### Chapter 9 (3rd edition)

#### 9.8 Lewis Structure

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| (a) | Cl
Cl—Al—Cl | 3 | Trigonal planar | 0 | Trigonal planar, AB$_3$ |
| (b) | Cl—Zn—Cl | 2 | Linear | 0 | Linear, AB$_2$ |
| (c) | Cl
Cl—Zn—Cl
Cl | 4 | Tetrahedral | 0 | Tetrahedral, AB$_4$ |
9.17 a. BrF$_5$ has the Lewis structure (The lone pairs on the F atoms are not shown.):

```
    F
   / \  \\
  F--Br--F
   \   / \\
    F
```

With six electron domains around the central atom, the electron-domain geometry is octahedral. Because one of the electron domains is a lone pair, we expect the molecular geometry to be a square pyramid (See Figure 9.2). The bond dipoles in this molecule do not sum to zero. Therefore, the molecule is **polar**.

b. BCl$_3$ has the Lewis structure (The lone pairs on the Cl atoms are not shown.):

```
    Cl
   /   \\
  Cl--B--Cl
```

Recall that boron is one of the elements that does not necessarily obey the octet rule (See Section 8.8 in the text). With three electron domains around the central atom, the electron-domain geometry is trigonal planar. Because none of the three electron domains on the central atom is a lone pair, the molecular geometry is also trigonal planar. The bonds are symmetrically distributed and the individual bond dipoles sum to zero. Therefore, the molecule is **nonpolar**.

9.39 Draw the Lewis structure of the molecule. Several resonance forms with formal charges
are shown.

\[
\begin{align*}
\left[ -\hat{\text{N}}\equiv\hat{\text{N}}\equiv-\hat{\text{N}} \right]^- & \quad \leftrightarrow \quad \left[ :\text{N}\equiv\text{N}$$^+$$\equiv-\hat{\text{N}}: \right]^- & \quad \leftrightarrow \quad \left[ 2^- :\hat{\text{N}}\equiv\text{N}$$^+$$\equiv\text{N} : \right]^- 
\end{align*}
\]

Count the number of electron domains around the central atom. Since there are two electron domains around N, the electron-domain geometry is linear (AB$_2$ w/no lone pairs) and we conclude that N is \textit{sp} hybridized.

Remember, a multiple bond is just \textit{one} electron domain.

\textbf{9.51} The electron configurations are listed. Refer to Figure 9.18 of the text for the molecular orbital diagram.

\begin{align*}
\text{Li}_2 &: \quad (\sigma_{1s})^2(\sigma_{1s}^*)^2(\sigma_{2s})^2 \quad \text{bond order} = 1 \\
\text{Li}_2^+ &: \quad (\sigma_{1s})^2(\sigma_{1s}^*)^2(\sigma_{2s})^1 \quad \text{bond order} = \frac{1}{2} \\
\text{Li}_2^- &: \quad (\sigma_{1s})^2(\sigma_{1s}^*)^2(\sigma_{2s})^2(\sigma_{2s}^*)^1 \quad \text{bond order} = \frac{1}{2}
\end{align*}

Order of increasing stability: \textit{Li}_2^- < \textit{Li}_2^+ < \textit{Li}_2

In reality, \textit{Li}_2^- is more stable than \textit{Li}_2 because there is less electrostatic repulsion in \textit{Li}_2^-.

\textbf{9.66} The ion contains 24 valence electrons. Of these, six are involved in three sigma bonds between the nitrogen and oxygen atoms. The hybridization of the nitrogen atom is \textit{sp}$_2$. There are 16 non-bonding electrons on the oxygen atoms. The remaining two electrons are in a delocalized pi molecular orbital which results from the overlap of the $p_z$ orbital of nitrogen and the $p_z$ orbitals of the three oxygen atoms. The molecular orbitals are similar to those of the carbonate ion (See Section 9.7 of the text).