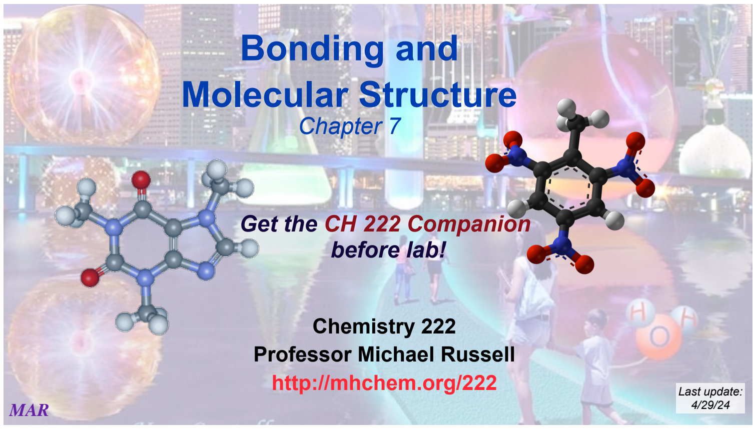
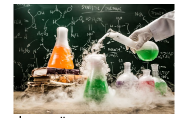


CH 222: Lectures and Labs

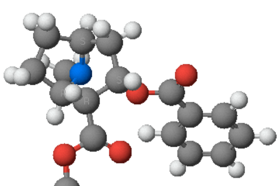


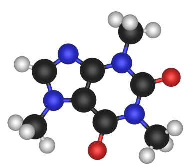
Lectures: **MWF** from **9 - 9:50 AM** in **AC 1303** *(this room)*
- Lectures recorded, available soon afterwards
- Lecture notes to print available (under "Problem Sets and Handouts", mhchem.org/222), get **CH 222 Companion** as soon as possible

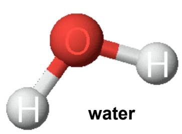
Labs (Section 01): **Mondays** from **1:10 - 5 PM**
- **Start** in **room AC 2501** (not AC 1303)
- Move to **AC 2507** ("the lab") around **3 PM**
- *For first day*, **bring a printed copy of the "Chromatography" Lab** (mhchem.org/222), a pair of **safety glasses** (Dollar store ok) and your **calculator**

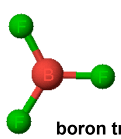
MAR *...more on Monday afternoon*

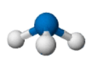
Chemical Bonding


cocaine


caffeine


water


boron trifluoride


ammonia

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Chemical Bonding

Problems and questions:

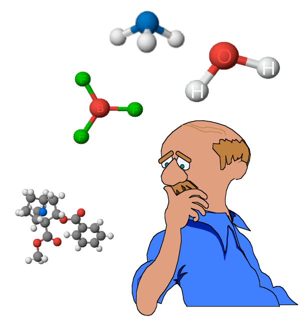
How is a molecule or polyatomic ion held together?

Why are atoms distributed at strange angles?

Why are molecules not flat?

Can we predict the structure?

How is structure related to chemical and physical properties?



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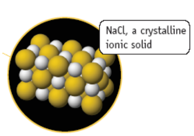
Two Extreme Forms of Chemical Bonds

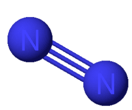


Ionic - complete transfer of electrons from one atom to another, metals + nonmetals

$$\mathrm{Na}\cdot + \cdot\ddot{\underset{..}{\mathrm{Cl}}}: \longrightarrow [\mathrm{Na}\cdot \quad \cdot\ddot{\underset{..}{\mathrm{Cl}}}:] \longrightarrow [\mathrm{Na}^{+} \quad :\ddot{\underset{..}{\mathrm{Cl}}}:^{-}]$$

metal atom — nonmetal atom — electron transfer from reducing agent to oxidizing agent — ionic compound. Ions have noble gas electron configurations.

Covalent - electrons shared between atoms, mostly nonmetals

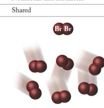
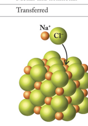
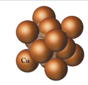
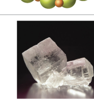


Most bonds are somewhere in between

Also* Metallic - *for metals, studied later

MAR *December 2018: Metavalent bonding (for metalloids!)*

Bonding Overview

Types of Chemical Bonds	Covalent	Ionic	Metallic
Elements involved:	Nonmetals and metalloids	Metals and nonmetals	Metals
Electron distribution:	Shared	Transferred	Pooled
Microview:			
Macroview:			
	Mostly just for nonmetals	*metals plus nonmetals*	*metals and alloys*

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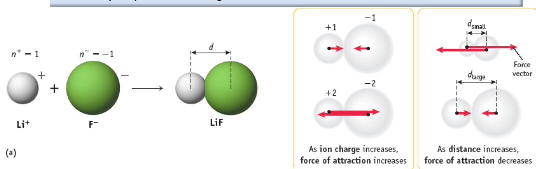
Ionic Forces - Coulomb's Law

CH 221 Flashback:

charge on + and − ions; charge on electron

$$\text{Force} = -k\frac{(n^{+}e)(n^{-}e)}{d^{2}}$$

proportionality constant; distance between ions



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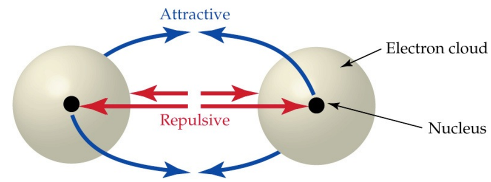
Higher ionic force, higher melting point, etc.

Covalent Bonding

Covalent bonds arise from the mutual attraction of 2 nuclei for the same electrons.



A covalent bond is a balance of attractive and repulsive forces.

Interatomic Interactions

H• •H

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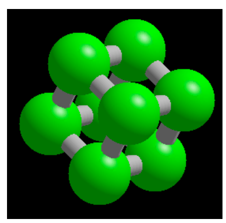
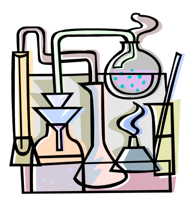
Covalent Bonding

- **Covalent bonding will be the focus of the first two chapters**
- **We will re-visit ionic bonding and Metallic bonding in a future chapter**
- **Important to know when a compound is ionic, covalent or metallic!**





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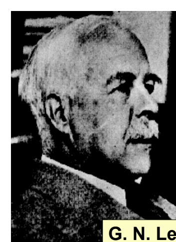
Electron Distribution in Molecules



G. N. Lewis 1875 - 1946

- **Electron distribution is depicted with Lewis electron dot structures**
- **Valence electrons are distributed as shared or BOND PAIRS and unshared or LONE PAIRS.**

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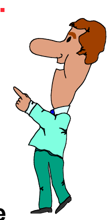
Bond and Lone Pairs

Valence electrons are distributed as shared or BOND PAIRS and unshared or LONE PAIRS.

H—Cl:

shared or bond pair

lone pair (LP)



This is called a LEWIS ELECTRON DOT structure

3 lone pairs + 1 bond pair = 4 pairs total

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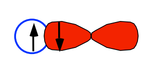
Bond Formation

A bond can result from a "head-to-head" overlap of atomic orbitals on neighboring atoms, making a sigma (σ) bond.

H• + •Cl: ⟶ H:Cl:



Overlap of H (1s) and Cl (3p)

Note that each atom has a single, unpaired electron in their atomic orbital.

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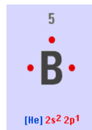
Valence Electrons

Electrons are divided between **core** and **valence electrons**

B $1s^2 2s^2 2p^1$

Core = [He] , valence = $2s^2 2p^1$

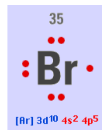
2 core e^-, 3 valence e^-



Br [Ar] $3d^{10} 4s^2 4p^5$

Core = [Ar] $3d^{10}$, valence = $4s^2 4p^5$

28 core e^-, 7 valence e^-



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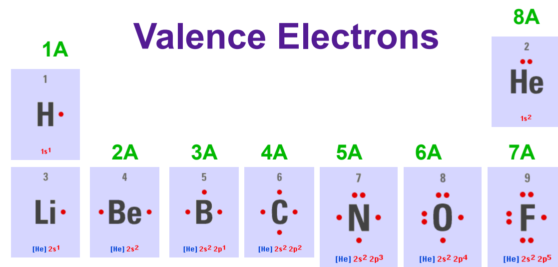
Valence Electrons



∴ Number of valence electrons is equal to the Group number

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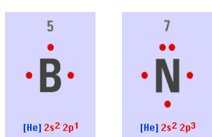
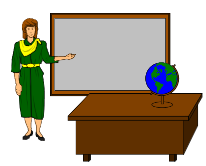
Building Lewis Structures

No. of valence electrons of a main group atom = Group number

For Groups 1A - 4A, no. of bond pairs = group number.

For Groups 5A - 7A, BPs = 8 - Grp. No.





MAR

Building a Lewis Dot Structure

No. of valence electrons of an atom = Group number

For Groups 1A - 4A (14), no. of bond pairs = group number

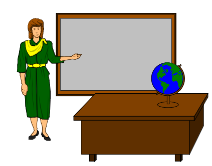
For Groups 5A (15) - 7A (17), BPs = 8 - Grp. No.

Except for H (and sometimes atoms of 3rd and higher periods),

BPs + LPs = 4

This observation is called the

OCTET RULE



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Building a Lewis Dot Structure

Ammonia, NH_3

1. Count valence electrons

   H = 1 and N = 5

   Total = (3 x 1) + 5

   = 8 electrons *or*

   4 pairs of electrons

2. Decide on the central atom; never H.

   Central atom is atom of lowest affinity for electrons.

   Therefore, N is central

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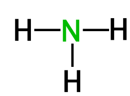
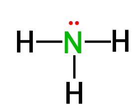
Building a Lewis Dot Structure

3. Form a sigma bond *(single bond)* between the central atom and surrounding atoms.

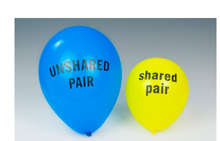


4. Remaining electrons form LONE PAIRS to complete octet as needed.



3 BOND PAIRS and 1 LONE PAIR. ***Note that N has a share in 4 pairs (8 electrons), while H shares 1 pair.***

See Building Lewis Structures handout



Unshared electron pairs ("lone pairs") take up more volume than shared electron pairs ("bonding pairs")

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Sulfite ion, SO_3^{2-}

Step 1. Central atom = S

Step 2. Count valence electrons

S = 6

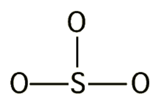
3 x O = 3 x 6 = 18

Negative charge = 2

TOTAL = 26 e- or 13 pairs

Step 3. Form sigma bonds

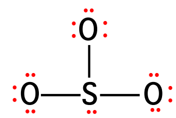
10 pairs of electrons are now left.



MAR

Sulfite ion, SO_3^{2-}

Remaining pairs become lone pairs, first on outside atoms and then on central atom.



Each atom is surrounded by an octet of electrons.

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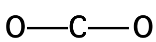
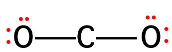
Carbon Dioxide, CO_2

1. Central atom = _______
2. Valence electrons = __ or __ pairs
3. Form sigma bonds.



This leaves 6 pairs.

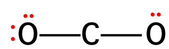
4. Place lone pairs on outer atoms.



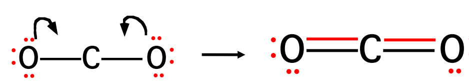
MAR

Carbon Dioxide, CO_2

4. Place lone pairs on outer atoms.



5. So that C has an octet, we shall form DOUBLE BONDS between C and O.



The second bonding pair forms a pi (π) bond.

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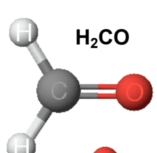
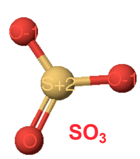
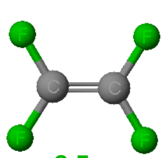
Double and even triple bonds are commonly observed for C, N, P, O, and S



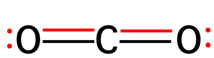
H_2CO



SO_3



C_2F_4



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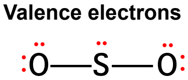
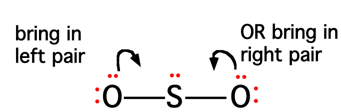
Sulfur Dioxide, SO_2

1. Central atom = S
2. Valence electrons = 18 or 9 pairs



3. Form pi (π) bond so that S has an octet - but note that there are two ways of doing this.



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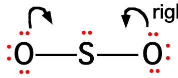
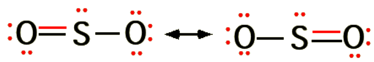
Sulfur Dioxide, SO_2

bring in left pair

OR bring in right pair



This leads to the following structures.



These equivalent structures are called **RESONANCE STRUCTURES**. The true electronic structure is a **HYBRID** of the two.

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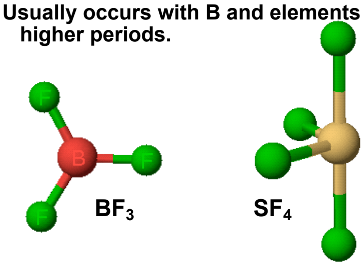
Violations of the Octet Rule

Usually occurs with B and elements of higher periods.



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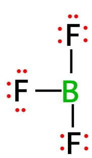
Boron Trifluoride

Central atom =

Valence electrons =
or electron pairs =

Assemble dot structure



The B atom has a share in only 6 electrons (or 3 pairs). B atom in many molecules is electron deficient.

Also common for Al and Be

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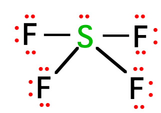
Sulfur Tetrafluoride, SF_4

Central atom =

Valence electrons = ___ or ___ pairs.

Form sigma bonds and distribute electron pairs.



5 pairs around the S atom. A common occurrence outside the 2nd period.

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Odd # of electrons: NO_2

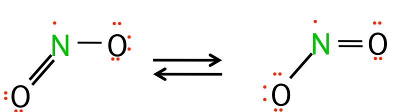
Paramagnetic compounds & free radicals

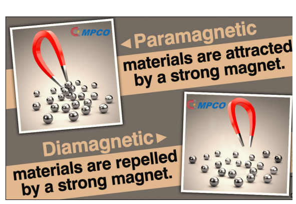
For NO_2, central atom = ___

Valence electrons = ___ or ___ pairs.

Odd e$^-$ occupies its own "space"

Form sigma bonds and distribute electron pairs.





Paramagnetic substances often more reactive than diamagnetic substances

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Formal Atom Charges

Atoms in molecules often bear a charge (+ or -).

The *predominant resonance structure* of a molecule is the one with *charges as close to 0* as possible.

Formal charge = Group number - $^1/_2$ (number bonding electrons) - (number lone pair electrons (lpe)),

or

FC = GN - bonds - lpe

Sum of all formal charges in a molecule **must equal ionic charge**

See Guide to Formal Charges

MAR

Formal Atom Charges

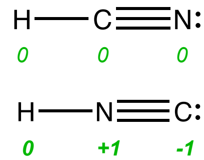
Formal charge = Group number - $^{1}/_{2}$ (number bonding electrons) - (number lone pair electrons)

or

FC = GN - bonds - lpe

Formal Charges In Isomers

more stable structure



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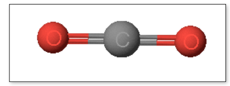
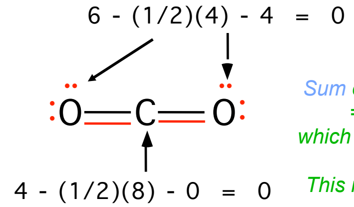
Carbon Dioxide, CO_2



FC = GN - bonds - lpe



*Sum of formal charges = = 0 + 0 + 0 = **0** which equals ionic charge on molecule*

This is a good structure!

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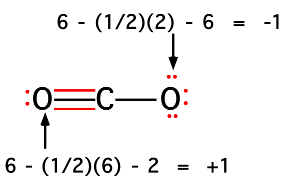
Formal Charge Comparison with CO_2

FC = GN - bonds - lpe



C atom formal charge is still 0.

***Sum** of formal charges = = -1 + 0 + 1 = **0** which equals ionic charge on molecule*

This is a valid resonance form, but not as good as previous slide

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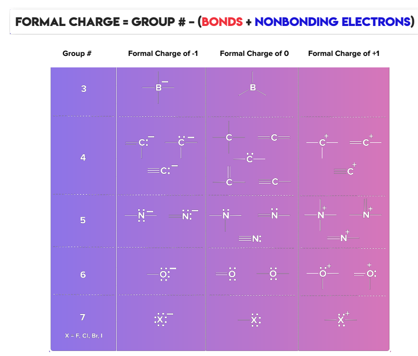
Formal Charge



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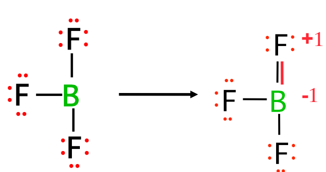
Boron Trifluoride, BF_3



What if we form a B-F double bond to satisfy the B atom octet?

F *never* makes double bonds!

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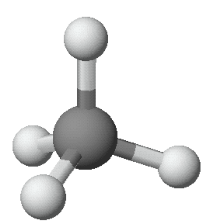
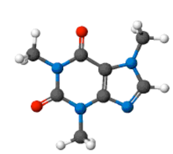
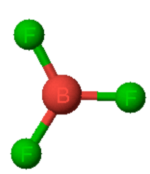
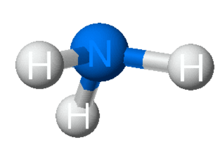
Molecular Geometry









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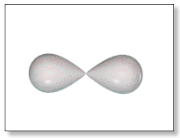
MOLECULAR GEOMETRY

VSEPR

Valence **S**hell **E**lectron **P**air **R**epulsion theory.

Most important factor in determining geometry is relative repulsion between electron pairs.



Molecule adopts the shape that minimizes the electron pair repulsions.

See Geometry and Polarity Guide

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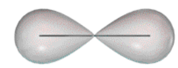
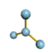
No. of e- Pairs Around Central Atom	Example	Geometry
2	F—Be—F 180˚	linear

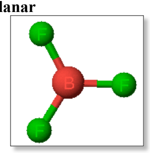


MAR See Geometry and Polarity Guide

VSEPR

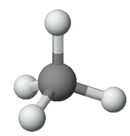
No. of e- Pairs Around Central Atom	Example	Electron Pair Geometry
2	F—Be—F 180˚	linear
3	BF_3 120˚	trigonal planar



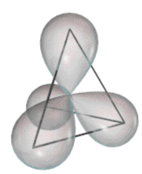


MAR See Geometry and Polarity Guide

VSEPR



No. of e- Pairs Around Central Atom	Example	Electron Pair Geometry
2	F—Be—F 180˚	linear
3	BF_3 120˚	trigonal planar
4	CH_4 109˚	tetrahedral



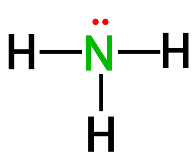
MAR See Geometry and Polarity Guide

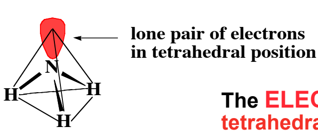
VSEPR

Structure Determination by VSEPR

Ammonia, NH_3

1. Draw electron dot structure
2. Count BPs and LPs = 4
3. The 4 electron pairs are at the corners of a tetrahedron.





The ELECTRON PAIR GEOMETRY is tetrahedral.

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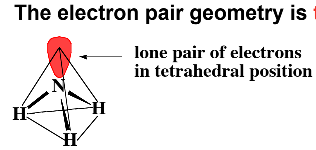
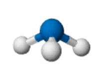
Structure Determination by VSEPR

Ammonia, NH_3

The electron pair geometry is tetrahedral.





The MOLECULAR GEOMETRY - the positions of the atoms - is TRIGONAL PYRAMID

See Geometry and Polarity Guide

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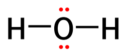
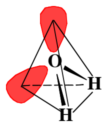
Structure Determination by VSEPR

Water, H_2O

1. Draw electron dot structure
2. Count BPs and LPs = 4
3. The 4 electron pairs are at the corners of a tetrahedron.





The electron pair geometry is TETRAHEDRAL

A caltrop

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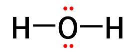
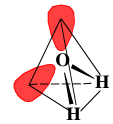
Structure Determination by VSEPR

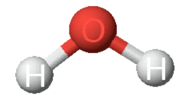
Water, H_2O





The electron pair geometry is TETRAHEDRAL

The molecular geometry is bent



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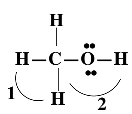
Structure Determination by VSEPR

Methanol, CH_3OH

1. Draw electron dot structure



2. Define bond angles 1 and 2

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Structure Determination by VSEPR

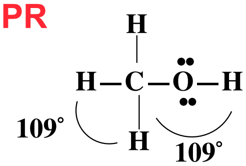
Methanol, CH_3OH

Define bond angles 1 and 2

Angle 1 = 109°

Angle 2 = 109°

In both cases the atom is surrounded by 4 electron pairs.



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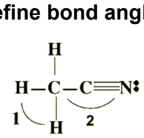
Structure Determination by VSEPR

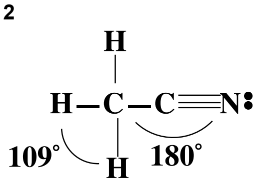
Acetonitrile, CH_3CN

Define bond angles 1 and 2



Angle 1 = 109°

Angle 2 = 180°



One C is surrounded by 4 electron "clouds" and the other by 2 "clouds"

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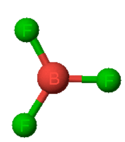
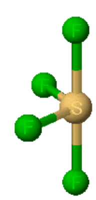
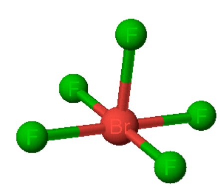
STRUCTURES WITH CENTRAL ATOMS THAT DO NOT OBEY THE OCTET RULE







Usually occurs with Group 3A elements and with those of 3rd period and higher.

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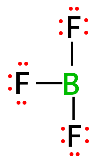
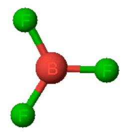
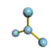
Compounds with 3 Pairs Around the Central Atom

The B atom is surrounded by only 3 electron pairs.
Bond angles are 120°



Geometry described as planar trigonal or trigonal planar





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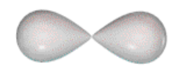
Compounds with 5 or 6 Pairs Around the Central Atom



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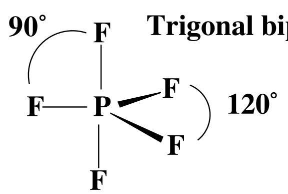
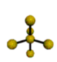
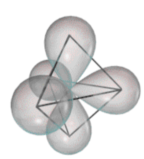
Compounds with 5 or 6 Pairs Around the Central Atom



Trigonal bipyramid

5 electron pairs





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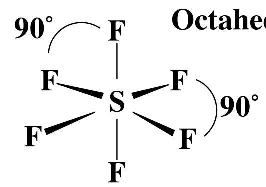
See *Geometry and Polarity Guide*

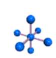
Compounds with 5 or 6 Pairs Around the Central Atom

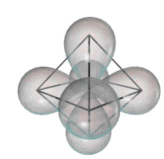


Octahedron

6 electron pairs

The game "Jacks"





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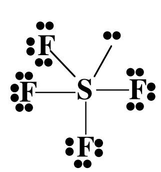
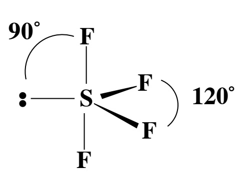
See *Geometry and Polarity Guide*

Sulfur Tetrafluoride, SF_4

Number of valence electrons = 34
Central atom = S
Dot structure



Electron pair geometry
= trigonal bipyramid
(because there are 5 pairs around the S)



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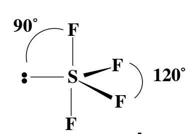
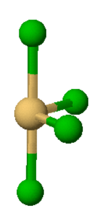
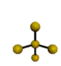
Sulfur Tetrafluoride, SF_4

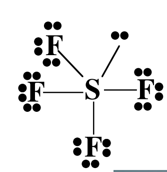
Molecular geometry = seesaw





Lone pair is in the equator because it requires more room.





Unshared electron pairs ("lone pairs") take up more volume than shared electron pairs ("bonding pairs")

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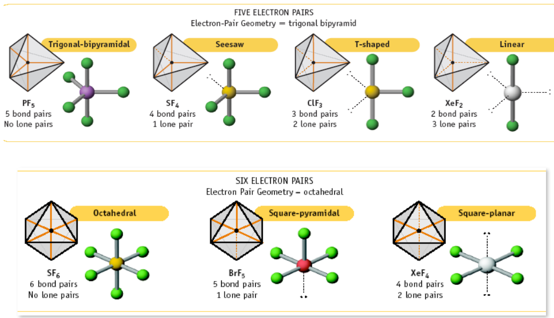
Other Molecular Geometries



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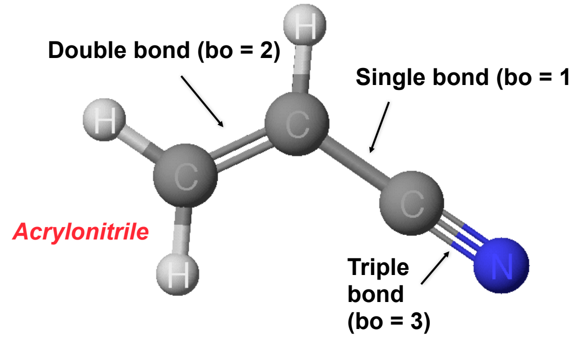
Bond Order

\# of bonds between a pair of atoms



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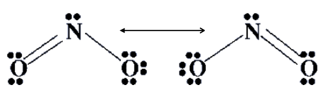
Bond Order

Fractional bond orders occur in molecules with resonance structures.

Consider NO_2^-



$$\text{Bond order} = \frac{\text{Total \# of e - pairs used for a type of bond}}{\text{Total \# of bonds of that type}}$$

$$\text{Bond order} = \frac{\text{3 e - pairs in N — O bonds}}{\text{2 N — O bonds}}$$

The N-O bond order = 1.5

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Bond Order

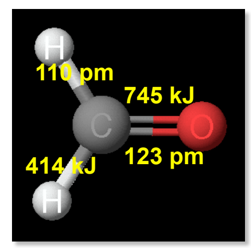
Bond order is related to two important bond properties:

(a) bond length

(b) bond energy

Bond length *is inversely proportional to bond order*

Bond energy *is* proportional *to bond order*



Bond lengths measured in pm (1 pm = 10^{-12} m) or Angstroms (1 Å = 10^{-10} m)

Bond energies measured in kJ (1 kJ = 10^3 J)

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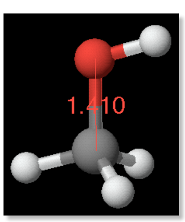
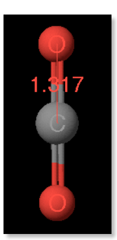
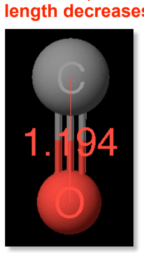
Bond Length





Bond length depends on **bond order**

As **bond order increases, bond length decreases**



Bond length is inversely proportional to bond order

Bond distances measured in Angstroms (1 Å = 10^{-10} m)

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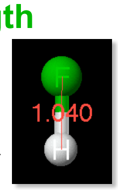
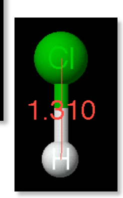
Bond Length

Bond length, the distance between two nuclei, also depends on the **size of bonded atoms**.

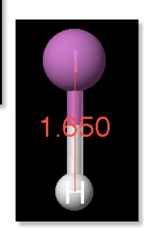
Bond length is inversely proportional to bond order



H-F



H-Cl



H-I

Bond distances measured in Angstroms (1 Å = 10^{-10} m)

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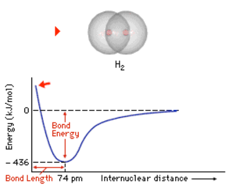
Bond Energy is measured by the energy required to break a bond

The **GREATER** the number of bonds (*bond order*) the **LARGER** the bond energy and the **SHORTER** the bond.



Average Bond Enthalpies (kJ/mol)

Single Bonds

C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146		
C—N	293	N—O	201	O—F	190	Cl—F	253
C—O	358	N—F	272	O—Cl	203	Cl—Cl	242
C—F	485	N—Cl	200	O—I	234		
C—Cl	328	N—Br	243			Br—F	237
C—Br	276			S—H	339	Br—Cl	218
C—I	240	H—H	436	S—F	327	Br—Br	193
C—S	259	H—F	567	S—Cl	253		
		H—Cl	431	S—Br	218	I—Cl	208
Si—H	323	H—Br	366	S—S	266	I—Br	175
Si—Si	226	H—I	299			I—I	151
Si—C	301						
Si—O	368						
Si—Cl	464						

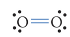
Multiple Bonds

C=C	614	N=N	418	O_2	495
C≡C	839	N≡N	941		
C=N	615	N=O	607	S=O	523
C≡N	891			S=S	418
C=O	799				
C≡O	1072				

Bond Energy

MAR

Bond Energy



121 pm

Bond order =2
498 kJ/mol

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Using Bond Energies

Estimate the energy of the reaction:

H-H + Cl-Cl → 2 H-Cl

Net energy = ΔH_{rxn} = energy required to break bonds - energy evolved when bonds formed

H-H = 436 kJ/mol
Cl-Cl = 243 kJ/mol
H-Cl = 431 kJ/mol

ΔH = bonds broken - bonds formed

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Using Bond Energies

ΔH = bonds broken - bonds formed

Estimate the energy of the reaction: H-H + Cl-Cl ----> 2 H-Cl

H-H = 436 kJ/mol
Cl-Cl = 243 kJ/mol
H-Cl = 431 kJ/mol

H—H

Cl—Cl

ΔH° = +679 kJ

"Bonds broken" *or* "Reactant bonds":
H-H + Cl-Cl bond energies = 436 kJ + 243 kJ = 679 kJ

"Bonds formed" *or* "Product bonds":
2 mol H-Cl bond energies = 2 x 431 kJ = 862 kJ

ΔH = bonds broken - bonds formed
ΔH = 679 kJ - 862 kJ = -183 kJ

***exothermic*!**

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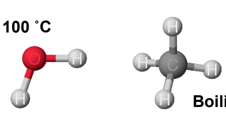
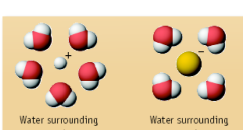
Molecular Polarity

Boiling point = 100 ˚C



Boiling point = -161 ˚C

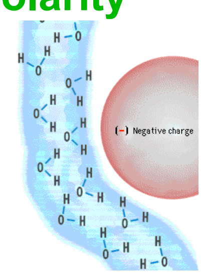


Why do ionic compounds dissolve in water?

Why do water and methane differ so much in their boiling points?

MAR

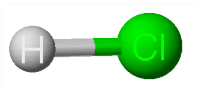
Molecular Polarity



Water molecules are attracted to balloons that have a static electric charge

MAR

Bond Polarity



HCl is **POLAR** because it has a positive end and a negative end (**dipoles**).

$+\delta$ $-\delta$
H–C̈l:

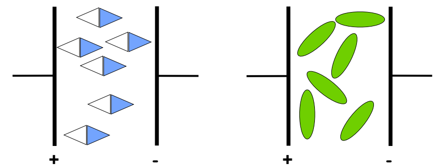
Cl has a greater share in bonding electrons than does H.

Cl has slight negative charge ($-\delta$) and H has slight positive charge ($+\delta$)

MAR

Bond Polarity

Dipole moment, μ, can measure dipole strength by placing molecules in electrical field. Polar molecules will align when the field is on. Nonpolar molecules will not.



Polar molecules *Nonpolar molecules*

MAR

Bond Polarity

$+\delta$ $-\delta$
H–C̈l:

Due to polarity, the H-Cl bond energy is GREATER than expected for a "pure" covalent bond.

BOND	ENERGY
"pure" bond	339 kJ/mol calc'd
real bond	432 kJ/mol measured

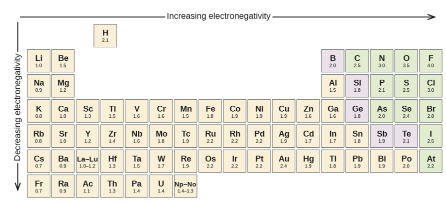
Difference = 92 kJ. This difference is proportional to the difference in **ELECTRONEGATIVITY, χ.**

See Polarity Guide

MAR

Electronegativity, χ

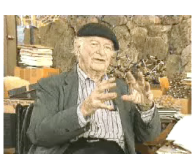
χ is a measure of the ability of an atom in a molecule to attract electrons to itself.

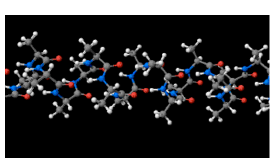


Electronegativities tend to increase up and to the right on the periodic table

MAR

Linus Pauling, 1901-1994



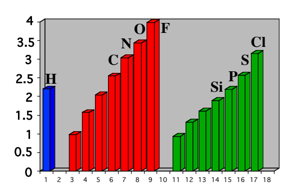


The only person to receive two unshared Nobel prizes (for **Peace** *and* **Chemistry**)

Chemistry areas: bonding, electronegativity, protein structure

A great Oregonian *and* a great Scientist

MAR



Electronegativity, χ

F has maximum χ.

Atom with lowest χ is the center atom in most molecules.

Relative values of χ determine **BOND POLARITY** (and point of attack on a molecule).

We are using "traditional" electronegativity values, but a new system has been introduced (January 2019)

MAR

Bond Polarity

Which bond is more polar (or DIPOLAR)?

	O-H	O-F
χ	3.5 - 2.1	3.5 - 4.0
$\Delta\chi$	1.4	0.5

$\therefore$ **OH is *more polar* than OF**

O—H   O—F

$-\delta$ $+\delta$ $+\delta$ $-\delta$

and polarity is "reversed"

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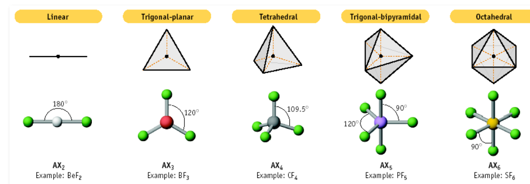
Molecular Polarity

Molecules will be polar *if*

a) bonds are polar

AND

b) the molecule is NOT "symmetric"



All above are symmetric and NOT polar (nonpolar)

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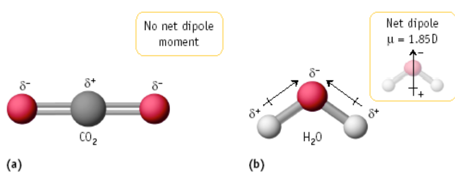
Polar or Nonpolar?

Compare CO_2 and H_2O. Which one is polar?

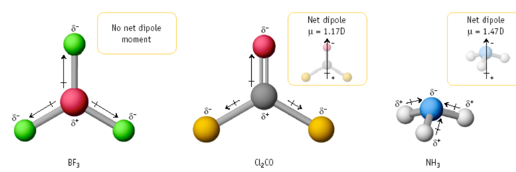


Polar bonds, nonpolar molecule   ***Polar bonds, polar molecule***

MAR

Polar or Nonpolar?

Consider AB_3 molecules: BF_3, Cl_2CO, and NH_3.



Polar and nonpolar molecules of the type AB_3.

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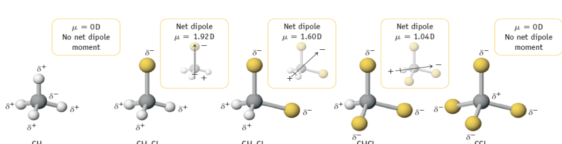
CH_4 through CCl_4 Polarity



Methane (CH_4) and carbon tetrachloride (CCl_4) are symmetrical and NOT polar.

All other compounds asymmetrical and polar.

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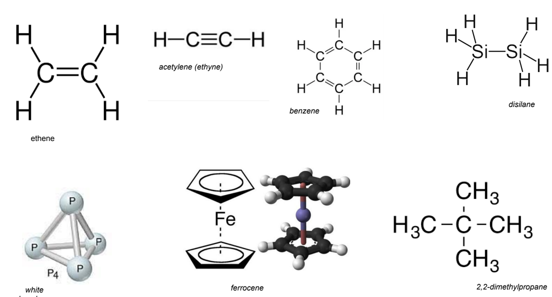
More on Molecular Polarity



***All* of these molecules are nonpolar due to their symmetry.**

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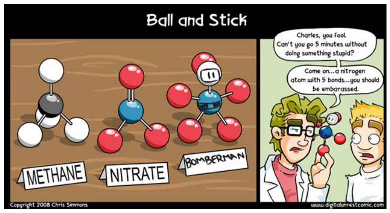
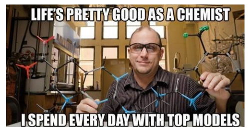
End of Chapter 7



See:

- *Chapter Seven Study Guide*
- *Chapter Seven Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*



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Important Equations, Constants, and Handouts from this Chapter:

- *know how to determine if* **ionic, covalent** *or* **metallic bonds** *are present*
- **ionic bond strength** *determined by* **Coulomb's Law**
- **# valence electrons = group number** *(US periodic table!)*
- *know the relationship between* **bond order, bond length** *and* **bond energy**
- *see* **Geometry and Polarity Guide** *and* **Bond Enthalpies and Electronegativities** *(handouts)*

Formal Charge = Group Number - bonds - lone pair electrons

FC = GN - bonds - lpe

ΔH_{rxn} = bonds broken - bonds formed

Lewis Structures / VSEPR: **bonding pairs, lone pairs, valence electrons, core electrons, total electrons, sigma bond, pi bond, VSEPR names (EPG & MG), formal charge, bond angles, polar, nonpolar, paramagnetic, diamagnetic, resonance structures, isomers**

$$\text{bond order (resonance)} = \frac{\#\ of\ e^-\ pairs\ used\ for\ a\ type\ of\ bond}{\#\ of\ bonds\ of\ that\ type}$$

MAR

End of Chapter Problems: *Test Yourself*

See practice problem set #1 and self quizzes for ***Lewis Structure / VSEPR*** *examples and practice*

1. Which of the following elements are capable of forming compounds in which the indicated atom has more than four valence electron pairs? N, As, C, O, Br, Be, S, Se
2. Which compound in each of the following pairs should require the higher temperature to melt?
   a. KBr or CsBr
   b. SrS or CaS
   c. LiF or BeO
3. Describe the EPG and MG around N in NH_2Cl.
4. Describe the EPG and MG around Cl in ClF_5.
5. Describe the EPG and MG around Te in TeF_4.
6. Which molecules are polar and which are nonpolar? H_2O, NH_3, CO_2, ClF, CCl_4
7. Give the bond order for each bond in the following molecules or ions: CH_2O, CO_2, NO_2^{+1}, CH_4
8. Oxygen difluoride is quite reactive with water, giving oxygen and HF:
   $OF_2(g) + H_2O(g) \rightarrow O_2(g) + 2\ HF(g)\quad \Delta H°_{rxn} = -318$ kJ
   Using bond energies, calculate the bond dissociation energy of the O-F bond in OF_2.

MAR

End of Chapter Problems: *Answers*

1. As, Br, S and Se
2. a. KBr b. CaS c. BeO
3. tetrahedral and trigonal pyramid
4. octahedral and square pyramid
5. trigonal bipyramid and seesaw
6. polar: H_2O, NH_3, ClF nonpolar: CO_2, CCl_4
7. **CH_2O** (2xBO=1 (C-H), 1xBO=2 (C=O)) , **CO_2** (2xBO=2 (C-O)), **NO_2^{+1}** (2xBO=2 (N-O)), **CH_4** (4xBO=1 (C-H))
8. D(O-F) = 195 kJ/mol

See practice problem set #1 and self quizzes for ***Lewis Structure / VSEPR*** *examples and practice*

MAR