Chapter 8

Chemical Bonding

1

Introduction

- How and why to atoms come together (bond) to form compounds?
- Why do different compounds have such different properties?
- What do molecules look like in 3 dimensions?

2

Chapter 8 Topics

1. **Types of Bonds**
2. **Ionic Bonding**
3. **Covalent Bonding**
4. **Shapes of Molecules**

3

Molecular and Ionic Compounds

TABLE 8.1 Properties of Two Carbon Compounds

Compound	Calcium Carbonate	Carbon Dioxide
Formula	$CaCO_3$	CO_2
Physical state	White solid	Colorless gas
Molar mass (g/mol)	100.1	44.01
Density (g/mL)	2.71	0.00198
Melting point (°C)	1339 (at high pressure)	–56.6 (at 5.11 atm)
Boiling point (°C)	Decomposes	Sublimes at 78.6
Electrical conductivity as solid	Very low	Very low
Electrical conductivity as liquid	High	Very low
Dissolves in	Acids	H_2O, CCl_4

4

Ionic and Covalent

TABLE 8.2 General Properties of Ionic and Covalent Substances

Ionic	Covalent
Crystalline solids	Gases, liquids, or solids
Hard and brittle solids	Brittle and weak solids, or soft and waxy solids
Very high melting point	Low melting point
Very high boiling point	Low boiling point
Good electrical conductor when molten or in solution	Poor conductor of electricity and heat
Often soluble in water but not in carbon tetrachloride	Often soluble in carbon tetrachloride but not in water

5

Ionic and Covalent

Calcite-CaCO3.ent

A. **In ionic compounds, ions are held together by electrostatic forces – forces between oppositely charged ions.**

A. **In molecular compounds, atoms are held together by covalent bonds in which electrons are shared.**

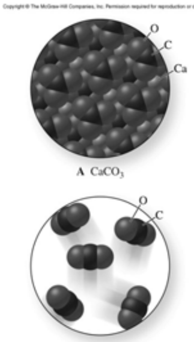


Figure 8.2

8.2 Ionic Bonding

- Lewis Symbols help us to focus on the valence electrons – those that can participate in bonding.

TABLE 8.3 Lewis Symbols of the Period 2 Elements

	Group							
	IA (1)	IIA (2)	IIIA (13)	IVA (14)	VA (15)	VIA (16)	VIIA (17)	VIIIA (18)
Electron configuration	$1s^22s^1$	$1s^22s^2$	$1s^22s^22p^1$	$1s^22s^22p^2$	$1s^22s^22p^3$	$1s^22s^22p^4$	$1s^22s^22p^5$	$1s^22s^22p^6$
Valence electrons	1	2	3	4	5	6	7	8
Lewis symbol	$\mathrm{Li}\cdot$	$\cdot\mathrm{Be}\cdot$	$\dot{\mathrm{B}}\cdot$	$\cdot\dot{\mathrm{C}}\cdot$	$:\dot{\mathrm{N}}\cdot$	$:\dot{\mathrm{O}}\cdot$	$:\ddot{\mathrm{F}}\cdot$	$:\ddot{\mathrm{Ne}}:$

7

EXAMPLE 8.4 Lewis Symbols of Ions

Write the Lewis symbols for the magnesium and sulfide ions. Then write a formula for a compound that would form between them, using their Lewis symbols.

Solution:

Magnesium can lose two valence electrons to achieve a noble-gas configuration, giving the Mg^{2+} cation with no remaining valence electrons:

$$\cdot\mathrm{Mg}\cdot \longrightarrow \mathrm{Mg}^{2+} + 2e^-$$

Sulfur has six valence electrons—two electron pairs and two unpaired electrons. It gains two electrons to form the S^{2-} anion:

$$:\dot{\underset{\cdot\cdot}{\mathrm{S}}}\cdot + 2e^- \longrightarrow :\ddot{\underset{\cdot\cdot}{\mathrm{S}}}:^{2-}$$

These ions associate in a 1:1 ratio to form MgS:

$$\mathrm{Mg}^{2+} + :\ddot{\underset{\cdot\cdot}{\mathrm{S}}}:^{2-} \longrightarrow (\mathrm{Mg}^{2+})(:\ddot{\underset{\cdot\cdot}{\mathrm{S}}}:^{2-}) \quad \text{or} \quad \mathrm{MgS}$$

8

Ionic Compounds

- Ions of like charge repel each other in ionic compounds, and opposite charged ions attract. This results in a 3-dimensional regular pattern called a crystal lattice.

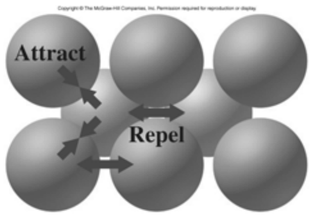


Figure 8.7

10

Structures of Ionic Compounds

Figure 8.8

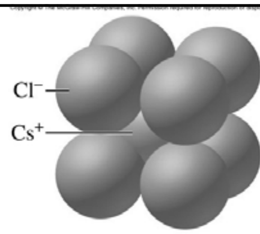
NaCl.ent

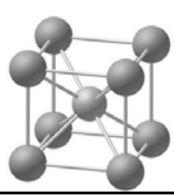
Sodium Chloride, NaCl, Crystal

11

Cesium Chloride Crystal

CsCl.ent





Calcium Fluoride Crystal

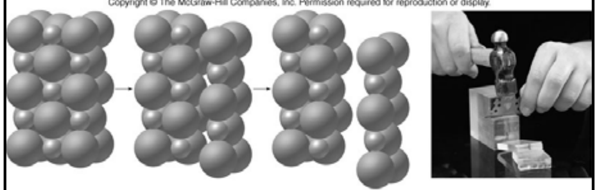
CaF2_.ent

Figure 8.10

13

Properties of Ionic Solids

- Why are crystalline ionic solids hard and brittle?
- Why are they poor conductors of electricity in the solid state?



8.3 Covalent Bonding

- **When two nonmetals form a bond, the bond is covalent. They are both close to the noble-gas electron configuration, so sharing will allow both to obtain it.**
- **In a covalent bond, each shared electron interacts simultaneously with two nuclei.**

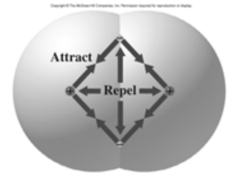


Figure 8.12

Carbon Dioxide – Covalent bonds

- **The atoms of CO_2 molecules are held together by strong covalent bonds. No bonds connect the molecules, so CO_2 molecules separate from each other into the gas state at room temperature.**

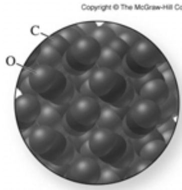
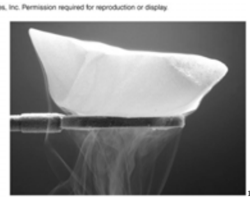


Figure 8.13



Dry_ice.ent

The Octet Rule

- **Just as in ionic bonding, covalent bonds are formed so that each atom can have the noble-gas electron configuration. Noble gases have 8 valence electrons, an <u>octet</u>.**

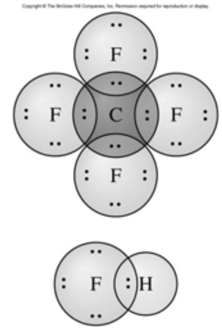


Figure 8.14

Lewis Formulas for the Diatomic Elements

- **How does hydrogen obtain a noble-gas electron configuration?**

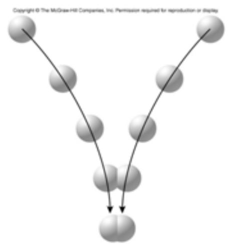


Figure 8.15

The Halogens

- **Do the atoms in each of these molecules have an octet?**
- **Why do the halogens exist as diatomic molecules?**

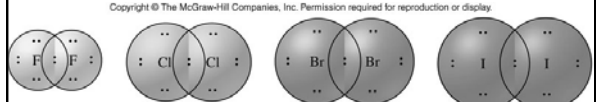


Figure 8.16

Multiple Bonds

- How many valence electrons does an oxygen atom have?
  - How many does it need to obtain an octet?
  - O_2 has a double bond, two pairs of shared electrons
- How many valence electrons does a nitrogen atom have?
  - How many does it need to obtain an octet?
  - N_2 has a triple bond, three pairs of shared electrons

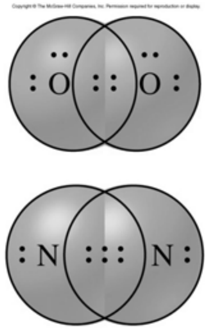


Figure 8.17

20

P_4 and S_8

- How do phosphorus and sulfur obtain an octet in P_4 and S_8?

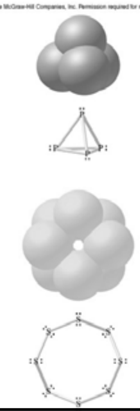


Figure 8.18

Valence Electrons and Number of Bonds

- How many bonds do each of the following atoms like to form?
  a) H
  b) Cl
  c) O
  d) N
  e) C

22

Practice

- Draw the Lewis structures for each of the following:
  a) C_2H_6
  b) C_2H_4
  c) C_2H_2
  d) HCN

23

Valence Electrons and Number of Bonds

- Draw a Lewis structure for each of the following based on how many bonds each likes to form. Remember to include the nonbonding electrons so that all have octets (except H).
  a) CO_2
  b) H_2CO
  c) NF_3
  d) H_2O

24

Steps for Writing <u>Lewis Structures</u>

1. Write an atomic skeleton.
2. Sum the valence electrons from each atom to get the total number of valence electrons.
3. Place 2 electrons (a single bond) between each pair of bonded atoms.
4. Add electrons to the outside atoms to make octets. Count electrons, and if you have enough, add to the central atom to give it an octet.
5. If you don't have enough electrons to give the central atom an octet, shift unshared electrons to bonding positions so that the central atom has an octet. This will give multiple bonds.

25

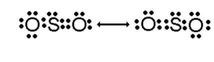
Writing Lewis Structures

- Write Lewis structures for each of the following:
  - CO_2
  - SO_2
  - NO_2^-
  - HNO_3

26

Resonance

- Sometimes Lewis structures do not give an accurate picture of the bonding.
- For example, there are two Lewis structures that represent the SO_2 molecule.



- The actual molecule is a blend of these two:
  - Both S-O bonds are the same length with partial double bond character

27

Resonance

- Like a single Lewis structure, each alone shows some aspect of the shell.
- Like a set of resonance structures, the pair of photos together gives a better representation of the shell.

Figure 8.19

28

Resonance Structures

- Which of the following have resonance?
- For those that do, how many resonance structures do they have?

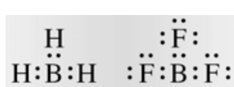
a) NO_3^-
b) HNO_3
c) CH_4
d) H_2CO

29

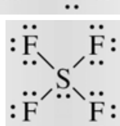
<u>Exceptions</u> to the Octet Rule

- Molecules with an odd number of electrons
  - E.g. NO

- <u>Incomplete</u> octets
  - E.g. BH_3 or BF_3



- Expanded octets
  - E.g. SF_4 or SF_6



30

ionic-covalent

Classification of Bonding

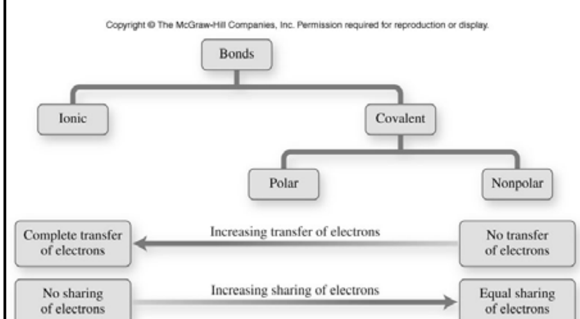


Figure 8.3

Polar Covalent Bonds

- In <u>polar covalent bonds</u>, electrons are not shared equally. This result in unequal sharing which can be described as partial electron transfer.
- In a polar bond there are partial charges on the atoms sharing electrons.

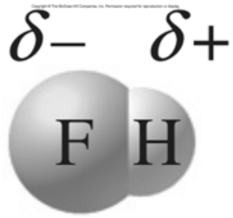
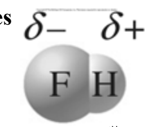


Figure 8.4

32

Electronegativity

- The electronegativity scale tells us which elements have a greater pull on electrons in a covalent bond.
  - The more electronegative element pulls bonding electrons more strongly and obtains a partial negative charge (δ–).
  - The less electronegative element loses electron density in the bond and therefore obtains a partial positive charge (δ+).



<u>Electronegativity Values</u>

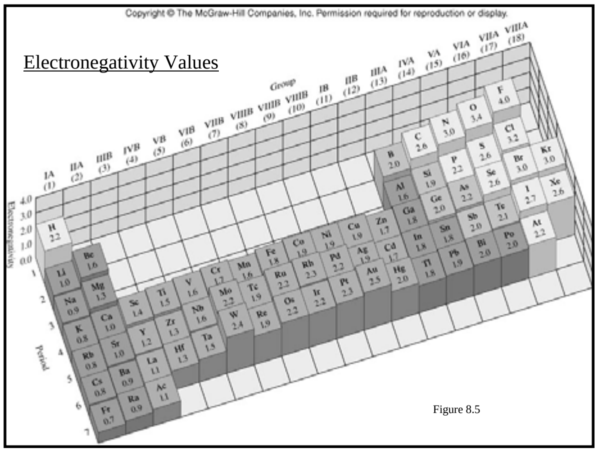


Figure 8.5

Electronegativity Trends

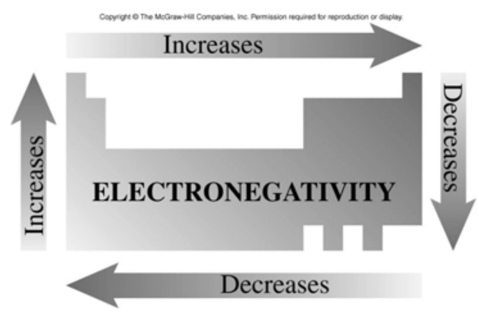


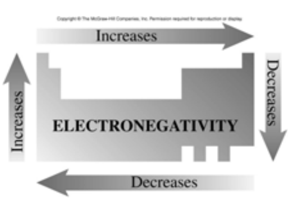
Figure 8.6

35

Electronegativity Trends

- The difference in electronegativity between metals and nonmetals is so large, that the electrons are transferred, not shared.
- The greater the electronegativity difference, the more polar the bond.
- Si-F > N-F> O-F >F-F

- What partial charges go on each atom?



8.4 Shapes of Molecules

- The shape of a molecule influences many of its properties including taste and smell.
- Sweet substances are often of a shape similar to that of glucose. It's –H and –OH groups fit into a taste receptor site on the tongue.

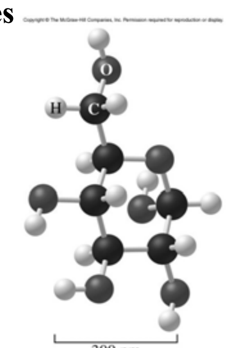


Figure 8.23

37

Odor Receptors

Molecule in Receptor Cavity							
Molecular Characteristics	Disk shape	Spherical shape	Rod shape	Wedge shape	Disk and tail shape	Attraction to negative centers	Attraction to positive centers
Primary Odor	Musky	Camphoric	Ethereal	Pepperminty	Floral	Pungent	Putrid
Chemical and Common Examples	Xylene Musk perfume/ aftershave	Camphor Mothballs	Diethyl ether Pears	Menthol Mint gum/ mouthwash	α-amyl pyridine Roses	Formic acid Vinegar	Hydrogen sulfide Rotten eggs

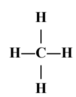
Figure 8.24

Predicting Shapes of Molecules

- A simple model used with Lewis structures allows us to predict shapes of molecules.
- Valence-Shell Electron-Pair Repulsion Theory (VSEPR theory) is based on the fact that negative charges repel one another.
  - Valence electron pairs (bonding or nonbonding) repel one another and take up positions that maximize their distance and angles between them.

VSEPR (CH_4)

- The methane molecule (CH_4) has four pairs of electrons around the central atom (all bonding):



- Maximizing their distance gives bond angles of 109.5°.

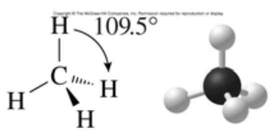
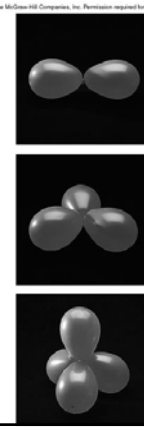


Figure p. 303

Parent Structures



VSEPR movie

Figure 8.25

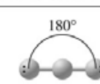
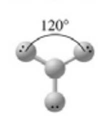
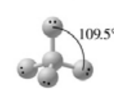
VSEPR program

Parent Structures

- When determining the parent structure and bond angles for a molecule or polyatomic ion, each of the following is counted as 1 set of electrons:
  - Nonbonding electron pair
  - Single bond
  - Double bond
  - Triple bond

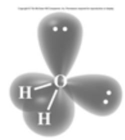
Multiple bonds don't repel and separate from each other so they are counted only once even though they contain more than one pair.

TABLE 8.5 Geometric Structures Arising from Different Numbers of Atoms

Number of Atoms or Electron Pairs	Parent Structure	Geometric Arrangement of Atoms or Electron Pairs
2	Linear	 180°
3	Trigonal planar	 120°
4	Tetrahedral	 109.5°

Parent Structures and Bond Angles

- **Determine the parent structure and bond angles for each of the following.**
  - a) NH_3
  - b) CO_2
  - c) SO_2
  - d) BH_3
  - e) H_2O



44

Shapes of Molecules and Ions

- **The actual shape of a molecule includes only the atoms.**
- **To envision the shape from a parent structure, make the nonbonding electrons invisible (but they are still there).**
- **What is the shape of a water molecule?**

H2o.ent

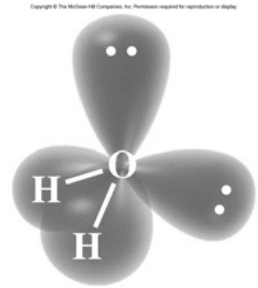


Figure 8.27

45

Shapes of Molecules and Ions

- **Why do CO_2 and SO_2 have different shapes?**

46

Figure 8.28

TABLE 8.6 Arrangement of Electron Pairs and Molecular Shapes

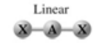
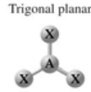
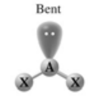
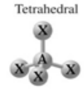
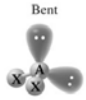
	General Formula	Number of Bonded Atoms	Number of Unshared Pairs	Molecular Shape	Examples
	Parent Structure: Linear				
2	AX_2	2	0	Linear 	$BeCl_2$, CO_2, HCN
	Parent Structure: Trigonal Planar				
3	AX_3	3	0	Trigonal planar 	BF_3, BH_3, SO_3, NO_3^-
	AX_2	2	1	Bent 	SO_2, NO_2^-

TABLE 8.6 Arrangement of Electron Pairs and Molecular Shapes *continued*

	Parent Structure: Tetrahedral				
4	AX_4	4	0	Tetrahedral 	CH_4, CH_2Cl_2, $SiCl_4$, $POCl_3$, BrO_4^-
	AX_3	3	1	Trigonal pyramidal 	NH_3, PF_3, NH_2Cl
	AX_2	2	2	Bent 	H_2O, F_2O, BrO_2, SO_2, SCl_2

Molecular Shapes

- **Trigonal planar and trigonal pyramidal molecular shapes are much different from one another.**

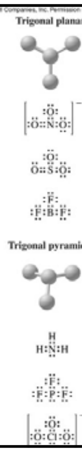


Figure 8.26

Describing Shapes of Molecules with More than One Central Atom

- Glycine: $H_2NCH_2CO_2H$
  1. Draw the Lewis Structure.
  2. Determine the parent structure, remembering to consider all sets of electrons, including nonbonding electrons.
  3. Determine the bond angles at each central atom (N, C, C, and O)

50

Describing Shapes of Molecules with More than One Central Atom

- Glycine: $H_2NCH_2CO_2H$
  1. Draw the Lewis Structure.
  2. Determine the parent structure, remembering to consider all sets of electrons, including nonbonding electrons.
  3. Determine the bond angles at each central atom (N, C, C, and O).

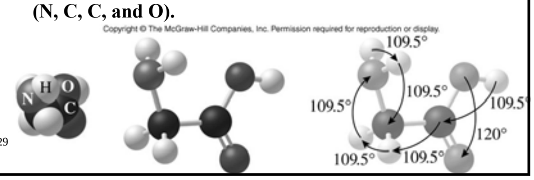


Figure 8.29

What are the bond angles?

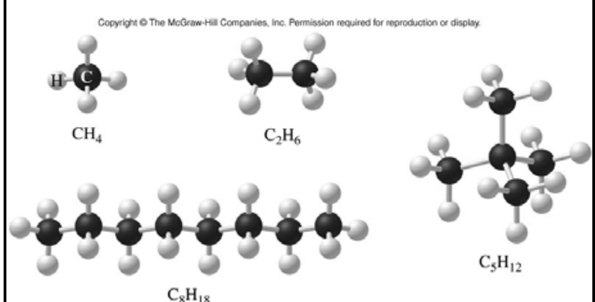


Figure 8.21

52

What are the bond angles?

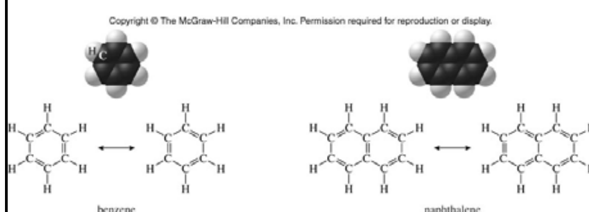


Figure 8.22

53

Molecular Polarity

- Which of these substances is a polar substance?

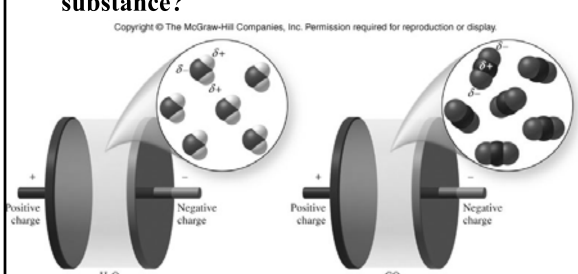
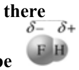


Figure 8.32

Molecular Polarity

- Molecules with polar bonds can be polar under certain conditions.
  - A diatomic molecule with a polar bond is polar since there is only one bond.
  - A molecule or polyatomic ion with polar bonds will be
    - polar if its geometry causes an asymmetric distribution of partial charge. (E.g. H_2O)
    - Nonpolar if its geometry is symmetrical enough to allow cancellation of the polar bonds. (E.g. CO_2)

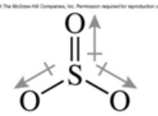


- A polyatomic molecule with nonpolar bonds cannot be polar. O_3 and NBr_3 have no polar bonds.

55

Nonpolar Molecules with Polar Bonds

- The molecules SO_3 and CCl_4 have polar bonds, but their perfect geometries with no nonbonding electrons on the central atom cause them to be nonpolar molecules.



Trigonal Planar

Figure 8.33

Tetrahedral

56

Which of the following are polar molecules?

- NH_3
- BH_3
- CS_2
- H_2S
- O_3
- CH_2Cl_2

57

All Hydrocarbons are Nonpolar

Figure 8.21

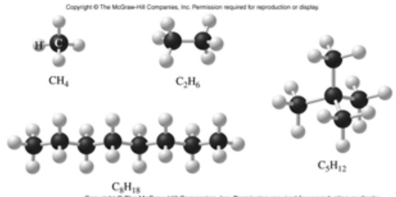
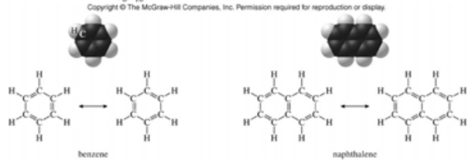
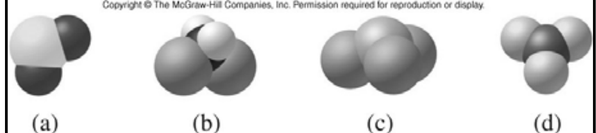


Figure 8.22



58

Which of the following are polar molecules?

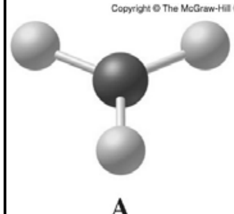


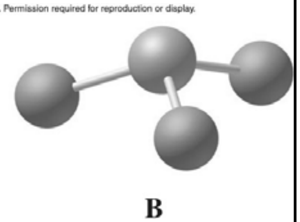
Question 8.116

- Polarity

59

Which is a polar molecule?





Question 8.119

60