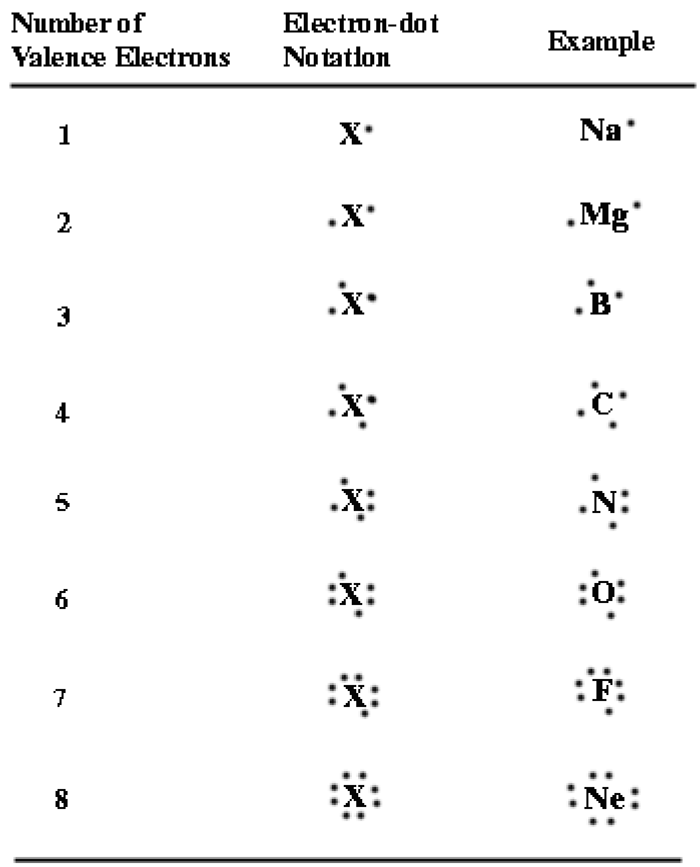
Unit 3 - Chemical Bonding and Molecular Structure

7.1 Ions

I. Valence Electrons – Outer energy level electrons

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



B. The Octet Rule – Ionic Compounds

1. Ionic compounds tend to form so that each atom, by gaining, or losing electrons, has an octet of electrons in its highest occupied energy level

II. Formation of Ions

A. Electron Configuration Changes

1. Cations LOSE their valence electrons to attain a noble-gas configuration
2. Formation of a sodium ion

$$Na = 1s^2 2s^2 2p^6 3s^1 \rightarrow Na^+ = 1s^2 2s^2 2p^6$$

3. Anions GAIN electrons to complete their valence shell noble-gas configuration
4. Formation of chloride ion

$$Cl = 1s^2 2s^2 2p^6 3s^2 3p^5 \rightarrow Cl^- = 1s^2 2s^2 2p^6 3s^2 3p^6$$

Common Ions and Their Charges

Monatomic Cations	Name
H^+	Hydrogen
Li^+	Lithium
Na^+	Sodium
K^+	Potassium
Mg^{2+}	Magnesium
Ca^{2+}	Calcium
Ba^{2+}	Barium
Al^{3+}	Aluminum

Monatomic Anions	Name
F^-	Fluoride
Cl^-	Chloride
Br^-	Bromide
I^-	Iodide
O^{2-}	Oxide
S^{2-}	Sulfide
N^{3-}	Nitride
P^{3-}	Phosphide

7.2 Ionic Bonds 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. $Na = 3s^1$ $Cl = 3s^23p^5$
     - b. $Na^+ = 2s^22p^6$ $Cl^- = 3s^23p^6$

Oppositely charged ions come together in a ratio that produces a net charge = 0

Na^+Cl^-

III. Properties of Ionic Compounds

	Ionic Compounds
Structure	Crystalline solids
Melting point	Generally high
Boiling Point	Generally high
Electrical Conductivity	Excellent conductors, molten and aqueous
Solubility in water	Generally soluble

7.3 Bonding in Metals

I. The Metallic Bond Model

- 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. Metallic Properties

- 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
- E. Metals atoms organized in compact, orderly crystalline patterns
- F. Different metallic elements (and carbon) can be mixed to form alloys
  1. Sterling silver
     - a. Ag = 92.5%, Cu = 7.5%
  2. Brass
     - a. Cu = 60%, Zn = 40%

8.1 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

8.2 The Nature of Covalent Bonding

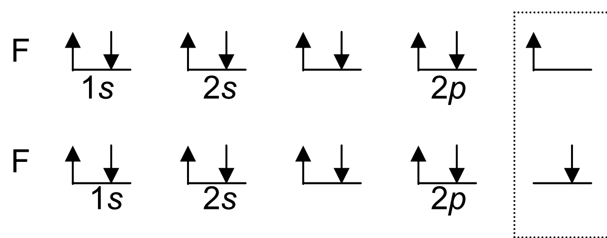
I. The Octet Rule in Covalent Bonding

A. Covalent compounds tend to form so that each atom, by sharing electrons, has an octet of electrons in its highest occupied energy level

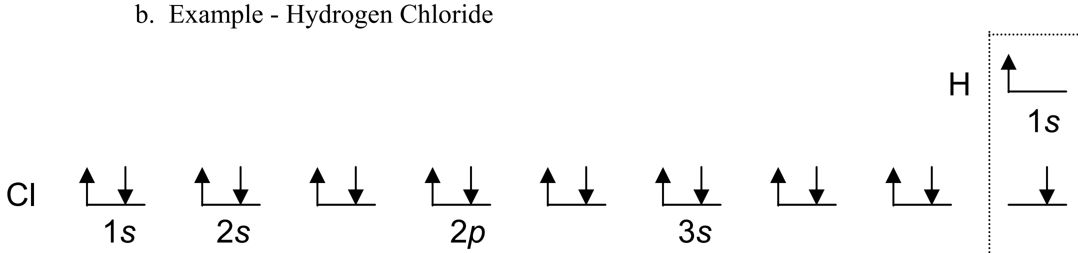
B. Single Covalent Bonds

1. One shared pair of electrons between two atoms

a. Example – Diatomic fluorine



b. Example - Hydrogen Chloride



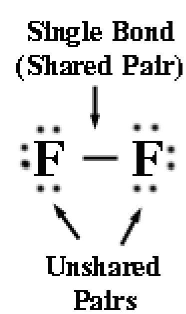
II. Lewis Structures

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

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

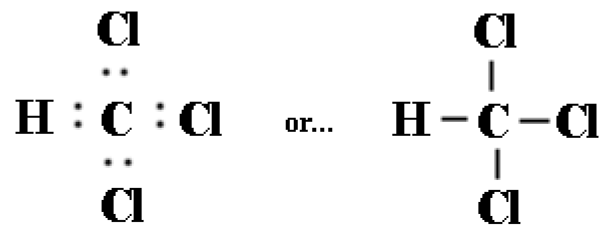
2. Write the electron dot notation for each type of atom in the molecule

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

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

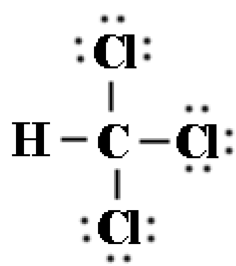
4. Arrange the atoms to form a skeleton structure for the molecule. If carbon is present, it is the central atom. Otherwise, the least electronegative element atom is central (except for hydrogen, which is never central). Then connect the atoms by electron-pair bonds

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

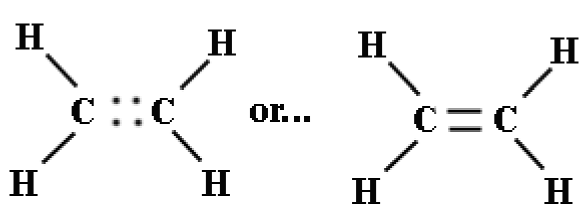


III. Multiple Covalent Bonds

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)

$$H-C\equiv C-H$$

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

8.3 Bonding Theories

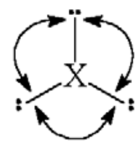
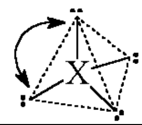
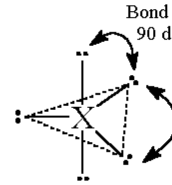
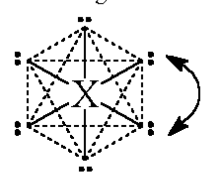
I. VSEPR (Valence Shell Electron Pair Repulsion) Theory

A. 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

B. 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	Bond angle = 180 degrees 
3	Trigonal Planar	All bond angles = 120 degrees 
4	Tetrahedral	All bond angles = 109.5 degrees 
5	Trigonal bipyramidal	Bond angles = 90 degrees  Bond angles = 120 degrees
6	Octahedral	All bond angles = 90 degrees or 180 degrees 

8.4 Polar Bonds and Molecules

I. <u>Bond Polarity</u>
   A. Nonpolar Covalent Bond
      1. 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
      1. A covalent bond in which the bonded atoms have an unequal attraction for the shared electrons and a resulting unbalanced distribution of charge

II. <u>Molecular Polarity</u>
      1. The uneven distribution of molecular charge
      2. Molecules with preferential orientation in an electric field