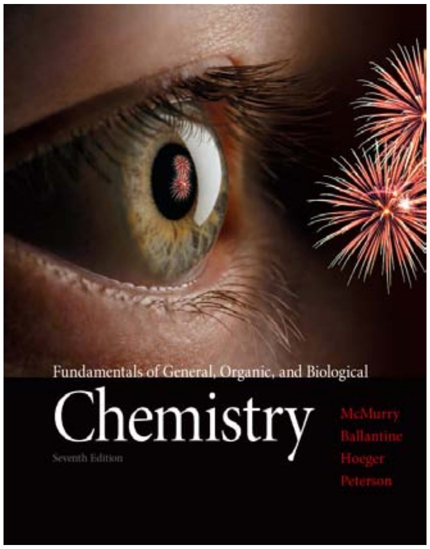
Chapter 4 Lecture



Fundamentals of General, Organic, and Biological Chemistry

7th Edition

McMurry, Ballantine, Hoeger, Peterson

Chapter Four

Molecular Compounds

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Outline

4.1 Covalent Bonds
4.2 Covalent Bonds and the Periodic Table
4.3 Multiple Covalent Bonds
4.4 Coordinate Covalent Bonds
4.5 Characteristics of Molecular Compounds
4.6 Molecular Formulas and Lewis Structures
4.7 Drawing Lewis Structures
4.8 The Shapes of Molecules
4.9 Polar Covalent Bonds and Electronegativity
4.10 Polar Molecules
4.11 Naming Binary Molecular Compounds

Goals

1. **What is a covalent bond?**

   Be able to describe the nature of covalent bonds and how they are formed.

2. **How does the octet rule apply to covalent bond formation?**

   Be able to use the octet rule to predict the numbers of covalent bonds formed by common main group elements.

3. **What are the major differences between ionic and molecular compounds?**

   Be able to compare the structures, compositions, and properties of ionic and molecular compounds.

Goals, *Continued*

4. **How are molecular compounds represented?**

   Be able to interpret molecular formulas and draw Lewis structures for molecules.

5. **What is the influence of valence-shell electrons on molecular shape?**

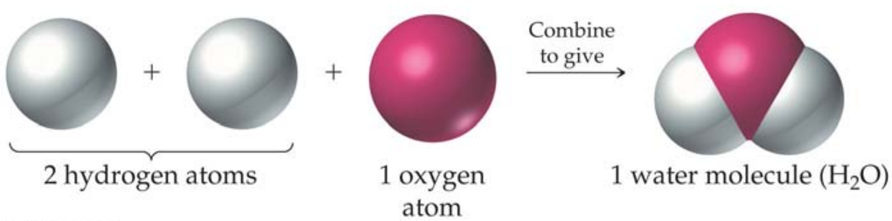
   Be able to use Lewis structures to predict molecular geometry.

6. **When are bonds and molecules polar?**

   Be able to use electronegativity and molecular geometry to predict bond and molecular polarity.

4.1 Covalent Bonds

- **Covalent bond**—A bond formed by sharing electrons between atoms
- **Molecule**—A group of atoms held together by covalent bonds



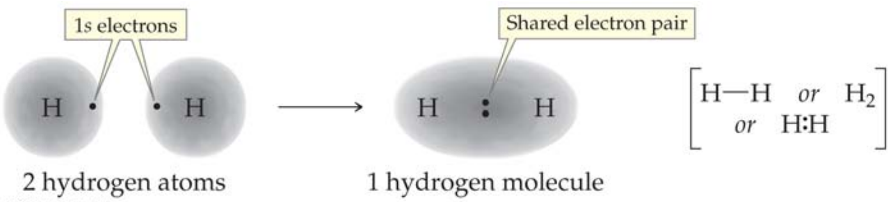
4.1 Covalent Bonds

- Main group elements undergo reactions that leave them with eight valence electrons (or two for hydrogen), so that they have a noble gas electron configuration.
- Nonmetals can achieve an electron octet by *sharing* an appropriate number of electrons in covalent bonds.

4.1 Covalent Bonds

Covalent bonding in hydrogen (H_2):

- Spherical 1*s* orbitals overlap to give an egg-shaped region.
- Two electrons between the nuclei, providing $1s^2$ configuration of helium.
- H-H, H:H and H_2 all represent a hydrogen molecule.

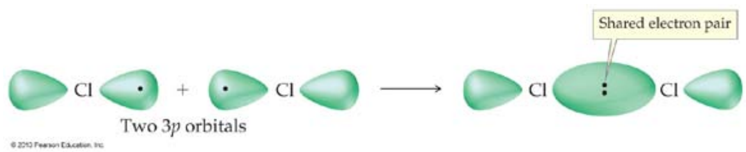


4.1 Covalent Bonds

- **Bond length** is the optimum distance between nuclei in a covalent bond.
  - If atoms are too far apart, attractive forces are small and no bond exists.
  - If atoms are too close, the repulsive interaction between nuclei is so strong that it pushes the atoms apart.
  - There is an optimum point where net attractive forces are maximized and where the molecule is most stable.
  - In the H_2 molecule, this optimum distance between nuclei is 74 pm.

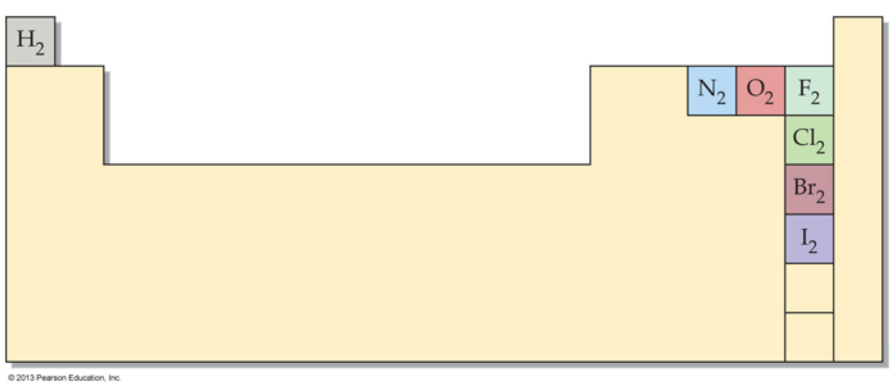
4.1 Covalent Bonds

- Chlorine also exists as a diatomic molecule due to the overlap of 3*p* orbitals.



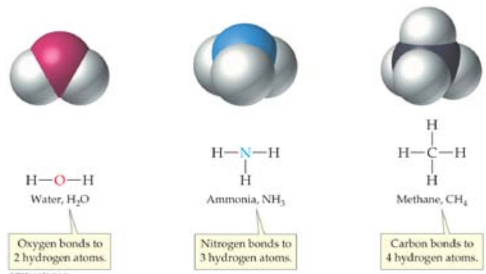
- There are seven diatomic elements: nitrogen, oxygen, hydrogen, flourine, chlorine, bromine, and iodine.

4.1 Covalent Bonds



4.2 Covalent Bonds and the Periodic Table

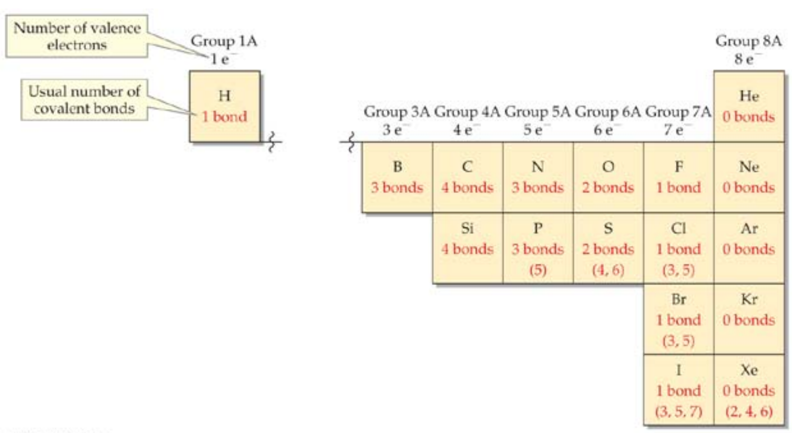
- A **molecular compound** is a compound that consists of molecules rather than ions.



In these compounds, each atom shares enough electrons to achieve a noble gas configuration or filled octet.

4.2 Covalent Bonds and the Periodic Table

FIGURE 4.3 Numbers of covalent bonds typically formed by main group elements to achieve octet configurations.

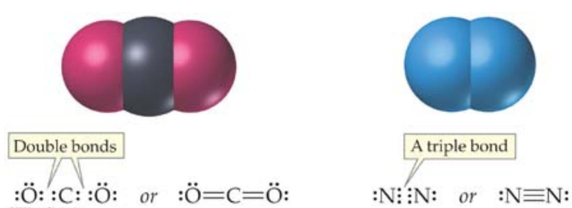


4.2 Covalent Bonds and the Periodic Table

- Exceptions to the Octet Rule:
  - Boron has only three electrons to share, and forms compounds with six electrons.
  - Elements in the third row and below in the periodic table have vacant *d* orbitals that can be used for bonding.

4.3 Multiple Covalent Bonds

The bonding in some molecules cannot be explained by the sharing of only two electrons between atoms.



The only way these molecules' atoms can have outer-shell electron octets is by sharing *more* than two electrons, resulting in the formation of *multiple* covalent bonds.

4.3 Multiple Covalent Bonds

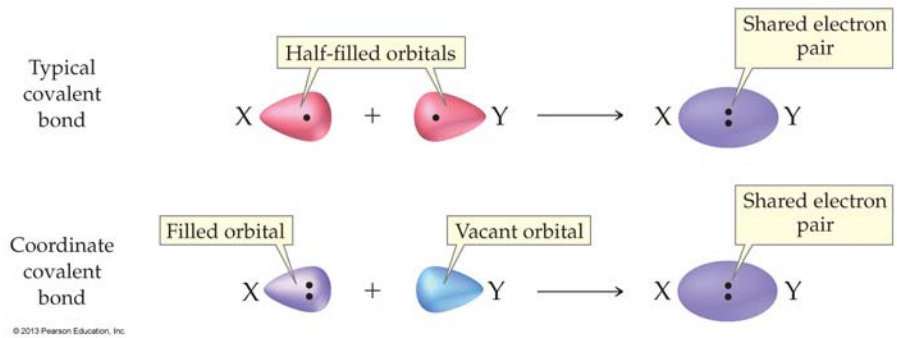
- **Single bond**—A covalent bond formed by sharing one electron pair.
  - Represented by a single line: H-H
- **Double bond**—A covalent bond formed by sharing two electron pairs.
  - Represented by a double line: O=O
- **Triple bond**—A covalent bond formed by sharing three electron pairs.
  - Represented by a triple line: N≡N

4.3 Multiple Covalent Bonds

- Carbon, nitrogen, and oxygen are the elements most often present in multiple bonds.
  - Carbon and nitrogen will form double and triple bonds.
  - Oxygen forms only double bonds.
- Multiple covalent bonding is particularly common in *organic* molecules, which consist predominantly of the element carbon.
- Note that in compounds containing multiple bonds, carbon still forms four covalent bonds, nitrogen still forms three covalent bonds, and oxygen still forms two covalent bonds.

4.4 Coordinate Covalent Bonds

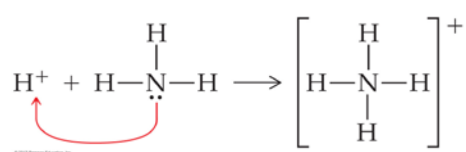
- A **coordinate covalent bond** is the covalent bond that forms when both electrons are donated by the same atom.



4.4 Coordinate Covalent Bonds

Once formed, a coordinate covalent bond is no different from any other covalent bond.

Coordinate covalent bonds often result in unusual bonding patterns, such as nitrogen with four covalent bonds, or oxygen with three bonds (H_3O^+).



4.5 Characteristics of Molecular Compounds

- Ionic compounds have high melting and boiling points because the attractive forces between oppositely charged ions are so strong.
- *Molecules* are neutral, so there is no strong attraction to hold them together.
- There are weaker forces between molecules, called *intermolecular forces.*

4.5 Characteristics of Molecular Compounds

- Very weak intermolecular forces = gas.
- Somewhat stronger intermolecular forces = liquid.
- Strongest intermolecular forces = molecular solid.
- Molecular solids have lower melting and boiling points than ionic compounds, are rarely soluble in water, and do not conduct electricity when melted.

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4.5 Characteristics of Molecular Compounds

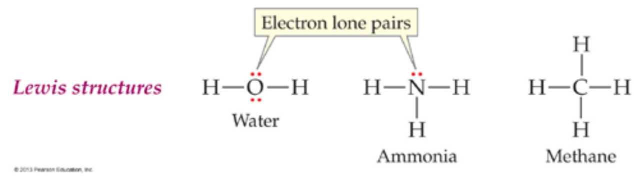
TABLE 4.1 A Comparison of Ionic and Molecular Compounds

Ionic Compounds	Molecular Compounds
Smallest components are ions (e.g., Na^+, Cl^-)	Smallest components are molecules (e.g., CO_2, H_2O)
Usually composed of metals combined with nonmetals	Usually composed of nonmetals with nonmetals
Crystalline solids	Gases, liquids, or low-melting-point solids
High melting points (e.g., NaCl = 801 °C)	Low melting points (H_2O = 0.0 °C)
High boiling points (above 700 °C) (e.g., NaCl = 1413 °C)	Low boiling points (e.g., H_2O = 100 °C; CH_3CH_2OH = 76 °C)
Conduct electricity when molten or dissolved in water	Do not conduct electricity
Many are water-soluble	Relatively few are water-soluble
Not soluble in organic liquids	Many are soluble in organic liquids

4.6 Molecular Formulas and Lewis Structures

- **Molecular formula**—A formula that shows the numbers and kinds of atoms in one molecule of a compound
- **Structural formula**—A molecular representation that shows the connections among atoms by using lines to represent covalent bonds
- **Lewis structure**—A molecular representation that shows both the connections among atoms and the locations of lone-pair valence electrons
- **Lone pair**—A pair of electrons that is not used for bonding

4.6 Molecular Formulas and Lewis Structures



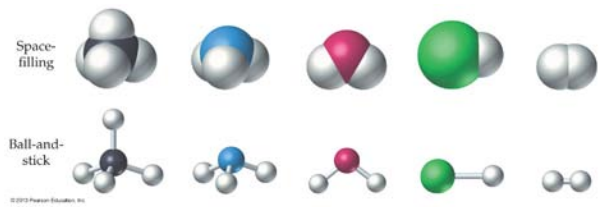
- A *molecular* formula gives the number of atoms that are combined in one molecule.
- An *ionic* formula gives only a ratio of ions.

4.7 Drawing Lewis Structures

- Lewis Structures for Molecules Containing C, N, O, X (Halogen), and H
  - C forms four covalent bonds and often bonds to other carbon atoms.
  - N forms three covalent bonds and has one lone pair of electrons.
  - O forms two covalent bonds and has two lone pairs of electrons.
  - Halogens (X) form one covalent bond and have three lone pairs of electrons.
  - H forms one covalent bond.

4.7 Drawing Lewis Structures

- Larger organic molecules are often written as *condensed structures* in which the bonds are not specifically shown.
- For clarity, *ball-and-stick* representations are often used. These are color-coded as follows: C is black, H is white, O is red, N is blue, S is yellow, P is dark blue, F is light green, Cl is green, Br is brownish red, and I is purple.



4.7 Drawing Lewis Structures

A General Method for Drawing Lewis Structures

1. **Find the total number of valence electrons of all atoms in the molecule or ion.** For an ion, add one electron for each negative charge or subtract one for each positive charge.
2. **Draw a line between each pair of connected atoms to represent the two electrons in a covalent bond.**
3. **Using the remaining electrons, add lone pairs so that each atom connected to the central atom (except H) gets an octet.**
4. **Place any remaining electrons in lone pairs on the central atom.**
5. **If the central atom does not have an octet after all electrons have been assigned, take a lone pair from a neighboring atom and form a multiple bond to the central atom.**

4.7 Drawing Lewis Structures

CO and NO: Pollutants or Miracle Molecules?

- CO is a poison and, NO, a pollutant in the environment.
- In 1992, it was also discovered that these molecules are key chemical messengers in the body.
- CO and NO are highly soluble and can diffuse from one cell to another, where they stimulate production of *guanylyl cyclase*.
- Guanylyl cyclase controls the production of *cyclic GMP*, which regulates many cellular functions.
- CO is associated with long-term memory. When CO production is blocked, long-term memories are no longer stored, and memories that previously existed are erased. When CO production is stimulated, memories are again laid down.
- NO fights infections and tumors, transmits messages between nerve cells and is associated with learning and memory, sleeping, and depression. It is also a *vasodilator*, a substance that allows blood vessels to relax and dilate.

4.8 The Shapes of Molecules

Valence-shell electron-pair repulsion (VSEPR) model

- VSEPR is a method for predicting molecular shape by noting how many electron charge clouds surround atoms and assuming that the clouds orient as far away from one another as possible.
- Constantly-moving valence electrons in bonds and lone pairs make up negatively charged clouds of electrons, which repel one another.
- The clouds tend to keep as far apart as possible, causing molecules to assume specific shapes.

4.8 The Shapes of Molecules

Applying the VSEPR model

STEP 1: Draw a Lewis structure of the molecule, and identify the atom whose geometry is of interest. This is usually the central atom.

STEP 2: Count the number of electron charge clouds surrounding the atom of interest. This is the total number of lone pairs plus connections to other atoms

STEP 3: Predict molecular shape by assuming that the charge clouds orient in space so that they are as far away from one another as possible

4.8 The Shapes of Molecules

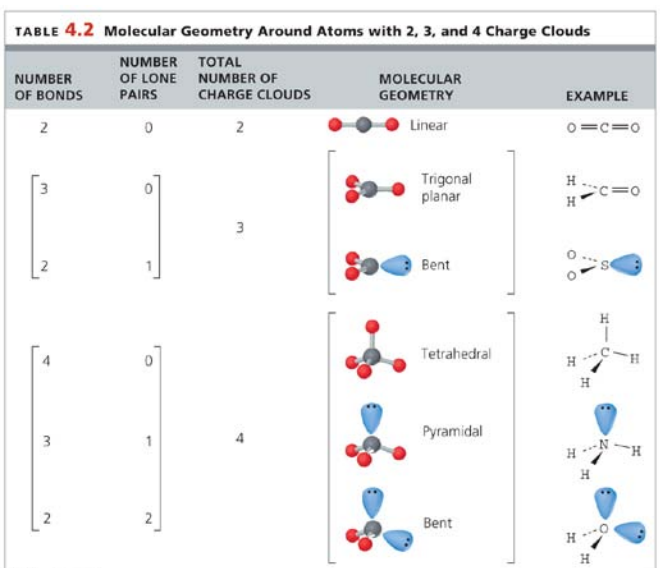


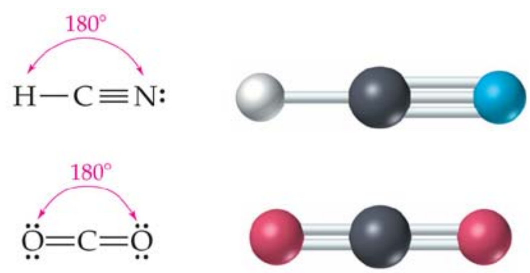
TABLE 4.2 Molecular Geometry Around Atoms with 2, 3, and 4 Charge Clouds

NUMBER OF BONDS	NUMBER OF LONE PAIRS	TOTAL NUMBER OF CHARGE CLOUDS	MOLECULAR GEOMETRY	EXAMPLE
2	0	2	Linear	O=C=O
3	0	3	Trigonal planar	$H_2C=O$
2	1	3	Bent	SO_2
4	0	4	Tetrahedral	CH_4
3	1	4	Pyramidal	NH_3
2	2	4	Bent	H_2O

4.8 The Shapes of Molecules

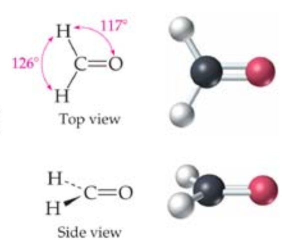
A **bond angle** is the angle formed by three adjacent atoms in a molecule.

- Two charge clouds are farthest apart when they point in opposite directions. Linear molecules result with **bond angles** of 180°.

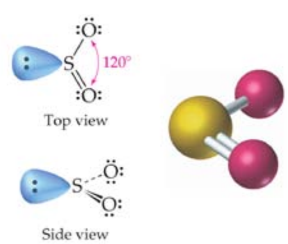


These molecules are linear, with bond angles of 180°.

4.8 The Shapes of Molecules



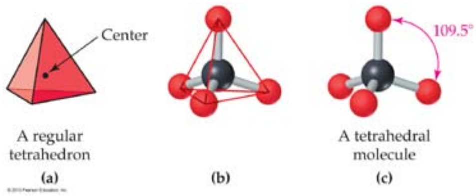
A formaldehyde molecule is planar triangular, with bond angles of roughly 120°.



An SO_2 molecule is bent, with a bond angle of approximately 120°.

- Three charge clouds are farthest apart if they lie in a plane and point to the corners of an equilateral triangle.
- Trigonal planar molecules result with **bond angles** of 120°.
- If one of the charge clouds is a lone pair, the molecule will be bent.

4.8 The Shapes of Molecules



- Four charge clouds extend to the corners of a *regular tetrahedron*, with bond angles of 109.5°.
- Shapes can be tetrahedral, pyramidal, or bent, depending on the number of lone pairs.
- Lone pairs repel strongly, reducing bond angles slightly.

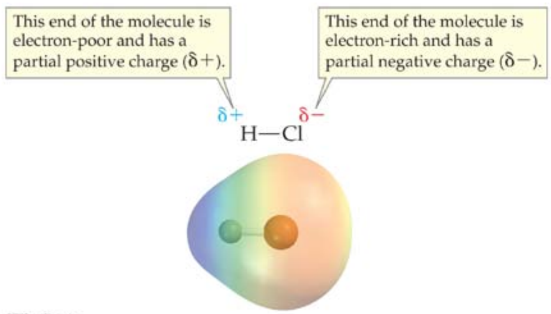
4.8 The Shapes of Molecules

VERY Big Molecules

- A polymer is formed of many repeating units connected in a long chain. Each unit is a simple molecule that has formed chemical bonds at both ends, linking it to other molecules.
- The units can be the same or different, connected with or without a pattern, and can be branched, unbranched, or cross-linked.
- Plastics, such as polyethylene, nylon, and Kevlar are all synthetic polymers.
- Carbohydrates, proteins, and DNA are all natural polymers.

4.9 Polar Covalent Bonds and Electronegativity

- When atoms are not identical, bonding electrons may be shared unequally.
- A **polar covalent bond** is one in which the electrons are attracted more strongly by one atom than by the other



4.9 Polar Covalent Bonds and Electronegativity

- The ability of an atom to attract electrons is called **electronegativity**.
- Fluorine, the most electronegative element, is assigned a value of 4, and less electronegative atoms are assigned lower values.
- Metallic elements have low electronegativities
- Halogens and reactive nonmetal elements have higher electronegativities.
- Electronegativity generally decreases going down the periodic table within a group.

4.9 Polar Covalent Bonds and Electronegativity

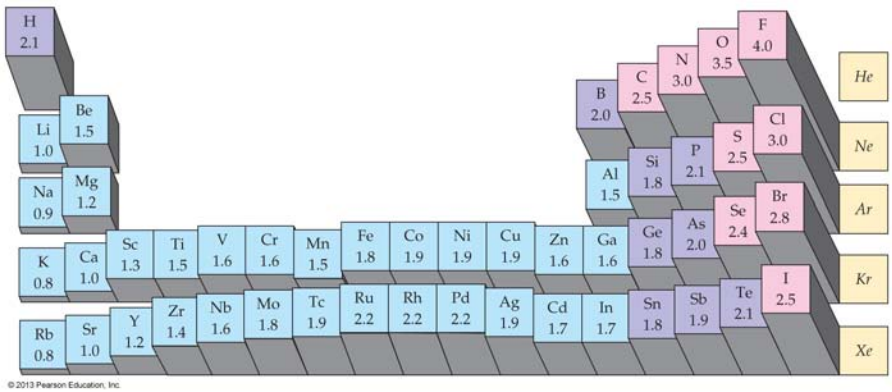


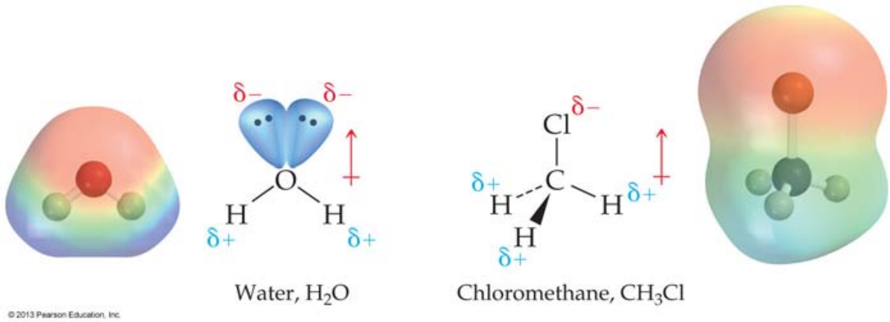
FIGURE 4.6 Electronegativities of several main-group and transition-metal elements.

4.9 Polar Covalent Bonds and Electronegativity

- Electronegativity differences of less than 0.5 result in nonpolar covalent bonds.
- Differences up to 1.9 indicate increasingly polar covalent bonds.
- Differences of 2 or more indicate substantially ionic bonds.
- There is no dividing line between covalent and ionic bonds; most bonds fall between these categories.

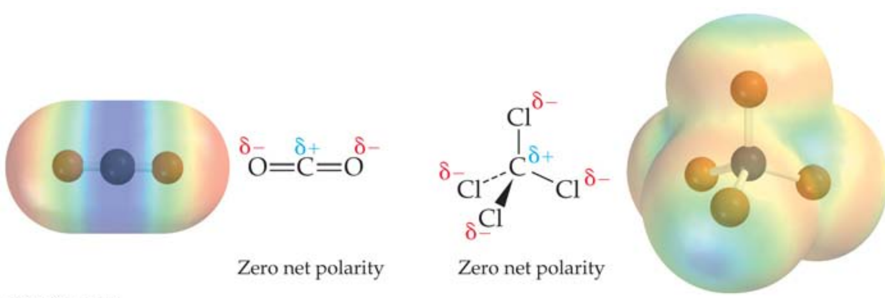
4.10 Polar Molecules

- Molecular polarity is due to individual bond polarities and lone-pair contributions. Electrons are displaced toward the more electronegative atom.



4.10 Polar Molecules

- Molecular polarity depends on the shape of the molecule.
- Symmetrical molecules can have polar bonds and be non-polar overall.



4.10 Polar Molecules

- Polarity has a dramatic effect on the physical properties of molecules particularly on melting points, boiling points, and solubilities.

4.11 Naming Binary Molecular Compounds

- When two different elements combine, they form a **binary compound**.
- The less-electronegative element is written first.
  - Metals are always written before nonmetals.
  - A nonmetal farther left on the periodic table generally comes before a nonmetal farther right.
- Molecular formulas require identifying exactly how many atoms of each element are included.

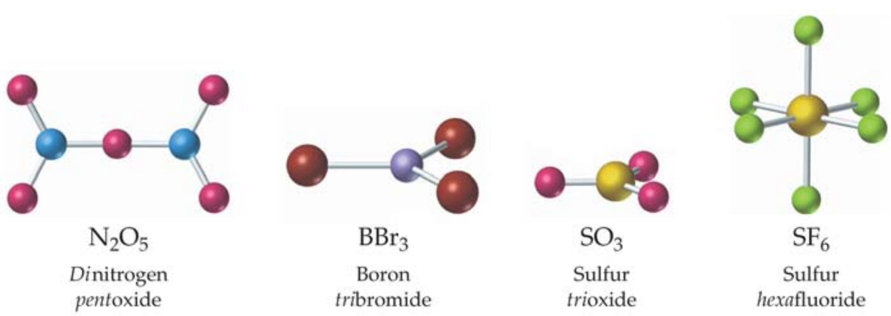
4.11 Naming Binary Molecular Compounds

TABLE 4.3 Numerical Prefixes Used in Chemical Names

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

1. Name the first element in the formula, using a prefix if needed.
2. Name the second element in the formula, using an *-ide* ending and a prefix if needed.

4.11 Naming Binary Molecular Compounds



- The prefix *mono*-, meaning one, is omitted except where needed to distinguish between two different compounds with the same elements, such as carbon monoxide and carbon dioxide.

4.11 Naming Binary Molecular Compounds

Damascenone by Any Other Name Would Smell as Sweet

- Chemical names are complex, as there are more than 19 million known compounds.
- A chemical name has to include enough information to tell chemists the composition and structure of the compound.
- Subtle differences in structure can result in significant differences in chemical or physical properties.
  - Geraniol ($C_{10}H_{18}O$), *3,7-dimethylocta-2,6-dien-1-ol* differs from citronellol ($C_{10}H_{20}O$, or *3,7-dimethyloct-6-en-1-ol*) by only one C—C double bond.
- Chemical names indicate nothing about toxicity or origin.

Chapter Summary

1. What is a covalent bond?

- A *covalent bond* is formed by the sharing of electrons.
- Atoms that share two electrons are joined by a *single bond*; atoms that share four electrons are joined by a *double bond*; and atoms that share six electrons are joined by a *triple bond*.
- The group of atoms held together by covalent bonds is called a *molecule*.
- Electron sharing occurs when singly-occupied valence orbitals *overlap*. The two electrons occupy both overlapping orbitals and belong to both atoms, bonding the atoms together.
- Electron sharing can occur when a filled orbital containing an unshared, *lone pair* overlaps a vacant orbital to form a *coordinate covalent bond*.

Chapter Summary, *Continued*

2. How does the octet rule apply to covalent bond formation?

- Atoms share enough electrons to reach a noble gas configuration.
- Hydrogen forms one covalent bond because it needs to share one more electron.
- Carbon and other group 4A elements form four covalent bonds.
- Nitrogen and other group 5A elements form three covalent bonds.
- Oxygen and other group 6A elements form two covalent bonds.
- Halogens (group 7A elements) form one covalent bond

Chapter Summary, *Continued*

3. What are the major differences between ionic and molecular compounds?

- *Molecular compounds* can be gases, liquids, or low-melting solids.
- They usually have lower melting points and boiling points than ionic compounds.
- Many are water insoluble.
- They do not conduct electricity when melted or dissolved.

Chapter Summary, *Continued*

4. How are molecular compounds represented?

- Formulas which show the numbers and kinds of atoms in a molecule, are called *molecular formulas*.
- *Lewis structures* show how atoms are connected in molecules.
  - Covalent bonds are indicated as lines between atoms.
  - Valence electron lone pairs are shown as dots.
- Lewis structures are drawn by counting the total number of valence electrons and then placing shared pairs and lone pairs so that all electrons are accounted for.

Chapter Summary, *Continued*

5. What is the influence of valence-shell electrons on molecular shape?

- Molecules have shapes that depend on the number of electron charge clouds (bonds and lone pairs) surrounding the various atoms.
- These shapes are predicted using the *valence-shell electron-pair repulsion* (VSEPR) model.
  - Atoms with two electron charge clouds adopt linear geometry.
  - Atoms with three charge clouds adopt trigonal planar geometry.
  - Atoms with four charge clouds adopt tetrahedral geometry.

Chapter Summary, *Continued*

6. When are bonds and molecules polar?

- Bonds between atoms are *polar covalent* if electrons are not shared equally.
- The ability of an atom to attract electrons is the atom's *electronegativity.*
  - It is highest for reactive nonmetal elements.
  - It is lowest for metals on the lower left.
- Comparing electronegativities allows a prediction of whether a bond is covalent, polar covalent, or ionic.
- Molecules can be polar if electrons are attracted more strongly to one part of the molecule than to another.