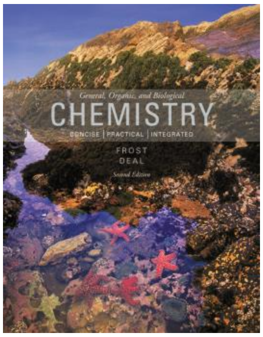

Lecture Presentation

Chapter 3

Compounds—Putting Particles Together

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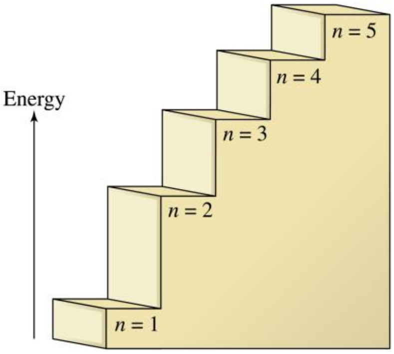
Outline

- 3.1 Electron Arrangements and the Octet Rule
- 3.2 In Search of an Octet, Part 1: Ion Formation
- 3.3 Ionic Compounds—Electron Give and Take
- 3.4 In Search of an Octet, Part 2: Covalent Bonding
- 3.5 The Mole: Counting Atoms and Compounds
- 3.6 Getting Covalent Compounds into Shape
- 3.7 Electronegativity and Molecular Polarity

3.1 Electron Arrangements and the Octet Rule

- The arrangement of electrons around an atom is one of the reasons elements form compounds.
- The exact location of any given electron outside the nucleus is difficult to determine, but most of them lie within an electron cloud.
- Because electrons are charged and are in constant motion, they possess energy.
- The electrons in an atom are found in distinct energy levels based on the amount of energy the electrons possess.
- The energy levels are an uneven staircase: the height of the steps decreases as you ascend.

3.1 Electron Arrangements and the Octet Rule



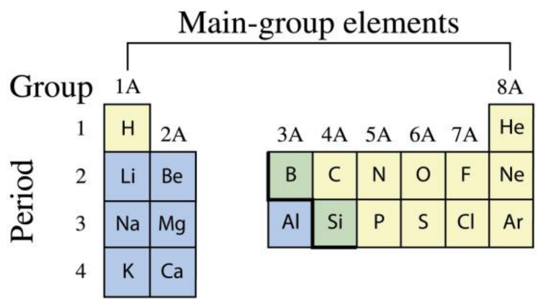
3.1 Electron Arrangements and the Octet Rule

- Electrons exist only at distinct energy levels (steps on the staircase), not in between.
- In general, electrons will occupy the lowest energy step first.
- The lowest energy level is found closest to the nucleus.
- Unlike a real staircase, the energy levels get closer together the farther you get from the nucleus.

3.1 Electron Arrangements and the Octet Rule

- The maximum number of electrons that can be found in any given energy level can be calculated by the formula $2n^2$, where n is the number of the energy level.
  - In the first energy level ($n = 1$), the maximum number of electrons present is 2.
  - In the second energy level ($n = 2$), the maximum number of electrons is 8.
  - In the third energy level ($n = 3$), 18, and so on.

3.1 Electron Arrangements and the Octet Rule



3.1 Electron Arrangements and the Octet Rule

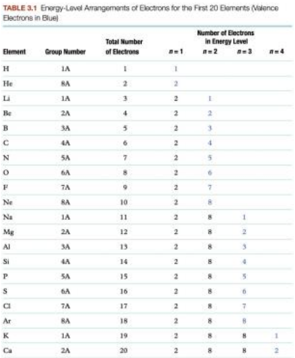


TABLE 3.1 Energy-Level Arrangements of Electrons for the First 20 Elements (Valence Electrons in Blue)

Element	Group Number	Total Number of Electrons	Number of Electrons in Energy Level $n = 1$	$n = 2$	$n = 3$	$n = 4$
H	1A	1	1			
He	8A	2	2			
Li	1A	3	2	1		
Be	2A	4	2	2		
B	3A	5	2	3		
C	4A	6	2	4		
N	5A	7	2	5		
O	6A	8	2	6		
F	7A	9	2	7		
Ne	8A	10	2	8		
Na	1A	11	2	8	1	
Mg	2A	12	2	8	2	
Al	3A	13	2	8	3	
Si	4A	14	2	8	4	
P	5A	15	2	8	5	
S	6A	16	2	8	6	
Cl	7A	17	2	8	7	
Ar	8A	18	2	8	8	
K	1A	19	2	8	8	1
Ca	2A	20	2	8	8	2

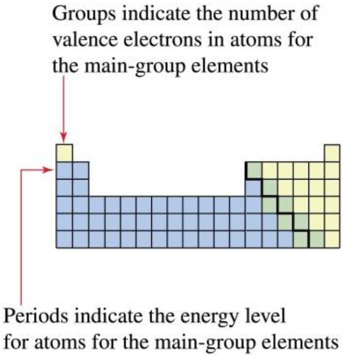
3.1 Electron Arrangements and the Octet Rule

- Examining the energy levels and number of electrons present in those energy levels provides insight into how the elements in a group or period are related to each other.
- Elements with the same number of electrons in their highest energy level are in the same group on the periodic table.
- Boron (B) and aluminum (Al) have three electrons in their highest energy level. Both are in Group 3A.

3.1 Electron Arrangements and the Octet Rule

- The highest energy level that contains electrons is the **valence shell**.
- The electrons residing in that energy level are called **valence electrons**.
- Valence electrons are farthest from the nucleus and are responsible for making compounds.
- Groups indicate the number of valence electrons in atoms for the main-group elements.
- Periods indicate the energy level for atoms.

3.1 Electron Arrangements and the Octet Rule



Groups indicate the number of valence electrons in atoms for the main-group elements

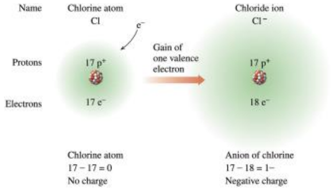
Periods indicate the energy level for atoms for the main-group elements

3.1 Electron Arrangements and the Octet Rule

- Elements in Group 8A, the **noble** or inert **gases**, are unreactive under all but extreme conditions.
- The noble gases (8A) have eight valence electrons (except for helium, which has only two valence electrons). Possessing eight valence electrons is a highly stable state for an atom.
- Most atoms will react with other atoms to achieve eight electrons in their valence shell.
- This is the **octet rule**.
- To achieve a valence octet, some atoms give away electrons, others will accept or take electrons, and others share electrons.
- These different modes of achieving a valence octet lead to different kinds of compounds.

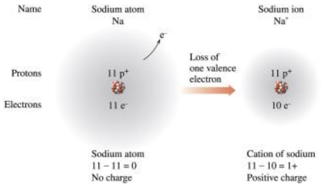
3.2 In Search of an Octet, Part 1: Ion Formation

- If an atom gains or loses electrons to achieve an octet, this newly formed particle would have an *unequal* number of protons and electrons and a net charge.
- These charged atoms are **ions**.
- Ions formed when atoms gain electrons and become negatively charged are called **anions**.

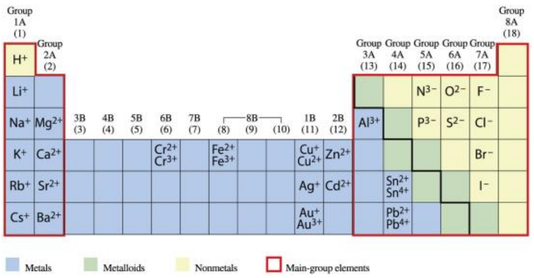


3.2 In Search of an Octet, Part 1: Ion Formation

- When an atom gains an electron to complete the octet in its valence shell, it becomes **isoelectronic** with its nearest noble gas.
- The ion is more stable than the atom because it has satisfied the octet rule.
- Ions formed when atoms lose electrons and become positively charged are called **cations**.



3.2 In Search of an Octet, Part 1: Ion Formation



3.2 In Search of an Octet, Part 1: Ion Formation

- **Predicting Charge on a Main-Group Ion**
  - Cations: Charge = group number
  - Anions: Charge = group number − 8
- Some transition metals form more than one ion: it is not possible to determine the charge of the ions formed by transition metals from their group number.
- **Polyatomic ions** consist of a group of nonmetals that form an ion. The charge is for the entire group of atoms.

3.2 In Search of an Octet, Part 1: Ion Formation

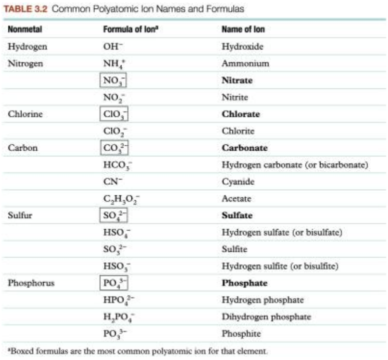


TABLE 3.2 Common Polyatomic Ion Names and Formulas

Nonmetal	Formula of Ion[a]	Name of Ion
Hydrogen	OH^-	Hydroxide
Nitrogen	NH_4^+	Ammonium
	$\boxed{NO_3^-}$	**Nitrate**
	NO_2^-	Nitrite
Chlorine	$\boxed{ClO_3^-}$	**Chlorate**
	ClO_2^-	Chlorite
Carbon	$\boxed{CO_3^{2-}}$	**Carbonate**
	HCO_3^-	Hydrogen carbonate (or bicarbonate)
	CN^-	Cyanide
	$C_2H_3O_2^-$	Acetate
Sulfur	$\boxed{SO_4^{2-}}$	**Sulfate**
	HSO_4^-	Hydrogen sulfate (or bisulfate)
	SO_3^{2-}	Sulfite
	HSO_3^-	Hydrogen sulfite (or bisulfite)
Phosphorus	$\boxed{PO_4^{3-}}$	**Phosphate**
	HPO_4^{2-}	Hydrogen phosphate
	$H_2PO_4^-$	Dihydrogen phosphate
	PO_3^{3-}	Phosphite

[a] Boxed formulas are the most common polyatomic ion for that element.

3.2 In Search of an Octet, Part 1: Ion Formation

Naming Ions

- For metal ions, add the word *ion* to the name of the metal.
- Transition metals that form more than one ion use a Roman numeral in parentheses following the name of the metal to indicate the charge on the ion.
- For nonmetals, the suffix *-ide* replaces the last few letters of the name of the element.
- Most polyatomic ions end in *-ate*. The ending *-ite* is used for related ions that have one less oxygen atom.

3.2 In Search of an Octet, Part 1: Ion Formation

Important Ions in the Body

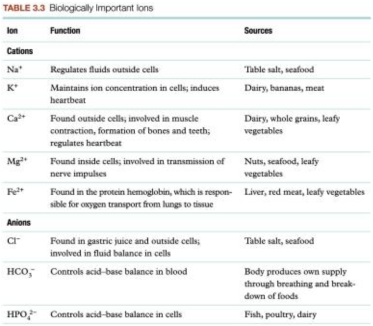


TABLE 3.3 Biologically Important Ions

Ion	Function	Sources
Cations		
Na^+	Regulates fluids outside cells	Table salt, seafood
K^+	Maintains ion concentration in cells; induces heartbeat	Dairy, bananas, meat
Ca^{2+}	Found outside cells; involved in muscle contraction, formation of bones and teeth; regulates heartbeat	Dairy, whole grains, leafy vegetables
Mg^{2+}	Found inside cells; involved in transmission of nerve impulses	Nuts, seafood, leafy vegetables
Fe^{2+}	Found in the protein hemoglobin, which is responsible for oxygen transport from lungs to tissue	Liver, red meat, leafy vegetables
Anions		
Cl^-	Found in gastric juice and outside cells; involved in fluid balance in cells	Table salt, seafood
HCO_3^-	Controls acid–base balance in blood	Body produces own supply through breathing and breakdown of foods
HPO_4^{2-}	Controls acid–base balance in cells	Fish, poultry, dairy

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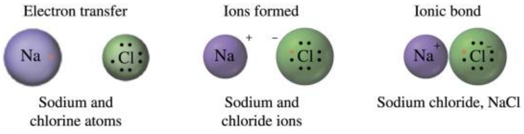
3.3 Ionic Compounds—Electron Give and Take

- When a metal and a nonmetal combine, electrons are transferred between the atoms, forming oppositely charged ions.
- The attraction of cation and anion is called an **ionic bond**.
- The result is an **ionic compound**.

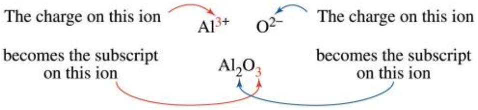


3.3 Ionic Compounds—Electron Give and Take

- The number of ions that combine to form an ionic compound is determined by the charge of the cation and the charge of the anion.
- Cations and anions combine so that the compound has a net charge of zero.
- When writing formulas for chemical compounds, we represent the number of each particle in the compound with a subscript.

3.3 Ionic Compounds—Electron Give and Take

- ***Steps to Check a Chemical Formula***
  - **Step 1:** Determine the charge of the ions from the periodic table if they are main-group elements.
  - **Step 2:** Combine the ions in a formula so that the total charge present sums to zero.
  - **Step 3:** Check your formula by examining the charges on the ions and subscripts in the formula.



The charge on this ion becomes the subscript on this ion

Al^{3+} O^{2-}

Al_2O_3

The charge on this ion becomes the subscript on this ion

- ***Steps to Predict Number and Charge in a Chemical Formula***
  - **Rule 1:** The formula of a ionic compound has no net charge (+ and – charges add up to zero).
  - **Rule 2:** The charge of the anion is known.

3.3 Ionic Compounds—Electron Give and Take

Naming Ionic Compounds

- To name an ionic compound, put the names of the two ions together, cation first.
- For transition metals, a Roman numeral is used to designate the charge.
- If a polyatomic ion is present, the name of the ion remains unchanged in the name of the compound.

3.4 In Search of an Octet, Part 2: Covalent Bonding

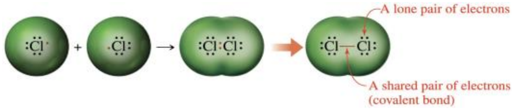
- Nonmetals combine by *sharing* valence electrons to achieve an octet.
- Sharing of electrons results in the formation of a **covalent bond**.
- The valence electrons in the bond belong to both of the atoms.
- When atoms share, the result is a **covalent compound**.
- The smallest (or fundamental) unit of a covalent compound is a **molecule**.

3.4 In Search of an Octet, Part 2: Covalent Bonding

- *The number of covalent bonds that an atom will form equals the number of electrons necessary to complete its octet*.
- The use of electron dot symbols was first developed by chemist G. N. Lewis to show how atoms share electrons in covalent bonding.
- The electron dot symbol for any atom consists of the elemental symbol plus a dot for each valence electron.

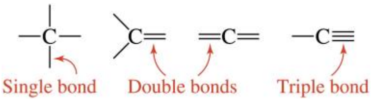
3.4 In Search of an Octet, Part 2: Covalent Bonding

- To complete its octet through a covalent bond, an atom must share its unpaired electrons with another atom.
- A shared pair of electrons, a **bonding pair** or simply a bond, is represented as a dash or line connecting the two electron dot symbols.
- Each of the atoms in the newly formed molecule now has an octet.



3.4 In Search of an Octet, Part 2: Covalent Bonding

- Carbon must form four covalent bonds to complete its octet.
  - It can form four **single bonds**, or
  - **double bonds** can form by sharing two electrons with one other atom, or
  - three electrons can be shared with one other atom to form a **triple bond**.
- In each case carbon has four bonding pairs—a complete octet—surrounding it.



3.4 In Search of an Octet, Part 2: Covalent Bonding

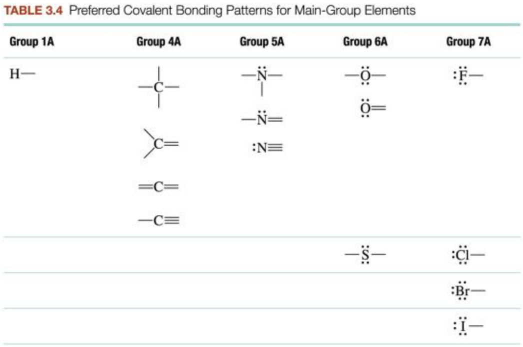


TABLE 3.4 Preferred Covalent Bonding Patterns for Main-Group Elements

Group 1A	Group 4A	Group 5A	Group 6A	Group 7A
H—	—C— >C═ ═C═ —C≡	—N̈— —N̈═ :N≡	—Ö— Ö═	:F̈—
			—S̈—	:C̈l—
				:B̈r—
				:Ï—

3.4 In Search of an Octet, Part 2: Covalent Bonding

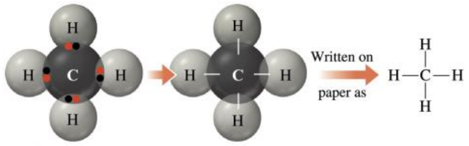
Formulas and Structures of Covalent Compounds

- The **molecular formula** for a covalent compound identifies *all* the components in a molecule.
- Glucose has the formula $C_6H_{12}O_6$. This tells us that a molecule of glucose has 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms.
- Molecular formulas do not reduce to the smallest whole number ratio.
- The formula does not show the structure. The electron dot symbols help us visualize the structure.

3.4 In Search of an Octet, Part 2: Covalent Bonding

Formulas and Structures of Covalent Compounds

- Methane, with a molecular formula of CH_4, is a covalent compound (all nonmetals).
- Carbon must make four bonds to complete its octet. Each hydrogen atom will make one bond.
- Connecting the atoms fulfills these requirements.
- This representation is called a **Lewis structure**.



3.4 In Search of an Octet, Part 2: Covalent Bonding

Naming Covalent Compounds

- **Binary compounds** are composed of only two elements and can be named by a three-step procedure:
  1. Name the first element in the formula.
  2. Name the second element in the formula and change the ending to *-ide*.
  3. Designate the number of each element present using one of the Greek prefixes.
- Indicating the number of each element present is important because nonmetals can combine with each other in multiple ways.
- Some compounds, such as water (H_2O) and ammonia (NH_3), are known by their traditional names.

3.4 In Search of an Octet, Part 2: Covalent Bonding

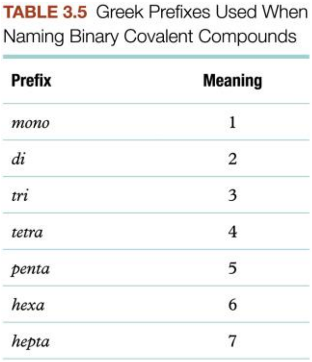
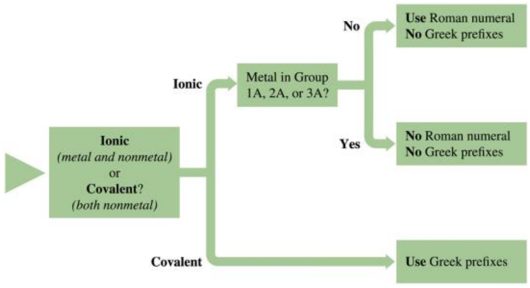


TABLE 3.5 Greek Prefixes Used When Naming Binary Covalent Compounds

Prefix	Meaning
mono	1
di	2
tri	3
tetra	4
penta	5
hexa	6
hepta	7

3.4 In Search of an Octet, Part 2: Covalent Bonding



3.5 The Mole: Counting Atoms and Compounds

- Chemists use a unit called the **mole** to relate the mass of an element in grams to the number of atoms it contains.
- The mole is a unit for counting atoms just as a *dozen* is a unit for counting things like eggs.
- The number of eggs in 1 dozen is 12. If each egg weighs 50 grams, we could determine how many eggs are present in a closed egg carton by weighing it.
- Similarly, 12 grams of carbon-12 contains a certain number of atoms, and this is called a mole.

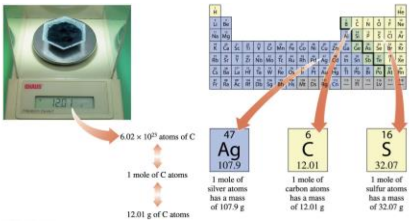
3.5 The Mole: Counting Atoms and Compounds

- One mole of any element has a **molar mass** in grams numerically equal to the atomic mass of that element.
- The molar mass tells us the mass of 1 mole of a substance and corresponds to the mass of a single atom in amu.
- The number of atoms present in one mole is about 602,000,000,000,000,000,000,000.
- The number of atoms in a mole can be better represented in scientific notation as 6.02×10^{23}.

3.5 The Mole: Counting Atoms and Compounds

- This is known as **Avogadro's number (N)** in honor of the Italian physicist Amedeo Avogadro.

  6.02×10^{23} atoms = 1 mole of atoms
- Just as a dozen eggs and a dozen bowling balls have different masses, a mole of silver atoms and a mole of carbon atoms do not have the same mass.



3.5 The Mole: Counting Atoms and Compounds

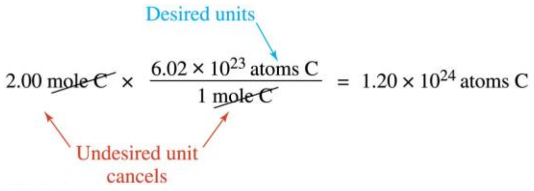
- Avogadro's number can be used as a conversion factor to convert between atoms and moles:

$$\frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mole}}$$

Steps to Converting between Units

Step 1: Examine the entire problem first, identifying the information you have and the information you are trying to determine.
Step 2: Establish the quantity you are trying to determine in the numerator so that your answer has those units.
Step 3: Determine conversion factors of the desired unit. Place the equivalent with the desired unit in the numerator and the undesired unit in the denominator so that the units cancel each other out, leaving the desired units.
Step 4: Set up the equation to cancel the given units (moles) and to provide the desired unit in the answer (atoms).
Step 5: Ask yourself if your answer makes sense. When solving a problem for the number of atoms, your answer should be a large number and have a large positive exponent.

3.5 The Mole: Counting Atoms and Compounds



Desired units

$$2.00 \cancel{\text{mole C}} \times \frac{6.02 \times 10^{23} \text{ atoms C}}{1 \cancel{\text{mole C}}} = 1.20 \times 10^{24} \text{ atoms C}$$

Undesired unit cancels

3.5 The Mole: Counting Atoms and Compounds

- The formula weight for a compound is the sum of the atomic masses.
- The molar mass of a compound is numerically equal to the formula weight with units of grams/mole.

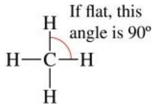
Finding the Number of Molecules in a Sample

- Step 1: Calculate the molar mass of the compound.
- Step 2: Apply conversion factors to reach the desired unit.

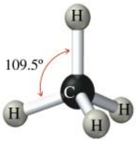
$$\text{grams} \times \frac{1 \text{ mole}}{\text{grams (F.W.)}} \times \frac{\text{molecules}}{\text{mole}} = \text{molecules}$$

3.6 Getting Covalent Compounds into Shape

- The Lewis structure for methane (CH_4) appears flat on paper, with 90° bond angles:



- If the electrons rearrange in three-dimensional space to get as far apart as possible, the molecule is **tetrahedral**.



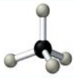
3.6 Getting Covalent Compounds into Shape

Determining the Shape of a Molecule

- To determine the shape (also called geometry) of a molecule using VSEPR, first determine the central atom. The central atom is represented with a capital *A* in the VSEPR form.
- Next, determine the number of charge clouds around the central atom. A single bond, a double bond, and a triple bond are all one charge cloud. Charge clouds in bonds are *bonding clouds* and are represented with a capital *B.*
- Charge clouds from lone pairs of electrons are called *nonbonding clouds* and are represented with a capital *N* in the VSEPR form.

3.6 Getting Covalent Compounds into Shape

TABLE 3.6 Predicting Molecular Shape Using VSEPR

VSEPR Form	Molecular Shape	Bond Angle	Example
AB_4	Tetrahedral	109.5°	
AB_3N	Pyramidal	< 109.5°	
AB_2N_2	Bent	< 109.5°	
AB_3	Trigonal planar	120°	
AB_2	Linear	180°	

3.6 Getting Covalent Compounds into Shape

Carbon

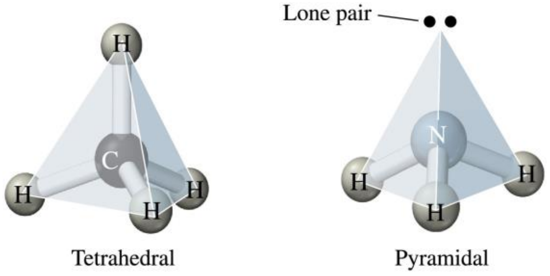
- If four atoms are bonded to a carbon, the shape is tetrahedral; if three atoms, the shape is trigonal planar; and if two atoms, the shape is linear.

TABLE 3.7 Predicting Molecular Shapes for Carbon

Molecular Shape	Number of Atoms Bonded Directly to Carbon
Tetrahedral	4
Trigonal Planar	3
Linear	2

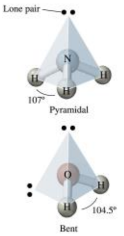
3.6 Getting Covalent Compounds into Shape

- Nonbonding pairs of electrons take up space but are invisible in the molecule's shape.



3.6 Getting Covalent Compounds into Shape

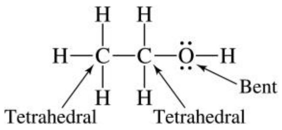
- **Nonbonded pairs** force bonded pairs closer together.
- A bonding pair of electrons is confined to the space directly between the two atoms being held. A nonbonded pair can move around more in the valence shell.
- The nonbonded pair repels the bonded pairs more than the bonded pairs repel each other, resulting in a smaller angle between the bonded pairs.



3.6 Getting Covalent Compounds into Shape

Molecular Shape of Larger Molecules

- Most molecules contain many carbon atoms with several charge clouds.
- Using VSEPR, the shape around any single atom that is bonded to at least two other atoms can be determined. Choose one or more atoms of interest in the molecule and determine the shape around those atoms.

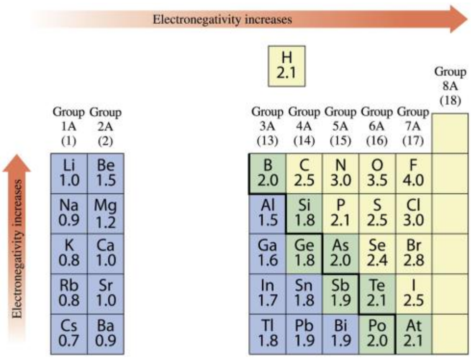


3.7 Electronegativity and Molecular Polarity

Electronegativity

- The ability of an atom to attract the bonding electrons of a covalent bond to itself is known as **electronegativity**.
- The element with the greatest electronegativity on the periodic table is fluorine.
- Electronegativity increases as you move closer to fluorine.

3.7 Electronegativity and Molecular Polarity



H 2.1

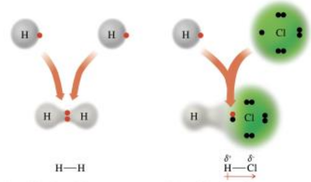
Group 1A (1)	Group 2A (2)	Group 3A (13)	Group 4A (14)	Group 5A (15)	Group 6A (16)	Group 7A (17)	Group 8A (18)
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	
Cs 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.1	

3.7 Electronegativity and Molecular Polarity

- Inequality occurs when two *different* atoms are involved in the covalent bond; the electrons are shared unequally.
- When two identical atoms share electrons to form a covalent bond, the electrons are shared equally.
- A covalent bond in which the electrons are not shared equally is a **polar covalent bond**.
- When the electrons are shared equally, the bond is a **nonpolar covalent bond**.

3.7 Electronegativity and Molecular Polarity

- Uneven sharing of electrons in a polar covalent bond is denoted by a lowercase Greek delta: δ.
- It can also be denoted by a dipole moment arrow. The arrow points toward the more electronegative element, and the plus on the arrow is at the partial positive end.



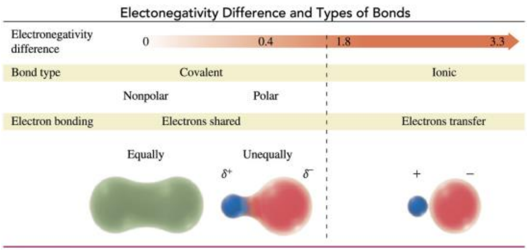
Equal sharing of electrons in a nonpolar covalent bond

Unequal sharing of electrons in a polar covalent bond (dipole)

3.7 Electronegativity and Molecular Polarity

- A covalent bond that does not share electrons equally has one end that is partially negative (δ^-) and one that is partially positive (δ^+).
- These bonds are *polar*.
- A bond that is “nonpolar” (sharing equally) has no ends or poles.
- The greater the electronegativity difference, the more polar the bond.

3.7 Electronegativity and Molecular Polarity



3.7 Electronegativity and Molecular Polarity

- Elements with an electronegativity difference of 1.8 or more will form an ionic bond from the *transfer* of electrons from the less electronegative element to the more electronegative element.
- Elements with an electronegativity difference of less than 1.8 will *share* electrons to form a covalent bond.
- The greater the difference in electronegativity of the two elements, the more polar the bond.

3.7 Electronegativity and Molecular Polarity

- It is possible to distinguish ionic bonds from covalent bonds by looking at the types of elements present (metals vs. nonmetals).
- To distinguish between nonpolar bonds, slightly polar bonds, and strongly polar bonds, remember two facts:
  1) Fluorine is the most electronegative of all the elements and
  2) electronegativities increase as you move toward fluorine on the periodic table.

3.7 Electronegativity and Molecular Polarity

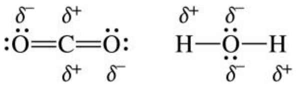
- To distinguish between nonpolar bonds, slightly polar bonds, and strongly polar bonds, remember two facts:
  1) Fluorine is the most electronegative of all the elements and
  2) electronegativities increase as you move toward fluorine on the periodic table.
- The *farther apart* two nonmetals are on the periodic table, the *greater* the polarity of the bond between them.
- One major exception to this rule is that a C—H bond is considered to be nonpolar.

3.7 Electronegativity and Molecular Polarity

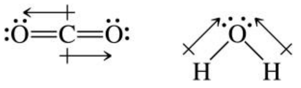
Molecular Polarity

- Molecules can be polar or nonpolar.
- A **polar molecule** has an uneven distribution of electrons over the molecule.
- A **nonpolar molecule** has an even distribution of electrons over the entire molecule.
- For molecules containing only two atoms connected by a covalent bond, the polarity of the molecule is the polarity of the bond.
- When a molecule is composed of three or more atoms connected by several bonds, we must consider both the electronegativity of the atoms involved *and* the shape of the molecule.

3.7 Electronegativity and Molecular Polarity

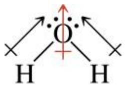


- Each molecule has two polar bonds, but carbon dioxide is a nonpolar molecule, whereas water is a polar molecule.
- Carbon dioxide is a linear molecule, but water is bent.



3.7 Electronegativity and Molecular Polarity

- In CO_2, the bonding electrons are being pulled equally toward the oxygen atoms, but in opposite directions. The dipoles cancel each other and the electrons are spread out equally over the entire molecule.
- Because of the bent shape of the water molecule, the bond dipoles add together to give an uneven distribution of electrons. This results in a molecular dipole for water.



- The interaction of electronegativity and molecular shape determines whether a molecule is polar or nonpolar.

Chapter Three Summary

- **3.1 Electron Arrangements and the Octet Rule**
  - Electrons have distinct energy levels, designated as n, where n = 1, 2, 3, and so on.
  - The electrons in the outermost energy level or shell are valence electrons.
  - Atoms react in order to become more stable.
  - The noble gases (Group 8A) are stable atoms: each has an octet of valence electrons in its valence shell.
  - An octet is eight electrons for all elements except hydrogen and helium, where a full valence shell is two electrons.

Chapter Three Summary

- **3.2 In Search of an Octet, Part 1: Ion Formation**
  - Ions gain or lose electrons to attain a valence octet.
  - An atom that gains electrons becomes a negatively charged anion, and an atom that loses electrons becomes a positively charged cation.
  - Nonmetal atoms form anions, and metal atoms form cations.
  - For main-group elements, the atom’s position on the periodic table determines its charge.
  - Polyatomic ions are groups of nonmetals that together have an ionic charge.
  - Cations have the same name as the element with the word *ion* added to the end.
  - Anions are named by changing the end of the element name to *-ide*.
  - Transition metals that form more than one cation use a Roman numeral in parentheses after the name of the metal to designate the charge.

Chapter Three Summary

- **3.3 Ionic Compounds—Give and Take**
  - An ionic compound is formed when an ionic bond is created between a metal and nonmetal.
  - Ionic bonds are formed by the attraction between an anion and a cation.
  - Anions and cations are formed when atoms gain or lose electrons.
  - Ionic compounds are neutral compounds. The total charge (number of ions times the charge of each) of the cations must equal the total charge of the anions in the compound.
  - To name an ionic compound, first name the cation (dropping the word *ion*) and then name the anion. If a transition metal is present, a Roman numeral is placed in parentheses following the name of the metal to designate its charge.
  - In writing ionic formulas, the metal ion always goes first. If more than one polyatomic ion is present, the number of ions present is represented outside a set of parentheses.

Chapter Three Summary

- **3.4 In Search of an Octet, Part 2: Covalent Bonding**
  - Covalent compounds form when nonmetals share electrons to achieve valence octets.
  - Sharing of electrons between two atoms is a covalent bond.
  - The number of covalent bonds that an atom will form is determined by the number of electrons necessary to complete its valence octet.
  - The molecular formula of a covalent compound gives the number of each type of atom in a molecule; the Lewis structure shows how those atoms are bonded together.
  - Covalent compounds containing only two elements are named by using the first element name, changing the second element ending to *-ide*, and using Greek prefixes to indicate the number of atoms of each in the compound.

Chapter Three Summary

- **3.5 The Mole: Counting Atoms and Compounds**
  - The molar mass relates the mass of a substance to a counting unit called the mole.
  - Avogadro’s number is a measure of the number of particles in a substance: 6.02 x 10^{23} per one mole.
  - It is possible to calculate a substance’s molar mass, which is equivalent to the atomic mass for an atom or the formal weight for a compound.
  - By using Avogadro’s number and molar mass as conversion factors, the number of moles or particles present in the mass of a compound can be determined.

Chapter Three Summary

- **3.6 Getting Covalent Compounds into Shape**
  - The arrangement of electrons determines the three-dimensional shape of covalent compounds.
  - A Lewis structure does not imply the three-dimensional shape of a molecule.
  - Valence shell electron pair repulsion (VSEPR) theory explains that the electrons around any given atom arrange themselves to get as far away from each other as possible.
  - For carbon atoms, this leads to shapes such as tetrahedral, trigonal planar, and linear. Each shape has a specific bond angle.
  - Nonbonding pairs reduce bond angles around that atom and alter the shape of the molecule.

Chapter Three Summary

- **3.7 Electronegativity and Molecular Polarity**
  - A molecule’s polarity is based on the individual bond polarities in the molecule.
  - Bond polarities are determined by comparing the electronegativities of the atoms in the bond.
  - Nonpolar covalent bonds form when atoms share electrons equally (electronegativities are equal).
  - Polar covalent bonds are formed when the sharing of electrons is unequal.
  - To determine polarity, consider the electronegativity of the atoms as well as the shape of the molecule.
  - The presence of polar covalent bonds in a molecule may result in a molecule that is polar, meaning that the electrons in a bond spend more time in some areas of the molecule than in others.
  - Polarity can be represented by a delta (δ) or a dipole moment arrow (+→).

Chapter Three Study Guide

- **3.1 Electron Arrangements and the Octet Rule**
  - Predict the number of valence electrons and energy levels for the main-group elements in the first four periods.
  - Recognize the unique stability associated with a valence shell containing eight electrons.
- **3.2 In Search of an Octet, Part 1: Ion Formation**
  - Predict the ionic charge of a main-group element using the periodic table.
  - Distinguish the name of an ion from its corresponding atom name.
  - Gain familiarity with polyatomic ions and their charges.

Chapter Three Study Guide (continued)

- **3.3 Ionic Compounds—Give and Take**
  - Predict the ionic charges present in an ionic compound.
  - Predict the ionic charge of a transition metal using the compound’s formula and the anionic charge.
  - Name ionic compounds given the formula.
  - Write the formula for ionic compounds given the name.

Chapter Three Study Guide (continued)

- **3.4 In Search of an Octet, Part 2: Covalent Bonding**
  - Distinguish ionic and covalent compounds.
  - Establish the relationship between the number of valence electrons present in the Period 1–3 nonmetals and Group 7A elements and the number of bonds that the atom typically makes in a molecule.
  - Draw Lewis structures for covalent compounds containing C, O, N, H, and the halogens (Group 7A).
  - Name binary covalent compounds given the formula.
  - Write the formula for a binary covalent compound given the name.

Chapter Three Study Guide (continued)

- **3.5 The Mole: Counting Atoms and Compounds**
  - Conceptualize the mole unit and Avogadro’s number.
  - Calculate molar mass for a compound.
  - Convert between the units of mole, number of particles, and gram.
- **3.6 Getting Covalent Compounds into Shape**
  - Predict the molecular shapes of small molecules using VSEPR.
  - Determine the effect of lone pair electrons on molecular shape.
- **3.7 Electronegativity and Molecular Polarity**
  - Predict covalent bond polarity based on electronegativity.
  - Predict molecular polarity from predicted bond polarities.