Chapter 3

Chemical Compounds and Bonding

Solutions for Practice Problems

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1. Determine ΔEN for each bond shown. Indicate whether each bond is ionic or covalent.
   (a) O–H **(b)** C–H **(c)** Mg–Cl **(d)** B–F
   (e) Cr–O **(f)** C–N **(g)** Na–I **(h)** Na–Br

   What Is Required?
   You have to determine the electronegativity difference between the bonded atoms listed and decide if it is an ionic or covalent bond.

   What Is Given?
   The identity of each element in the chemical bond is given.

   Plan Your Strategy
   Look up the electronegativity of each given element in Figure 3.6, student textbook page 71. For each bond combination, subtract the lower electronegativity value from the higher value to get the ΔEN for the bond. From Figure 3.8, student textbook page 73, you can see that a ΔEN value between 0.0 and 1.7 will likely be covalent, while a value between 1.7 and 3.3 will likely be ionic.

   Act on Your Strategy
   (a) O–H, $\Delta EN = EN\text{O} - EN\text{H} = 3.44 - 2.20 = 1.24$, covalent
   (b) C–H, $\Delta EN = EN\text{C} - EN\text{H} = 2.55 - 2.20 = 0.35$, covalent
   (c) Mg–Cl, $\Delta EN = EN\text{Cl} - EN\text{Mg} = 3.16 - 1.31 = 1.85$, ionic
   (d) B–F, $\Delta EN = EN\text{F} - EN\text{B} = 3.98 - 2.04 = 1.94$, ionic
   (e) Cr–O, $\Delta EN = EN\text{O} - EN\text{Cr} = 3.44 - 1.66 = 1.78$, ionic
   (f) C–N, $\Delta EN = EN\text{N} - EN\text{C} = 3.04 - 2.55 = 0.49$, covalent
   (g) Na–I, $\Delta EN = EN\text{I} - EN\text{Na} = 2.66 - 0.93 = 1.73$, ionic
   (h) Na–Br, $\Delta EN = EN\text{Br} - EN\text{Na} = 2.96 - 0.93 = 2.03$, ionic

   Check Your Solution
   An ionic bond will fall in the ΔEN range of 1.7–3.3. A covalent bond will fall in the ΔEN range of 0.0–1.7. Double check your electronegativity values from Figure 3.6. None of your answers should be negative values.

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2. **Problem**
   For each bond below, determine ΔEN. Is the bond ionic or covalent?
   (a) Ca–O **(b)** K–Cl **(c)** K–F
   (d) Li–F **(e)** Li–Br **(f)** Ba–O

What Is Required?
You have to determine the electronegativity difference between the bonded atoms listed and decide if it is an ionic or covalent bond.

What Is Given?
The identity of each element in the chemical bond is given.

Plan Your Strategy
Look up the electronegativity of each given element in Figure 3.6, student textbook page 71. For each bond combination, subtract the lower electronegativity value from the higher value to get the ΔEN for the bond. From Figure 3.8, student textbook page 73, you can see that a ΔEN value between 0.0 and 1.7 will likely be covalent, while a value between 1.7 and 3.3 will likely be ionic.

Act on Your Strategy
(a) Ca–O, $\Delta EN = EN\text{O} - EN\text{Ca} = 3.44 - 1.00 = 2.44$, ionic
(b) K–Cl, $\Delta EN = EN\text{Cl} - EN\text{K} = 3.16 - 0.82 = 2.34$, ionic
(c) K–F, $\Delta EN = EN\text{F} - EN\text{K} = 3.98 - 0.82 = 3.16$, ionic
(d) Li–F, $\Delta EN = EN\text{F} - EN\text{Li} = 3.98 - 0.98 = 3.00$, ionic
(e) Li–Br, $\Delta EN = EN\text{Br} - EN\text{Li} = 2.96 - 0.98 = 1.98$, ionic
(f) Ba–O, $\Delta EN = EN\text{O} - EN\text{Ba} = 3.44 - 0.89 = 2.45$, ionic

Check Your Solution
An ionic bond will fall in the ΔEN range of 1.7–3.3. A covalent bond will fall in the ΔEN range of 0.0–1.7. Double check your electronegativity values from Figure 3.6. None of your answers should be negative values.

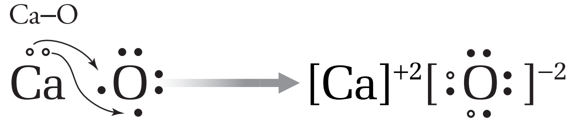
3. Problem
Draw Lewis structures to represent the formation of each bond in question 2.

What Is Required?
You have to draw Lewis structures for the bonds listed in Problem 2.

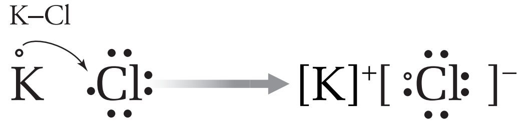
What Is Given?
The identity of each element in the chemical bond is given.

Plan Your Strategy
Apply the octet rule for each element in the bond. Determine the number of valence electrons in the outermost shell of each element in the bond. The electrons from the less electronegative element will be transferred to the element with the higher electronegativity to achieve a stable noble gas configuration.

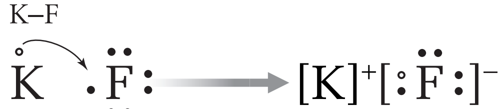
Act on Your Strategy
(a) Ca–O



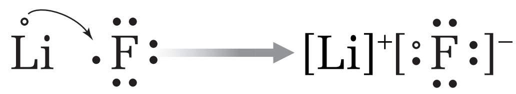
(b) K–Cl



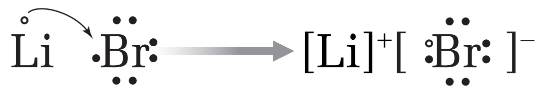
(c) K–F



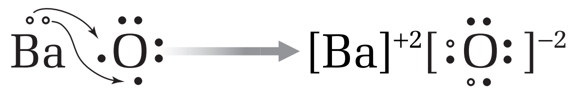
(d) Li–F



(e) Li–Br



(f) Ba–O



Check Your Solution

In each case, after the bond has been made, the element with the higher electronegativity should have a total of 8 electrons around it. It should have a net negative charge. The element with the lower electronegativity loses all of its valence electrons and should carry a net positive charge. The charge number of the two elements in the bond should be the same.

Solutions for Practice Problems

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4. Problem

For each pair of elements, determine ΔEN.

(a) magnesium and chlorine
(b) calcium and chlorine
(c) lithium and oxygen
(d) sodium and oxygen
(e) potassium and sulfur
(f) calcium and bromide

What Is Required?

You have to determine the electronegativity difference between the bonded atoms listed.

What Is Given?

The identity of each element in the chemical bond is given.

Plan Your Strategy

Look up the electronegativity of each given element in Figure 3.6, student textbook page 71. For each bond combination, subtract the lower electronegativity value from the higher value to get the ΔEN for the bond.

Act on Your Strategy

(a) Mg–Cl, $\Delta EN = EN\text{Cl} - EN\text{Mg} = 3.16 - 1.31 = 1.85$
(b) Ca–Cl, $\Delta EN = EN\text{Cl} - EN\text{Ca} = 3.16 - 1.00 = 2.16$
(c) Li–O, $\Delta EN = EN\text{O} - EN\text{Li} = 3.44 - 0.98 = 2.46$
(d) Na–O, $\Delta EN = EN\text{O} - EN\text{Na} = 3.44 - 0.93 = 2.51$
(e) K–S, $\Delta EN = EN\text{S} - EN\text{K} = 2.58 - 0.82 = 1.76$
(f) Ca–Br, $\Delta EN = EN\text{Br} - EN\text{Ca} = 2.96 - 1.00 = 1.96$

Check Your Solution

Double check your electronegativity values from Figure 3.6. None of your answers should be negative values.

5. Problems

Draw Lewis structures to show how each pair of elements in question 4 forms bonds to achieve a stable octet.

What Is Required?
You have to draw Lewis structures for the bonds listed in Problem 4.

What Is Given?
The identity of each element in the chemical bond is given.

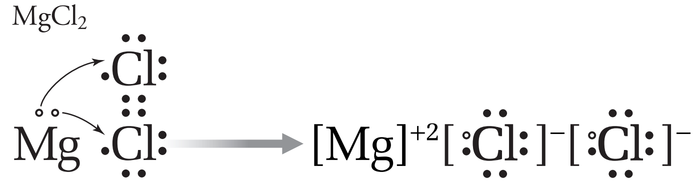
Plan Your Strategy
Apply the octet rule for each element in the bond. Determine the number of valence electrons in the outermost shell of each element in the bond. The electrons from the less electronegative element will be transferred to the element with the higher electronegativity to achieve a stable noble gas configuration.

(a) If there are too many electrons from the lower *EN* element than the stable structure of the higher *EN* element can take, then add more atoms from the element with the higher *EN* value until all their noble gas configurations are fulfilled.

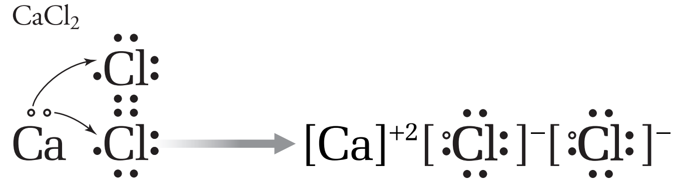
(b) Alternatively, if there are too few electrons of the less electronegative element available to fill the outer shell of the higher electronegative element, then add more atoms of the lower *EN* element until the noble gas configuration of the high *EN* element is fulfilled.

Act on Your Strategy

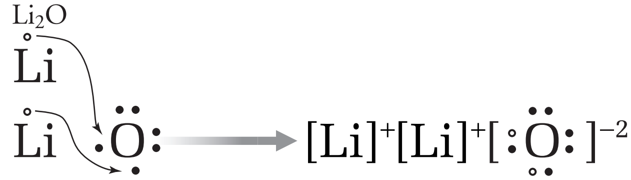
(a) $MgCl_2$



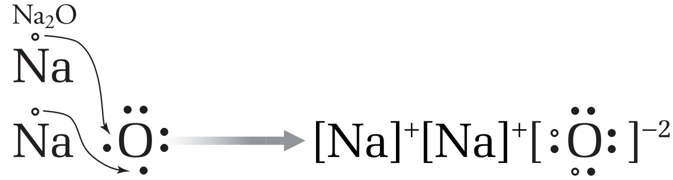
(b) $CaCl_2$



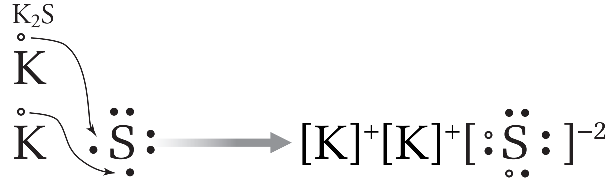
(c) Li_2O



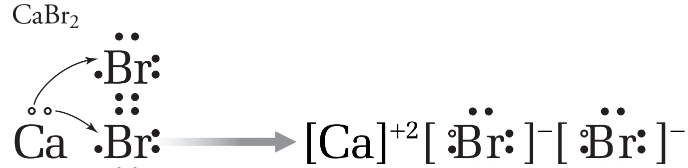
(d) Na_2O



(e) K_2S



(f) $CaBr_2$



Check Your Solution
In each case, after the bond has been made, the element(s) with the higher electronegativity should have a total of 8 electrons around it. It should have a net negative charge. The element(s) with the lower electronegativity loses all of its valence electrons and should carry a net positive charge. The total charge number(s) of the positively-charged element(s) should be the same as the total charge number(s) of the negatively-charged element(s) in each bond.

Solutions for Practice Problems

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6. Problem
Show the formation of a covalent bond between two atoms of each diatomic element.
(a) iodine **(b)** bromine **(c)** hydrogen **(d)** fluorine

What Is Required?
You have to draw the formation of the diatomic elements listed, using Lewis structures.

What Is Given?
The identity of the diatomic element is given.

Plan Your Strategy
Determine the number of valence electrons in the outermost shell of each element. Draw the Lewis structure twice (two atoms) for each element and fit them together, joined side by side by one pair of electrons. The final joined diagram should allow each element in the pair to have its own stable noble gas configuration of electrons around it, as per the octet rule.

Act on Your Strategy
(a) I_2

(b) Br_2

(c) H_2

(d) F_2

Check Your Solution
Draw a circle around each atom and its surrounding electrons in the bond, enclosing the joined electron pair as well. Each circled atom in the pair should have its own stable noble gas configuration of electrons around it, as per the octet rule.

7. Problem
Use Lewis structures to show the simplest way in which each pair of elements forms a covalent bond, according to the octet rule.
(a) hydrogen and oxygen **(b)** chlorine and oxygen
(c) carbon and hydrogen **(d)** iodine and hydrogen

(e) nitrogen and hydrogen

(f) hydrogen and rubidium

What Is Required?

You have to draw the final molecular structure that arises from the combination of the elements listed, using Lewis structures and the octet rule.

What Is Given?

The identity of the elements in the molecules is given.

Plan Your Strategy

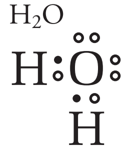
Step 1 Determine the number of valence electrons in the outermost shell of each element.

Step 2 Draw the Lewis structure only once first for each element and fit them together, joined side by side by one pair of electrons. Remember that in a Lewis structure with more than 4 electrons, the pairs of electrons must fill consecutively around the atom, not arbitrarily.

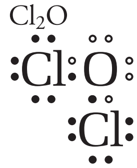
Step 3 Determine which one of the elements in the pair has unpaired (lone) electrons still in its outer shell. For each unpaired electron it has, add an atom of the other element to it until all the unpaired electrons have been paired.

Act on Your Strategy

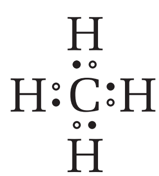
(a) H_2O



(b) Cl_2O

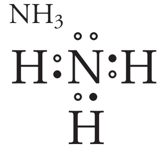


(c) CH_4



(d) HI

(e) NH_3



(f) RbH

Check Your Solution

Draw a circle around each atom and its surrounding electrons in the bond, enclosing the joined electron pair as well. Each circled atom in the bond should have its own stable noble gas configuration of electrons around it, as per the octet rule.

Solutions for Practice Problems

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8. Problem

One carbon atom is bonded to two sulfur atoms. Use a Lewis structure represent the bonds.

What Is Required?

You have to draw the Lewis configuration of the CS_2 molecule.

What Is Given?

The elements and their total number in the molecule are given.

Plan Your Strategy

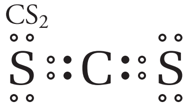
Step 1 Determine the number of valence electrons in the outermost shell of each element.

Step 2 Draw the Lewis structure of each atom and fit them together, joined side by side by one pair of electrons. Remember that in a Lewis structure with more than 4 electrons, the pairs of electrons must fill consecutively around the atom, not arbitrarily.

Step 3 For those elements still left with unpaired electrons, join them together again into double or triple bonds until all their electrons are filled.

Act on Your Strategy

CS_2



Check Your Solution

Draw a circle around each atom and its surrounding electrons in the bond, enclosing all joined electron pairs as well. Each circled atom in the bond should have its own stable noble gas configuration of electrons around it, as per the octet rule.

9. Problem

A molecule contains one hydrogen atom bonded to a carbon atom, which is bonded to a nitrogen atom. Use a Lewis structure represent the bonds.

What Is Required?

You have to draw the Lewis configuration of the HCN molecule.

What Is Given?

The elements and their total number in the molecule are given.

Plan Your Strategy

Step 1 Determine the number of valence electrons in the outermost shell of each element.

Step 2 Draw the Lewis structure of each atom and fit them together, joined side by side by one pair of electrons. Remember that in a Lewis structure with more than 4 electrons, the pairs of electrons must fill consecutively around the atom, not arbitrarily.

Step 3 For those elements still left with unpaired electrons, join them together again into double or triple bonds until all their electrons are filled.

Act on Your Strategy

HCN

Check Your Solution
Draw a circle around each atom and its surrounding electrons in the bond, enclosing all joined electron pairs as well. Each circled atom in the bond should have its own stable noble gas configuration of electrons around it, as per the octet rule.

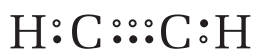
10. Problem
Two carbon atoms and two hydrogen atoms bond together, forming a molecule. Each atom achieves a full outer electron level. Use a Lewis structure to represent the bonds.

What Is Required?
You have to draw the Lewis configuration of the C_2H_2 molecule.

What Is Given?
The elements and their total number in the molecule are given.

Plan Your Strategy
Step 1 Determine the number of valence electrons in the outermost shell of each element.
Step 2 Draw the Lewis structure of each atom and fit them together, joined side by side by one pair of electrons. Remember that in a Lewis structure with more than 4 electrons, the pairs of electrons must fill consecutively around the atom, not arbitrarily.
Step 3 For those elements still left with unpaired electrons, join them together again into double or triple bonds until all their electrons are filled.

Act on Your Strategy
C_2H_2



Check Your Solution
Draw a circle around each atom and its surrounding electrons in the bond, enclosing all joined electron pairs as well. Each circled atom in the bond should have its own stable noble gas configuration of electrons around it, as per the octet rule.

Solutions for Practice Problems

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11. Problem
Predict whether each bond will be covalent, polar covalent, or ionic.

(a) C–F **(b)** O–N **(c)** Cl–Cl **(d)** Cu–O
(e) Si–H **(f)** Na–F **(g)** Fe–O **(h)** Mn–O

What Is Required?
You have to determine the electronegativity difference between the bonded atoms listed and decide if it is an ionic, covalent, or polar covalent bond.

What Is Given?
The identity of each element in the chemical bond is given.

Plan Your Strategy
Look up the electronegativity of each given element in Figure 3.6, student textbook page 71. For each bond combination, subtract the lower electronegativity value from the higher value to get the ΔEN for the bond. From Figure 3.8, student textbook page 73, you can see that a ΔEN value between 0.0 and 0.5 will likely be covalent, a value between 0.5 and 1.7 will likely be polar covalent, and a value between 1.7 and 3.3 will likely be ionic.

Act on Your Strategy

(a) C–F, $\Delta EN = EN\text{F} - EN\text{C} = 3.98 - 2.55 = 1.43$, polar covalent
(b) O–N, $\Delta EN = EN\text{O} - EN\text{N} = 3.44 - 3.04 = 0.40$, covalent
(c) Cl–Cl, $\Delta EN = EN\text{Cl} - EN\text{Cl} = 3.16 - 3.16 = 0.00$, covalent
(d) Cu–O, $\Delta EN = EN\text{O} - EN\text{Cu} = 3.44 - 1.90 = 1.54$, polar covalent
(e) Si–H, $\Delta EN = EN\text{H} - EN\text{Si} = 2.20 - 1.90 = 0.30$, covalent
(f) Na–F, $\Delta EN = EN\text{F} - EN\text{Na} = 3.98 - 0.93 = 3.05$, ionic
(g) Fe–O, $\Delta EN = EN\text{O} - EN\text{Fe} = 3.44 - 1.83 = 1.61$, polar covalent
(h) Mn–O, $\Delta EN = EN\text{O} - EN\text{Mn} = 3.44 - 1.55 = 1.89$, ionic

Check Your Solution

An ionic bond will fall in the ΔEN range of 1.7–3.3. A polar covalent bond will fall in the ΔEN range of 0.5–1.7. A covalent bond will fall in the ΔEN range of 0.0–0.5. Double check your electronegativity values from Figure 3.6. None of your answers should be negative values.

12. Problem

For each polar covalent bond in problem 11, indicate the locations of the partial charges.

What Is Required?

You have to show the partial charges in the polar covalent bonds you identified in problem 11 above.

What Is Given?

You have identified the polar covalent bonds in problem 11 above.

Plan Your Strategy

For each of the polar covalent bond pairs, the element with the most valence electrons around it (higher *EN* value) will be more partially negative. The element with the least number of valence electrons around (lower *EN* value) it will be more partially positive.

Act on Your Strategy

The polar covalent molecules identified in problem 11 are C–F, Cu–O, and Fe–O.

δ^+ C—F δ^-
δ^+ C—O δ^-
δ^+ Fe—O δ^-

Check Your Solution

Draw the Lewis configurations of each to bond pair. Circle each atom individually, enclosing all of its surrounding paired electrons as well as the shared pair in the bond. The partially positive side is the element with the least number of paired electrons. The partially negative side is the element with the most number of paired electrons.

13. Problem

Arrange the bonds in each set in order of increasing polarity. (A completely polarized bond is an ionic bond.)

(a) H–Cl, O–O, N–O, Na–Cl

(b) C–Cl, Mg–Cl, P–O, N–N

What Is Required?

You have to order each bond pair in each set according to increasing polarity.

What Is Given?

The identity of the element in each bond pair is given

Plan Your Strategy
First determine the ΔEN value for each pair. The greater the value, the greater is the bond polarity.

Act on Your Strategy

(a)	Order: O–O,	N–O,	H–Cl,	Na–Cl
	ΔEN: 0.00,	0.40,	0.96,	2.23
(b)	Order: N–N,	C–Cl,	P–O,	Mg–Cl
	ΔEN: 0.00,	0.61,	1.25,	1.85

Check Your Solution
Double check your electronegativity values from Figure 3.6. None of your answers should be negative values.

Solutions for Practice Problems

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14. Problem
Use the periodic table to predict the most common valences of the atoms in Groups 16 (VIA) and 17 (VIIA).

What Is Required?
You need to predict the most common valences of atoms in Groups VIA and VIIA.

What Is Given?
Group 16 (VIA) and Group 17 (VIIA) on the periodic table.

Plan Your Strategy
The group A number is usually a quick indication of the most common valence of the elements in that group.
IA indicates +1, IIA indicates +2, IIIA indicates +3, IVA indicates +4, VA indicates +/–3, VIA indicates –2, and VIIA indicates –1.

Act on Your Strategy

Group 16, common valence	–2
Group 17, common valence	–1

Check Your Solution
Pick any element that falls under Groups 16 and 17 and draw out its electronic configuration or Lewis structure. Your diagram will show the element under Group 16 requires 2 more electrons to reach a stable noble gas configuration. The element under group 17 requires 1 more electron to reach a stable noble gas configuration.

15. Problem
If you had to assign a valence to the noble gases, what would it be? Explain your answer.

What Is Required?
You have to assign a valence to the noble gases.

What Is Given?
The group is that of the noble gases.

Plan your Strategy
Recall that the valence is an oxidation number. That is, it is a charge reflecting the number of electrons available from, or needed by, an element to allow it to reach a stable noble gas configuration in a bond.

Act on Your Strategy
One would assign zero valence for noble gases, because they are already in a stable configuration according to the octet rule. Wrap the periodic table into a cylinder. The noble gases are between the alkali metals, + 1 and halogens −1, showing that the noble gases would neither lose nor gain electrons.

Check Your Solution
Pick any element that falls under Groups 18 and draw out its electronic configuration or Lewis structure. Your diagram will show the element has 8 valence electrons in its outer shell, that is, it is already in a stable configuration according to the octet rule.

Solutions for Practice Problems

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16. Problem
Write a balanced formula for a compound that contains sulfur and each of the following elements. Use a valence of −2 for sulfur.

(a) sodium **(b)** calcium **(c)** barium
(d) aluminum **(e)** rubidium **(f)** hydrogen

What Is Required?
You have to combine sulfur with the elements listed and give its balanced formula.

What Is Given?
The name of the element to combine with sulfur is given. The valence of sulfur is −2.

Plan Your Strategy
Step 1 Write the unbalanced formula, putting the metal element before the non-metal element. Sulfur is a non-metal, so it will appear at the back of all the formulas.
Step 2 Determine the valence of each element and put it on top of its name.
Step 3 Use arrows to connect the valence number (without its sign) of the metal element to sulfur, making it the subscript number of sulfur.
Step 4 Repeat step 3, this time connecting the valence number of sulfur to the name of the metal.
Step 5 Check the subscripts and remove any that are "1." If the subscripts on both are the same, also remove the subscripts.

Act on Your Strategy

(a) $\begin{matrix} +1 & -2 \\ Na & S \end{matrix} \longrightarrow Na_2S$

(b) $\begin{matrix} +2 & -2 \\ Ca & S \end{matrix} \longrightarrow CaS$

(c) $\begin{matrix} +2 & -2 \\ Ba & S \end{matrix} \longrightarrow BaS$

(d) $\begin{matrix} +3 & -2 \\ Al & S \end{matrix} \longrightarrow Al_2S_3$

(e) +1 −2

Rb S $\longrightarrow$ Rb_2S

(f) +1 −2

H S $\longrightarrow$ H_2S

Check Your Solution

Draw a Lewis structure of the molecule. Your diagram will show stable configurations of each atom according to the octet rule.

17. Problem

Write a balanced formula for a compound that contains calcium and each of the following elements.

(a) oxygen **(b)** sulfur **(c)** carbon
(d) bromine **(e)** phosphorus **(f)** fluorine

What Is Required?

You have to combine calcium with the elements listed and give its balanced formula.

What Is Given?

The name of the element to combine with calcium is given. The valence of calcium is +2.

Plan Your Strategy

Step 1 Write the unbalanced formula, putting the metal element before the non-metal element. In this case, calcium is a metal, so it will appear in the front of all the formulas.
Step 2 Determine the valence of each element and put it on top of its name.
Step 3 Use arrows to connect the valence number (without its sign) of calcium to the name of the non-metal element, making it the subscript number of the non-metal element.
Step 4 Repeat step 3, this time connecting the valence number of the non-metal element to calcium.
Step 5 Check the subscripts and remove any that are "1." If the subscripts on both are the same, also remove the subscripts.

Act on Your Strategy

(a) +2 −2

Ca O $\longrightarrow$ Ca_2O_2 $\longrightarrow$ CaO

(b) +2 −2

Ca S $\longrightarrow$ Ca_2S_2 $\longrightarrow$ CaS

(c) +2 −4

Ca C $\longrightarrow$ Ca_4C_2 $\longrightarrow$ Ca_2C

[Note: Calcium carbiole used to produce acetylene has a formula, CaC2. This is misnamed and should be called calcium acetylide.

(d) +2 −1

Ca Br $\longrightarrow$ $CaBr_2$ $\longrightarrow$ $CaBr_2$

(e) +2 −3

Ca P $\longrightarrow$ Ca_3P_2 $\longrightarrow$ Ca_3P_2

(f) +2 −1

Ca F $\longrightarrow$ CaF_2 $\longrightarrow$ CaF_2

Check Your Solution
Draw a Lewis structure of the molecule. Your diagram will show stable configurations of each atom according to the octet rule.

Solutions for Practice Problems

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8. Problem
Use the information in Table 3.4 to write a chemical formula for a compound that contains sodium and each of the following polyatomic ions.

(a) nitrate **(b)** phosphate **(c)** sulfite
(d) acetate **(e)** thiosulfate **(f)** carbonate

What Is Required?
You have to combine sodium with the polyatomic ions listed and give its balanced formula.

What Is Given?
The name of the polyatomic ion to combine with sodium is given. From Table 3.4, the valence of sodium is +1.

Plan Your Strategy
Step 1 Write the unbalanced formula, putting the sodium metal before the polyatomic ion. Place a bracket around the polyatomic ion.
Step 2 Determine the valence of each component and put it on top of its name.
Step 3 Use arrows to connect the valence number (without its sign) of sodium to the name of the polyatomic ion, making it the subscript number of the polyatomic ion.
Step 4 Repeat step 3, this time connecting the valence number of the polyatomic ion to sodium.
Step 5 Check the subscripts and remove any that are "1," but outside a polyatomic ion bracket only. If the subscripts outside the polyatomic ion bracket and of sodium's are the same, remove them as well.

Act on Your Strategy

(a) +1 −1

Na (NO_3) $\longrightarrow$ $NaNO_3$

(b) +1 −3

Na (PO_4) $\longrightarrow$ Na_3PO_4

(c) $+1$ -2

Na (SO_3) $\longrightarrow$ Na_2SO_3

(d) $+1$ -1

Na (CH_3COO) $\longrightarrow$ $NaCH_3COO$

(e) $+1$ -2

Na (S_2O_3) $\longrightarrow$ $Na_2S_2O_3$

(f) $+1$ -2

Na (CO_3) $\longrightarrow$ Na_2CO_3

Check Your Solution

Draw a Lewis structure of the molecule. Your diagram will show stable configurations of each component of the molecule according to the octet rule.

19. Problem

Repeat question 18 using magnesium instead of sodium.

What Is Required?

You have to combine magnesium with the polyatomic ions listed in problem 18 and give its balanced formula.

What Is Given?

The name of the polyatomic ion to combine with magnesium is given. From Table 3.4, the valence of magnesium is +2.

Plan Your Strategy

Step 1 Write the unbalanced formula, putting the magnesium metal before the polyatomic ion. Place a bracket around the polyatomic ion.

Step 2 Determine the valence of each component and put it on top of its name.

Step 3 Use arrows to connect the valence number (without its sign) of magnesium to the name of the polyatomic ion, making it the subscript number of the polyatomic ion.

Step 4 Repeat step 3, this time connecting the valence number of the polyatomic ion to magnesium.

Step 5 Check the subscripts and remove any that are "1," but *outside* a polyatomic ion bracket only. If the subscripts outside the polyatomic ion bracket and of magnesium's are the same, remove them as well.

Act on Your Strategy

(a) $+2$ -1

Mg (NO_3) $\longrightarrow$ $Mg(NO_3)_2$

(b) $+2$ -3

Mg (PO_4) $\longrightarrow$ $Ma_3(PO_4)_2$

(c) $\overset{+2}{\text{Mg}}\ \ \overset{-2}{(SO_3)} \longrightarrow Mg_2(SO_3)_2 \longrightarrow MgSO_3$

(d) $\overset{+2}{\text{Mg}}\ \ \overset{-1}{(CH_3COO)} \longrightarrow Mg(CH_3COO)_2$

(e) $\overset{+2}{\text{Mg}}\ \ \overset{-2}{(S_2O_3)} \longrightarrow Mg_2(S_2O_3)_2 \longrightarrow MgS_2O_3$

(f) $\overset{+2}{\text{Mg}}\ \ \overset{-2}{(CO_3)} \longrightarrow Mg_2(CO_3)_2 \longrightarrow MgCO_3$

Check Your Solution
Draw a Lewis structure of the molecule. Your diagram will show stable configurations of each component of the molecule according to the octet rule.

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20. Problem
Use the IUPAC name for each compound.

(a) Al_2O_3 **(b)** $CaBr_2$ **(c)** Na_3P
(d) Cu_2S **(e)** MgN_3 **(f)** HgI_2

What Is Required?
You need to give the IUPAC name for the compounds listed.

What Is Given?
The chemical formula of each compound is given.

Plan Your Strategy
Step 1 Write the name of the metal first and the non-metal next.
Step 2 Change the suffix (ending) of the non-metal to *-ide*.
Step 3 Put the names together.

Act on Your Strategy
(a) Al_2O_3 = aluminum oxide
(b) $CaBr_2$ = calcium bromide
(c) Na_3P = sodium phosphide
d) Cu_2S = copper(I) sulphide
(e) Mg_3N_2 = magnesium nitride
(f) HgI_2 = mercury(II) iodide

Check Your Solution
Try finding the names in other chemistry reference books or on the Internet and see if the formulas and names are the same.

21. Problem
Write the formula of each compound.

(a) iron(II) sulfide
(b) stannous oxide
(c) chromium(II) oxide
(d) colbaltous chloride
(e) manganese(III) iodide
(f) zinc oxide

What Is Required?
You need to give the formula for the compounds listed.

What Is Given?
The IUPAC name of each compound is given.

Plan Your Strategy

Step 1 Write the symbol of the metal first and the non-metal next.

Step 2 Check the valence of the metal and non-metal and put it on top of its respective symbol.

Step 3 Using the arrow crossing method, cross the valence number of the metal to the non-metal symbol as its subscript. Cross the valence number of the non-metal to the metal symbol as its subscript.

Step 4 Check the subscripts and remove any that are "1." If the subscripts are the same, remove them as well.

Act on Your Strategy

(a) iron(II) sulfide

$$\overset{+2}{\mathrm{Fe}} \quad \overset{-2}{\mathrm{S}} \longrightarrow \mathrm{Fe_2S_2} \longrightarrow \mathrm{FeS}$$

(b) stannous oxide

$$\overset{+2}{\mathrm{Sn}} \quad \overset{-2}{\mathrm{O}} \longrightarrow \mathrm{Sn_2O_2} \longrightarrow \mathrm{SnO}$$

(c) chromium(II) oxide

$$\overset{+2}{\mathrm{Cr}} \quad \overset{-2}{\mathrm{O}} \longrightarrow \mathrm{Cr_2O_2} \longrightarrow \mathrm{CrO}$$

(d) cobaltous chloride

$$\overset{+2}{\mathrm{Co}} \quad \overset{-1}{\mathrm{Cl}} \longrightarrow \mathrm{CoCl_2}$$

(e) magnesium(II) iodide

$$\overset{+2}{\mathrm{Mg}} \quad \overset{-1}{\mathrm{I}} \longrightarrow \mathrm{MgI_2}$$

(f) zinc oxide

$$\overset{+2}{\mathrm{Zn}} \quad \overset{-2}{\mathrm{O}} \longrightarrow \mathrm{Zn_2O_2} \longrightarrow \mathrm{ZnO}$$

Check Your Solution
Draw a Lewis structure of the molecule. Your diagram will show stable configurations of each component of the molecule according to the octet rule. Also try finding the names in other chemistry reference books or on the Internet and see if the formulas and names are the same.

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22. Problem

Write the IUPAC name for each compound.

(a) H_2Se **(b)** HCl **(c)** HF
(d) LiH **(e)** CaH_2 **(f)** PH_3

What Is Required?

You need to give the IUPAC name for the compounds listed.

What Is Given?

The chemical formula of each compound is given.

Plan Your Strategy

Check the position of H in the formula.

(a) If H appears in front of the formula, the name starts with "hydrogen..." The non-metal beside it will end with the suffix *-ide*.

(b) If H appears at the end of the formula, the name ends with "hydride" and the metal before the hydrogen will bear its normal name. If there are subscripts to the hydrogen, use the numerical prefix to describe it, that is, *di-*, *tri-* etc.

Act on Your Strategy

(a) H_2Se = hydrogen selenide
(b) HCl = hydrogen chloride
(c) HF = hydrogen floride
(d) LiH = lithium hydride
(e) CaH_2 = calcium hydride
(f) PH_3 = phosphorus(III) hydride or phosphorus trihydride.

Check Your Solution

Try finding the names in other chemistry reference books or on the Internet and see if the formulas and names are the same.

Solutions for Practice Problems

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23. Problem

Write the IUPAC name for each compound.

(a) $(NH_4)_2SO_3$ **(b)** $Al(NO_2)_3$ **(c)** Li_2CO_3
(d) $Ni(OH)_2$ **(e)** $Ag_3(PO_4)$ **(f)** $Cu(CH_3COO)_2$

What Is Required?

You need to give the IUPAC name for the compounds listed.

What Is Given?

The chemical formula of each compound is given.

Plan Your Strategy

Step 1 Write the name of the polyatomic cation first and the polyatomic anion next.

Step 2 The name of the cation is usually the full name of a metal or the polyatomic cation (check the valence from the subscript outside the bracket).

Step 3 The name of the polyatomic anion depends on its valence (check the subscript outside the bracket for this) and as well as the number of oxygen atoms it has (in the case of an oxy-anion).

Step 4 Put the names together as you would for a binary compound.

Act on Your Strategy

(a) $(NH_4)_2SO_3$ = ammonium sulfite

(b) $Al(NO_2)_3$ = aluminum nitrite
(c) Li_2CO_3 = lithium carbonate
(d) $Ni(OH)_2$ = nickel(II) hydroxide
(e) $Ag_3(PO_4)$ = silver phosphate
(f) $Cu(CH_3COO)_2$ = copper(III) acetate

Check Your Solution
Try finding the names in other chemistry reference books or on the Internet and see if the formulas and names are the same.

24. Problem
Write the IUPAC name for each compound.
(a) SF_6 **(b)** N_2O_5 **(c)** PCl_5 **(d)** CF_4

What Is Required?
You need to give the IUPAC name for the compounds listed.

What Is Given?
The chemical formula of each compound is given.

Plan Your Strategy
In the case of naming compounds of two non-metals, a prefix is used to indicate the number of atoms of each element.
Step 1 Write the normal names of the two non-metals in the order they appear in the formula.
Step 2 Change the suffix of the second non-metal to *-ide*.
Step 3 Add the appropriate prefix (*di-, tri-*, etc.) to each non-metal, according to the subscript they each carry.
Step 4 Put the names together.

Act on Your Strategy
(a) SF_6 = sulfur hexafluoride
(b) N_2O_5 = dinitrogen pentoxide
(c) PCl_5 = phosphorus pentachloride
(d) CF_4 = carbon tetrafluoride

Check Your Solution
Try finding the names in other chemistry reference books or on the Internet and see if the formulas and names are the same.