WORKED EXAMPLE 4.1 Predicting Ionic and Molecular Compounds

Which of the following compounds would you expect to be ionic and which molecular (covalent)?

(a) BaF_2 **(b)** SF_4 **(c)** PH_3 **(d)** CH_3OH

Strategy

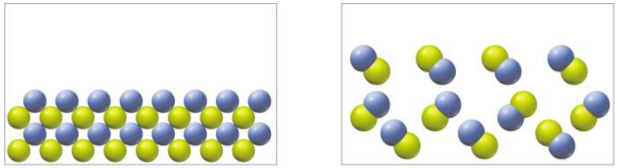
Remember that covalent bonds generally form between nonmetal atoms, yielding covalent molecular compounds, whereas ionic bonds form between metal and nonmetal atoms.

Solution

Compound **(a)** is composed of a metal (barium) and a nonmetal (fluorine) and is likely to be ionic. Compounds **(b)–(d)** are composed entirely of nonmetals and therefore are probably molecular.

Problem 4.1 Which of the following compounds would you expect to be ionic and which molecular (covalent)?
(a) LiBr **(b)** $SiCl_4$ **(c)** BF_3 **(d)** CaO

Key Concept Problem 4.2 Which of the following drawings is most likely to represent an ionic compound and which a molecular compound? Explain.



(a) (b)

WORKED EXAMPLE 4.2 Ionization Energies

Arrange the elements Se, Cl, and S in order of increasing ionization energy.

Strategy

Ionization energy generally increases from left to right across a row of the periodic table and decreases from top to bottom down a group. Chlorine should have a larger E_i than its neighbor sulfur, and selenium should have a smaller E_i than sulfur.

Solution

The order is Se < S < Cl.

Problem 4.7 Using the periodic table as your guide, predict which element in each of the following pairs has the larger ionization energy:

(a) K or Br **(b)** S or Te **(c)** Ga or Se **(d)** Ne or Sr

WORKED EXAMPLE 4.3 Higher Ionization Energies

Which has the larger fifth ionization energy, Ge or As?

Strategy

Look at their positions in the periodic table. The group 4A element germanium has four valence-shell electrons and thus has four relatively low ionization energies, whereas the group 5A element arsenic has five valence-shell electrons and has five low ionization energies.

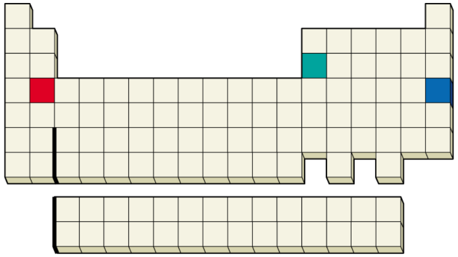
Solution

Germanium has a larger E_{i5} than arsenic.

Problem 4.8
(a) Which has the larger third ionization energy, Be or N?
(b) Which has the larger fourth ionization energy, Ga or Ge?

Problem 4.9 Three atoms have the following electron configurations:
(a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$ **(b)** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ **(c)** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
Which of the three has the largest E_{i1}? Which has the smallest E_{i4}?

Key Concept Problem 4.10 Order the indicated three elements according to the ease with which each is likely to lose its third electron:



WORKED EXAMPLE 4.4 Electron Affinities

Why does nitrogen have a less favorable (more positive) E_{ea} than its neighbors on either side, C and O?

Strategy and Solution

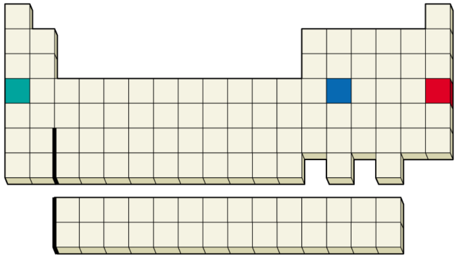
The magnitude of an element's E_{ea} depends on the element's valence-shell electron configuration. The electron configurations of C, N, and O are

Carbon: $1s^2\ 2s^2\ 2p_x{}^1 2p_y{}^1$ Nitrogen: $1s^2\ 2s^2\ 2p_x{}^1 2p_y{}^1 2p_z{}^1$
Oxygen: $1s^2\ 2s^2\ 2p_x{}^2 2p_y{}^1 2p_z{}^1$

Carbon has only two electrons in its $2p$ subshell and can readily accept another in its vacant $2p_z$ orbital. Nitrogen, however, has a half-filled $2p$ subshell, so the additional electron must pair up in a $2p$ orbital where it feels a repulsion from the electron already present. Thus, the E_{ea} of nitrogen is less favorable than that of carbon. Oxygen also must add an electron to an orbital that already has one electron, but the additional stabilizing effect of increased Z_{eff} across the periodic table counteracts the effect of electron repulsion, resulting in a more favorable E_{ea} for O than for N.

Problem 4.11 Explain why manganese (atomic number 25) has a less favorable E_{ea} than its neighbors on either side.

Key Concept Problem 4.12 Which of the indicated three elements has the least favorable E_{ea}? Which has the most favorable E_{ea}?



WORKED EXAMPLE 4.5 Chemical Reactions and the Octet Rule

Lithium metal reacts with nitrogen to yield Li_3N. What noble gas configuration does the nitrogen atom in Li_3N have?

Strategy and Solution

The compound Li_3N contains three lithium ions, each formed by the loss of a 2*s* electron from lithium metal (group 1A). The nitrogen atom in Li_3N must therefore gain three electrons over the neutral atom, making it triply negative and giving it a valence-shell octet with the neon configuration:

N configuration $(1s^2 2s^2 2p^3)$ N^{3-} configuration: $(1s^2 2s^2 2p^6)$

Problem 4.13 What noble gas configurations are the following elements likely to adopt in reactions when they form ions?

(a) Rb **(b)** Ba **(c)** Ga **(d)** F

Problem 4.14 What are group 6A elements likely to do when they form ions—gain electrons or lose them? How many?

WORKED EXAMPLE 4.6 **Lattice Energies**

Which has the larger lattice energy, NaCl or CsI?

Strategy

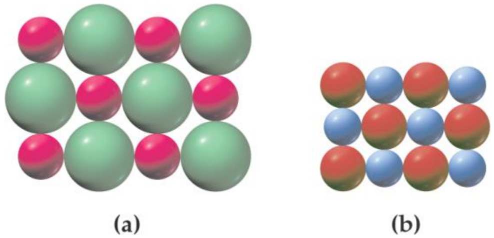
The magnitude of a substance's lattice energy is affected both by the charges on its constituent ions and by the sizes of those ions. The higher the charges on the ions and the smaller the sizes of the ions, the larger the lattice energy. All four ions—Na^+, Cs^+, $C1^-$, and I^-—are singly charged, but they differ in size.

Solution

Because Na^+ is smaller than Cs^+ and $C1^-$ is smaller than I^-, the distance between ions is smaller in NaCl than in CsI. Thus, NaCl has the larger lattice energy.

WORKED KEY CONCEPT EXAMPLE 4.7 Lattice Energies

Which of the following alkali metal halides has the larger lattice energy? Which has the smaller lattice energy? Explain.



Strategy

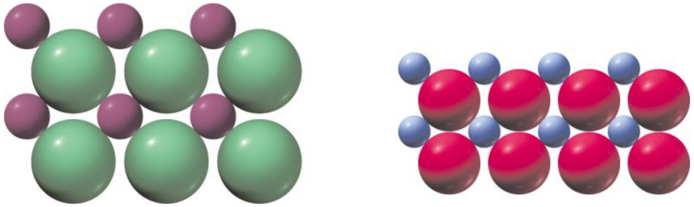
The magnitude of a lattice energy depends directly on the charge on the ions and inversely on the distance between ions (that is, on the radii of the ions). In this instance, all the ions in both drawings are singly charged, so only the size of the ions is important.

Solution

The ions in drawing **(b)** are smaller than those in drawing **(a)**, so **(b)** has the larger lattice energy.

Problem 4.16 Which substance in each of the following pairs has the larger lattice energy?
(a) KCl or RbCl **(b)** CaF_2 or BaF_2 **(c)** CaO or KI

Key Concept Problem 4.17 One of the following pictures represents NaCl and one represents MgO. Which is which? Which has the larger lattice energy?



General Chemistry: Atoms First
By John E. McMurry and Robert C. Fay

WORKED EXAMPLE 4.8 Naming Binary Ionic Compounds

Give systematic names for the following compounds:

(a) $BaCl_2$ **(b)** $CrCl_3$ **(c)** PbS **(d)** Fe_2O_3

Strategy

Try to figure out the number of positive charges on each cation by counting the number of negative charges on the associated anion(s). Refer to Figures 4.11 and 4.12 if you are unsure.

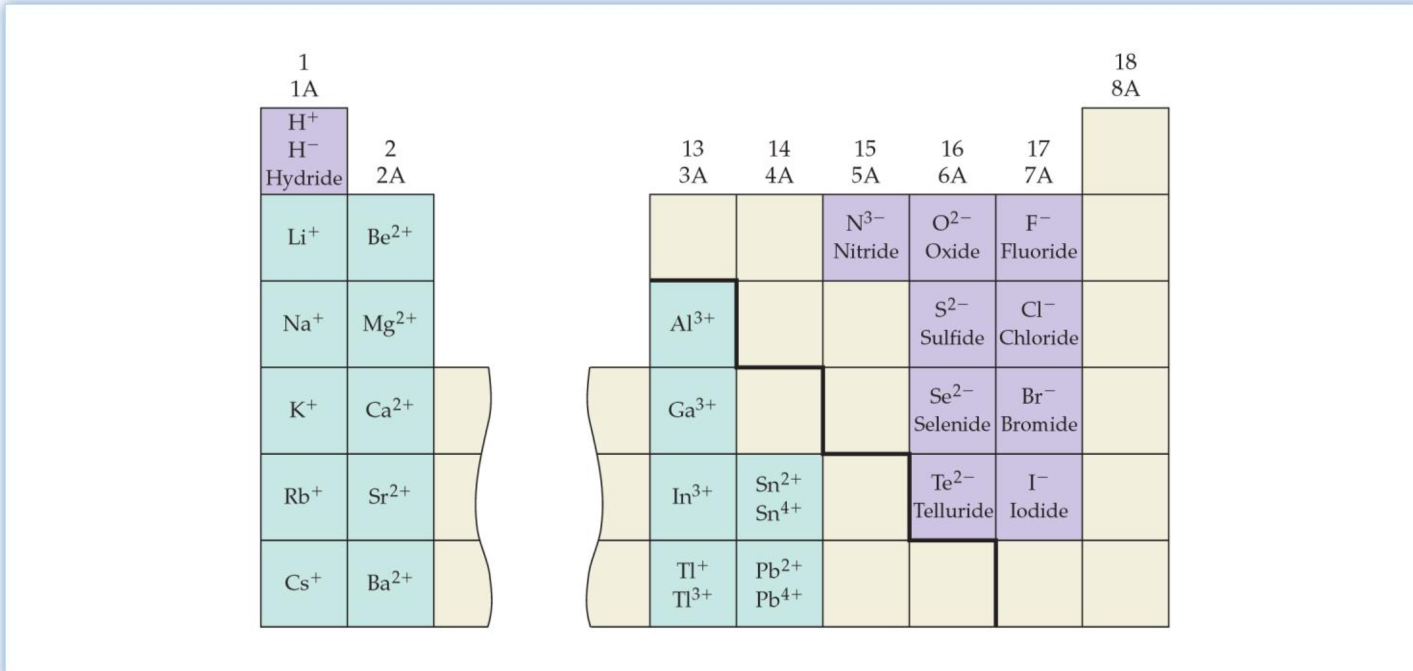


FIGURE 4.11
Main-group cations (green) and anions (purple). A cation bears the same name as the element it is derived from; an anion name has an *-ide* ending.

WORKED EXAMPLE 4.8 Naming Binary Ionic Compounds

Continued

3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B
Sc^{3+}	Ti^{3+}	V^{3+}	Cr^{2+} Cr^{3+}	Mn^{2+}	Fe^{2+} Fe^{3+}	Co^{2+}	Ni^{2+}	Cu^{+} Cu^{2+}	Zn^{2+}
Y^{3+}					Ru^{3+}	Rh^{3+}	Pd^{2+}	Ag^{+}	Cd^{2+}
									Hg^{2+}

FIGURE 4.12
Common transition-metal ions. Only ions that exist in aqueous solution are shown.

Solution

(a) Barium chloride — No Roman numeral is necessary because barium, a group 2A element, forms only Ba^{2+}.

(b) Chromium(III) chloride — The Roman numeral III is necessary to specify the +3 charge on chromium (a transition metal).

(c) Lead(II) sulfide — The sulfide anion (s^{2-}) has a double negative charge, so the lead cation must be doubly positive.

(d) Iron(III) oxide — The three oxide anions (O^{2-}) have a total negative charge of –6, so the two iron cations must have a total charge of +6. Thus, each is Fe(III).

WORKED EXAMPLE 4.9 Converting Names into Formulas

Write formulas for the following compounds:

(a) Magnesium fluoride **(b)** Tin(IV) oxide **(c)** Iron(III) sulfide

Strategy

For transition-metal compounds, the charge on the cation is indicated by the Roman numeral in the name. Knowing the number of positive charges, you can then figure out the number of necessary negative charges for the associated anions.

Solution

(a) MgF_2 Magnesium (group 2A) forms only a 2+ cation, so there must be two fluoride ions (F^-) to balance the charge.

(b) SnO_2 Tin(IV) has a +4 charge, so there must be two oxide ions (O^{2-}) to balance the charge.

(c) Fe_2S_3 Iron(III) has a +3 charge and sulfide ion a –2 charge (s^{2-}), so there must be two irons and three sulfurs.

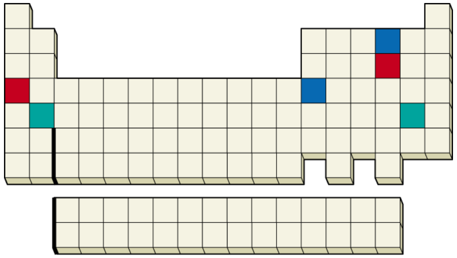
Problem 4.18 Give systematic names for the following compounds:
(a) CsF **(b)** K_2O **(c)** CuO **(d)** BaS **(e)** $BeBr_2$

Problem 4.19 Write formulas for the following compounds:
(a) Vanadium(III) chloride **(b)** Manganese(IV) oxide
(c) Copper(II) sulfide **(d)** Aluminum oxide

WORKED EXAMPLE 4.9 Converting Names into Formulas

Continued

Key Concept Problem 4.20 Three binary ionic compounds are represented on the following periodic table: red with red, green with green, and blue with blue. Name each, and give its likely formula.



WORKED EXAMPLE 4.10 Naming Compounds with Polyatomic Ions

Give systematic names for the following compounds:

(a) $LiNO_3$ **(b)** $KHSO_4$ **(c)** $CuCO_3$ **(d)** $Fe(C1O_4)_3$

Strategy

The names and charges of the common polyatomic ions must be memorized. Refer to Table 4.4 if you need help.

TABLE 4.4 Some Common Polyatomic Ions

Formula	Name	Formula	Name
Cation		**Singly charged anions (continued)**	
NH_4^+	Ammonium	NO_2^-	Nitrite
		NO_3^-	Nitrate
Singly charged anions			
$CH_3CO_2^-$	Acetate	**Doubly charged anions**	
CN^-	Cyanide	CO_3^{2-}	Carbonate
ClO^-	Hypochlorite	CrO_4^{2-}	Chromate
ClO_2^-	Chlorite	$Cr_2O_7^{2-}$	Dichromate
ClO_3^-	Chlorate	O_2^{2-}	Peroxide
ClO_4^-	Perchlorate	HPO_4^{2-}	Hydrogen phosphate
$H_2PO_4^-$	Dihydrogen phosphate	SO_3^{2-}	Sulfite
HCO_3^-	Hydrogen carbonate (or bicarbonate)	SO_4^{2-}	Sulfate
		$S_2O_3^{2-}$	Thiosulfate
HSO_4^-	Hydrogen sulfate (or bisulfate)		
		Triply charged anion	
OH^-	Hydroxide	PO_4^{3-}	Phosphate
MnO_4^-	Permanganate		

WORKED EXAMPLE 4.10 Naming Compounds with Polyatomic Ions

Continued

Solution

(a) Lithium nitrate	Lithium (group 1A) forms only the Li^+ ion and does not need a Roman numeral.
(b) Potassium hydrogen sulfate	Potassium (group 1A) forms only the K^+ ion.
(c) Copper(II) carbonate	The carbonate ion has a –2 charge, so copper must be +2. A Roman numeral is needed because copper, a transition metal, can form more than one ion.
(d) Iron(III) perchlorate	There are three perchlorate ions, each with a –1 charge, so the iron must have a +3 charge.

WORKED EXAMPLE 4.11 Writing Formulas of Compounds with Polyatomic Ions

Write formulas for the following compounds:

(a) Potassium hypochlorite **(b)** Silver(I) chromate **(c)** Iron(III) carbonate

Strategy

(a) KC1O — Potassium forms only the K^+ ion, so only one $C1O^-$ is needed.

(b) Ag_2CrO_4 — The polyatomic chromate ion has a –2 charge, so two Ag^+ ions are needed.

(c) $Fe_2(CO_3)_3$ — Iron(III) has a +3 charge and the polyatomic carbonate ion has a –2 charge, so there must be two iron ions and three carbonate ions. The polyatomic carbonate ion is set off in parentheses to indicate that there are three of them.

Problem 4.21 Give systematic names for the following compounds:

(a) $Ca(ClO)_2$ **(b)** $Ag_2S_2O_3$ **(c)** NaH_2PO_4
(d) $Sn(NO_3)_2$ **(e)** $Pb(CH_3CO_2)_4$ **(f)** $(NH_4)_2SO_4$

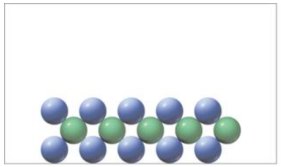
Problem 4.22 Write formulas for the following compounds:

(a) Lithium phosphate **(b)** Magnesium hydrogen sulfate
(c) Manganese(II) nitrate **(d)** Chromium(III) sulfate

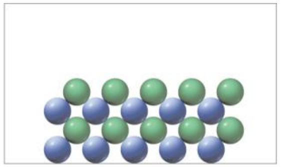
WORKED EXAMPLE 4.11 Writing Formulas of Compounds with Polyatomic Ions

Continued

Key Concept Problem 4.23 The following drawings are those of solid ionic compounds, with green spheres representing the cations and blue spheres representing the anions in each.



(1)



(2)

Which of the following formulas are consistent with each drawing?

(a) LiBr **(b)** $NaNO_2$ **(c)** $CaCl_2$ **(d)** K_2CO_3 **(e)** $Fe_2(SO4)_3$