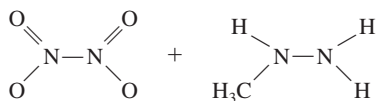
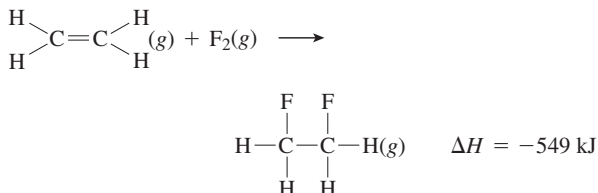


Use bond energies to estimate  $\Delta H$  for this reaction. The structures for the reactants are:



61. Consider the following reaction:



Estimate the carbon–fluorine bond energy given that the C–C bond energy is 347 kJ/mol, the C=C bond energy is 614 kJ/mol, and the F–F bond energy is 154 kJ/mol.

62. Consider the following reaction:



The bond energy for  $\text{A}_2$  is one-half the amount of the AB bond energy. The bond energy of  $\text{B}_2 = 432 \text{ kJ/mol}$ . What is the bond energy of  $\text{A}_2$ ?

63. Compare your answers from parts a and b of Exercise 53 with  $\Delta H$  values calculated for each reaction using standard enthalpies of formation in Appendix 4. Do enthalpy changes calculated from bond energies give a reasonable estimate of the actual values?
64. Compare your answer from Exercise 56 to the  $\Delta H$  value calculated from standard enthalpies of formation in Appendix 4. Explain any discrepancies.
65. The standard enthalpies of formation for  $\text{S}(\text{g})$ ,  $\text{F}(\text{g})$ ,  $\text{SF}_4(\text{g})$ , and  $\text{SF}_6(\text{g})$  are +278.8, +79.0, -775, and -1209 kJ/mol, respectively.
- Use these data to estimate the energy of an S–F bond.
  - Compare your calculated values to the value given in Table 8.4. What conclusions can you draw?
  - Why are the  $\Delta H_f^\circ$  values for  $\text{S}(\text{g})$  and  $\text{F}(\text{g})$  not equal to zero, since sulfur and fluorine are elements?
66. Use the following standard enthalpies of formation to estimate the N–H bond energy in ammonia:  $\text{N}(\text{g})$ , 472.7 kJ/mol;  $\text{H}(\text{g})$ , 216.0 kJ/mol;  $\text{NH}_3(\text{g})$ , -46.1 kJ/mol. Compare your value to the one in Table 8.4.

### Lewis Structures and Resonance

67. Write Lewis structures that obey the octet rule for each of the following.
- |                    |                          |                  |
|--------------------|--------------------------|------------------|
| a. HCN             | d. $\text{NH}_4^+$       | g. $\text{CO}_2$ |
| b. $\text{PH}_3$   | e. $\text{H}_2\text{CO}$ | h. $\text{O}_2$  |
| c. $\text{CHCl}_3$ | f. $\text{SeF}_2$        | i. HBr           |
- Except for HCN and  $\text{H}_2\text{CO}$ , the first atom listed is the central atom. For HCN and  $\text{H}_2\text{CO}$ , carbon is the central atom.
68. Write Lewis structures that obey the octet rule for each of the following molecules and ions. (In each case the first atom listed is the central atom.)
- $\text{POCl}_3$ ,  $\text{SO}_4^{2-}$ ,  $\text{XeO}_4$ ,  $\text{PO}_4^{3-}$ ,  $\text{ClO}_4^-$
  - $\text{NF}_3$ ,  $\text{SO}_3^{2-}$ ,  $\text{PO}_3^{3-}$ ,  $\text{ClO}_3^-$

c.  $\text{ClO}_2^-$ ,  $\text{SCl}_2$ ,  $\text{PCl}_2^-$

d. Considering your answers to parts a, b, and c, what conclusions can you draw concerning the structures of species containing the same number of atoms and the same number of valence electrons?

69. One type of exception to the octet rule are compounds with central atoms having fewer than eight electrons around them.  $\text{BeH}_2$  and  $\text{BH}_3$  are examples of this type of exception. Draw the Lewis structures for  $\text{BeH}_2$  and  $\text{BH}_3$ .
70. Lewis structures can be used to understand why some molecules react in certain ways. Write the Lewis structures for the reactants and products in the reactions described below.
- Nitrogen dioxide dimerizes to produce dinitrogen tetroxide.
  - Boron trihydride accepts a pair of electrons from ammonia, forming  $\text{BH}_3\text{NH}_3$ .
- Give a possible explanation for why these two reactions occur.
71. The most common type of exception to the octet rule are compounds or ions with central atoms having more than eight electrons around them.  $\text{PF}_5$ ,  $\text{SF}_6$ ,  $\text{ClF}_3$  and  $\text{Br}_3^-$  are examples of this type of exception. Draw the Lewis structure for these compounds or ions. Which elements, when they have to, can have more than eight electrons around them? How is this rationalized?
72.  $\text{SF}_6$ ,  $\text{ClF}_5$ , and  $\text{XeF}_4$  are three compounds whose central atoms do not follow the octet rule. Draw Lewis structures for these compounds.
73. Write Lewis structures for the following. Show all resonance structures where applicable.
- $\text{NO}_2^-$ ,  $\text{NO}_3^-$ ,  $\text{N}_2\text{O}_4$  ( $\text{N}_2\text{O}_4$  exists as  $\text{O}_2\text{N}-\text{NO}_2$ .)
  - $\text{OCN}^-$ ,  $\text{SCN}^-$ ,  $\text{N}_3^-$  (Carbon is the central atom in  $\text{OCN}^-$  and  $\text{SCN}^-$ .)
74. Some of the important pollutants in the atmosphere are ozone ( $\text{O}_3$ ), sulfur dioxide, and sulfur trioxide. Write Lewis structures for these three molecules. Show all resonance structures where applicable.
75. Benzene ( $\text{C}_6\text{H}_6$ ) consists of a six-membered ring of carbon atoms with one hydrogen bonded to each carbon. Write Lewis structures for benzene, including resonance structures.
76. Borazine ( $\text{B}_3\text{N}_3\text{H}_6$ ) has often been called “inorganic” benzene. Write Lewis structures for borazine. Borazine contains a six-membered ring of alternating boron and nitrogen atoms with one hydrogen bonded to each boron and nitrogen.
77. An important observation supporting the concept of resonance in the localized electron model was that there are only three different structures of dichlorobenzene ( $\text{C}_6\text{H}_4\text{Cl}_2$ ). How does this fact support the concept of resonance (see Exercise 75)?
78. Consider the following bond lengths:
- $$\text{C—O} \quad 143 \text{ pm} \quad \text{C=O} \quad 123 \text{ pm} \quad \text{C}\equiv\text{O} \quad 109 \text{ pm}$$
- In the  $\text{CO}_3^{2-}$  ion, all three C–O bonds have identical bond lengths of 136 pm. Why?
79. Place the species below in order of shortest to longest nitrogen–nitrogen bond.
- $$\text{N}_2 \quad \text{N}_2\text{F}_4 \quad \text{N}_2\text{F}_2$$
- ( $\text{N}_2\text{F}_4$  exists as  $\text{F}_2\text{N}-\text{NF}_2$ , and  $\text{N}_2\text{F}_2$  exists as  $\text{FN}-\text{NF}$ .)