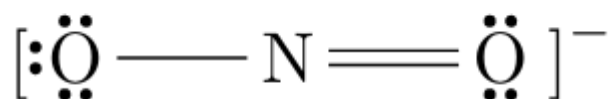7.42-7.50, 7.52-7.59, 7.61-7.62, 7.81

$$[:\ddot{\text{O}} \text{---} \text{N} =\!\!= \ddot{\text{O}}\,]^{-}$$

7.42

$$:\text{N} \equiv \text{N} \text{---} \ddot{\text{O}}:$$

$$\left[\text{H} \text{---} \text{C} =\!\!= \ddot{\text{O}} \atop \underset{:\ddot{\text{O}}:}{|} \right]^{-}$$
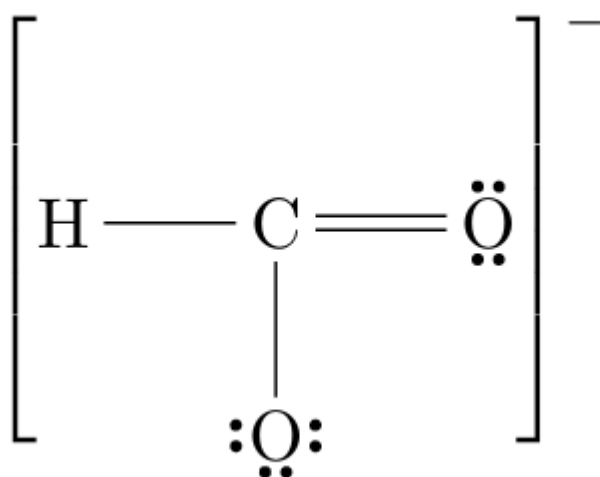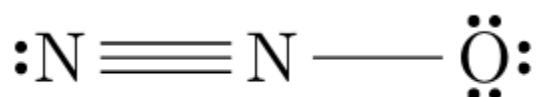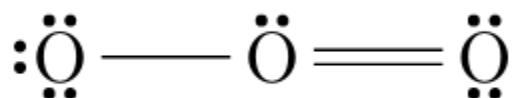
7.43 When the structure of a molecule exhibits resonance, the term ``resonance hybrid''
is justified because the actual structure is an average of all valid Lewis structures.

$$\ddot{\text{O}} =\!\!= \ddot{\text{O}} \text{---} \ddot{\text{O}}:$$

7.44

$$:\ddot{\text{O}} \text{---} \ddot{\text{O}} =\!\!= \ddot{\text{O}}$$
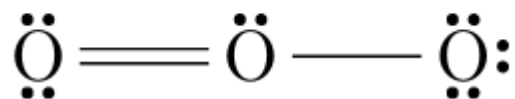
Since both resonance
structures contribute to the resonance hybrid, then we would expect both bonds to be
equal in length. Furthermore, we would expect these bond lengths to be intermediate
between a single bond and a double bond.

7.45 The existence of resonance structures for a molecule such as benzene results in

unusually unreactive chemical behavior, because the resulting resonance hybrid contains delocalized electrons that lowers the energy of the molecule and stabilizes it.
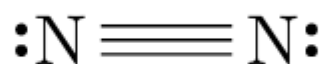
7.46 NO_3^- is predicted to have the longest bonds because it has the smallest bond order. NO_2^+ is predicted to have the shortest bonds because it has the largest bond order.

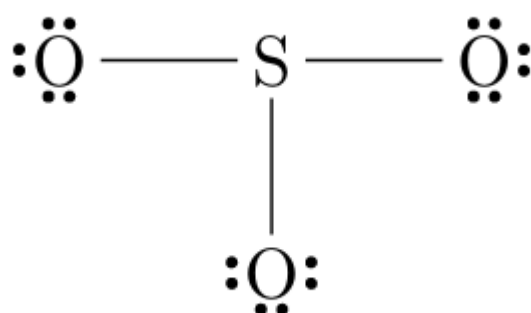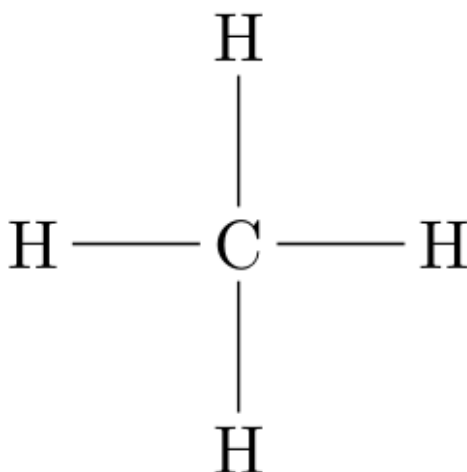4.47 PH_3

4.48 PART A: too many valence electrons; PART B: too many valence electrons

4.49 PART A: too many valence electrons; PART B: too few valence electrons
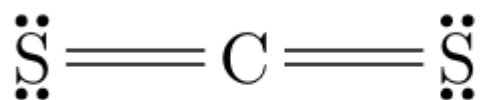
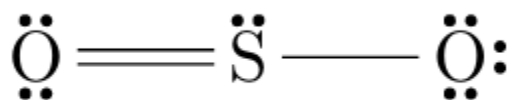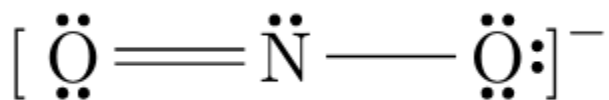$$:N\equiv\!\!\equiv\!\!\equiv N:$$

4.50

7.52 Orbital overlap explains the buildup of electron density between nuclei in a chemical bond, because the electron waves in this orbital overlap occupy the same region of space and interact with one another through constructive interference.

7.53 Sigma bonds overlap in an end-to-end fashion along a line connecting two nuclei, known as the bonding axis. On the other hand, pi bonds overlap on a side-by-side fashion above and below the bonding axis. Sigma and pi bonds are similar because they

both involve the buildup of electron density between nuclei.

7.54 $CH_3OH < CO_3^{2-} < CO_2 < CO$

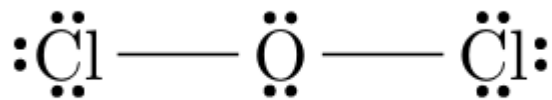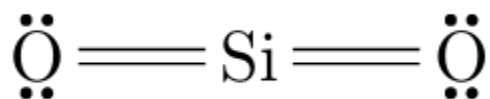$$\ddot{\underset{\cdot\cdot}{S}}=\!\!=C=\!\!=\ddot{\underset{\cdot\cdot}{S}}$$

7.55

$$H$$
$$|$$
$$H-C-\overset{\cdot\cdot}{\underset{\cdot\cdot}{C}}l\!:$$
$$|$$
$$H$$

$$[\,\overset{\cdot\cdot}{\underset{\cdot\cdot}{O}}=\!\!=\ddot{N}-\overset{\cdot\cdot}{\underset{\cdot\cdot}{O}}\!:]^{-}$$

$$\overset{\cdot\cdot}{\underset{\cdot\cdot}{O}}=\!\!=\ddot{\underset{\cdot\cdot}{S}}-\overset{\cdot\cdot}{\underset{\cdot\cdot}{O}}\!:$$
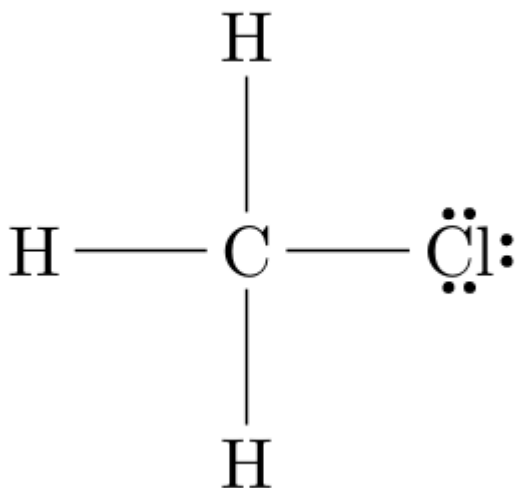
PART A: two pi bonds; PART B: zero pi bonds; PART C: one pi bond; PART D: one pi bond

$$:\overset{\cdot\cdot}{\underset{\cdot\cdot}{C}}l-\overset{\cdot\cdot}{\underset{\cdot\cdot}{O}}-\overset{\cdot\cdot}{\underset{\cdot\cdot}{C}}l\!:$$

7.56

H   H
| |
C = C
| |
H   H

  H $\ddot{N}$
  | ‖
  | ‖
  C ≡ C — C

$\ddot{O} = Si = \ddot{O}$
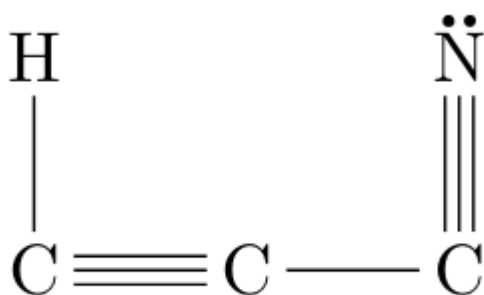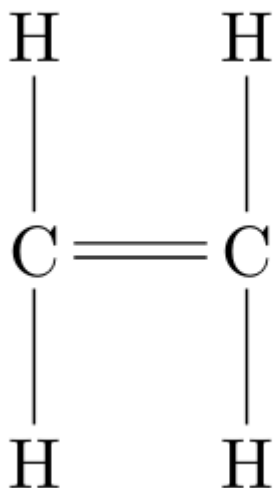
PART A: zero pi bonds; PART
B: one pi bond; PART C: four pi bonds; PART D: two pi bonds
7.57 The need for explanations of molecular shapes and bond angles are the observations that compels us to consider the concept of hybridization. It is intended to reconcile the notion of orbital overlap with these observations.
7.58 The Pauli exclusion principle suggests a reason why the number of hybrid orbitals formed must always be equal to the number of atomic orbitals combined, since no more than two electrons can be assigned to the same orbital and that these electrons must have opposite spins.
7.59 sp3 orbitals are the hybrid orbitals that are generated by combining one s and three p valence orbitals of an atom. A total of four sp3 hybrid orbitals are generated.
7.61 (A) four sp3 hybrid orbitals; (B) four sp3 hybrid orbitals; (C) two sp hybrid orbitals; (D) three sp2 hybrid orbitals
7.62 (A) two sp hybrid orbitals; (B) two sp hybrid orbitals; (C) three sp2 hybrid orbitals; (D) four sp4 hybrid orbitals

7.81 (A) sp, sp2, and sp3; (B) sp and sp2; (C) sp