

## Sample Exercise 8.1 Magnitudes of Lattice Energies

Without consulting Table 8.2, arrange the ionic compounds NaF, CsI, and CaO in order of increasing lattice energy.

### Solution

**Analyze** From the formulas for three ionic compounds, we must determine their relative lattice energies.

**Plan** We need to determine the charges and relative sizes of the ions in the compounds. We then use Equation 8.4 qualitatively to determine the relative energies, knowing that (a) the larger the ionic charges, the greater the energy and (b) the farther apart the ions are, the lower the energy.

$$E_{el} = \frac{\kappa Q_1 Q_2}{d} \quad [8.4]$$

**Solve** NaF consists of  $\text{Na}^+$  and  $\text{F}^-$  ions, CsI of  $\text{Cs}^+$  and  $\text{I}^-$  ions, and CaO of  $\text{Ca}^{2+}$  and  $\text{O}^{2-}$  ions. Because the product  $Q_1 Q_2$  appears in the numerator of Equation 8.4, the lattice energy increases dramatically when the charges increase. Thus, we expect the lattice energy of CaO, which has 2+ and 2− ions, to be the greatest of the three.

The ionic charges are the same in NaF and CsI. As a result, the difference in their lattice energies depends on the difference in the distance between ions in the lattice. Because ionic size increases as we go down a group in the periodic table (Section 7.3), we know that  $\text{Cs}^+$  is larger than  $\text{Na}^+$  and  $\text{I}^-$  is larger than  $\text{F}^-$ . Therefore, the distance between  $\text{Na}^+$  and  $\text{F}^-$  ions in NaF is less than the distance between the  $\text{Cs}^+$  and  $\text{I}^-$  ions in CsI. As a result, the lattice energy of NaF should be greater than that of CsI. In order of increasing energy, therefore, we have  $\text{CsI} < \text{NaF} < \text{CaO}$ .

## Sample Exercise 8.1 Magnitudes of Lattice Energies

Continued

**Check** Table 8.2 confirms our predicted order is correct.

**TABLE 8.2 • Lattice Energies for Some Ionic Compounds**

Compound	Lattice Energy (kJ/mol)	Compound	Lattice Energy (kJ/mol)
LiF	1030	MgCl <sub>2</sub>	2326
LiCl	834	SrCl <sub>2</sub>	2127
LiI	730		
NaF	910	MgO	3795
NaCl	788	CaO	3414
NaBr	732	SrO	3217
NaI	682		
KF	808	ScN	7547
KCl	701		
KBr	671		
CsCl	657		
CsI	600		

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### Practice Exercise

Which substance do you expect to have the greatest lattice energy, MgF<sub>2</sub>, CaF<sub>2</sub>, or ZrO<sub>2</sub>?

**Answer:** ZrO<sub>2</sub>

## Sample Exercise 8.2 Charges on Ions

Predict the ion generally formed by (a) Sr, (b) S, (c) Al.

### Solution

**Analyze** We must decide how many electrons are most likely to be gained or lost by atoms of Sr, S, and Al.

**Plan** In each case we can use the element's position in the periodic table to predict whether the element forms a cation or an anion. We can then use its electron configuration to determine the most likely ion formed.

**Solve** (a) Strontium is a metal in group 2A and therefore forms a cation. Its electron configuration is  $[\text{Kr}]5s^2$ , and so we expect that the two valence electrons can be lost easily to give an  $\text{Sr}^{2+}$  ion. (b) Sulfur is a nonmetal in group 6A and will thus tend to be found as an anion. Its electron configuration ( $[\text{Ne}]3s^23p^4$ ) is two electrons short of a noble-gas configuration.

**Check** The ionic charges we predict here are confirmed in Tables 2.4 and 2.5.

TABLE 2.4 • Common Cations\*

Charge	Formula	Name	Formula	Name
1+	$\text{H}^+$	hydrogen ion	$\text{NH}_4^+$	ammonium ion
	$\text{Li}^+$	lithium ion	$\text{Cu}^+$	copper(I) or cuprous ion
	$\text{Na}^+$	sodium ion		
	$\text{K}^+$	potassium ion		
	$\text{Cs}^+$	cesium ion		
	$\text{Ag}^+$	silver ion		
2+	$\text{Mg}^{2+}$	magnesium ion	$\text{Co}^{2+}$	cobalt(II) or cobaltous ion
	$\text{Ca}^{2+}$	calcium ion	$\text{Cu}^{2+}$	copper(II) or cupric ion
	$\text{Sr}^{2+}$	strontium ion	$\text{Fe}^{2+}$	iron(II) or ferrous ion
	$\text{Ba}^{2+}$	barium ion	$\text{Mn}^{2+}$	manganese(II) or manganous ion
	$\text{Zn}^{2+}$	zinc ion	$\text{Hg}_2^{2+}$	mercury(I) or mercurous ion
	$\text{Cd}^{2+}$	cadmium ion	$\text{Hg}^{2+}$	mercury(II) or mercuric ion
			$\text{Ni}^{2+}$	nickel(II) or nickelous ion
			$\text{Pb}^{2+}$	lead(II) or plumbous ion
			$\text{Sn}^{2+}$	tin(II) or stannous ion
3+	$\text{Al}^{3+}$	aluminum ion	$\text{Cr}^{3+}$	chromium(III) or chromic ion
			$\text{Fe}^{3+}$	iron(III) or ferric ion

\*The ions we use most often in this course are in boldface. Learn them first.  
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TABLE 2.5 • Common Anions\*

Charge	Formula	Name	Formula	Name
1−	$\text{H}^-$	hydride ion	$\text{CH}_3\text{COO}^-$ (or $\text{C}_2\text{H}_3\text{O}_2^-$ )	acetate ion
	$\text{F}^-$	fluoride ion	$\text{ClO}_3^-$	chlorate ion
	$\text{Cl}^-$	chloride ion	$\text{ClO}_4^-$	perchlorate ion
	$\text{Br}^-$	bromide ion	$\text{NO}_3^-$	nitrate ion
	$\text{I}^-$	iodide ion	$\text{MnO}_4^-$	permanganate ion
	$\text{CN}^-$	cyanide ion		
	$\text{OH}^-$	hydroxide ion		
2−	$\text{O}^{2-}$	oxide ion	$\text{CO}_3^{2-}$	carbonate ion
	$\text{O}_2^{2-}$	peroxide ion	$\text{CrO}_4^{2-}$	chromate ion
	$\text{S}^{2-}$	sulfide ion	$\text{Cr}_2\text{O}_7^{2-}$	dichromate ion
			$\text{SO}_4^{2-}$	sulfate ion
3−	$\text{N}^{3-}$	nitride ion	$\text{PO}_4^{3-}$	phosphate ion

\*The ions we use most often are in boldface. Learn them first.  
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## Sample Exercise 8.2 Charges on Ions

Continued

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### Practice Exercise

Predict the charges on the ions formed when magnesium reacts with nitrogen.

**Answer:**  $\text{Mg}^{2+}$  and  $\text{N}^{3-}$

## Sample Exercise 8.3 Lewis Structure of a Compound

Given the Lewis symbols for nitrogen and fluorine in Table 8.1, predict the formula of the stable binary compound (a compound composed of two elements) formed when nitrogen reacts with fluorine and draw its Lewis structure.

TABLE 8.1 • Lewis Symbols

Group	Element	Electron Configuration	Lewis Symbol	Element	Electron Configuration	Lewis Symbol
1A	Li	[He]2s <sup>1</sup>	Li·	Na	[Ne]3s <sup>1</sup>	Na·
2A	Be	[He]2s <sup>2</sup>	·Be·	Mg	[Ne]3s <sup>2</sup>	·Mg·
3A	B	[He]2s <sup>2</sup> 2p <sup>1</sup>	·B·	Al	[Ne]3s <sup>2</sup> 3p <sup>1</sup>	·Al·
4A	C	[He]2s <sup>2</sup> 2p <sup>2</sup>	·C·	Si	[Ne]3s <sup>2</sup> 3p <sup>2</sup>	·Si·
5A	N	[He]2s <sup>2</sup> 2p <sup>3</sup>	·N·	P	[Ne]3s <sup>2</sup> 3p <sup>3</sup>	·P·
6A	O	[He]2s <sup>2</sup> 2p <sup>4</sup>	·O·	S	[Ne]3s <sup>2</sup> 3p <sup>4</sup>	·S·
7A	F	[He]2s <sup>2</sup> 2p <sup>5</sup>	·F·	Cl	[Ne]3s <sup>2</sup> 3p <sup>5</sup>	·Cl·
8A	Ne	[He]2s <sup>2</sup> 2p <sup>6</sup>	·Ne·	Ar	[Ne]3s <sup>2</sup> 3p <sup>6</sup>	·Ar·

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### Solution

**Analyze** The Lewis symbols for nitrogen and fluorine reveal that nitrogen has five valence electrons and fluorine has seven.

**Plan** We need to find a combination of the two elements that results in an octet of electrons around each atom. Nitrogen requires three additional electrons to complete its octet, and fluorine requires one. Sharing a pair of electrons between one N atom and one F atom will result in an octet of electrons for fluorine but not for nitrogen. We therefore need to figure out a way to get two more electrons for the N atom.

**Solve** Nitrogen must share a pair of electrons with three fluorine atoms to complete its octet. Thus, the binary compound these two elements form must be NF<sub>3</sub>:



## Sample Exercise 8.3 Lewis Structure of a Compound

Continued

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**Check** The Lewis structure in the center shows that each atom is surrounded by an octet of electrons. Once you are accustomed to thinking of each line in a Lewis structure as representing *two* electrons, you can just as easily use the structure on the right to check for octets.

### Practice Exercise

Compare the Lewis symbol for neon with the Lewis structure for methane,  $\text{CH}_4$ . In what important way are the electron arrangements about neon and carbon alike? In what important way are they different?

**Answer:** Both atoms have an octet of electrons. However, the electrons about neon are unshared electron pairs, whereas those about carbon are shared with four hydrogen atoms.

## Sample Exercise 8.4 Bond Polarity

In each case, which bond is more polar: (a) B—Cl or C—Cl, (b) P—F or P—Cl ? Indicate in each case which atom has the partial negative charge.

### Solution

**Analyze** We are asked to determine relative bond polarities, given nothing but the atoms involved in the bonds.

**Plan** Because we are not asked for quantitative answers, we can use the periodic table and our knowledge of electronegativity trends to answer the question.

### Solve

(a) The chlorine atom is common to both bonds. Therefore, the analysis reduces to a comparison of the electronegativities of B and C. Because boron is to the left of carbon in the periodic table, we predict that boron has the lower electronegativity. Chlorine, being on the right side of the table, has a higher electronegativity. The more polar bond will be the one between the atoms having the lowest electronegativity (boron) and the highest electronegativity (chlorine). Consequently, the B—Cl bond is more polar; the chlorine atom carries the partial negative charge because it has a higher electronegativity.

(b) In this example phosphorus is common to both bonds, and the analysis reduces to a comparison of the electronegativities of F and Cl. Because fluorine is above chlorine in the periodic table, it should be more electronegative and will form the more polar bond with P. The higher electronegativity of fluorine means that it will carry the partial negative charge.

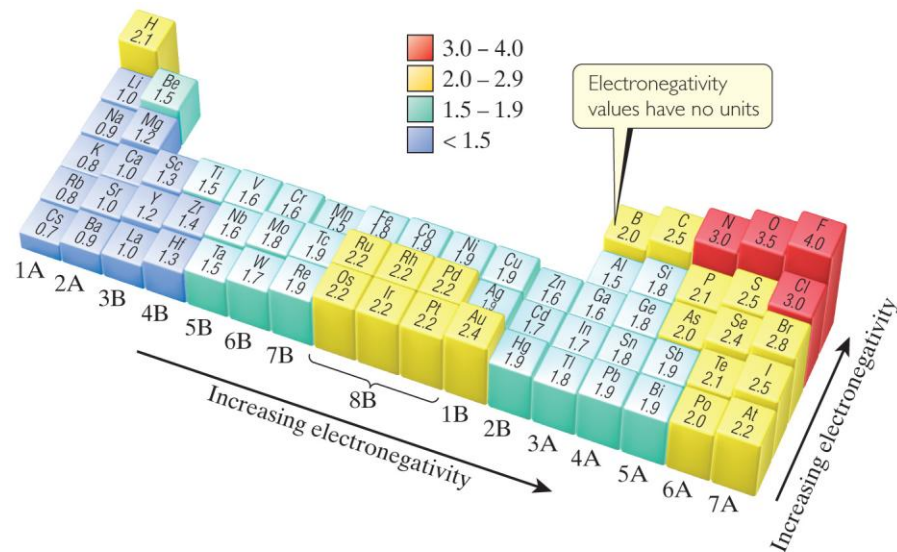
## Sample Exercise 8.4 Bond Polarity

Continued

### Check

(a) Using Figure 8.7: The difference in the electronegativities of chlorine and boron is  $3.0 - 2.0 = 1.0$ ; the difference between chlorine and carbon is  $3.0 - 2.5 = 0.5$ . Hence, the B—Cl bond is more polar, as we had predicted.

(b) Using Figure 8.7: The difference in the electronegativities of chlorine and phosphorus is  $3.0 - 2.1 = 0.9$ ; the difference between fluorine and phosphorus is  $4.0 - 2.1 = 1.9$ . Hence, the P—F bond is more polar, as we had predicted.



### Practice Exercise

Which of the following bonds is most polar: S—Cl, S—Br, Se—Cl, or Se—Br ?

**Answer:** Se—Cl



## Sample Exercise 8.5 Dipole Moments of Diatomic Molecules

The bond length in the HCl molecule is 1.27 Å. **(a)** Calculate the dipole moment, in debyes, that results if the charges on the H and Cl atoms were and 1+ and 1−, respectively. **(b)** The experimentally measured dipole moment of HCl(g) is 1.08 D. What magnitude of charge, in units of  $e$ , on the H and Cl atoms leads to this dipole moment?

### Solution

**Analyze and Plan** We are asked in **(a)** to calculate the dipole moment of HCl that would result if there were a full charge transferred from H to Cl. We can use Equation 8.11 to obtain this result. In **(b)**, we are given the actual dipole moment for the molecule and will use that value to calculate the actual partial charges on the H and Cl atoms.

**(a)** The charge on each atom is the electronic charge,  
 $e = 1.60 \times 10^{-19} \text{ C}$ .  
The separation is 1.27 Å. The dipole moment is therefore

$$\mu = Qr = (1.60 \times 10^{-19} \text{ C})(1.27 \text{ Å}) \left( \frac{10^{-10} \text{ m}}{1 \text{ Å}} \right) \left( \frac{1 \text{ D}}{3.34 \times 10^{-30} \text{ C-m}} \right) = 6.08 \text{ D}$$

**(b)** We know the value of  $\mu$ , 1.08 D, and the value of  $r$ , 1.27 Å. We want to calculate the value of  $Q$ :

$$Q = \frac{\mu}{r} = \frac{(1.08 \text{ D}) \left( \frac{3.34 \times 10^{-30} \text{ C-m}}{1 \text{ D}} \right)}{(1.27 \text{ Å}) \left( \frac{10^{-10} \text{ m}}{1 \text{ Å}} \right)} = 2.84 \times 10^{-20} \text{ C}$$

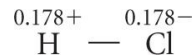
$$\text{Charge in } e = (2.84 \times 10^{-20} \text{ C}) \left( \frac{1 e}{1.60 \times 10^{-19} \text{ C}} \right) = 0.178e$$

We can readily convert this charge to units of  $e$ :

## Sample Exercise 8.5 Dipole Moments of Diatomic Molecules

Continued

Thus, the experimental dipole moment indicates that the charge separation in the HCl molecule is



Because the experimental dipole moment is less than that calculated in part (a), the charges on the atoms are much less than a full electronic charge. We could have anticipated this because the H—Cl bond is polar covalent rather than ionic.

### Practice Exercise

The dipole moment of chlorine monofluoride, ClF(*g*), is 0.88 D. The bond length of the molecule is 1.63 Å. **(a)** Which atom is expected to have the partial negative charge? **(b)** What is the charge on that atom in units of *e*?

**Answers:** (a) F, (b) 0.11–

## Sample Exercise 8.6 Drawing a Lewis Structure

Draw the Lewis structure for phosphorus trichloride,  $\text{PCl}_3$ .

### Solution

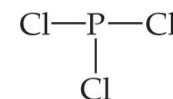
**Analyze and Plan** We are asked to draw a Lewis structure from a molecular formula. Our plan is to follow the five-step procedure just described.

### Solve

First, we sum the valence electrons. Phosphorus (group 5A) has five valence electrons, and each chlorine (group 7A) has seven. The total number of valence electrons is therefore

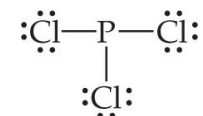
$$5 + (3 \times 7) = 26$$

Second, we arrange the atoms to show which atom is connected to which, and we draw a single bond between them. There are various ways the atoms might be arranged. In binary compounds, however, the first element in the chemical formula is generally surrounded by the remaining atoms. Thus, we begin with a skeleton structure that shows a single bond between the P atom and each Cl atom:



(It is not crucial that the Cl atoms be left of, right of, and below the P atom—any structure that shows each of the three Cl atoms bonded to P will work.)

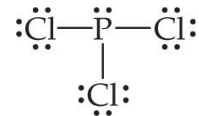
Third, we complete the octets on the atoms bonded to the central atom. Placing octets around each Cl atom accounts for 24 electrons (remember, each line in our structure represents *two* electrons):



## Sample Exercise 8.6 Drawing a Lewis Structure

Continued

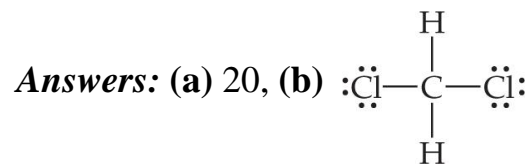
Fourth, recalling that our total number of electrons is 26, we place the remaining two electrons on the central atom, completing the octet around it:



This structure gives each atom an octet, so we stop at this point. (In checking for octets, remember to count both electrons in a single bond twice, once for each atom in the bond.)

### Practice Exercise

- (a) How many valence electrons should appear in the Lewis structure for  $\text{CH}_2\text{Cl}_2$ ?  
(b) Draw the Lewis structure.

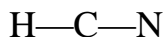


## Sample Exercise 8.7 Lewis Structure with a Multiple Bond

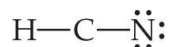
Draw the Lewis structure for HCN.

### Solution

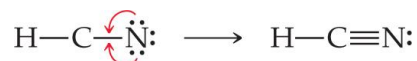
Hydrogen has one valence electron, carbon (group 4A) has four, and nitrogen (group 5A) has five. The total number of valence electrons is, therefore,  $1 + 4 + 5 = 10$ . In principle, there are different ways in which we might choose to arrange the atoms. Because hydrogen can accommodate only one electron pair, it always has only one single bond associated with it. Therefore, C—H—N is an impossible arrangement. The remaining two possibilities are H—C—N and H—N—C. The first is the arrangement found experimentally. You might have guessed this because (a) the formula is written with the atoms in this order and (b) carbon is less electronegative than nitrogen. Thus, we begin with the skeleton structure



The two bonds account for four electrons. The H atom can have only two electrons associated with it, and so we will not add any more electrons to it. If we place the remaining six electrons around N to give it an octet, we do not achieve an octet on C:



We therefore try a double bond between C and N, using one of the unshared pairs we placed on N. Again we end up with fewer than eight electrons on C, and so we next try a triple bond. This structure gives an octet around both C and N:



The octet rule is satisfied for the C and N atoms, and the H atom has two electrons around it. This is a correct Lewis structure.

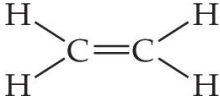
## Sample Exercise 8.7 Lewis Structure with a Multiple Bond

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### Practice Exercise

Draw the Lewis structure for (a)  $\text{NO}^+$  ion, (b)  $\text{C}_2\text{H}_4$ .

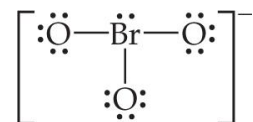
**Answers:** (a)  $[\text{:N}\equiv\text{O:}]^+$ , (b) 

## Sample Exercise 8.8 Lewis Structure for a Polyatomic Ion

Draw the Lewis structure for the  $\text{BrO}_3^-$  ion.

### Solution

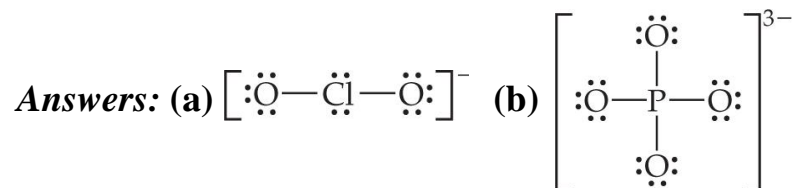
Bromine (group 7A) has seven valence electrons, and oxygen (group 6A) has six. We must add one more electron to our sum to account for the  $1-$  charge of the ion. The total number of valence electrons is, therefore,  $7 + (3 \times 6) + 1 = 16$ . For oxyanions— $\text{BrO}_3^-$ ,  $\text{SO}_4^{2-}$ ,  $\text{NO}_3^-$ ,  $\text{CO}_3^{2-}$ , and so forth—the oxygen atoms surround the central nonmetal atom. After following this format and then putting in the single bonds and distributing the unshared electron pairs, we have



Notice that the Lewis structure for an ion is written in brackets and the charge is shown outside the brackets at the upper right.

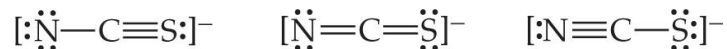
### Practice Exercise

Draw the Lewis structure for (a)  $\text{ClO}_2^-$ , (b)  $\text{PO}_4^{3-}$



## Sample Exercise 8.9 Lewis Structures and Formal Charges

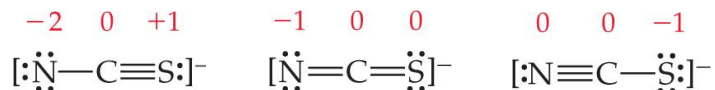
Three possible Lewis structures for the thiocyanate ion,  $\text{NCS}^-$ , are



(a) Determine the formal charges in each structure. (b) Based on the formal charges, which Lewis structure is the dominant one?

### Solution

(a) Neutral N, C, and S atoms have five, four, and six valence electrons, respectively. We can determine the formal charges in the three structures by using the rules we just discussed:



As they must, the formal charges in all three structures sum to  $1-$ , the overall charge of the ion. (b) The dominant Lewis structure generally produces formal charges of the smallest magnitude (guideline 1). That rules out the left structure as the dominant one. Further, as discussed in Section 8.4, N is more electronegative than C or S. Therefore, we expect any negative formal charge to reside on the N atom (guideline 2). For these two reasons, the middle Lewis structure is the dominant one for  $\text{NCS}^-$ .

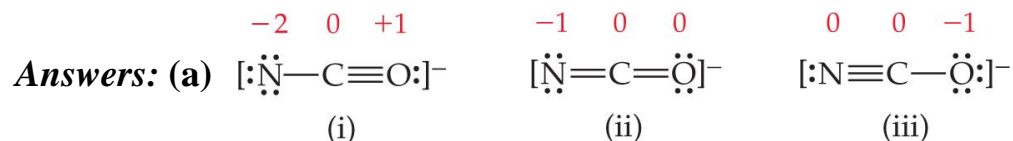


## Sample Exercise 8.9 Lewis Structures and Formal Charges

Continued

### Practice Exercise

The cyanate ion,  $\text{NCO}^-$ , has three possible Lewis structures. **(a)** Draw these three structures and assign formal charges in each. **(b)** Which Lewis structure is dominant?



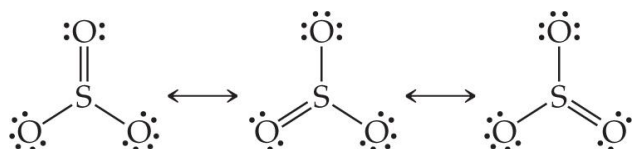
**(b)** Structure (iii), which places a negative charge on oxygen, the most electronegative element in the ion, is the dominant Lewis structure.

## Sample Exercise 8.10 Resonance Structures

Which is predicted to have the shorter sulfur–oxygen bonds,  $\text{SO}_3$  or  $\text{SO}_3^{2-}$ ?

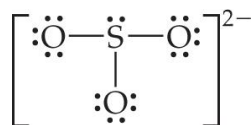
### Solution

The sulfur atom has six valence electrons, as does oxygen. Thus,  $\text{SO}_3$  contains 24 valence electrons. In writing the Lewis structure, we see that three equivalent resonance structures can be drawn:



As with  $\text{NO}_3^-$ , the actual structure of  $\text{SO}_3$  is an equal blend of all three. Thus, each S–O bond length should be about one-third of the way between the length of a single bond and the length of a double bond. That is, they should be shorter than single bonds but not as short as double bonds.

The  $\text{SO}_3^{2-}$  ion has 26 electrons, which leads to a dominant Lewis structure in which all the S–O bonds are single:



Our analysis of the Lewis structures leads us to conclude that  $\text{SO}_3$  should have the shorter S–O bonds and  $\text{SO}_3^{2-}$  the longer ones. This conclusion is correct: The experimentally measured S–O bond lengths are 1.42 Å in  $\text{SO}_3$  and 1.51 Å in  $\text{SO}_3^{2-}$ .

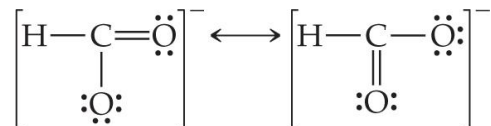
## Sample Exercise 8.10 Resonance Structures

Continued

### Practice Exercise

Draw two equivalent resonance structures for the formate ion,  $\text{HCO}_2^-$ .

**Answer:**



## Sample Exercise 8.11 Lewis Structure for an Ion with More than an Octet of Electrons

Draw the Lewis structure for  $\text{ICl}_4^-$ .

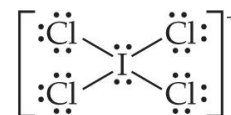
### Solution

Iodine (group 7A) has seven valence electrons. Each chlorine atom (group 7A) also has seven. An extra electron is added to account for the 1- charge of the ion. Therefore, the total number of valence electrons is

$$7 + (4 \times 7) + 1 = 36$$

The I atom is the central atom in the ion. Putting eight electrons around each Cl atom (including a pair of electrons between I and each Cl to represent the single bond between these atoms) requires  $8 \times 4 = 32$  electrons.

We are thus left with  $36 - 32 = 4$  electrons to be placed on the larger iodine:



Iodine has 12 valence electrons around it, four more than needed for an octet.

## Sample Exercise 8.11 Lewis Structure for an Ion with More than an Octet of Electrons

Continued

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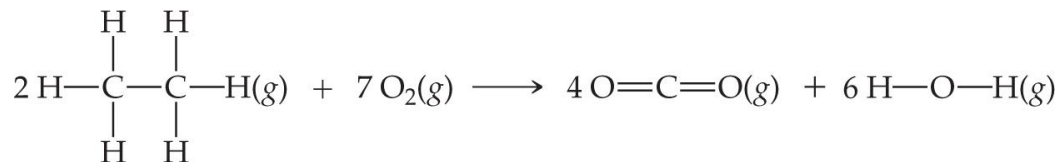
### Practice Exercise

(a) Which of the following atoms is never found with more than an octet of valence electrons around it: S, C, P, Br? (b) Draw the Lewis structure for  $\text{XeF}_2$ .

**Answers:** (a) C, (b)  $\text{:}\ddot{\text{F}}\text{—}\ddot{\text{Xe}}\text{—}\ddot{\text{F}}\text{:}$

## Sample Exercise 8.12 Using Average Bond Enthalpies

Using data from Table 8.4, estimate  $\Delta H$  for the reaction



### Solution

**Analyze** We are asked to estimate the enthalpy change for a chemical reaction by using average bond enthalpies for the bonds broken and formed.

**Plan** In the reactants, we must break twelve C—H bonds and two C—C bonds in the two molecules of  $\text{C}_2\text{H}_6$  and seven  $\text{O}_2$  bonds in the seven  $\text{O}_2$  molecules. In the products, we form eight C=O bonds (two in each  $\text{CO}_2$ ) and twelve O—H bonds (two in each  $\text{H}_2\text{O}$ ).

TABLE 8.4 • Average Bond Enthalpies (kJ/mol)

#### Single Bonds

C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146		
C—N	293	N—O	201	O—F	190	Cl—F	253
C—O	358	N—F	272	O—Cl	203	Cl—Cl	242
C—F	485	N—Cl	200	O—I	234		
C—Cl	328	N—Br	243			Br—F	237
C—Br	276			S—H	339	Br—Cl	218
C—I	240	H—H	436	S—F	327	Br—Br	193
C—S	259	H—F	567	S—Cl	253		
		H—Cl	431	S—Br	218	I—Cl	208
Si—H	323	H—Br	366	S—S	266	I—Br	175
Si—Si	226	H—I	299			I—I	151
Si—C	301						
Si—O	368						
Si—Cl	464						

#### Multiple Bonds

C=C	614	N=N	418	O <sub>2</sub>	495
C≡C	839	N≡N	941		
C=N	615	N=O	607	S=O	523
C≡N	891			S=S	418
C=O	799				
C≡O	1072				

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## Sample Exercise 8.12 Using Average Bond Enthalpies

Continued

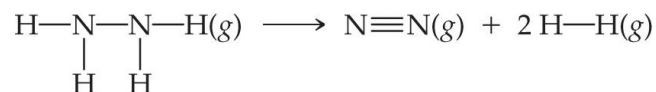
**Solve** Using Equation 8.12 and data from Table 8.4, we have

$$\begin{aligned}\Delta H &= [12D(\text{C—H}) + 2D(\text{C—C}) + 7D(\text{O}_2)] - [8D(\text{C=O}) + 12D(\text{O—H})] \\ &= [12(413 \text{ kJ}) + 2(348 \text{ kJ}) + 7(495 \text{ kJ})] - [8(799 \text{ kJ}) + 12(463 \text{ kJ})] \\ &= 9117 \text{ kJ} - 11948 \text{ kJ} \\ &= -2831 \text{ kJ}\end{aligned}$$

**Check** This estimate can be compared with the value of  $-2856 \text{ kJ}$  calculated from more accurate thermochemical data; the agreement is good.

### Practice Exercise

Using Table 8.4, estimate  $\Delta H$  for the reaction



**Answer:**  $-86 \text{ kJ}$

TABLE 8.4 • Average Bond Enthalpies (kJ/mol)

Single Bonds							
C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146		
C—N	293	N—O	201	O—F	190	Cl—F	253
C—O	358	N—F	272	O—Cl	203	Cl—Cl	242
C—F	485	N—Cl	200	O—I	234		
C—Cl	328	N—Br	243			Br—F	237
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C—I	240	H—H	436	S—F	327	Br—Br	193
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Si—Si	226	H—I	299			I—I	151
Si—C	301						
Si—O	368						
Si—Cl	464						
Multiple Bonds							
C=C	614	N=N	418	O <sub>2</sub>	495		
C≡C	839	N≡N	941				
C=N	615	N=O	607	S=O	523		
C≡N	891			S=S	418		
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C≡O	1072						

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