\text{H}_6$. (Benzene consists of a ring of six carbon atoms with one hydrogen bonded to each carbon.)

Section 8.13

36. Predict the structure of each of the following molecules or ions:
   a. $\text{SeF}_6$  
   b. $\text{N}_2\text{O}$  
   c. $\text{ClF}_4^+$  
   d. $\text{ClO}^-$  
   e. $\text{CF}_3\text{Cl}$ (carbon is central atom)

37. Predict the structure of HCN.

38. Using the VSEPR model, determine the molecular geometry for each of the following molecules:
   a. $\text{SCl}_4$  
   b. $\text{H}_2\text{Se}$  
   c. $\text{IF}_4^-$  
   d. $\text{SnCl}_5^-$  
   e. $\text{TiCl}_5^+$

39. Which of the molecules or ions in Problem 36 contain polar covalent bonds? Are polar?
Section 9.1

1. What geometry do the following hybrid bonds possess?
   a. $sp$
   b. $sp^2$
   c. $sp^3$
   d. $sp^3d$
   e. $sp^3d^2$

2. Predict the type of hybrid orbital that the central atoms of each of the following compounds display:
   a. SiH$_4$
   b. H$_3$O$^+$
   c. PCl$_5$
   d. NCl$_3$
   e. AsH$_3$
   f. SiF$_6$
   g. CH$_3^+$

3. Predict the geometries of the following compounds:
   a. SF$_2$
   b. SF$_4$
   c. XeF$_2$
   d. XeF$_4$
   e. IF$_5$
   f. ClF$_3$

4. Predict the geometry about the indicated atom, and identify the hybridization of each atom.
   a. the two carbon atoms and the nitrogen atom of glycine
   
   \[
   \text{H} - \text{C} - \text{C} \quad \text{O} \\
   \text{H} \quad \text{a} \quad \text{b} \\
   \text{NH}_2 \\
   \text{O}
   \]

   b. the carbon atom in CF$_2$Cl$_2$
   c. the phosphorous atom in PCl$_5$
   d. the nitrogen atom in NH$_2^-$

Section 9.2

5. Determine the number of sigma and pi bonds in each of the following:
   a. \[
   \text{H} - \text{C} = \text{C} - \text{OH} \\
   \text{H}
   \]
   b. \[
   \text{H} - \text{C} = \text{C} \equiv \text{N} \\
   \text{H}
   \]
   c. \[
   \text{H} - \text{C} = \text{C} = \text{C} = \text{H} \\
   \text{H}
   \]

6. The structure of urea is
   \[
   \text{H} - \text{N} - \text{C} - \text{N} - \text{H} \\
   \text{H} \quad \text{O} \\
   \text{H}
   \]

   a. How many $\sigma$ bonds are there?
   b. How many $\pi$ bonds?
   c. What is the hybridization at the carbon?
   d. How are the nitrogen atoms hybridized?
   e. What is the N-C-N bond angle expected to be?
   f. How many lone pairs of electrons are there?
### Answers to Exercises

**Ch. 8**

1. a. polar covalent  
   b. ionic  
   c. polar covalent  
   d. covalent  

2. \(-7.36 \times 10^{-19}\) J

3. \(E = -1160\) kJ/mole

4. a. Fr, Rb, Mg  
   b. Al, B, C, N  
   c. Ga, As, P, O  
   d. P, S, Cl

5. \(\text{Rb} < \text{Ca} < \text{Nb} < \text{Si} < \text{Pt} < \text{N} < \text{F}\)

6. a. 0  
   b. 2.0  
   c. 1.7  
   d. 1.4  
   e. 0.4

7. \(\text{I}_2 < \text{CN}^- < \text{H}_2\text{O} < \text{NaBr} < \text{MnO}_2\)

8. \(\text{O}_3, \text{P}_8 < \text{H}_2\text{S}, \text{CH}_4 < \text{NO} < \text{CO}_2\)

9. HF; NO; CO; HCl

10. The difference in the electronegativity between Be and F is higher than the one between Be and Cl.
    - \(3.0 - 1.5 = 1.5\) (covalent bond)
    - \(4.0 - 1.5 = 2.5\) (ionic)

11. a. no dipole  
    b. negative toward F  
    c. negative toward O  
    d. no dipole

12. all

13. b. The opposing bond polarities in a tetrahedral structure cancel out. Thus \(\text{CF}_4\) has no dipole moment.

14. a. \(\text{O}^{2-}\), isoelectronic with neon  
    b. \(\text{Be}^{2+}\), isoelectronic with helium  
    c. \(\text{I}^{-}\), isoelectronic with xenon  
    d. \(\text{Te}^{2-}\), isoelectronic with xenon  
    e. \(\text{Na}^{+}\), isoelectronic with neon

15. \(\text{Ca}^{2+}, \text{K}^+, \text{Cl}^-, \text{S}^{2-}\)

16. \(\text{Sr}^{2+} < \text{Rb}^+ < \text{Br}^- < \text{Se}^{2-}\)

17. a. CaO  
    b. KCl  
    c. Rb, S  
    d. Ba, P, \(\text{P}_2\)
18. a. Mg₃N₂  
    b. NaF  
    c. CaS  
    d. SrTe

19. a. [Ne], [Ne]  
    b. [Ne], [Ne]  
    c. [Ar], [Ar]  
    d. [Kr], [Xe]

20. a. Ba²⁺ < Cs⁺ < I⁻ < Te²⁻; all can be written as [Xe]  
    b. K⁺ < Cs⁺ < O²⁻ < S²⁻; K⁺ = [Ar], Cs⁺ = [Xe], O²⁻ = [Ne], S²⁻ = [Ar]

21. a. Ca²⁺ < Sr²⁺ < F⁻ < Cl⁻; Ca²⁺ = [Ar], Sr²⁺ = [Kr], F⁻ = [Ne], Cl⁻ = [Ar]  
    b. Be³⁺ < Li⁺ < Mg²⁺ < Na⁺; Be³⁺ = [He], Li⁺ = [He], Mg²⁺ = [Ne], Na⁺ = [Ne]

22. a. -509 kJ  
    b. -2881 kJ  
    c. -158 kJ

23. a. -169 kJ  
    b. -1091 kJ  
    c. +81 kJ

24. a. 6 kJ difference  
    b. 7 kJ difference

25. ΔH° = -136.7 kJ/mole

26. a. • Sr •  
    b. [: Br :]⁻  
    c. [: I — C≡N :]  
    d. • Ga •  
    e. [: Cl :]⁻  
    f. [: P :]⁻  
    g. [H — N — H]⁻  
    h. Se≡C≡Se

27. a. [H]⁺  
    b. • C •  
    c. • P •  
    d. [P]⁵⁺

28. a. [: F — As — F :]⁻  
    b. [: O—O—O :]⁻  
    c. [H—O—H]⁺  
    d. [H—B—H]⁻  
    e. [H—N—H]⁺  
    f. O≡O
29. a. \[
\begin{array}{c}
\text{Cl} \\
\text{B} \\
\text{Cl}
\end{array}
\]

b. \[
\begin{array}{c}
\text{F} \\
\text{As} \\
\text{F}
\end{array}
\]

c. \[
\begin{array}{c}
\text{O} \\
\text{Br} \\
\text{O}
\end{array}
\]

30. a. \[
\begin{array}{c}
\text{Cl} \\
\text{Sb} \\
\text{Cl}
\end{array}
\]

b. \[
\begin{array}{c}
\text{F} \\
\text{Al} \\
\text{F}
\end{array}
\]

c. \[
\begin{array}{c}
\text{Cl} \\
\text{P} \\
\text{Cl}
\end{array}
\]

31. a. 0
b. 0
c. -1
d. +1
e. +1
f. 0

32. \[
\begin{array}{c}
\text{O} \\
\text{N} \\
\text{O}
\end{array}
\]

33. \[
\text{O} = \text{C}
\]
is the only reasonable structure. The carbon has a formal charge of -1, and oxygen has a +1 formal charge. It is possible to draw a structure such as \[
\text{O} = \text{C}
\] in which both C and O have a zero formal charge, but the carbon would be electron deficient (lacks an octet).

34. The only reasonable structure has an N–N single bond. All formal charges = 0.

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

Each nitrogen has a formal charge of +1 and the two hydrogens that are single have formal charges of -1.

35. \[
\begin{array}{c}
\text{H} \\
\text{C} \\
\text{H}
\end{array}
\]

36. a. octahedral
b. linear
c. see-saw
d. linear
e. tetrahedral

37. \[
\text{H} = \text{N}
\] is linear (H–C≡N)
38. a. see-saw  
    b. bent
    c. square planar  
    d. trigonal bipyramidal  
    e. linear

39. All contain polar covalent bonds; b, c, d, and e are polar. Negative is toward: O in b; equatorial fluorine in c; O in d; fluorine in e.

---

Ch. 9.

1. a. linear  
    b. trigonal planar
    c. tetrahedral  
    d. trigonal bipyramidal  
    e. octahedral

2. a. $sp^3$
    b. $sp^3$
    c. $dsp^3$
    d. $sp^3$
    e. $sp^3$
    f. $d^2sp^3$
    g. $sp^2$

3. a. angular (like water)  
    b. see-saw
    c. linear  
    d. square planar  
    e. square pyramidal  
    f. T-shaped

4. a. carbon a-$sp^3$, tetrahedral; carbon b-$sp^3$, trigonal planar; nitrogen-$sp^3$, based on tetrahedron-trigonal pyramid  
    b. $sp^3$-tetrahedral  
    c. $dsp^3$-trigonal bipyramid  
    d. $sp^3$-based on tetrahedron-bent

5. a. 8 sigma, 2 pi
    b. 5 sigma, 2 pi
    c. 7 sigma, 3 pi
    d. 1 sigma, 1 pi
    e. 120°
    f. 4