UNIT (3) COMPOUNDS

Substances are either elements or compounds. In *unit 2* we studied elements, and in this unit we will study compounds.
A **compound** is a substance that consists of two or more *different* elements.
The elements in a compound are not just mixed together. Their atoms are bonded together in a specific way.
The forces that hold atoms together in a compound are called **chemical bonds**.
We will study ionic and covalent bonds.
An **ionic bond** involves the transfer of electrons from a metal to a nonmetal.
A **covalent bond** consists of a pair of electrons shared between two nonmetals.

Metals lose their valence electrons and nonmetals gain electrons to satisfy the octet rule. Electron sharing satisfies the octet rule.

3.1 The Octet Rule (Rule of 8)

In the formation of either ionic bond or covalent bond, atoms **lose, gain, or share** electrons to achieve an electron configuration identical to the noble gas nearest them in the periodic table. These noble gas configurations have eight electrons in their valence shells (except for helium, which has two electrons).

The **octet rule** states that atoms tend to combine in such a way that each has eight electrons in their valence shells identical to the noble gas nearest them in the periodic table.

3.2 Ions and the Octet Rule

As you studied in *unit 2*, atoms are neutral because they have equal numbers of electrons and protons. By losing or gaining one or more electrons, an atom can be converted into a charged particle called an **ion**.

The loss of electron(s) from a neutral atom gives a *positively* charged ion called **cation** (pronounced cat-ion).
The gain of electron(s) by a neutral atom gives a *negatively* charged ion called **anion** (pronounced an-ion).

For most of s block and p block elements, the charge on an ion can be predicted from the position of the element on the periodic table.
The metals (on the left-hand side of the table) lose electrons to form cations.
The Group IA (lose ONE electron), Group IIA (lose TWO electrons), and Group IIIA (lose THREE electrons).

The nonmetals (on the right-hand side of the table) gain electrons to form anions.
The Group VIIA (gain ONE electron), Group VIA (gain TWO electrons), and Group VA (gain THREE electrons).
Some transition metals and metals in Group IVA have variable charges (more than one positive ion). See Figure 3.1.

Figure 3.1
Some common ions and their locations on the periodic table are given. The table lists only the elements that you need to memorize.

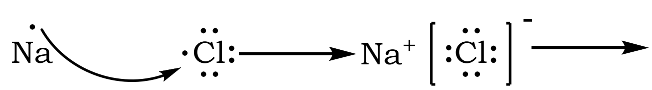
IA	IIA											IIIA	IVA	VA	VIA	VIIA	
H^+																	
Li^+	Be^{2+}													N^{3-}	O^{2-}	F^-	
Na^+	Mg^{2+}											Al^{3+}		P^{3-}	S^{2-}	Cl^-	
K^+	Ca^{2+}						Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+}	Cu^+ Cu^{2+}	Zn^{2+}					Br^-	
Rb^+	Sr^{2+}									Ag^+			Sn^{2+} Sn^{4+}			I^-	
Cs^+	Ba^{2+}										Hg_2^{2+} Hg^{2+}		Pb^{2+} Pb^{4+}				

(Note: charges are written numbers first then sign).

3.3 Ionic Bond Formation

An ionic bond forms by transfer of electron(s) from the metals to the nonmetals.
The result is a formation of an **ionic compound**.
Lewis structures (electron-dot symbols) are helpful in visualizing the formation of ionic compounds.

Using Lewis symbols, the formation of the ionic compound NaCl from the elements sodium and chlorine can be shown as follows:



Formula (NaCl)

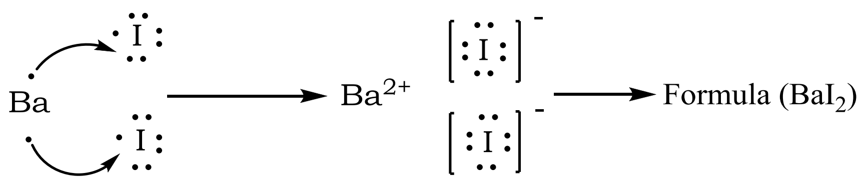
Sodium needed to lose one electron for octet formation (the neon electron configuration), chlorine needed to gain one electron for octet formation (the argon electron configuration). The electron transfer required 1:1 ratio of reacting atoms 1 Na to 1 Cl.

Worked Example 3-1

Use the electron-dot symbols to write the equation for the formation of the ionic compound formed between barium and iodine.

Solution

Barium has to lose two electrons for octet formation (the xenon electron configuration).
Iodine has to gain one electron for octet formation (the xenon electron configuration).
The transfer of two electrons from barium requires the acceptance of those two electrons by two iodine atoms.
The electron transfer requires 1:2 ratio of reacting atoms 1 Ba to 2 I.



Practice 3-1

Use the electron-dot symbols to write the equation for the formation of the ionic compound formed between aluminum and fluorine.

Answer

3.4 Writing Formulas for Ionic Compounds

Ionic compounds are electrically neutral. Therefore, when writing formulas, the cations (positive) and anions (negative) must combine to produce a net charge of zero. Formulas for ionic compounds are called *formula units*.

The correct combining ratio when Na^+ ions and Cl^- ions combine is: NaCl (one to one).
The correct combining ratio when Na^+ ions and O^{2-} ions combine is: Na_2O (two to one).
The correct combining ratio when Na^+ ions and P^{3-} ions combine is: Na_3P (three to one).

Worked Example 3-2

Write the formula for the ionic compound that is formed when each of the following pairs of ions interact:

a) K^+ and S^{2-}
b) Mg^{2+} and O^{2-}
c) Ca^{2+} and I^-
d) Li^+ and N^{3-}
e) Al^{3+} and S^{2-}

Solution

a) The cation has a charge of 1+ and anion has a charge of 2-. Thus two positive ions are required for each negative ion in a neutral formula unit.
The formula is **K_2S**.
b) The cation has a charge of 2+ and anion has a charge of 2-. The ratio is 1:1.
The formula is **MgO**.
c) The cation has a charge of 2+ and anion has a charge of 1-. Two negative ions are required for each positive ion. The formula is **CaI_2**.
d) The cation has a charge of 1+ and anion has a charge of 3-. Three positive ions are required for each negative ion. The formula is **Li_3N**.
e) The cation has a charge of 3+ and anion has a charge of 2-. Two positive ions are required for three negative ions. The formula is **Al_2S_3**.

K^+ and S^{2-}	K_2S
Mg^{2+} and O^{2-}	MgO
Ca^{2+} and I^-	CaI_2
Li^+ and N^{3-}	Li_3N
Al^{3+} and S^{2-}	Al_2S_3

3.5 Naming Ions

Names of cations and anions are formed by a system developed by the International Union of Pure and Applied Chemistry (IUPAC).

A) Names of Cations From Metals That Form Only One Type of Positive Ion:
Elements in Groups IA, IIA, and IIIA and some transition elements form only one type of cations. For these ions the name of the cation is the name of the metal followed by the word "ion":

Na^{+}	sodium ion	K^{+}	potassium ion	Mg^{2+}	magnesium ion
Al^{3+}	aluminum ion	Ag^{+}	silver ion	Zn^{2+}	zinc ion

B) Names of Cations From Metals That Form Two different Positive ions.
Metals in Group IVA and most transition metals form more than one type of cation, so the name of the cation must show its charge. For these ions the charge on the ion is given as a Roman numeral in parentheses right after (**with no space**) the metal name.

Sn^{2+}	tin(II)	Sn^{4+}	tin(IV)
Pb^{2+}	lead(II)	Pb^{4+}	lead(IV)
Cu^{+}	copper(I)	Cu^{2+}	copper(II)
Fe^{2+}	iron(II)	Fe^{3+}	iron(III)
Co^{2+}	cobalt(II)	Co^{3+}	cobalt(III)
Hg_2^{2+}	mercury(I)	Hg^{2+}	mercury(II)

C) Names of Anions:
Anions are named by replacing the ending of the element name with *–ide* followed by the word "ion":

F^{-}	fluoride ion	Cl^{-}	chloride ion	Br^{-}	bromide ion	I^{-}	iodide ion
O^{2-}	oxide ion	S^{2-}	sulfide ion	N^{3-}	nitride ion	P^{3-}	phosphide ion

D) Names of Polyatomic Ions:
A **polyatomic ion** is an ion that contains two or more elements. You must **memorize** the names and the formulas of the following polyatomic ions:

NH_4^{+}	ammonium	SO_3^{2-}	sulfite
CN^{-}	cyanide	SO_4^{2-}	sulfate
OH^{-}	hydroxide	HSO_3^{-}	hydrogen sulfite
$C_2H_3O_2^{-}$	acetate	HSO_4^{-}	hydrogen sulfate
CrO_4^{2-}	chromate	PO_3^{3-}	phosphite
$Cr_2O_7^{2-}$	dichromate	PO_4^{3-}	phosphate
MnO_4^{-}	permanganate	HPO_4^{2-}	hydrogen phosphate
NO_2^{-}	nitrite	ClO^{-}	hypochlorite
NO_3^{-}	nitrate	ClO_2^{-}	chlorite
CO_3^{2-}	carbonate	ClO_3^{-}	chlorate
HCO_3^{-}	hydrogen carbonate	ClO_4^{-}	perchlorate

The common name for HCO_3^-, HSO_3^-, and HSO_4^- are **bicarbonate**, **bisulfite**, and **bisulfate** respectively.

3.6 Naming Ionic Compounds

I: Binary ionic compounds from metals that form only one type of positive ion:

These compounds contain only two elements a metal ion and a nonmetal ion.
The chemical name is composed of the name of the metal followed by the name of the nonmetal, which has been modified with the suffix *–ide*.

Worked Example 3-3

Name the following binary ionic compounds:
$NaCl$ $MgBr_2$ AlP K_2S SrF_2 ZnI_2

Solution

$NaCl$	sodium chloride	K_2S	potassium sulfide
$MgBr_2$	magnesium bromide	SrF_2	strontium fluoride
AlP	aluminum phosphide	ZnI_2	zinc iodide

Practice 3-2

Name the following binary ionic compounds:
BaO Ca_3P_2 Sr_3N_2 Ag_2S $LiBr$ $NiCl_2$

Answer

II: Binary ionic compounds from metals that form two different positive ions:

To name these compounds we must include the charge on the cation as a Roman numeral in parentheses right after (**with no space**) the metal name, followed by the name of the anion.

Worked Example 3-4

Name the following binary ionic compounds:

$FeBr_3$ CoF_2 SnO PbI_4 HgS Cu_3P

Solution

$FeBr_3$	iron(III) bromide	PbI_4	lead(IV) iodide
CoF_2	cobalt(II) fluoride	HgS	mercury(II) sulfide
SnO	tin(II) oxide	Cu_3P	copper(I) phosphide

Practice 3-3

Name the following binary ionic compounds:

SnS_2 PbI_2 Hg_2O $CuCl_2$ FeN Co_2O_3

Answer

III: ionic compounds that include polyatomic ions

Naming these compounds is similar to naming binary compounds. The cation is named first, followed by the name for the negative polyatomic ion.

Worked Example 3-5

Name the following polyatomic ionic compounds:

$Ca(NO_3)_2$ $ZnSO_4$ NH_4CN

Li_3PO_4 Na_2CO_3 $Mg(HCO_3)_2$

Solution

$Ca(NO_3)_2$	calcium nitrate
$ZnSO_4$	zinc sulfate
NH_4CN	ammonium cyanide
Li_3PO_4	lithium phosphate
Na_2CO_3	sodium carbonate
$Mg(HCO_3)_2$	magnesium hydrogen carbonate

Practice 3-4

Name the following polyatomic ionic compounds:

$Sr(ClO_4)_2$ Na_2CrO_4 KH_2PO_4 Ag_2SO_4 $Co(OH)_3$ $Cu(C_2H_3O_2)_2$

Answer

3.7 Covalent Bond Formation and Lewis Structure

The Covalent Bond Model (Molecules)

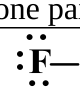
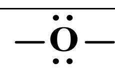
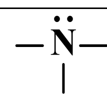
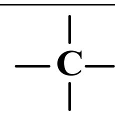
- Molecular compounds are compounds formed between two or more nonmetals.
- Molecular compounds are made up of discrete units called molecules.
- Within the molecule nonmetal atoms are held together by covalent bonds.

A **covalent bond** forms when electron pairs are shared between two nonmetals.
In general, the number of covalent bonds that a nonmetal atom forms is the same as the number of electrons it needs to have an octet.

Nonmetals may share all or some of their valence electrons. The shared electrons are called *bonding electrons* and the unshared electrons are called *lone pairs*. The bonding electrons are shown as dashes and lone pairs as dots.

According to the **octet rule** the total number of (bonds + lone pairs) should equal four, which is a total of 8 electrons.

The following table summarizes the typical number of bonds and lone pairs for the atoms of interest:

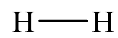
Atoms	Number of bonds	Number of lone pairs	Bonds and lone pairs
F, Cl, Br, I	1	3	
O, S	2	2	
N, P	3	1	
C, Si	4	0	
H	1	0	

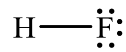
A molecular representation that shows both the connections among atoms and the locations of lone pairs are called **Lewis structure**.

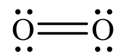
Worked Example 3-6

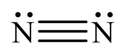
Draw a Lewis structure for each of the following: H_2, HF, O_2, and N_2.

Solution

H_2 H (1 bond) 

HF H (1 bond), F(1 bond, 3 lone pairs) 

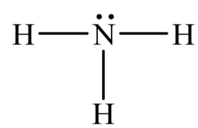
O_2 O (2 bonds, 2 lone pairs) 

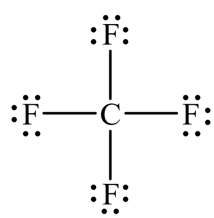
N_2 N (3 bonds, 1 lone pair) 

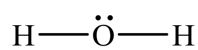
Worked Example 3-7

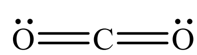
Draw Lewis structure for each of the following: NH_3, CF_4, H_2O, and CO_2.

Solution

NH_3 N (3 bonds, 1 lone pair), H(1 bond) 

CF_4 C (4 bonds), F (1 bond, 3 lone pairs) 

H_2O H (1 bond), O (2 bonds, 2 lone pairs) 

CO_2 C (4 bonds), O (2 bonds, 2 lone pairs) 

Notice that the central atom in Lewis structures is the atom that appears only once in the formula.

3.8 Naming Molecular Compounds

We will only consider binary (two element) molecular compounds.

The names of binary compounds (molecules) are written as two words.
First word: Full name of the first nonmetal in the formula; a Greek numerical prefix is used to show the number of atoms.
Second word: The stem of the name of the second nonmetal in the formula with the suffix-ide; a Greek prefix is used to show the number of atoms.

Greek prefixes

1 (mono-)	2 (di-)	3 (tri-)	4 (tetra-)	5 (penta-)
6 (hexa-)	7 (hepta-)	8 (octa-)	9 (nona-)	10 (deca-)

Worked Example 3-8

Name the following binary molecular compounds:
N_2O_5 CO_2 P_4S_3 XeF_6 ICl NH_3 I_4O_9 CO H_2O H_2O_2

Solution

N_2O_5	dinitrogen pentoxide	NH_3	nitrogen trihydride (ammonia)
CO_2	*carbon dioxide	I_4O_9	tetraiodine nonoxide
P_4S_3	tetraphosphorous trisulfide	CO	*carbon monoxide
XeF_6	*xenon hexafluoride	H_2O	dihydrogen monoxide (water)
ICl	iodine monochloride	H_2O_2	dihydrogen dioxide

*When only one atom of the first nonmetal is present, omit the initial prefix mono-. That is, the prefix mono- is never used with the cation.
Oxides of carbon are named by using “mono-” and “di-” to distinguish between the two oxides. Notice that we say *mon*oxide rather than *mono*oxide.

When the prefix ends in a or o and the element name begins with a or o, for the ease of pronunciation, the vowel of the prefix is dropped.

Practice 3-5

Name each of the following

S_2F_{10} SiI_4 P_2O_5 B_4Cl_4 P_4S_7 NBr_3 I_2Cl_6 SI_5

Answer

3.9 The Shape of Molecules: Molecular Geometry

We can predict the molecular geometry using the **V**alence **S**hell **E**lectron **P**air **R**epulsion (**VSEPR**) model:
According to VSEPR, "electron groups", as bonding pairs or lone pairs, stay as far apart as possible so that electron-electron repulsions are at minimized.

We will examine only five cases:

I) 2 electron groups (2 bonding pairs with no lone pairs)

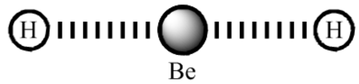
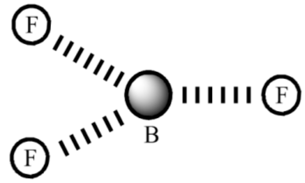
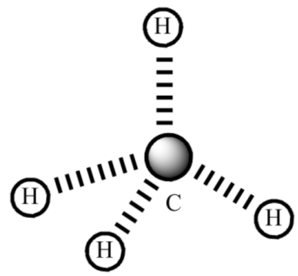
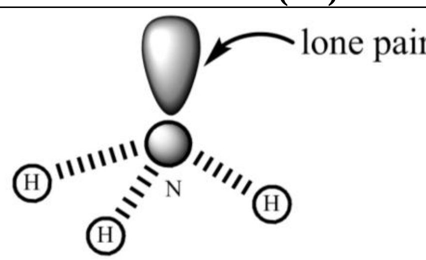
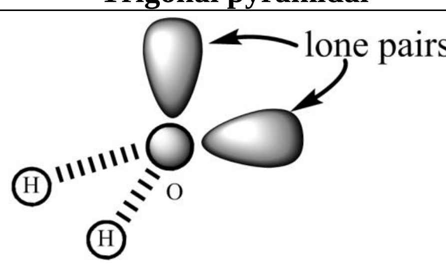
II) 3 electron groups (3 bonding pairs with no lone pairs)

III) 4 electron groups (4 bonding pairs with no lone pairs)

IV) 4 electron groups (3 bonding pairs and 1 lone pair)

V) 4 electron groups (2 bonding pairs and 2 lone pairs)

See the table:

molecules	bonded atoms	lone pairs	lone pairs + bonded	molecular shape
BeH_2	**2**	**0**	**2**	 **Linear**
BF_3	**3**	**0**	**3**	 **Trigonal planar**
CH_4	**4**	**0**	**4**	 **Tetrahedral (Td)**
NH_3	**3**	**1**	**4**	 **Trigonal pyramidal**
H_2O	**2**	**2**	**4**	 **Angular (bent)**

3.10 Electronegativity

Electronegativity is the measure of the ability of an atom to attract bonding electrons. Electronegativity displays a periodic trend. The element fluorine (F) is the most electronegative atom and is assigned 4.0, and all other elements are assigned values in relation to fluorine. Electronegativity generally increases from left to right across a row (period) of the periodic table and increases from bottom to top within a column (group). In general, the closer an element is to fluorine, the greater its electronegativity.

element	H	C	N	O	F
electronegativity	2.1	2.5	3.0	3.5	4.0
element		Si	P	S	Cl
electronegativity		1.8	2.2	2.5	3.0

3.11 Bond Polarity

The bonding electrons are shared equally between two nonmetal atoms of identical electronegativity. When electrons in a covalent bond are shared by atoms with different electronegativities, the atoms "pull" on the electrons with different strengths very similar to a tug-of-war. The slight winner will be the more electronegative atom (one closer to fluorine). This unequal sharing of electrons gives the bond a partially positive end (at the less electronegative atom) and a partially negative end (at the more electronegative atom).

Nonpolar covalent bond: bonding electrons are shared equally.
Examples of nonpolar bonds: H_2, N_2, O_2, F_2, Br_2, CH_4.

Polar covalent bond: bonding electrons are not shared equally due to differences in electronegativity. The unequal sharing means that the bonding electrons spend more time near the more electronegative atom and less time near the less electronegative atom. As a result the more electronegative atom carries a partial negative charge, represented by δ- (delta minus), and the less electronegative atom carries a partial positive charge, represented by δ+.

Examples of polar bonds for HF and H_2O molecules:

$$\overset{\delta+}{\mathrm{H}}-\overset{\delta-}{\mathrm{F}} \qquad \overset{\delta+}{\mathrm{H}}-\overset{\delta-}{\mathrm{O}}-\overset{\delta+}{\mathrm{H}}$$

The distinctions between nonpolar covalent, polar covalent, and ionic bonds are not always clear. In general, we find guidance by examining the difference in electronegativity between the bonded atoms.

Nonpolar covalent: Electronegativity difference between 0 to 0.4

Polar covalent: Electronegativity difference between 0.5 to 1.9
Ionic bond: Electronegativity difference larger than 1.9

Worked Example 3-9

Determine whether each of the following covalent bonds is polar or nonpolar.

a) N – H bond b) S – O bond c) C – S bond

Solution

a) N – H polar bond An electronegativity difference is (3.0 - 2.1 = 0.9).
b) S – O polar bond An electronegativity difference is (3.5 - 2.5 = 1.0).
c) C – S nonpolar bond An electronegativity difference is (2.5 - 2.5 = 0).

3.12 Molecular Polarity

Molecules can exhibit polarity similar to bond polarity. A molecule may be nonpolar despite the presence polar bonds.
We must consider both the geometry of the molecule and the polarity of bonds to determine whether or not a molecule is polar.
If a molecule has **no** polar bonds, then the molecule is nonpolar (nonpolar bond / nonpolar molecule).
If a molecule has polar bond(s), then the molecule is polar only if the centers of positive and negative charges don't coincide.
In a polar molecule, one side has a partial positive charge and the other has a partial negative charge.

Worked Example 3-10

Determine whether H_2, HF, and CO_2 are polar molecules.

Solution

H_2 is a diatomic element; therefore the H – H bond is nonpolar and the H_2 molecule is **nonpolar**.

The linear HF molecule contains a polar bond (electronegativity difference of 1.9). The hydrogen side of molecule is positive and the fluorine side of the molecule is negative. **Thus, HF is a polar molecule.**

The CO_2 molecule contains polar bonds (electronegativity difference of 1.0). However, the linear shape of molecule causes the two polar bonds to oppose and cancel one another and the molecule is **nonpolar.**

Homework Problems

3.1 Use the electron-dot symbols to write the equation for the formation of the ionic compounds from each of the following pairs:

a. K and F
b. Na and O
c. Ca and P
d. Al and S

3.2 Write the formula for the ionic compound that is formed from each of the following pairs of ions:

a. Na^{+} and F^{-}
b. Ca^{2+} and OH^{-}
c. Pb^{4+} and O^{2-}
d. Zn^{2+} and $C_2H_3O_2^{-}$
e. NH_4^{+} and S^{2-}
f. Al^{3+} and HPO_4^{2-}
g. Mg^{2+} and CO_3^{2-}
h. Pb^{2+} and I^{-}
i. Cu^{2+} and N^{3-}
j. Fe^{3+} and SO_3^{2-}

3.3 Write chemical formula for the following compounds:

a. barium nitrate
b. strontium chlorate
c. ammonium phosphate
d. cobalt(II) sulfite
e. mercury(II) iodide
f. copper(I) cyanide
g. magnesium phosphide
h. potassium sulfide
i. zinc hydroxide
j. silver chromate
k. iron(III) oxide
l. lead(II) permanganate

3.4 Name each of the following compounds:

a. $KClO_4$
b. Co_3N_2
c. NiF_2
d. NH_4OH
e. $NaNO_2$
f. $Sn(C_2H_3O_2)_2$
g. $Ca(MnO_4)_2$
h. $FeCr_2O_7$
i. $CuHPO_4$
j. $Al(HCO_3)_3$
k. MgH_2
l. PbS

3.5 Name the following molecular compounds:

a. CI_4
b. P_2I_4
c. Br_3O_8
d. N_2O_3
e. BCl_3
f. N_2O_5
g. P_4O_6
h. O_2F_2
i. IF_7
j. $SiBr_4$
k. H_2S
l. P_4S_{10}

3.6 For the SCl_2 molecule:

a. draw the Lewis structure
b. use VSEPR to predict the shape of the molecule
c. are there any polar bonds? Explain.
d. is the molecule polar? Explain.