Lattice energy	−690. kJ/mol
Ionization energy for K	419 kJ/mol
Electron affinity of Cl	−349 kJ/mol
Bond energy of Cl_2	239 kJ/mol
Enthalpy of sublimation for K	64 kJ/mol

46. Use the following data to estimate ΔH_f° for magnesium fluoride.

$$Mg(s) + F_2(g) \longrightarrow MgF_2(s)$$

Lattice energy	−3916 kJ/mol
First ionization energy of Mg	735 kJ/mol
Second ionization energy of Mg	1445 kJ/mol
Electron affinity of F	−328 kJ/mol
Bond energy of F_2	154 kJ/mol
Enthalpy of sublimation of Mg	150. kJ/mol

47. Consider the following energy changes:

	ΔH (kJ/mol)
$Mg(g) \rightarrow Mg^+(g) + e^-$	735
$Mg^+(g) \rightarrow Mg^{2+}(g) + e^-$	1445
$O(g) + e^- \rightarrow O^-(g)$	−141
$O^-(g) + e^- \rightarrow O^{2-}(g)$	878

Magnesium oxide exists as $Mg^{2+}O^{2-}$ and not as Mg^+O^-. Explain.

48. Compare the electron affinity of fluorine to the ionization energy of sodium. Is the process of an electron being "pulled" from the sodium atom to the fluorine atom exothermic or endothermic? Why is NaF a stable compound? Is the overall formation of NaF endothermic or exothermic? How can this be?

49. LiI(*s*) has a heat of formation of −272 kJ/mol and a lattice energy of −753 kJ/mol. The ionization energy of Li(*g*) is 520. kJ/mol, the bond energy of $I_2(g)$ is 151 kJ/mol, and the electron affinity of I(*g*) is −295 kJ/mol. Use these data to determine the heat of sublimation of Li(*s*).

50. Use the following data to estimate ΔH for the reaction $S^-(g) + e^- \rightarrow S^{2-}(g)$. Include an estimate of uncertainty.

	ΔH_f°	**Lattice Energy**	**I.E. of M**	**ΔH_{sub} of M**
Na_2S	−365	−2203	495	109
K_2S	−381	−2052	419	90
Rb_2S	−361	−1949	409	82
Cs_2S	−360	−1850	382	78

$$S(s) \longrightarrow S(g) \qquad \Delta H = 277 \text{ kJ/mol}$$
$$S(g) + e^- \longrightarrow S^-(g) \qquad \Delta H = -200 \text{ kJ/mol}$$

Assume that all values are known to ±1 kJ/mol.

51. Rationalize the following lattice energy values:

Compound	**Lattice Energy (kJ/mol)**
CaSe	−2862
Na_2Se	−2130
CaTe	−2721
Na_2Te	−2095

52. The lattice energies of $FeCl_3$, $FeCl_2$, and Fe_2O_3 are (in no particular order) −2631, −5359, and −14,774 kJ/mol. Match the appropriate formula to each lattice energy. Explain.

Bond Energies

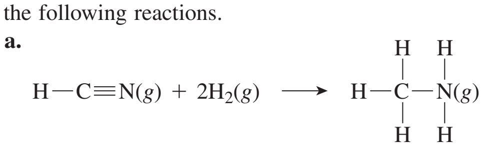
53. Use bond energy values (Table 8.4) to estimate ΔH for each of the following reactions in the gas phase.

a. $H_2 + Cl_2 \rightarrow 2HCl$

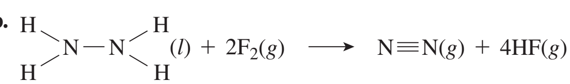
b. $N\equiv N + 3H_2 \rightarrow 2NH_3$

54. Use bond energy values (Table 8.4) to estimate ΔH for each of the following reactions.

a.



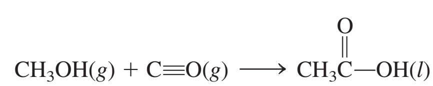
b.



55. Use bond energies (Table 8.4) to predict ΔH for the isomerization of methyl isocyanide to acetonitrile:

$$CH_3N\equiv C(g) \longrightarrow CH_3C\equiv N(g)$$

56. Acetic acid is responsible for the sour taste of vinegar. It can be manufactured using the following reaction:



Use tabulated values of bond energies (Table 8.4) to estimate ΔH for this reaction.

57. Use bond energies to predict ΔH for the combustion of ethanol:

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

58. Use bond energies to estimate ΔH for the combustion for one mole of acetylene:

$$C_2H_2(g) + \tfrac{5}{2}O_2(g) \longrightarrow 2CO_2(g) + H_2O(g)$$

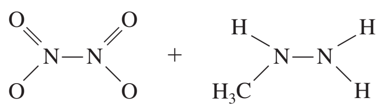
59. Use bond energies to estimate ΔH for the following reaction:

$$H_2O_2(aq) + CH_3OH(aq) \longrightarrow H_2CO(aq) + 2H_2O(l)$$

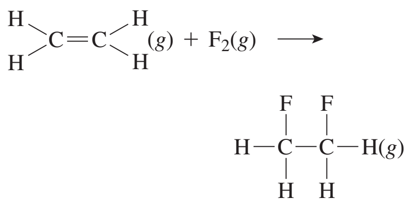
60. The space shuttle orbiter utilizes the oxidation of methyl hydrazine by dinitrogen tetroxide for propulsion:

$$5N_2O_4(l) + 4N_2H_3CH_3(l) \longrightarrow 12H_2O(g) + 9N_2(g) + 4CO_2(g)$$

Use bond energies to estimate ΔH for this reaction. The structures for the reactants are:



61. Consider the following reaction:



$\Delta H = -549$ kJ

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:

$$A_2 + B_2 \longrightarrow 2AB \qquad \Delta H = -285 \text{ kJ}$$

The bond energy for A_2 is one-half the amount of the AB bond energy. The bond energy of B_2 = 432 kJ/mol. What is the bond energy of A_2?

63. Compare your answers from parts a and b of Exercise 53 with ΔH 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 ΔH value calculated from standard enthalpies of formation in Appendix 4. Explain any discrepancies.

65. The standard enthalpies of formation for S(*g*), F(*g*), $SF_4(g)$, and $SF_6(g)$ are +278.8, +79.0, −775, and −1209 kJ/mol, respectively.
  a. Use these data to estimate the energy of an S—F bond.
  b. Compare your calculated values to the value given in Table 8.4. What conclusions can you draw?
  c. Why are the ΔH_f° values for S(*g*) and F(*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: N(*g*), 472.7 kJ/mol; H(*g*), 216.0 kJ/mol; $NH_3(g)$, −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.** NH_4^+	**g.** CO_2
b. PH_3	**e.** H_2CO	**h.** O_2
c. $CHCl_3$	**f.** SeF_2	**i.** HBr

Except for HCN and H_2CO, the first atom listed is the central atom. For HCN and H_2CO, 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.)
  a. $POCl_3$, SO_4^{2-}, XeO_4, PO_4^{3-}, ClO_4^-
  b. NF_3, SO_3^{2-}, PO_3^{3-}, ClO_3^-
  c. ClO_2^-, SCl_2, 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. BeH_2 and BH_3 are examples of this type of exception. Draw the Lewis structures for BeH_2 and 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.
  a. Nitrogen dioxide dimerizes to produce dinitrogen tetroxide.
  b. Boron trihydride accepts a pair of electrons from ammonia, forming BH_3NH_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. PF_5, SF_4, ClF_3 and 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. SF_6, ClF_5, and 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.
  a. NO_2^-, NO_3^-, N_2O_4 (N_2O_4 exists as $O_2N—NO_2$.)
  b. OCN^-, SCN^-, N_3^- (Carbon is the central atom in OCN^- and SCN^-.)

74. Some of the important pollutants in the atmosphere are ozone (O_3), sulfur dioxide, and sulfur trioxide. Write Lewis structures for these three molecules. Show all resonance structures where applicable.

75. Benzene (C_6H_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 ($B_3N_3H_6$) 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 ($C_6H_4Cl_2$). How does this fact support the concept of resonance (see Exercise 75)?

78. Consider the following bond lengths:

C—O 143 pm C═O 123 pm C≡O 109 pm

In the 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.

N_2 N_2F_4 N_2F_2

(N_2F_4 exists as $F_2N—NF_2$, and N_2F_2 exists as FN—NF.)