Chapter 6 Notes - Chemical Bonding

Chemical bond - A mutual electrical attraction between the nuclei and valence electrons of different atoms that binds the atoms together

6-1 Introduction to Chemical Bonding

I. Types of Chemical Bonding
  - A. Ionic Bonding
    - 1. Chemical bonding that results from the electrical attraction between large numbers of cations and anions
    - 2. Electrons are transferred in pure ionic bonding
  - B. Covalent Bonding
    - 1. Results from the sharing of electron pairs between two atoms
      - a. Nonpolar Covalent Bond - A covalent bond in which the bonding electrons are shared equally by the bonded atoms, resulting in a balanced distribution of charge
      - b. Polar Covalent Bond - A covalent bond in which the bonded atoms have an unequal attraction for the shared electrons and a resulting unbalanced distribution of charge

General Rules for Determining Bond Character

Bond Type	*Electronegativity Difference*	*Percent Ionic Character*
Ionic	Greater than 1.7	Greater than 50%
Polar Covalent	Between 0.3 and 1.7	Between 5% and 50%
Nonpolar Covalent	0.3 and less	Less than 5%

6-2 Covalent Bonding and Molecular Compounds

I. Important Definitions
  - A. Molecule
    - 1. A neutral group of atoms that are held together by covalent bonds
  - B. Diatomic Molecule
    - 1. A molecule containing only two atoms
  - C. Molecular Compound
    - 1. A chemical compound whose simplest units are molecules
  - D. Chemical Formula
    - 1. Indicates the relative numbers of atoms of each kind of a chemical compound by using atomic symbols and numerical subscripts
  - E. Molecular Formula
    - 1. Shows the types and numbers of atoms combined in a single molecule of a molecular compound

II. Formation of a Covalent Bond

A. Interatomic Forces
   1. As atoms approach one another, there are a number of forces at work
      a. repulsion of one nucleus for another
      b. repulsion of electrons for other electrons
      c. attraction of electrons to nuclei

B. Energy Considerations
   1. Atoms approach each other
      a. Potential energy decreases as attractive forces dominate over repulsive forces
   2. Atoms get too close
      a. Potential energy begins to increase as repulsive forces dominate over attractive forces

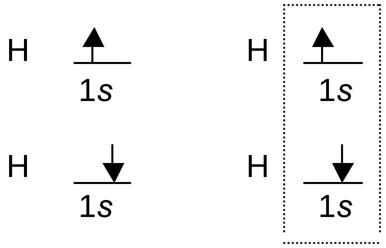
III. Characteristics of the Covalent Bond

A. Bond Length
   1. The distance between two bonded atoms at their minimum potential energy

B. Bond Energy
   1. The energy required to break of chemical bond and form neutral isolated atoms

$$H_2 \ + \ 436\ \text{kJ} \rightarrow 2H$$

   2. Covalent Bonding in Hydrogen



Hydrogen Atoms

Hydrogen Molecules

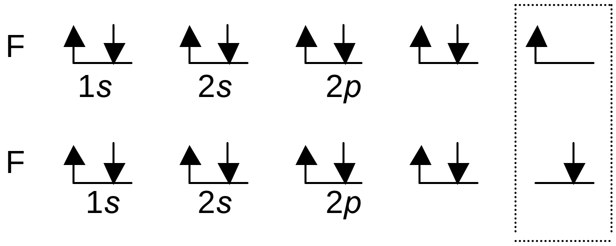
Bond Lengths and Bond Energies for Selected Covalent Bonds					
Bond	*Length (pm)*	*Energy (kJ/mol)*	*Bond*	*Length (pm)*	*Energy (kJ/mol)*
H - H	74	436	C - C	154	346
F - F	141	159	C - N	147	305
Cl - Cl	199	243	C - O	143	358
Br - Br	228	193	C - H	109	418
I - I	267	151	C - Cl	177	327
H - F	92	569	C - Br	194	285
H - Cl	127	432	N - N	145	180

H - Br	141	366	N - H	101	386
H - I	161	299	O - H	96	459

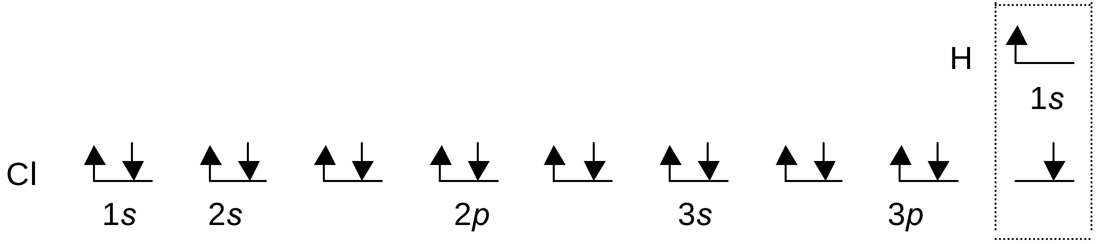
IV. The Octet Rule

A. The Octet Rule
   1. Chemical compounds tend to form so that each atom, by gaining, losing or sharing electrons, has an octet of electrons in its highest occupied energy level

B. Diatomic Fluorine



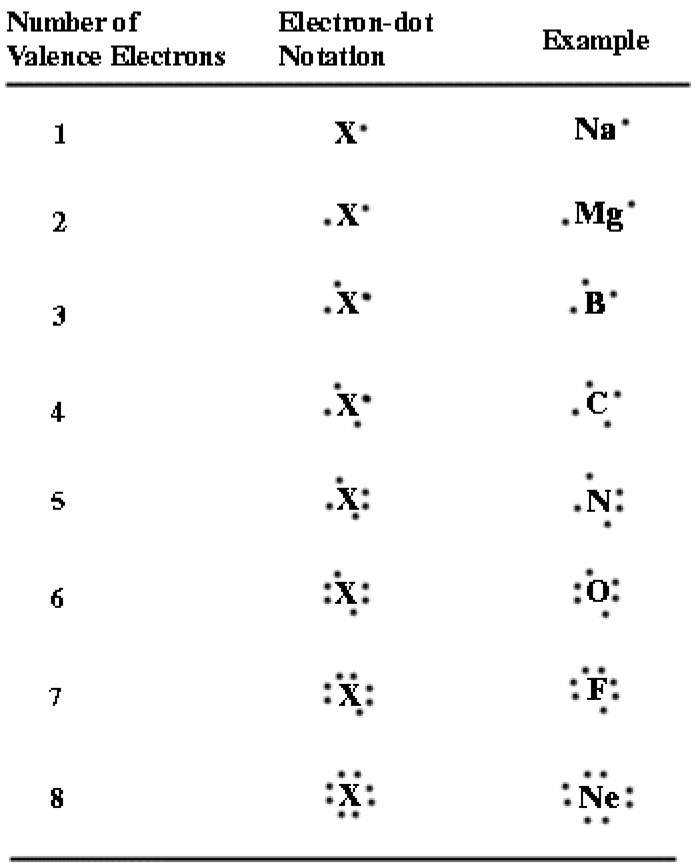
C. Hydrogen Chloride



D. Exceptions to the Octet Rule
   1. Hydrogen Bonding
      a. Two valence electrons (helium configuration)
   2. Expanded Valence
      a. More than eight electrons
      b. *d* orbitals involved as well as *s* and *p*

V. Electron-Dot Notation

A. Electron-dot Notation
   1. An electron-configuration notation in which only the valence electrons of an atom of a particular element are shown, indicated by dots placed around the element's symbol
   2. Inner shell electrons are not shown
   3. These are the electrons usually involved in the formation of covalent bonds



Number of Valence Electrons	Electron-dot Notation	Example
1	X	Na
2	X	Mg
3	X	B
4	X	C
5	X	N
6	X	O
7	X	F
8	X	Ne

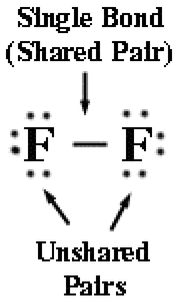
VI. <u>Lewis Structures</u>

A. Unshared Pairs (Lone Pairs)

1. A pair of electrons that is not involved in bonding and that belongs exclusively to one atom

B. Lewis Structures

1. Formulas in which atomic symbols represent nuclei and inner-shell electrons, dot pairs or dashes between two atomic symbols represent electron pairs in covalent bonds, and dots adjacent to only one atomic symbol represent unshared electrons

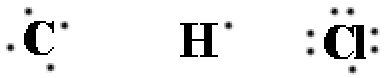


C. Structural Formula
   1. Formulas indicating the kind, number, arrangement, and bonds but not unshared pairs of the atoms in a molecule

$$\mathrm{F-F}$$

D. Drawing Lewis Structures (trichloromethane, $CHCl_3$ as an example)
   1. Determine the type and number of atoms in the molecule

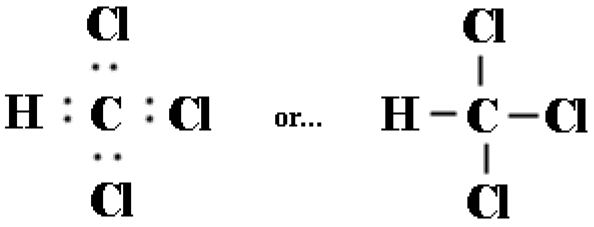
      1 x C, 1 x H, 3 x Cl
   2. Write the electron dot notation for each type of atom in the molecule



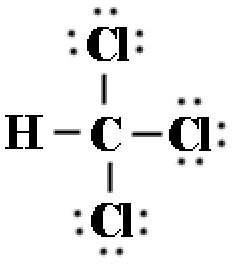
   3. Determine the total number of valence electrons to be combined

C	1 x $4e^-$ =	$4e^-$
H	1 x $1e^-$ =	$1e^-$
Cl	3 x $7e^-$ =	$21e^-$
		$26e^-$

   4. Arrange the atoms to form a skeleton structure for the molecule. If carbon is present, it is the central atom. Otherwise, the least electronegative element atom is central (except for hydrogen, which is never central). Then connect the atoms by electron-pair bonds



   5. Add unshared pairs of electrons so that each hydrogen atom shares a pair of electrons and each other nonmetal is surrounded by eight electrons



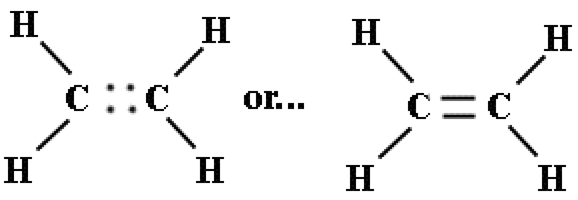
   6. Count the electrons in the structure to be sure that the number of valence electrons used equals the number available

VII. <u>Multiple Covalent Bonds</u>

A. Double Bonds

1. A covalent bond produced by the sharing of two pairs of electrons between two atoms

ethene



2. Higher bond energy and shorter bond length than single bonds

B. Triple Bonds

1. A covalent bond produced by the sharing of three pairs of electrons between two atoms

ethyne (acetylene)

2. Higher bond energy and shorter bond length than single or double bonds

Bond Lengths and Bond Energies for Single and Multiple Covalent Bonds					
Bond	*Length (pm)*	*Energy (kJ/mol)*	*Bond*	*Length (pm)*	*Energy (kJ/mol)*
C - C	154	346	C - O	143	358
C=C	134	612	C=O	120	799
C≡C	120	835	C≡O	113	1072
C - N	147	305	N - N	145	180
C=N	132	615	N=N	125	418
C≡N	116	887	N≡N	110	942

VIII. <u>Resonance Structures</u>

A. Resonance

1. Resonance refers to bonding in molecules or ions that cannot be correctly represented by a single Lewis structure

Ozone (O_3)

2. Neither structure is "correct"
3. Electrons do not move "back and forth"
4. The two bonds are equal

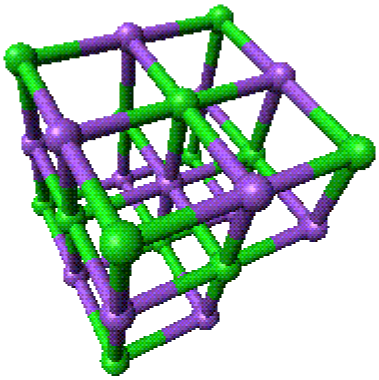
IX. Covalent-Network Bonding
- A. Covalent-Network Compounds
  - 1. Many covalently bonded compounds do not form individual molecules
    - a. Diamond
    - b. Quartz

6-3 Ionic Bonding and Ionic Compounds

I. Introduction
- A. Ionic Compounds
  - 1. A compound composed of positive and negative ions that are combined so that the numbers of positive and negative charges are equal
    - a. Most are crystalline solids
    - b. Examples include NaCl, $MgBr_2$, Na_2O
- B. Formula Unit
  - 1. The simplest collection of atoms from which an ionic compound's formula can be established

II. Formation of Ionic Compounds
- A. Electron Configuration Changes
  - 1. Electrons are transferred from the highest energy level of one atom to the highest energy level of a second atom, creating noble gas configurations in all atoms involved
  - 2. Formation of sodium chloride
    - a. $Na = 3s^1$ $Cl = 3s^2 3p^5$
    - b. $Na^+ = 2s^2 2p^6$ $Cl^- = 3s^2 3p^6$
- B. Characteristics of Ionic Bonding
  - 1. Lattice Energy
    - a. The energy released when one mole of an ionic crystalline compound is formed brom gaseous ions
      $Na^+(g) + Cl^-(g) \rightarrow NaCl(s) + 787.5\ kJ$
    - b. Formation of ionic compounds is ALWAYS exothermic
  - 2. Lattice Structure
    - a. Three dimensional arrangements vary depending upon the sizes and charges of the ions

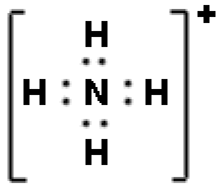
NaCl



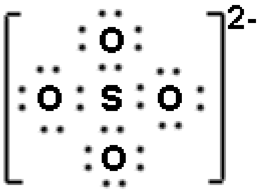
- C. A Comparison of Ionic and Molecular Compounds

	Ionic Compounds	*Molecular Compounds*
Melting point	Generally high	Generally low
Boiling Point	Generally high	Generally low
Electrical Conductivity	Excellent conductors, molten and aqueous	Poor conductors, except aqueous acids
Solubility in water	Generally soluble	Polar covalent molecules are soluble

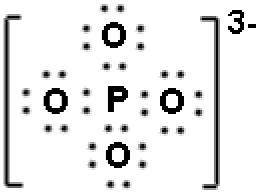
- D. Polyatomic Ions
  - 1. A charged group of covalently bonded atoms
    - a. Creation of octets results in an excess or deficit of electrons



Ammonium Ion



Sulfate Ion



Phosphate Ion

6-4 Metallic Bonding

I. <u>The Metallic Bond Model</u>

- A. Metallic Bonding
  - 1. The chemical bonding that results from the attraction between metal atoms and the surrounding sea of electrons
- B. Electron Delocalization in Metals
  - 1. Vacant *p* and *d* orbitals in metal's outer energy levels overlap, and allow outer electrons to move freely throughout the metal
  - 2. Valence electrons do not belong to any one atom

II. <u>Metallic Properties</u>

- A. Metals are good conductors of heat and light
- B. Metals are shiny
  - 1. Narrow range of energy differences between orbitals allows electrons to be easily excited, and emit light upon returning to a lower energy level
- C. Metals are Malleable
  - 1. Can be hammered into thin sheets
- D. Metals are ductile
  - 1. Ability to be drawn into wire
    - a. Metallic bonding is the same in all directions, so metals tend not to be brittle

III. <u>Metallic Bond Strength</u>

- A. Heat of Vaporization
  - 1. The ease with which atoms in a metallic solid can be separated from one another into individual gaseous atoms is related to bond strength

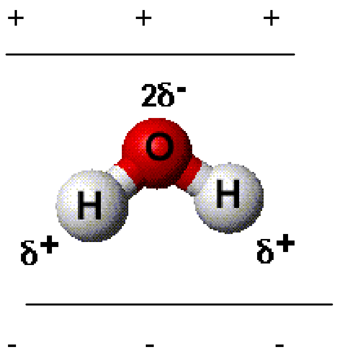
Table 6-4 **Heats of Vaporization of Some Metals (kJ/mol)**

Period	*Element*		
<u>Second</u>	Li 147	Be 297	
<u>Third</u>	Na 97	Mg 128	Al 294
<u>Fourth</u>	K 77	Ca 155	Sc 333
<u>Fifth</u>	Rb 76	Sr 137	Y 365
<u>Sixth</u>	Cs 64	Ba 140	La 402

6-5 Molecular Geometry

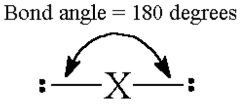
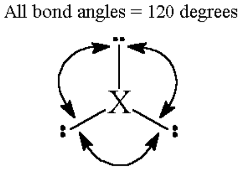
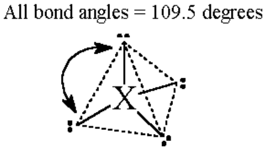
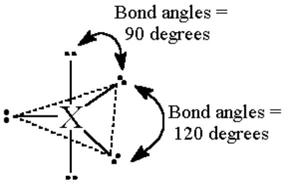
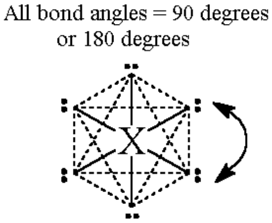
I. <u>VSEPR (Valence Shell Electron Pair Repulsion) Theory</u>

- A. Molecular Polarity
  - 1. The uneven distribution of molecular charge
  - 2. Molecules with preferential orientation in an electric field

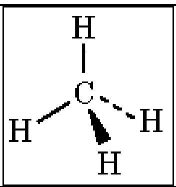
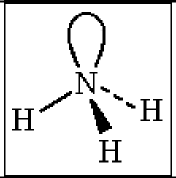
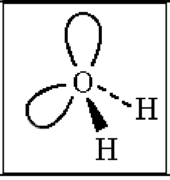


- B. VSEPR Theory
  - 1. Repulsion between the sets of valence-level electrons surrounding an atom causes these sets to be oriented as far apart as possible
- C. VSEPR and Unshared Electron Pairs
  - 1. Unshared pairs take up positions in the geometry of molecules just as atoms do
  - 2. Unshared pairs have a relatively greater effect on geometry than do atoms
  - 3. Lone (unshared) electron pairs require more room than bonding pairs (they have greater repulsive forces) and tend to compress the angles between bonding pairs
  - 4. Lone pairs do not cause distortion when bond angles are 120° or greater

Arrangement of Electron Pairs Around an Atom Yielding Minimum Repulsion

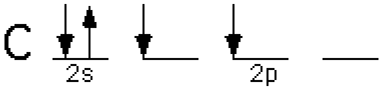
# of Electron Pairs	Shape	Arrangement of Electron Pairs
2	Linear	
3	Trigonal Planar	
4	Tetrahedral	
5	Trigonal bipyramidal	
6	Octahedral	

Comparison of Tetrahedral Bond Angles

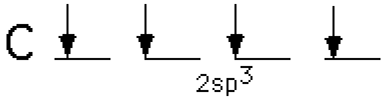
Compound	Structure	Angle between Hydrogens
Methane		109.5°
Ammonia		107°
Water		104.5°

II. Hybridization

- A. Hybridization
  - 1. The mixing of two or more atomic orbitals of similar energies on the same atom to produce new orbitals of equal energies
- B. Hybrid Orbitals
  - 1. Orbitals of equal energy produced by the combination of two or more orbitals on the same atom



Carbon's expected configuration



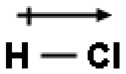
Carbon's hybridized configuration

Table 6-6 **Geometry of Hybrid Orbitals**

Atomic Orbitals	Type of hybridization	# of hybrid orbitals	Geometry
s, p	*sp*	2	Linear
s, p, p	sp^2	3	Trigonal-planar
s, p, p, p	sp^3	4	Tetrahedral

III. Intermolecular Forces

- A. Intermolecular Forces
  - 1. Forces of attraction between molecules
    - a. Generally weaker than bonds that join atoms in molecules
    - b. Boiling point gives a rough estimate of intermolecular forces
      - (1) high bp = large attractive forces
      - (2) low bp = small attractive forces
- B. Molecular Polarity and Dipole-Dipole Forces
  - 1. Dipole
    - a. Created by equal but opposite charges that are separated by a short distance
    - b. A dipole is represented by an arrow with a head pointing toward the negative pole and a crossed tail situated at the positive pole



  - 2. Dipole-Dipole forces
    - a. The negative region of one molecule is attracted to the positive region of another molecule
    - b. A polar molecule can induce a dipole in a nonpolar molecule by temporarily attracting its electrons
- C. Hydrogen Bonding
  - 1. The intermolecular force in which a hydrogen atom that is bonded to a highly electronegative atom is attracted to an unshared pair of electrons of an electronegative atom in a nearby molecule

        a. Usually represented by dotted lines

D. London Dispersion Forces
    1. The intermolecular attractions resulting from the constant motion of electrons and the creation of instantaneous dipoles
        a. Electrons moving in an atom may create temporary unbalanced distribution of charge
        b. All molecules experience London forces
        c. London forces are the only forces of attraction among noble-gas atoms, nonpolar, and slightly polar molecules
    2. London forces increase with the number of electrons in an atom or molecule
        a. Greater mass = Greater London forces

Table 6-7 **Boiling Points and Bonding Types**

Bonding Type	*Substance*	*Bp (1 atm, ° C)*
Nonpolar covalent	H_2	-253
	O_2	-183
	Cl_2	-34
	Br_2	59
	CH_4	-164
	CCl_4	77
	C_6H_6	80
Polar covalent	PH_3	-88
	NH_3	-33
	H_2S	-61
	H_2O	100
	HF	20
	HCl	-85
	ICl	97
Ionic	NaCl	1413
	MgF_2	2239
Metallic	Cu	2567
	Fe	2750
	W	5660