Chapter 5. Chemical Bonding: The Covalent Bond Model

Introduction to Inorganic Chemistry
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Online Tests on Following days

March 24, 2017: Test 1 (Chapters 1-3)
April 7, 2017 : Test 2 (Chapters 4-5)
April 28, 2017: Test 3 (Chapters 6,7 &8)
May 12, 2017 : Test 4 (Chapters 9, 10 &11)
May 15, 2017: Make Up Exam: Chapters 1-11)

.

Chapter 5 Chemical Bonding: The Covalent Bond Model

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Section 5.1

The Covalent Bond Model

Key Differences Between Ionic and Covalent Bonding

1. Ionic bonds form between a metal and nonmetal. Covalent bonds usually form between nonmetals.
2. Ionic bonds involve electron transfer. Covalent bonds involve electron sharing.
3. Ionic compounds do not contain discrete molecules. Covalent compounds has a molecule as its basic structural unit.

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Section 5.1

The Covalent Bond Model

Key Differences Between Ionic and Covalent Bonding

4. All ionic compounds are solids at room temperature. Covalent compounds are varied exits as gases and liquids as well.
5. Soluble ionic solids form aqueous solutions that conduct electricity. Soluble covalent compounds usually produce a non-conducting aqueous solution.

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Section 5.1

The Covalent Bond Model

Covalent Bond

- A chemical bond resulting from two nuclei attracting the same shared electrons.

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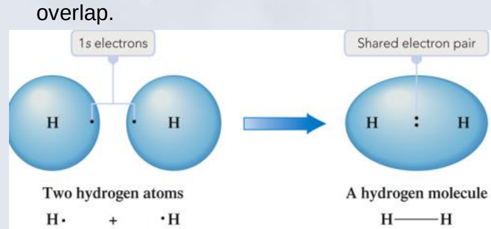
Section 5.1

The Covalent Bond Model

A Hydrogen Molecule

- Electron sharing can occur only when electron orbitals from two different atomic orbitals overlap.



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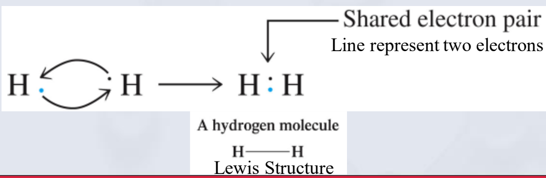
Section 5.1

The Covalent Bond Model

Lewis Notation

- The two shared electrons do double duty, helping each hydrogen atom achieve a helium noble-gas configuration.



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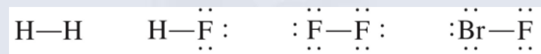
Section 5.2

Lewis Structures for Molecular Compounds

Bonding Electrons

- Pairs of valence electrons that are shared between atoms in a covalent bond.
- Represent the shared electron pairs with dashes.



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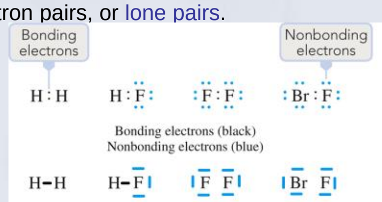
Section 5.2

Lewis Structures for Molecular Compounds

Nonbonding Electrons

- Pairs of valence electrons on an atom that are not involved in electron sharing.
- Also referred to as unshared electron pairs, lone electron pairs, or lone pairs.



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Section 5.3

Single, Double, and Triple Covalent Bonds

Single Covalent Bond

- A covalent bond in which two atoms share one pair of electrons.

H–H

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Section 5.3

Single, Double, and Triple Covalent Bonds

Double Covalent Bond

- A covalent bond in which two atoms share two pairs of electrons.

O=C=O

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Section 5.3

Single, Double, and Triple Covalent Bonds

Triple Covalent Bond

- A covalent bond in which two atoms share three pairs of electrons.

N º N

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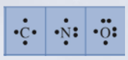
Section 5.4

Valence Electrons and Number of Covalent Bonds Formed

- There is a strong tendency for atoms of nonmetallic elements to form a specific number of covalent bonds.
- The number of bonds formed is equal to the number of electrons the nonmetallic atom must share to obtain an octet of electrons.



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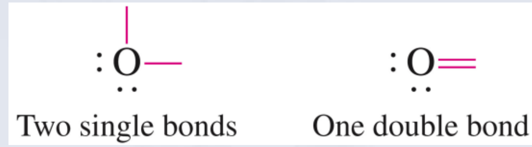
Section 5.4

Valence Electrons and Number of Covalent Bonds Formed

Oxygen (6 Valence Electrons, 2 Octet Vacancies)



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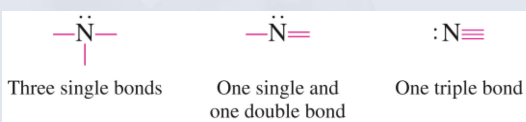
Section 5.4

Valence Electrons and Number of Covalent Bonds Formed

Nitrogen (5 Valence Electrons, 3 Octet Vacancies)



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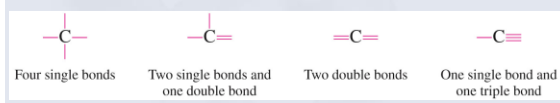
Section 5.4

Valence Electrons and Number of Covalent Bonds Formed

Carbon (4 Valence Electrons, 4 Octet Vacancies)



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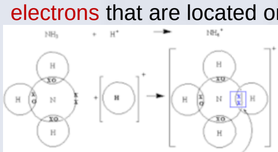
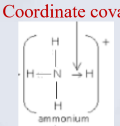
Section 5.5

Coordinate Covalent Bonds

- A covalent bond in which both electrons of a shared pair come from one of the two atoms involved in the bond is called Coordinate covalent bond.
- Enables an atom that has two or more vacancies in its valence shell to share a pair of nonbonding electrons that are located on another atom.



Coordinate covalent bond



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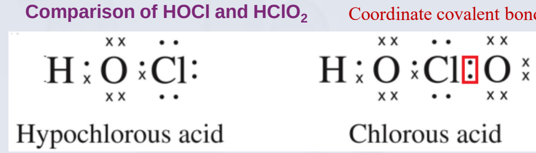
Section 5.5

Coordinate Covalent Bonds

Comparison of HOCl and $HClO_2$

Coordinate covalent bond



- In HOCl, all the bonds are "ordinary" covalent bonds.
- In $HClO_2$, the "new" chlorine-oxygen bond is a coordinate covalent bond.

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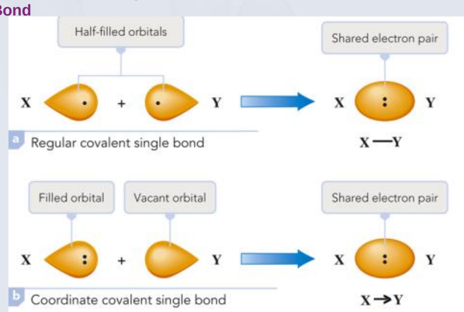
Section 5.5

Coordinate Covalent Bonds

Formation of a "Regular" Covalent Bond vs a Coordinate Covalent Bond



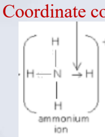
Section 5.5

Coordinate Covalent Bonds

- Atoms participating in coordinate covalent bonds generally do not form their normal number of covalent bonds.
- Once a coordinate covalent bond forms, it is indistinguishable from other covalent bonds in a molecule.

5 Valence Electrons, 3 Octet Vacancies

Coordinate covalent bond



Section 5.6

Systematic Procedures for Drawing Lewis Structures

Steps for Writing Lewis Structures

1. Calculate the total number of valence electrons available in the molecule by adding together the valence electron counts for all atoms in the molecule. (Use the periodic table.)

   Example: H_2O

   $2 (1\ e^-) + 6\ e^- = 8\ e^-$ total

Section 5.6

Systematic Procedures for Drawing Lewis Structures

Steps for Writing Lewis Structures

2. Write the chemical symbols of the atoms in the molecule in the order in which they are bonded to one another, and then place a single covalent bond, involving two electrons, between each pair of bonded atoms.
   - Determine central atom – usually atom that appears only once in the formula or bigger atom.

   Example: H_2O

   H–O–H

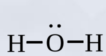
Section 5.6

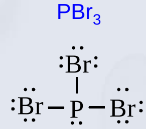
Systematic Procedures for Drawing Lewis Structures

Steps for Writing Lewis Structures

3. Add nonbonding electron pairs to the structure such that each atom bonded to the central atom has an octet of electrons. Remember that for hydrogen, an "octet" is only 2 ("duet") electrons.

   Examples: H_2O and PBr_3





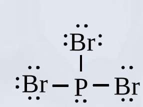
Section 5.6

Systematic Procedures for Drawing Lewis Structures

Steps for Writing Lewis Structuresa

4. Place any remaining electrons on the central atom of the structure.

   Example: PBr_3 = 26 valence e^-



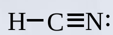
Section 5.6

Systematic Procedures for Drawing Lewis Structures

Steps for Writing Lewis Structures

5. If there are not enough electrons to give the central atom an octet, then use one or more pairs of nonbonding electrons on the terminal atoms bonded to the central atom to form double or triple bonds.

   Example: HCN



Section 5.6

Systematic Procedures for Drawing Lewis Structures

Steps for Writing Lewis Structures

6. Count the total number of electrons in the completed Lewis structure to make sure it is equal to the total number of valence electrons available for bonding, as calculated in Step 1. (Serves as a double-check.)

Section 5.6

Systematic Procedures for Drawing Lewis Structures

Concept Check

Draw a Lewis structure for each of the following molecules:

H_2

F_2

HF

Section 5.6

Systematic Procedures for Drawing Lewis Structures

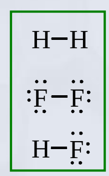
Concept Check

Draw a Lewis structure for each of the following molecules:

H_2

F_2

HF



Section 5.6

Systematic Procedures for Drawing Lewis Structures

Concept Check

Draw a Lewis structure for each of the following molecules:

NH_3

CO_2

CCl_4

Section 5.6

Systematic Procedures for Drawing Lewis Structures

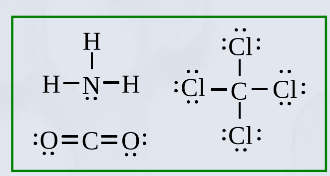
Concept Check

Draw a Lewis structure for each of the following molecules:

NH_3

CO_2

CCl_4



Section 5.7

Bonding in Compounds with Polyatomic Ions Present

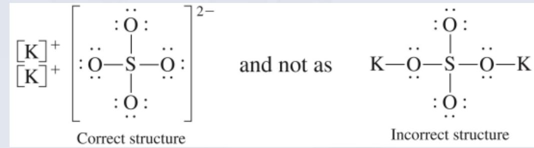
Ionic Compounds Containing Polyatomic Ions

- Covalent bonding exists within the polyatomic ion and ionic bonding exists between it and ions of opposite charge.

Section 5.7

Bonding in Compounds with Polyatomic Ions Present

Lewis Structure of Potassium Sulfate



- Polyatomic ion charge is not localized on a particular atom but rather is associated with the ion as a whole.
- It is customary to use brackets and show ionic charge outside the brackets.

Section 5.8

Molecular Geometry

- A description of the three-dimensional arrangement of atoms within a molecule.
- An important factor in determining the physical and chemical properties of a substance.

Section 5.8

Molecular Geometry

VSEPR Theory

- VSEPR: Valence Shell Electron-Pair Repulsion.
- A set of procedures for predicting the molecular geometry of a molecule using the information contained in the molecule's Lewis structure.
- The structure around a given atom is determined principally by minimizing electron pair repulsions.

Section 5.8

Molecular Geometry

VSEPR Electron Group

- A collection of valence electrons present in a localized region about the central atom in a molecule.
- The four electrons in a double bond or the six electrons in a triple bond are localized in the region between two bonded atoms in a manner similar to the two electrons of a single bond.

Section 5.8

Molecular Geometry

Steps to Apply the VSEPR Model

1. Draw the a Lewis structure for the molecule and identify the specific atom for which geometrical information is desired (usually central atom).

Section 5.8

Molecular Geometry

Steps to Apply the VSEPR Model

2. Determine the number of VSEPR electron groups present about the central atom.
   a. No distinction is made between bonding and nonbonding electron groups. Both are counted.
   b. Single, double, and triple bonds are all counted equally as "one electron group" because each takes up only one region of space about a central atom.

Section 5.8

Molecular Geometry

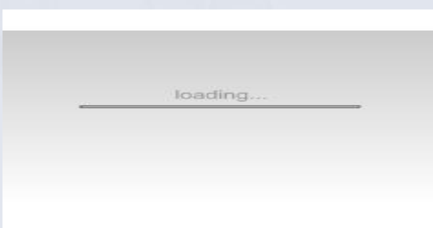
Steps to Apply the VSEPR Model

3. Predict the VSEPR electron group arrangement about the atom by assuming that the electron groups orient themselves in a manner that minimizes repulsions.

Section 5.8

Molecular Geometry

VSEPR: Two Electron Pairs



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Section 5.8

Molecular Geometry

VSEPR: Two Electron Groups

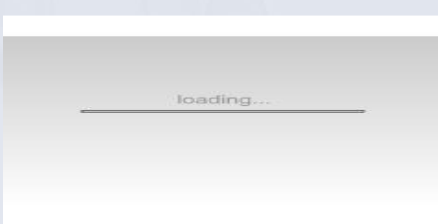
VSEPR electron group arrangement on central atom

2	2
linear	linear

Section 5.8

Molecular Geometry

VSEPR: Three Electron Pairs



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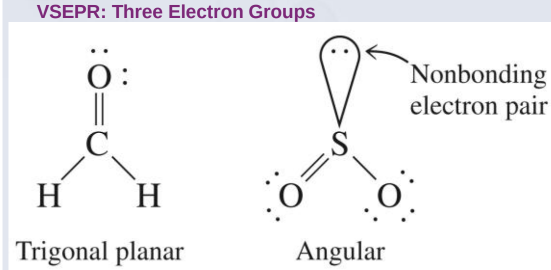
Section 5.8

Molecular Geometry

VSEPR: Three Electron Groups



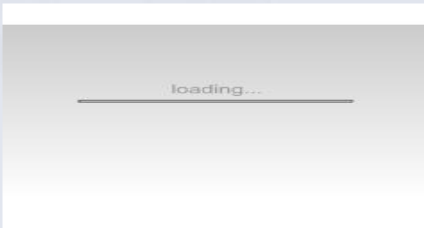
VSEPR electron group arrangement on central atom

3	3

Section 5.8

Molecular Geometry

VSEPR: Four Electron Pairs



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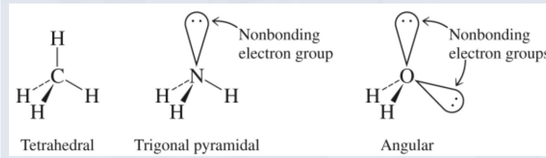
Section 5.8

Molecular Geometry

VSEPR: Four Electron Groups



VSEPR electron group arrangement on central atom
4 4 4

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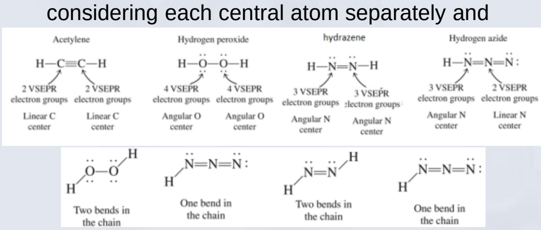
Section 5.8

Molecular Geometry

Molecules with More Than One Central Atom

- The molecular shape of molecules that contain more than one central atom can be obtained by considering each central atom separately and



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Section 5.8

Molecular Geometry

Molecular Geometries and Angles

Linear	180
Trigonal Planar	120
Tetrahedral	109.5
Triangular Bipyrimidal	120, 90
Octahedral	90
Angular	120
Angular	109.5
Trigonal Pyramidal	109.5

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Section 5.8

Molecular Geometry

Concept Check

Determine the shape for each of the following molecules, and include bond angles:

HCN
PH_3
O_3

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Section 5.8

Molecular Geometry

Concept Check

Determine the shape for each of the following molecules, and include bond angles:

HCN – linear, 180°
PH_3 – trigonal pyramid, 107°
O_3 – bent, 120°

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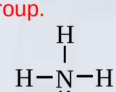
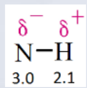
Section 5.9

Electronegativity

- The ability of an atom in a molecule to attract shared electrons to itself.
- A measure of the relative attraction that an atom has for the shared electrons in a bond.
- On the periodic table, electronegativity generally increases across a period and decreases down a group.

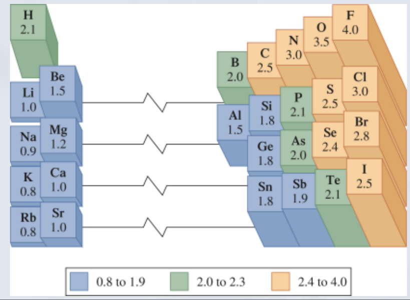




Section 5.9

Electronegativity

Pauling Electronegativity Values



Section 5.9

Electronegativity

Concept Check

What is the general trend for electronegativity across rows and down columns on the periodic table?

Explain the trend.

Section 5.9

Electronegativity

Concept Check

If lithium and fluorine react, which has more attraction for an electron? Why?

In a bond between fluorine and iodine, which has more attraction for an electron? Why?

Section 5.9

Electronegativity

Concept Check

If lithium and fluorine react, which has more attraction for an electron? Why?

In a bond between fluorine and iodine, which has more attraction for an electron? Why?

Section 5.10

Bond Polarity

Nonpolar Covalent Bond in Homo-nuclear

- A covalent bond in which there is equal sharing of electrons between two atoms.

H_2: H – H

N_2: N≡N

Section 5.10

Bond Polarity

Polar Covalent Bond in Hetero-nuclear

- A covalent bond in which there is unequal sharing of electrons between two atoms.

Electronegativity Cl = 3.0 H = 2.1 O = 2.5

H-Cl, C≡O

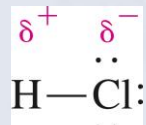
Electronegativity difference 0.9 0.4

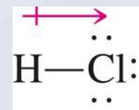
Section 5.10

Bond Polarity

Polar Covalent Bond

- It creates fractional positive and negative charges on atoms.
- Electrons spend more time near the more electronegative atom of the bond and less time near the less electronegative atom of the bond.





Section 5.10

Bond Polarity

Bond Polarity

- A measure of the degree of inequality in the sharing of electrons between two atoms in a chemical bond.
- The greater the electronegativity difference between the two bonded atoms, the greater the polarity of the bond.

Section 5.10

Bond Polarity

Bond Polarity

Bond Type	Electronegativity Difference
Nonpolar Covalent	0.4 or less
Polar Covalent	Greater than 0.4 to 1.5
Polar Covalent	Between 1.5 and 2.0 (between nonmetals)
Ionic	Between 1.5 and 2.0 (metal and nonmetal)
Ionic	Greater than 2.0

Section 5.10

Bond Polarity

Exercise

Arrange the following bonds from most to least polar:

a) N-F O-F C-F

b) C-F N-O Si-F

c) Cl-Cl B-Cl S-Cl

Section 5.10

Bond Polarity

Exercise

Arrange the following bonds from most to least polar:

a) N-F O-F C-F
a) C-F, N-F, O-F

b) C-F N-O Si-F
` b) Si-F, C-F, N-O

c) Cl-Cl B-Cl S-Cl
c) B-Cl, S-Cl, Cl-Cl

Section 5.10

Bond Polarity

Concept Check

Which of the following bonds would be the least polar yet still be considered polar covalent?

Mg-O   C-O   O-O   Si-O   N-O

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Section 5.10

Bond Polarity

Concept Check

Which of the following bonds would be the least polar yet still be considered polar covalent?

Mg-O   C-O   O-O   Si-O   **N-O**

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Section 5.10

Bond Polarity

Concept Check

Which of the following bonds would be the most polar without being considered ionic?

Mg-O   C-O   O-O   Si-O   N-O

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Section 5.10

Bond Polarity

Concept Check

Which of the following bonds would be the most polar without being considered ionic?

Mg-O   C-O   O-O   **Si-O**   N-O

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Section 5.11

Molecular Polarity

- A measure of the degree of inequality in the attraction of bonding electrons to various locations within a molecule.
- Polar molecule – a molecule in which there is an unsymmetrical distribution of electron charge.
- Nonpolar molecule – a molecule in which there is a symmetrical distribution of electron charge.

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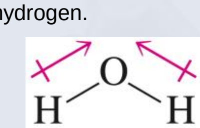
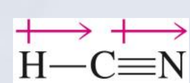
Section 5.11

Molecular Polarity

Polar Molecules: H_2O and HCN

- For H_2O, the bond polarities associated with the two hydrogen-oxygen bonds do not cancel one another because of the nonlinearity of the molecule.
- For HCN, nitrogen is more electronegative than hydrogen.





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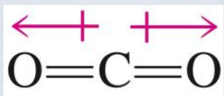
Section 5.11

Molecular Polarity

Nonpolar Molecule: CO_2

- The effects of the two polar bonds are canceled as a result of the oxygen atoms being arranged symmetrically around the carbon atom.



Section 5.11

Molecular Polarity

Concept Check

True or false:

A molecule that has polar bonds will always be polar.

- If true, explain why.
- If false, provide a counter-example.

Section 5.11

Molecular Polarity

Concept Check

True or false:

A molecule that has polar bonds will always be polar.

- If true, explain why.
- If false provide a counter-example.

Section 5.11

Molecular Polarity

Let's Think About It

- Draw the Lewis structure for SiO_2.
- Does SiO_2 contain polar bonds?
- Is the molecule polar or nonpolar overall? Why?

Section 5.11

Molecular Polarity

Concept Check

Which of the following molecules are polar?

F_2
HF
NH_3
SO_2
CCl_4

Section 5.11

Molecular Polarity

Concept Check

Which of the following molecules are polar?

F_2
HF
NH_3
SO_2
CCl_4

Section 5.12

Naming Binary Molecular Compounds

Binary Molecular Compound

- A molecular compound in which only two nonmetallic elements are present.
- The full name of the nonmetal of lower electronegativity is given first, followed by a separate word containing the stem of the name of the more electronegative nonmetal and the suffix *–ide*.
- Numerical prefixes, giving numbers of atoms, precede the names of both nonmetals.

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Binary Covalent Compounds

- Examples:

CO_2	Carbon dioxide
SF_6	Sulfur hexafluoride
N_2O_4	Dinitrogen tetroxide

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Common Numerical Prefixes

Table 5.1 Numerical Prefixes for the Numbers 1 Through 10

Number	Numerical Prefix	Example of Prefix Use*	
1	mono-	CO	carbon monoxide
2	di-	O_2F_2	dioxygen difluoride
3	tri-	NCl_3	nitrogen trichloride
4	tetra-	S_4N_4	tetrasulfur tetranitride
5	penta-	ClF_5	chlorine pentafluoride
6	hexa-	I_2F_6	diiodine hexafluoride
7	hepta-	IF_7	iodine heptafluoride
8	octa-	P_4O_8	tetraphosphorus octoxide
9	nona-	P_4S_9	tetraphosphorus nonasulfide
10	deca-	P_4Se_{10}	tetraphosphorus decaselenide

* When the prefix ends in "a" or "o" and the element name begins with "a" or "o," the final vowel of the prefix is usually dropped for ease of pronunciation. For example, "monoxide" is used instead of "monooxide," and "hexoxide" is used instead of "hexaoxide."

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Some Binary Molecular Compounds That Have Common Names

Table 5.2 Selected Binary Molecular Compounds That Have Common Names

Compound Formula	Accepted Common Name
H_2O	water
H_2O_2	hydrogen peroxide
NH_3	ammonia
N_2H_4	hydrazine
CH_4	methane
C_2H_6	ethane
PH_3	phosphine
AsH_3	arsine

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Exercise

Which of the following compounds is named incorrectly?

a) NO_2	nitrogen dioxide
b) P_2O_5	phosphorus pentoxide
c) PCl_3	phosphorus trichloride
d) SO_3	sulfur trioxide
e) ICl	iodine monochloride

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Exercise

Which of the following compounds is named incorrectly?

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