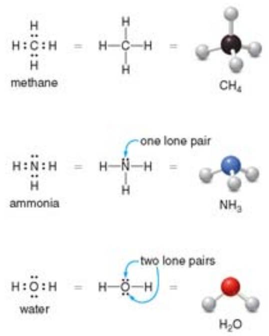
Covalent Molecules: Structures and Properties

Covalent bond: sharing of electrons between two atoms

Lewis Dot Structure: Specifies an element and uses dots to show only the valence electrons

Examples: Mg: Na·

Lewis Dot Structures of Covalent Molecules



Drawing Lewis Dot Structures

1. **Establish a skeletal structure**
   a. least electronegative atom in center
   b. H, F often at ends, seldom in center
   c. carbon likes to form C-C chains

2. **Determine the total number of valence electrons in the entire compound**
   a. add up valence electrons for each atom
   b. add an electron for each negative charge
   c. subtract an electron for each positive charge

3. **Place a pair of electrons between bonded atoms**

4. **Move electron pairs around as needed to satisfy the octet rule for the central atom (double or triple bonds may be necessary)**

5. **Double check that <u>each</u> atom has a full octet <u>and</u> that the total number of electrons shown matches that calculated in step #2**

Covalent Bond: A <u>sharing</u> of electrons between 2 atoms to fulfill the octet rule

Examples: H· + H· → H_2; HF, H_2O

We can also use a dash (-) to represent a covalent bond between atoms: H-F, H-O-H

Double Bond: Bond resulting from the sharing of <u>2</u> electrons; stronger than a single bond

Example: CO_2 O=C=O

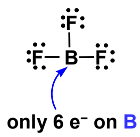
Triple Bond: Bond resulting from the sharing of <u>3</u> electrons; stronger than a double bond

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

•Most of the common elements generally follow the octet rule.

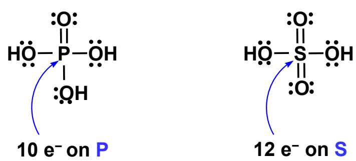
•H is a notable exception, because it needs only 2 e^- in bonding.

•Elements in group 3A do not have enough valence e^- to form an octet in a neutral molecule.



•Elements in the third row have empty *d* orbitals available to accept electrons.

•Thus, elements such as P and S may have more than 8 e^- around them.



Naming Covalent Compounds

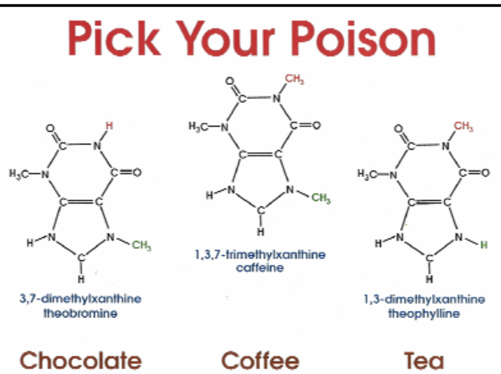
Covalent bonds usually occur between non-metals

- Name of the less electronegative element first
- Stem of the name of the more electronegative element plus the "ide" suffix
- Indicate the number of each type of atom by using the Greek prefix (mono, di, tri, etc.) (omit "mono" from the first element)

Examples: N_2O_5 = Dinitrogen pentoxide

CO_2 = Carbon Dioxide, CCl_4 = Carbon tetrachloride

Practice: BF_3 = ? SO_3 = ?



Molecular Geometry: VSEPR Theory

Covalent compounds exist as 3-dimensional structures; we can predict these shapes based on electron arrangement

VSEPR Theory: Valence-Shell Electron Pair Repulsion theory; electron pairs in the valence shell repel each other and try to get as far away from each other as possible

Central Atom: any atom in a molecule or ion that is bonded to 2 or more other atoms

Electron pairs around a central atom orient themselves as far away from each other as possible

Linear: 2 pairs of electrons around central atom, orient on opposite sides of the central atom (180^{o} apart)

Trigonal Planar: 3 pairs of electrons around central atom, orient in a triangle around the central atom (120^{o} apart)

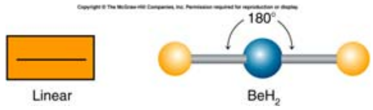
Tetrahedral: 4 bonded pairs of electrons orient around the central atom in the middle (109.5^{o} apart)

Trigonal pyramidal: 3 bonded pairs and 1 lone pair of electrons orient in a tetrahedron around the central atom (107^{o} apart)

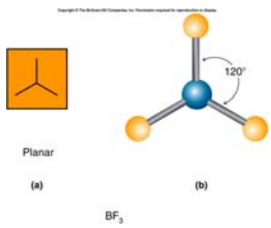
Bent/Angular: 2 bonded pairs of electrons and 2 lone pairs of electrons orient in a tetrahedron around the central atom (104.5^{o} apart)

Examples of Covalent Molecular Shapes

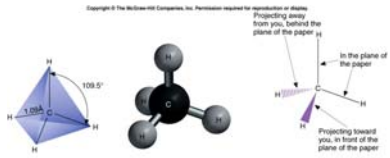
Linear



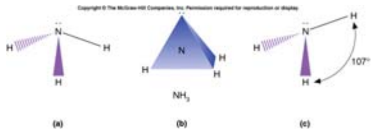
Trigonal planar



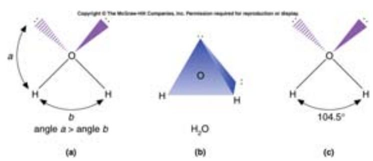
Tetrahedral



Trigonal Pyramid



Bent/Angular



When counting the number of electron pairs surrounding a central atom, remember:

- All valence shell electron pairs count, whether they are bonding pairs or lone pairs
- Double or triple bonds are counted like a single pair of electrons when predicting shapes

Determining Molecular Shape from Lewis Dot Structure

Table 4.5 Molecular Structure: The Geometry of a Molecule Is Affected by the Number of Nonbonded Electron Pairs Around the Central Atom and the Number of Bonded Atoms

Bonded Atoms	Nonbonding Electron Pairs	Bond Angle	Molecular Structure	Example
2	0	180°	Linear	CO_2
3	0	120°	Trigonal planar	SO_3
2	1	120°	Angular	SO_2
4	0	~109°	Tetrahedral	CH_4
3	1	<107°	Trigonal pyramidal	NH_3
2	2	<104.5°	Angular	H_2O

<u>Polarity of Covalent Bonds</u>

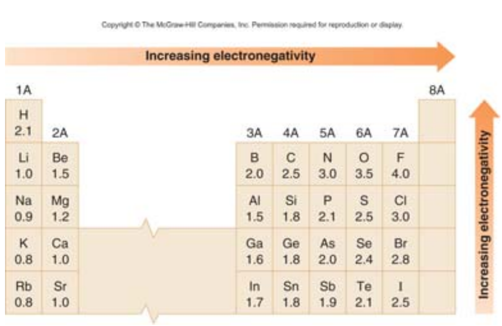
Not all atoms attract electrons to the same degree

Electronegativity (E_n): Tendency of an atom to attract shared electrons of a covalent bond

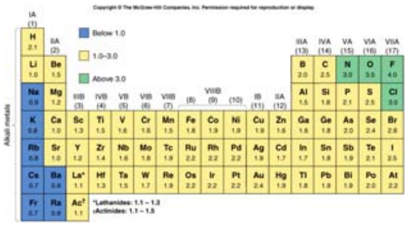
Electronegativity increases <u>across</u> Periodic Table,

decreases <u>down</u> Periodic Table

Electronegativity and the Periodic Table



Electronegativity Values



Nonpolar Covalent Bond: Covalent bond in which the bonding electrons are shared equally by both atoms

Example: Cl-Cl, H-H

Polar Covalent Bond: Covalent bond in which the bonded electrons are shared unequally

Example: H-Cl, H-O-H

This polarization (unequal distribution of electrons) results in "partial charges" on the atoms within the covalently bonded molecule, while the net charge on the molecule remains zero.

The more electronegative atom acquires a partial negative charge (δ^-); the less electronegative atom acquires a partial positive charge (δ^+)

Example: H-Cl

H has less electronegativity than Cl

Therefore electrons are more attracted to Cl atom

δ^+ δ^-
H-Cl

Practice: I-Cl

The greater the difference in electronegativity between bonded atoms, the more polar the bond

TABLE 4.3 Electronegativity Difference and Bond Type

Electronegativity Difference	Bond Type	Electron Sharing
Less than 0.5 units	Nonpolar	Electrons are equally shared.
0.5–1.9 units	Polar covalent	Electrons are unequally shared; they are pulled towards the more electronegative element.
Greater than 1.9 units	Ionic	Electrons are transferred from the less electronegative element to the more electronegative element.

Lewis Dot Structures and Polarity of Molecules

Whole molecules can also be described as either polar or nonpolar

Polar molecules: molecules with uneven charge distribution; a positive end and a negative end to the molecule (dipole)

Nonpolar molecules: molecules whose charge distribution is symmetrical within the molecule; no positive or negative part

Examples: H_2 is a nonpolar molecule

HCl is a polar molecule

H_2O is a very polar molecule

Polarity affects many properties (solubility, melting point, boiling point)

Intermolecular forces: interactions between different molecules

Intramolecular forces: interactions within a molecule; the bonds between atoms that hold a molecule together

Molecules with more intermolecular forces have higher melting and boiling points (ionic, polar covalent, nonpolar covalent)

Summary and Key Points

Ionic bond vs. Covalent bond

Polar vs. nonpolar covalent bond

Electronegativity

Lewis Dot structures!

Writing formulas for ionic compounds

Naming ionic and covalent compounds

VSEPR theory, molecular shape, polarity