CHAPTER 8 Solutions to suggested end-of-chapter problems

27. Using the periodic table, the general trend for electronegativity is:
   
   (1) Increase as we go from left to right across a period
   (2) Decrease as we go down a group

   Using these trends, the expected orders are:
   
   a. C < N < O   b. Se < S < Cl   c. Sn < Ge < Si   d. Tl < Ge < S


29. The most polar bond will have the greatest difference in electronegativity between the two atoms. From positions in the periodic table, we would predict:
   
   a. Ge–F   b. P–Cl   c. S–F   d. Ti–Cl


31. The general trends in electronegativity used in Exercises 27 and 29 are only rules of thumb. In this exercise, we use experimental values of electronegativities and can begin to see several exceptions. The order of EN from Figure 8.3 is:
   
   a. C (2.5) < N (3.0) < O (3.5) same as predicted
   b. Se (2.4) < S (2.5) < Cl (3.0) same
   c. Si = Ge = Sn (1.8) different
   d. Tl (1.8) = Ge (1.8) < S (2.5) different

   Most polar bonds using actual EN values:
   
   a. Si–F and Ge–F have equal polarity (Ge–F predicted).
   b. P–Cl (same as predicted)
   c. S–F (same as predicted)  
   d. Ti–Cl (same as predicted)

32. The order of EN from Figure 8.3 is:
   
   a. Rb (0.8) = K (0.8) < Na (0.9), different  b. Ga (1.6) < B (2.0) < O (3.5), same
   c. Br (2.8) < Cl (3.0) < F (4.0), same  d. S (2.5) < O (3.5) < F (4.0), same

   Most polar bonds using actual EN values:
   
   a. C–H most polar (Sn–H predicted)
b. Al–Br most polar (Tl–Br predicted).

c. Si–O (same as predicted).

d. Each bond has the same polarity, but the bond dipoles point in opposite directions. Oxygen is the positive end in the O–F bond dipole, and oxygen is the negative end in the O–Cl bond dipole (O–F predicted).

33. Use the electronegativity trend to predict the partial negative end and the partial positive end of the bond dipole (if there is one). To do this, you need to remember that H has electronegativity between B and C and identical to P. Answers b, d, and e are incorrect. For d (Br₂), the bond between two Br atoms will be a pure covalent bond, where there is equal sharing of the bonding electrons, and no dipole moment exists. For b and e, the bond polarities are reversed. In Cl–I, the more electronegative Cl atom will be the partial negative end of the bond dipole, with I having the partial positive end. In O–P, the more electronegative oxygen will be the partial negative end of the bond dipole, with P having the partial positive end. In the following, we used arrows to indicate the bond dipole. The arrow always points to the partial negative end of a bond dipole (which always is the most electronegative atom in the bond).

34. See Exercise 33 for a discussion on bond dipoles. We will use arrows to indicate the bond dipoles. The arrow always points to the partial negative end of the bond dipole, which will always be to the more electronegative atom. The tail of the arrow indicates the partial positive end of the bond dipole.

a. C–O

b. P–H is a pure covalent (nonpolar) bond because P and H have identical electronegativities.

c. H–Cl

d. Br–Te

e. Se–S The actual electronegativity difference between Se and S is so small that this bond is probably best characterized as a pure covalent bond having no bond dipole.

35. Bonding between a metal and a nonmetal is generally ionic. Bonding between two nonmetals is covalent, and in general, the bonding between two different nonmetals is usually polar covalent. When two different nonmetals have very similar electronegativities, the bonding is pure covalent or just covalent.

a. ionic b. covalent c. polar covalent

d. ionic e. polar covalent f. covalent

42. Te²⁻: [Kr]5s²⁴d¹⁰⁵p⁶; Cl⁻: [Ne]3s²³p⁶; Sr²⁺: [Ar]4s²³d¹⁰⁴p⁶; Li⁺: 1s²
44. a. \( \text{Al}^{3+} \) and \( \text{Cl}^- \); AlCl\(_3\), aluminum chloride  
b. \( \text{Na}^+ \) and \( \text{O}^{2-} \); Na\(_2\)O, sodium oxide  
c. \( \text{Sr}^{2+} \) and \( \text{F}^- \); SrF\(_2\), strontium fluoride  
d. \( \text{Ca}^{2+} \) and \( \text{S}^{2-} \); CaS, calcium sulfide

46. a. \( \text{Sr}^{2+} \): \([\text{Ar}]4s^23d^{10}4p^6\)  
b. \( \text{P}^{3-} \) and \( \text{S}^{2-} \): \([\text{Ne}]3s^23p^6\)  
c. \( \text{In}^+ \): \([\text{Kr}]5s^24d^{10}\)  
d. \( \text{Pb}^{2+} \): \([\text{Xe}]6s^24f^{14}5d^{10}\)

48. a. Cs\(_2\)S is composed of Cs\(^+\) and S\(^2-\). Cs\(^+\) has the same electron configuration as Xe, and S\(^2-\) has the same configuration as Ar.  
b. SrF\(_2\); Sr\(^{2+}\) has the Kr electron configuration, and F\(^-\) has the Ne configuration.  
c. Ca\(_3\)N\(_2\); Ca\(^{2+}\) has the Ar electron configuration, and N\(^3-\) has the Ne configuration.  
d. AlBr\(_3\); Al\(^3+\) has the Ne electron configuration, and Br\(^-\) has the Kr configuration.

54. a. \( \text{V} > \text{V}^{2+} > \text{V}^{3+} \)  
b. \( 
\text{Cs}^+ > \text{Rb}^+ > \text{K}^+ > \text{Na}^+ \)  
c. \( \text{Te}^{2-} > \Gamma > \text{Cs}^+ > \text{Ba}^{2+} \)

56. a. LiF; Li\(^+\) is smaller than Cs\(^+\).  
b. NaBr; Br\(^-\) is smaller than I\(^-\).  
c. BaO; O\(^2-\) has a greater charge than Cl\(^-\).  
d. CaSO\(_4\); Ca\(^{2+}\) has a greater charge than Na\(^+\).  
e. K\(_2\)O; O\(^2-\) has a greater charge than F\(^-\).  
f. Li\(_2\)O; both ions are small in Li\(_2\)O.

58. 
\[
\text{Mg}(s) \rightarrow \text{Mg}(g) \quad \Delta H = 150. \text{kJ} \quad \text{(sublimation)}
\]
\[
\text{Mg}(g) \rightarrow \text{Mg}^+(g) + e^- \quad \Delta H = 735 \text{kJ} \quad \text{(IE\(_1\))}
\]
\[
\text{Mg}^+(g) \rightarrow \text{Mg}^{2+}(g) + e^- \quad \Delta H = 1445 \text{kJ} \quad \text{(IE\(_2\))}
\]
\[
\text{F}_2(g) \rightarrow 2 \text{F}(g) \quad \Delta H = 154 \text{kJ} \quad \text{(BE)}
\]
\[
2 \text{F}(g) + 2 e^- \rightarrow 2 \text{F}^-(g) \quad \Delta H = 2(-328) \text{kJ} \quad \text{(EA)}
\]
\[
\text{Mg}^{2+}(g) + 2 \text{F}^-(g) \rightarrow \text{MgF}_2(s) \quad \Delta H = -2913 \text{kJ} \quad \text{(LE)}
\]
\[
\text{Mg}(s) + \text{F}_2(g) \rightarrow \text{MgF}_2(s) \quad \Delta H^\circ = -1085 \text{kJ/mol}
\]

60. 
\[
\text{Na}(g) \rightarrow \text{Na}^+(g) + e^- \quad \Delta H = \text{IE} = 495 \text{kJ} \text{ (Table 7.5)}
\]
\[
\text{F}(g) + e^- \rightarrow \text{F}^-(g) \quad \Delta H = \text{EA} = -327.8 \text{kJ} \text{ (Table 7.7)}
\]
\[
\text{Na}(g) + \text{F}(g) \rightarrow \text{Na}^+(g) + \text{F}^-(g) \quad \Delta H = 167 \text{kJ}
\]

The described process is endothermic. What we haven’t accounted for is the extremely favorable lattice energy. Here, the lattice energy is a large negative (exothermic) value, making the overall formation of NaF a favorable exothermic process.

61. Use Figure 8.11 as a template for this problem.

\[
\text{Li}(s) \rightarrow \text{Li}(g) \quad \Delta H_{\text{sub}} = ?
\]
\[
\text{Li}(g) \rightarrow \text{Li}^+(g) + e^- \quad \Delta H = 520 \text{ kJ}
\]
\[
\frac{1}{2} \text{I}_2(g) \rightarrow \text{I}(g) \quad \Delta H = 151/2 \text{ kJ}
\]
\[
\text{I}(g) + e^- \rightarrow \text{I}^-(g) \quad \Delta H = -295 \text{ kJ}
\]
\[
\text{Li}^+(g) + \Gamma(g) \rightarrow \text{LiI}(s) \quad \Delta H = -753 \text{ kJ}
\]
\[
\text{Li}(s) + \frac{1}{2} \text{I}_2(g) \rightarrow \text{LiI}(s) \quad \Delta H = -292 \text{ kJ}
\]
\[
\Delta H_{\text{sub}} + 520 + 151/2 - 295 - 753 = -292, \quad \Delta H_{\text{sub}} = 161 \text{ kJ}
\]

64. Lattice energy is proportional to the charge of the cation times the charge of the anion $Q_1Q_2$.

<table>
<thead>
<tr>
<th>Compound</th>
<th>$Q_1Q_2$</th>
<th>Lattice Energy</th>
</tr>
</thead>
<tbody>
<tr>
<td>FeCl$_2$</td>
<td>(+2)(-1) = -2</td>
<td>-2631 kJ/mol</td>
</tr>
<tr>
<td>FeCl$_3$</td>
<td>(+3)(-1) = -3</td>
<td>-5359 kJ/mol</td>
</tr>
<tr>
<td>Fe$_2$O$_3$</td>
<td>(+3)(-2) = -6</td>
<td>-14,744 kJ/mol</td>
</tr>
</tbody>
</table>

65. a. $\text{H} = \text{H} + \text{Cl} = \text{Cl} \quad \rightarrow \quad 2 \text{H} = \text{Cl}$

Bonds broken: \begin{align*}
1 \text{H} = \text{H} (432 \text{ kJ/mol}) \\
1 \text{Cl} = \text{Cl} (239 \text{ kJ/mol})
\end{align*}

Bonds formed: \begin{align*}
2 \text{H} = \text{Cl} (427 \text{ kJ/mol})
\end{align*}

\[\Delta H = \Sigma D_{\text{broken}} - \Sigma D_{\text{formed}}, \quad \Delta H = 432 \text{ kJ} + 239 \text{ kJ} - 2(427) \text{ kJ} = -183 \text{ kJ}\]

b. $\equiv\equiv\equiv\equiv + 3 \text{H} = \text{H} \quad \rightarrow \quad 2 \text{H} = \equiv\equiv\equiv\equiv$

Bonds broken: \begin{align*}
1 \equiv\equiv\equiv\equiv (941 \text{ kJ/mol}) \\
3 \text{H} = \text{H} (432 \text{ kJ/mol})
\end{align*}

Bonds formed: \begin{align*}
6 \equiv\equiv\equiv\equiv (391 \text{ kJ/mol})
\end{align*}

\[\Delta H = 941 \text{ kJ} + 3(432) \text{ kJ} - 6(391) \text{ kJ} = -109 \text{ kJ}\]

66. Sometimes some of the bonds remain the same between reactants and products. To save time, only break and form bonds that are involved in the reaction.

a. $\equiv\equiv\equiv\equiv + 2 \text{H} = \text{H} \quad \rightarrow \quad \equiv\equiv\equiv\equiv + \equiv\equiv\equiv\equiv$

Bonds broken: \begin{align*}
1 \equiv\equiv\equiv\equiv (891 \text{ kJ/mol}) \\
2 \text{H} = \text{H} (432 \text{ kJ/mol})
\end{align*}

Bonds formed: \begin{align*}
1 \equiv\equiv\equiv\equiv (305 \text{ kJ/mol}) \\
2 \equiv\equiv\equiv\equiv (413 \text{ kJ/mol}) \\
2 \equiv\equiv\equiv\equiv (391 \text{ kJ/mol})
\end{align*}
\[ \Delta H = 891 \text{ kJ} + 2(432 \text{ kJ}) - [305 \text{ kJ} + 2(413 \text{ kJ}) + 2(391 \text{ kJ})] = -158 \text{ kJ} \]

b. \[
\begin{array}{c}
\text{H}_2\text{N} - \text{N} - \text{H} + 2 \text{F} = \text{4 H} - \text{F} + \text{N} = \text{N}
\end{array}
\]

Bonds broken:
- 1 N–N (160. kJ/mol)
- 4 N–H (391 kJ/mol)
- 2 F–F (154 kJ/mol)

Bonds formed:
- 4 H–F (565 kJ/mol)
- 1 N≡N (941 kJ/mol)

\[ \Delta H = 160. \text{ kJ} + 4(391 \text{ kJ}) + 2(154 \text{ kJ}) - [4(565 \text{ kJ}) + 941 \text{ kJ}] = -1169 \text{ kJ} \]

67.

\[
\begin{array}{c}
\text{H}_2\text{C} - \text{N} = \text{C} + \text{H}_2\text{C} = \text{N}\text{H}
\end{array}
\]

Bonds broken: 1 C–N (305 kJ/mol)  Bonds formed: 1 C–C (347 kJ/mol)

\[ \Delta H = \Sigma D_{\text{broken}} - \Sigma D_{\text{formed}}, \quad \Delta H = 305 - 347 = -42 \text{ kJ} \]

*Note:* Sometimes some of the bonds remain the same between reactants and products. To save time, only break and form bonds that are involved in the reaction.

68.

\[
\begin{array}{c}
\text{H}_2\text{C} - \text{O} - \text{H} + \text{C} = \text{O} \rightarrow \text{H}_2\text{C} - \text{C} = \text{O} - \text{H}
\end{array}
\]

Bonds broken:
- 1 C≡O (1072 kJ/mol)
- 1 C–O (358 kJ/mol)

Bonds formed:
- 1 C–C (347 kJ/mol)
- 1 C≡O (745 kJ/mol)
- 1 C–O (358 kJ/mol)

\[ \Delta H = 1072 + 358 - [347 + 745 + 358] = -20. \text{ kJ} \]

69.

\[
\begin{array}{c}
\text{H}_2\text{S} - \text{H} + 3 \text{F} = \text{F} - \text{S} - \text{F} + 2 \text{H} - \text{F}
\end{array}
\]
Bonds broken:  Bonds formed:

2 S–H (347 kJ/mol)  4 S–F (327 kJ/mol)
3 F–F (154 kJ/mol)  2 H–F (565 kJ/mol)

\[ \Delta H = 2(347) + 3(154) - [4(327) + 2(565)] = -1282 \text{ kJ} \]

73. \[
\begin{array}{c}
\text{H} \quad \text{H} \\
\text{C} \quad \text{C} \\
\text{H} \quad \text{H}
\end{array}
+ \quad \begin{array}{c}
\text{F} \quad \text{F}
\end{array}
\rightarrow \begin{array}{c}
\text{H} \quad \text{C} \quad \text{C} \\
\text{H} \quad \text{H} \\
\text{F} \quad \text{F}
\end{array} \quad \Delta H = -549 \text{ kJ}
\]

Bonds broken:  Bonds formed:

1 C=C (614 kJ/mol)  1 C–C (347 kJ/mol)
1 F–F (154 kJ/mol)  2 C–F (D_{CF} = C–F bond energy)

\[ \Delta H = -549 \text{ kJ} = 614 \text{ kJ} + 154 \text{ kJ} - [347 \text{ kJ} + 2D_{CF}], \ 2D_{CF} = 970, \ D_{CF} = 485 \text{ kJ/mol} \]

74. Let \( x \) = bond energy for \( A_2 \), so \( 2x \) = bond energy for \( AB \).

\[ \Delta H = -285 \text{ kJ} = x + 432 \text{ kJ} - [2(2x)], \ 3x = 717, \ x = 239 \text{ kJ/mol} \]

The bond energy for \( A_2 \) is 239 kJ/mol.

78. \( \text{NH}_3(\text{g}) \rightarrow \text{N}(\text{g}) + 3 \text{ H(\text{g})} \) \[ \Delta H^\circ = 3D_{NH} = 472.7 + 3(216.0) - (-46.1) = 1166.8 \text{ kJ} \]

\[ D_{NH} = \frac{1166.8 \text{ kJ}}{3 \text{ mol NH bonds}} = 388.93 \text{ kJ/mol} = 389 \text{ kJ/mol} \]

\( D_{\text{calc}} = 389 \text{ kJ/mol} \) as compared with 391 kJ/mol in Table 8.4. There is good agreement.

80. \( 1/2 \text{N}_2(\text{g}) + 1/2 \text{ O}_2(\text{g}) \rightarrow \text{NO}(\text{g}) \) \[ \Delta H = 90. \text{ kJ} \]

Bonds broken:  Bonds formed:

1/2 N≡N (941 kJ/mol)  1 NO (\( D_{NO} = \text{NO bond energy} \))
1/2 O=O (495 kJ/mol)

\[ \Delta H = 90. \text{ kJ} = 1/2(941) + 1/2(495) - (D_{NO}), \ D_{NO} = 628 \text{ kJ/mol} \]

From this data, the calculated NO bond energy is 628 kJ/mol.

83. Drawing Lewis structures is mostly trial and error. However, the first two steps are always the same. These steps are (1) count the valence electrons available in the molecule/ion, and (2) attach all atoms to each other with single bonds (called the skeletal structure). Unless noted otherwise, the atom listed first is assumed to be the atom in the middle, called the
central atom, and all other atoms in the formula are attached to this atom. The most notable exceptions to the rule are formulas that begin with H, e.g., H₂O, H₂CO, etc. Hydrogen can never be a central atom since this would require H to have more than two electrons. In these compounds, the atom listed second is assumed to be the central atom.

After counting valence electrons and drawing the skeletal structure, the rest is trial and error. We place the remaining electrons around the various atoms in an attempt to satisfy the octet rule (or duet rule for H).

a. CCl₄ has 4 + 4(7) = 32 valence electrons.

Cl—C—Cl

Skeletal structure

b. NC₃ has 5 + 3(7) = 26 valence electrons.

Cl—N—Cl

Skeletal structure

Lewis structure

SeCl₂ has 6 + 2(7) = 20 valence electrons.

Se—Cl

Skeletal structure

Lewis structure
d. ICl has 7 + 7 = 14 valence electrons.

Cl—I

Skeletal structure

Lewis structure

a. POCl₃ has 5 + 6 + 3(7) = 32 valence electrons.

POCl₃

Skeletal structure

Lewis structure

Note: This structure uses all 32 e⁻ while satisfying the octet rule for all atoms. This is a valid Lewis structure.

SO₄²⁻ has 6 + 4(6) + 2 = 32 valence electrons.

Note: A negatively charged ion will have additional electrons to those that come from the valence shell of the atoms. The magnitude of the negative charge indicates the number of extra electrons to add in.

XeO₄, 8 + 4(6) = 32 e⁻

PO₄³⁻, 5 + 4(6) + 3 = 32 e⁻
ClO$_4^-$ has $7 + 4(6) + 1 = 32$ valence electrons

\[
\begin{align*}
\text{Skeletal structure} & \quad \text{Lewis structure} \\
\begin{array}{c}
\dfrac{\text{O}}{\text{Cl}} \quad \dfrac{\text{O}}{\dfrac{\text{O}}{\dfrac{\text{O}}{\text{Xe}}}}
\end{array}
& \quad \begin{array}{c}
\dfrac{\text{O}}{\text{O}} \quad \dfrac{\text{O}}{\text{O}} \quad \dfrac{\text{O}}{\text{O}}
\end{array}
\end{align*}
\]

Note: All of these species have the same number of atoms and the same number of valence electrons. They also have the same Lewis structure.

b. NF$_3$ has $5 + 3(7) = 26$ valence electrons. \( \text{SO}_3^{2-}, 6 + 3(6) + 2 = 26 \text{ e}^- \)

\[
\begin{align*}
\text{Skeletal structure} & \quad \text{Lewis structure} \\
\begin{array}{c}
\dfrac{\text{F}}{\text{N}} \quad \dfrac{\text{F}}{\text{F}} \\
\dfrac{\text{F}}{\text{F}}
\end{array}
& \quad \begin{array}{c}
\dfrac{\text{F}}{\text{N}} \quad \dfrac{\text{F}}{\text{F}} \\
\dfrac{\text{F}}{\text{F}}
\end{array}
\end{align*}
\]

PO$_3^{3-}$, $5 + 3(6) + 3 = 26 \text{ e}^-$ \quad ClO$_3^-$, $7 + 3(6) + 1 = 26 \text{ e}^-$

\[
\begin{align*}
\text{Skeletal structure} & \quad \text{Lewis structure} \\
\begin{array}{c}
\dfrac{\text{O}}{\text{P}} \quad \dfrac{\text{O}}{\text{O}}
\end{array}
& \quad \begin{array}{c}
\dfrac{\text{O}}{\text{O}} \quad \dfrac{\text{O}}{\text{O}}
\end{array}
\end{align*}
\]

Note: Species with the same number of atoms and valence electrons have similar Lewis structures.

c. ClO$_2^-$ has $7 + 2(6) + 1 = 20$ valence

\[
\begin{align*}
\text{Skeletal structure} & \quad \text{Lewis structure} \\
\text{O} & \quad \text{O} \quad \text{Cl}
\end{align*}
\]

\[
\begin{align*}
\text{Skeletal structure} & \quad \text{Lewis structure} \\
\begin{array}{c}
\dfrac{\text{Cl}}{\text{S}} \quad \dfrac{\text{Cl}}{\text{Cl}}
\end{array}
& \quad \begin{array}{c}
\dfrac{\text{Cl}}{\text{P}} \quad \dfrac{\text{Cl}}{\text{Cl}}
\end{array}
\end{align*}
\]

Note: All of these species have the same number of atoms and the same number of valence electrons. They also have the same Lewis structure.
Note: Species with the same number of atoms and valence electrons have similar Lewis structures.

d. Molecules ions that have the same number of valence electrons and the same number of atoms will have similar Lewis structures.

85. BeH₂, \(2 + 2(1) = 4\) valence electrons
   \(\text{BH}_3, 3 + 3(1) = 6\) valence electrons

   ![BeH2 Structure](image)

   ![B(3)H3 Structure](image)

87. PF₅, \(5 + 5(7) = 40\) valence electrons
   SF₄, \(6 + 4(7) = 34\) e⁻

   ![PF5 Structure](image)

   ![SF4 Structure](image)

   ClF₃, \(7 + 3(7) = 28\) e⁻
   Br₃⁻, \(3(7) + 1 = 22\) e⁻

   ![ClF3 Structure](image)

   ![Br3- Structure](image)

Row 3 and heavier nonmetals can have more than 8 electrons around them when they have to. Row 3 and heavier elements have empty d orbitals that are close in energy to valence s and p orbitals. These empty d orbitals can accept extra electrons.

For example, P in PF₅ has its five valence electrons in the 3s and 3p orbitals. These s and p orbitals have room for three more electrons, and if it has to, P can use the empty 3d orbitals for any electrons above 8.

90. Ozone: \(O_3\) has \(3(6) = 18\) valence electrons.

   ![Ozone Structure](image)
Sulfur dioxide: $\text{SO}_2$ has $6 + 2(6) = 18$ valence electrons.

\[ \ddot{\text{O}} \dddot{\text{S}} \dddot{\text{O}} : \quad \rightleftharpoons \ddot{\text{O}} \dddot{\text{S}} = \dddot{\text{O}} \]

Sulfur trioxide: $\text{SO}_3$ has $6 + 3(6) = 24$ valence electrons.

\[ :\dddot{\text{O}} : \quad \rightleftharpoons \dddot{\text{O}} : \dddot{\text{S}} : \quad \rightleftharpoons \dddot{\text{O}} \ \dddot{\text{S}} = \ddot{\text{O}} \]

94. $\text{CO}_3^{2-}$ has $4 + 3(6) + 2 = 24$ valence electrons.

\[ \begin{array}{c}
\text{[ } \dddot{\text{O}} \ddot{\text{C}} \dddot{\text{O}} \text{]}^{2-} \\
\rightleftharpoons \text{[ } \dddot{\text{O}} \ddot{\text{C}} \ddot{\text{O}} : \dddot{\text{O}} \text{]}^{2-} \\
\rightleftharpoons \text{[ } \dddot{\text{O}} \dddot{\text{C}} = \dddot{\text{O}} \text{]}^{2-}
\end{array} \]

Three resonance structures can be drawn for $\text{CO}_3^{2-}$. The actual structure for $\text{CO}_3^{2-}$ is an average of these three resonance structures. That is, the three C–O bond lengths are all equivalent, with a length somewhere between a single and a double bond. The actual bond length of 136 pm is consistent with this resonance view of $\text{CO}_3^{2-}$.

98. $\text{H}_2\text{NOH} \ (14 \text{ e}^-)$

\[ \text{H} \dddot{\text{N}} \dddot{\text{O}} \ddot{\text{H}} \quad \text{Single bond between N and O} \]
From the Lewis structures, the order from shortest $\rightarrow$ longest N–O bond is:

$$\text{NO}^+ < \text{N}_2\text{O} < \text{NO}_2^- < \text{NO}_3^- < \text{H}_2\text{NOH}$$

100. **Carbon:** $\text{FC} = 4 - 2 - 1/2(6) = -1$; **oxygen:** $\text{FC} = 6 - 2 - 1/2(6) = +1$

Electronegativity predicts the opposite polarization. The two opposing effects seem to partially cancel to give a much less polar molecule than expected.

101. See Exercise 84 for the Lewis structures of $\text{POCl}_3$, $\text{SO}_4^{2-}$, $\text{ClO}_4^-$ and $\text{PO}_4^{3-}$. All these compounds/ions have similar Lewis structures to those of $\text{SO}_2\text{Cl}_2$ and $\text{XeO}_4$ shown below. Formal charge $= \text{[number of valence electrons on free atom]} - \text{[number of lone pair electrons on atom]} + 1/2(\text{number of shared electrons of atom})$.

a. $\text{POCl}_3$: P, $\text{FC} = 5 - 1/2(8) = +1$

b. $\text{SO}_4^{2-}$: S, $\text{FC} = 6 - 1/2(8) = +2$

c. $\text{ClO}_4^-$: Cl, $\text{FC} = 7 - 1/2(8) = +3$

d. $\text{PO}_4^{3-}$: P, $\text{FC} = 5 - 1/2(8) = +1$

e. $\text{SO}_2\text{Cl}_2$, $6 + 2(6) + 2(7) = 32\ e^-$

f. $\text{XeO}_4$, $8 + 4(6) = 32\ e^-$

g. $\text{ClO}_3^-$, $7 + 3(6) + 1 = 26\ e^-$

h. $\text{NO}_4^{3-}$, $5 + 4(6) + 3 = 32\ e^-$
103. \( \text{O}_2\text{F}_2 \) has \( 2(6) + 2(7) = 26 \) valence e\(^-\). The formal charge and oxidation number (state) of each atom is below the Lewis structure of \( \text{O}_2\text{F}_2 \).

\[
\begin{array}{c}
\text{Cl, FC} = 7 - 2 - \frac{1}{2}(6) = +2 \\
\text{N, FC} = 5 - \frac{1}{2}(8) = +1
\end{array}
\]

Oxidation states are more useful when accounting for the reactivity of \( \text{O}_2\text{F}_2 \). We are forced to assign +1 as the oxidation state for oxygen due to the bonding to fluorine. Oxygen is very electronegative, and +1 is not a stable oxidation state for this element.

106. The nitrogen-nitrogen bond length of 112 pm is between a double (120 pm) and a triple (110 pm) bond. The nitrogen-oxygen bond length of 119 pm is between a single (147 pm) and a double bond (115 pm). The third resonance structure shown below doesn’t appear to be as important as the other two since there is no evidence from bond lengths for a nitrogen-oxygen triple bond or a nitrogen-nitrogen single bond as in the third resonance form. We can adequately describe the structure of \( \text{N}_2\text{O} \) using the resonance forms:

\[
\begin{align*}
\text{N} &= \text{N} = \text{O} \\
\text{N} &= \text{N} = \text{O}
\end{align*}
\]

Assigning formal charges for all three resonance forms:

\[
\begin{align*}
\text{N} &= \text{N} = \text{O} \\
\text{N} &= \text{N} = \text{O} \\
\text{N} &= \text{N} = \text{O}
\end{align*}
\]

For:
We should eliminate N≡O because it has a formal charge of +1 on the most electronegative element (O). This is consistent with the observation that the N–N bond is between a double and triple bond and that the N–O bond is between a single and double bond.

114. a. PCl$_3$ has 5 + 3(7) = 26 valence electrons.

\[ \text{Trigonal pyramid; all angles are } <109.5^\circ. \]

b. SCl$_2$ has 6 + 2(7) = 20 valence electrons.

\[ \text{V-shaped; angle is } <109.5^\circ. \]

c. SiF$_4$ has 4 + 4(7) = 32 valence electrons.

\[ \text{Tetrahedral; all angles are } 109.5^\circ. \]

Note: In PCl$_3$, SCl$_2$, and SiF$_4$, there are four pairs of electrons about the central atom in each case in this exercise. All of the structures are based on a tetrahedral geometry, but only SiF$_4$ has a tetrahedral structure. We consider only the relative positions of the atoms when describing the molecular structure.

115. a. XeCl$_2$ has 8 + 2(7) = 22 valence electrons.
There are five pairs of electrons about the central Xe atom. The structure will be based on a trigonal bipyramid geometry. The most stable arrangement of the atoms in XeCl$_2$ is a linear molecular structure with a 180° bond angle.

b. ICl$_3$ has 7 + 3(7) = 28 valence electrons.

\[
\begin{align*}
\text{Cl}: & = 90^\circ \\
\text{I}: & = 90^\circ \\
\text{T-shaped; the CIICl angles are } & \approx 90^\circ. \text{ Since the lone pairs will take up more space, the CIICl bond angles will probably be slightly less than 90°.}
\end{align*}
\]

c. TeF$_4$ has 6 + 4(7) = 34 valence electrons.

d. PCl$_5$ has 5 + 5(7) = 40 valence electrons.

All the species in this exercise have five pairs of electrons around the central atom. All the structures are based on a trigonal bipyramid geometry, but only in PCl$_5$ are all the pairs, bonding pairs. Thus PCl$_5$ is the only one for which we describe the molecular structure as trigonal bipyramid. Still, we had to begin with the trigonal bipyramid geometry to get to the structures (and bond angles) of the others.

118. All have polar bonds; in SiF$_4$, the individual bond dipoles cancel when summed together, and in PCl$_3$ and SC1$_2$, the individual bond dipoles do not cancel. Therefore, SiF$_4$ has no net dipole moment (is nonpolar), and PCl$_3$ and SC1$_2$ have net dipole moments (are polar). For PCl$_3$, the negative end of the dipole moment is between the more electronegative chlorine atoms, and the positive end is around P. For SC1$_2$, the negative end is between the more electronegative Cl atoms, and the positive end of the dipole moment is around S.

119. All have polar bonds, but only TeF$_4$ and ICl$_3$ have dipole moments. The bond dipoles from the five P–Cl bonds in PCl$_5$ cancel each other when summed together, so PCl$_5$ has no net dipole moment. The bond dipoles in XeCl$_2$ also cancel:

\[
\begin{align*}
\text{:Cl} & \quad \text{Xe} \quad \text{:Cl}
\end{align*}
\]

Because the bond dipoles from the two Xe–Cl bonds are equal in magnitude but point in opposite directions, they cancel each other, and XeCl$_2$ has no net dipole moment (is
nonpolar). For TeF$_4$ and ICl$_3$, the arrangement of these molecules is such that the individual bond dipoles do not all cancel, so each has an overall net dipole moment (is polar).

121. Molecules that have an overall dipole moment are called polar molecules, and molecules that do not have an overall dipole moment are called nonpolar molecules.

a. OCl$_2$, $6 + 2(7) = 20$ e$^-$  

V-shaped, polar; OCl$_2$ is polar because the two O–Cl bond dipoles don’t cancel each other. The resulting dipole moment is shown in the drawing.

KrF$_2$, $8 + 2(7) = 22$ e$^-$  

Linear, nonpolar; the molecule is nonpolar because the two Kr–F bond dipoles cancel each other.

b. SO$_3$, $6 + 3(6) = 24$ e$^-$  

V-shaped, polar; the S–O bond dipoles do not cancel, so SO$_3$ is polar (has a net dipole moment). Only one resonance structure is shown.

NF$_3$, $5 + 3(7) = 26$ e$^-$  

Trigonal pyramid, polar; bond dipoles do not cancel.

Note: All four species contain three atoms. They have different structures because the number of lone pairs of electrons around the central atom are different in each case.
resonance structure is shown.

IF$_3$ has $7 + 3(7) = 28$ valence electrons.

\[
\begin{array}{c}
\text{IF$_3$} \\
\text{T-shaped, polar; bond dipoles do not cancel.}
\end{array}
\]

*Note*: Each molecule has the same number of atoms but different structures because of differing numbers of lone pairs around each central atom.

c. CF$_4$, $4 + 4(7) = 32$ e$^-$

\[
\begin{array}{c}
\text{CF$_4$} \\
\text{Tetrahedral, nonpolar; bond dipoles cancel.}
\end{array}
\]

SeF$_4$, $6 + 4(7) = 34$ e$^-$

\[
\begin{array}{c}
\text{SeF$_4$} \\
\text{See-saw, polar; bond dipoles do not cancel.}
\end{array}
\]

KrF$_4$, $8 + 4(7) = 36$ valence electrons

\[
\begin{array}{c}
\text{KrF$_4$} \\
\text{Square planar, nonpolar; bond dipoles cancel.}
\end{array}
\]

*Note*: Again, each molecule has the same number of atoms but different structures because of differing numbers of lone pairs around the central atom.

d. IF$_5$, $7 + 5(7) = 42$ e$^-$

\[
\begin{array}{c}
\text{IF$_5$} \\
\text{Square pyramid, polar; bond dipoles do not cancel.}
\end{array}
\]

AsF$_5$, $5 + 5(7) = 40$ e$^-$

\[
\begin{array}{c}
\text{AsF$_5$} \\
\text{Trigonal bipyramid, nonpolar; bond dipoles cancel.}
\end{array}
\]

*Note*: Yet again, the molecules have the same number of atoms but different structures because of the presence of differing numbers of lone pairs.
125. All these molecules have polar bonds that are symmetrically arranged about the central atoms. In each molecule, the individual bond dipoles cancel each other out to give no net overall dipole moment. All these molecules are nonpolar even though they all contain polar bonds.

138. C≡O (1072 kJ/mol); N≡N (941 kJ/mol); CO is polar, whereas N₂ is nonpolar. This may lead to a great reactivity for the CO bond.