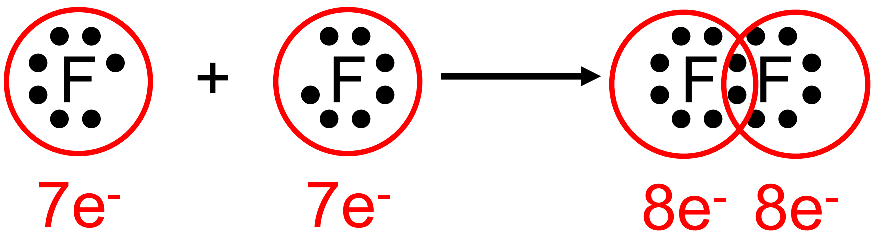
Chemical Bonding

Lewis Structures

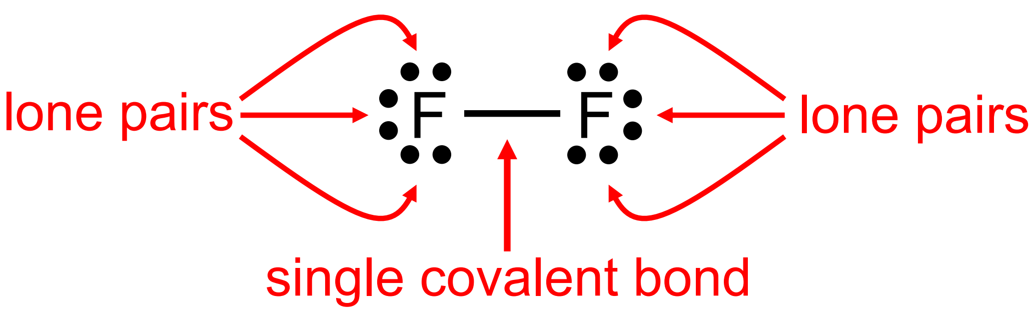
Lewis structures represent covalent bond formation

Want to create 8 electrons around each atom: octet



Bonding Pairs: Shared electrons count for both atoms

Lone Pairs: Non-shared electrons count for 1 atom

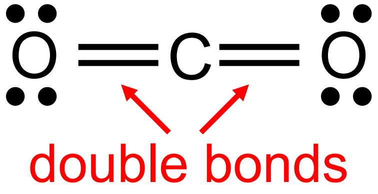


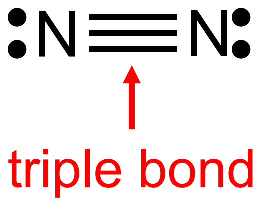
Multiple Bonds

More than one pair of electrons is shared between atoms so each atom can form an octet.

Single Bond:	1 shared pair:	1 dash (-)
Double bonds:	2 shared pairs:	2 dashes (=)
Triple bonds:	3 shared pairs:	3 dashes (≡)

Allows atoms in the Lewis structure to share extra electrons if there are not enough for the central atom





Writing Lewis Structures: General Information

Electronegativity

Central atom usually has the *lowest* electronegativity
(atom lower or to the left in periodic table)

Terminal atoms (except H) have *higher* electronegativities

Terminal Atoms

Bonded to only one other atom

Hydrogen atoms are terminal atoms

Bonding

Hydrogen atoms are bonded to oxygen atoms in oxoacids

Make the molecule as symmetrical as possible

Write the Lewis Structure of HNO

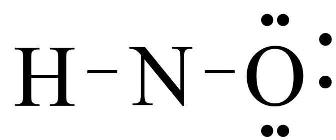
1. Add up the valence electrons in the structure

1 (H) + 5 (N) + 6 (O) = 12 valence electrons

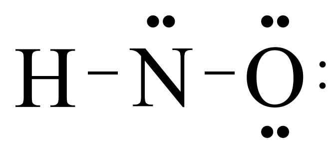
2. Arrange the atoms & place bonding electrons

H - N – O nitrogen less electronegative, put in the center

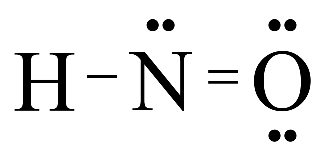
3. Place e- pairs around terminal atoms



4. Place remaining electron pairs on central atom



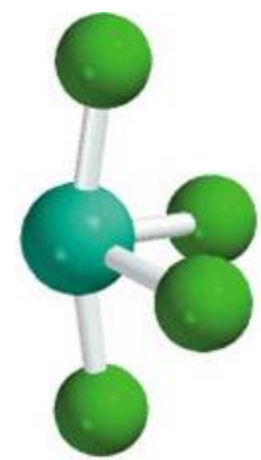
5. Add double bond to finish nitrogen octet



The Valence-Shell Electron-Pair Repulsion (VSEPR)

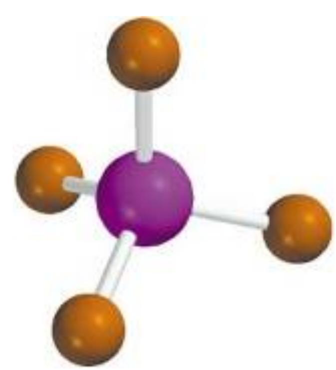
Method based on the idea that pairs of valence electrons in bonded atoms repel one another.

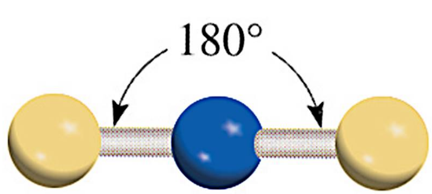
Assumes electron pairs try to get as far apart as possible
Each electron pair or bond takes up ~ same amount of space
# of bonds or pairs determines molecular geometry



Molecular Geometry:

The shape of a molecule that describes the location of nuclei & the connections between them.

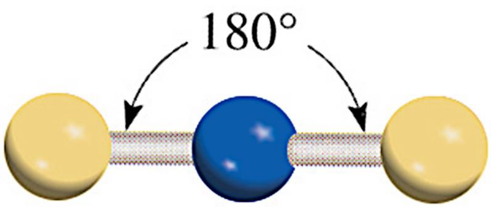




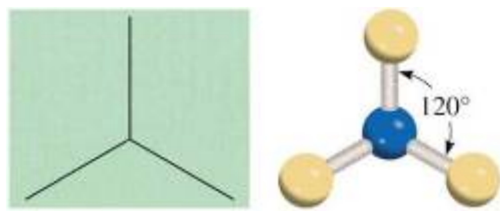
Molecules with No Lone Pairs

Bond angles due to # of repulsions
Each bond takes up space of 1 electron pair

AB_2
Linear

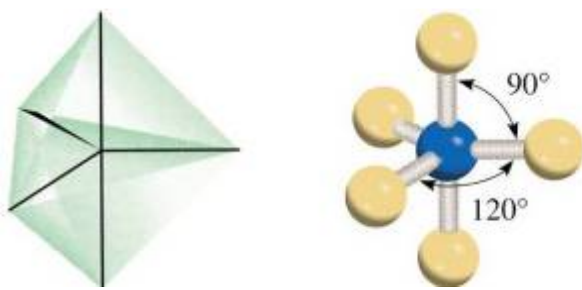


AB_3
Trigonal planar

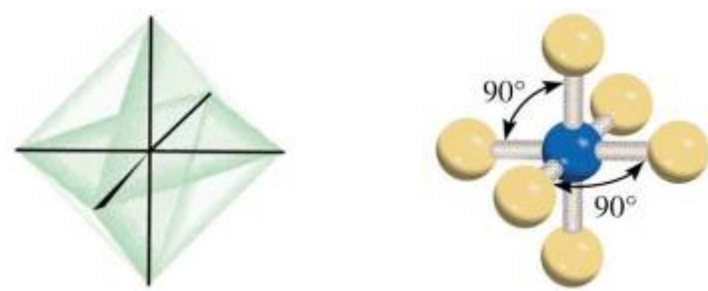


AB_4
Tetrahedral

AB_5
Trigonal bipyramid



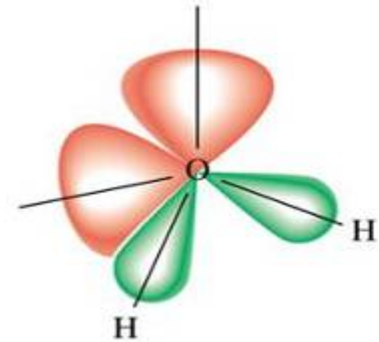
AB_6
Octahedral



Molecules with Lone Pairs

Lone pair electrons not seen but take up space

- Act as “invisible bond”
- Count electrons as E’s
- Single, double or triple bonds count as 1 bond



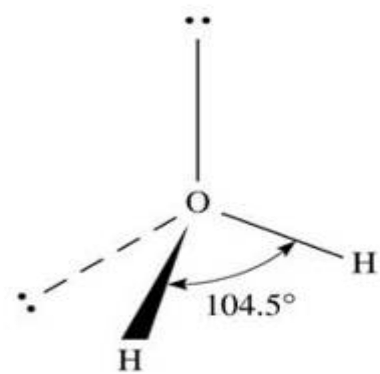
To determine molecular geometry

- Add up all the B’s and E’s on the molecule

$$\mathrm{H{-}\ddot{\underset{\cdot\cdot}{O}}{-}H} \qquad \mathrm{B{-}\ddot{\underset{\cdot\cdot}{A}}{-}B} \qquad AB_2E_2$$

- The sum equals number of spaces needed

$2B + 2E = 4$ # spaces = 4



- Match to table of geometries without lone pairs

Electron pair geometry: Tetrahedral

Molecular Geometry: Bent

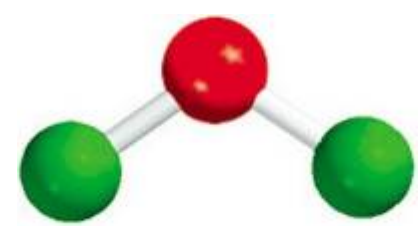
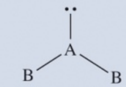
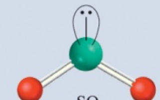
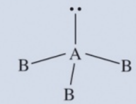
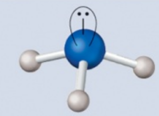
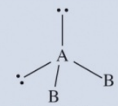
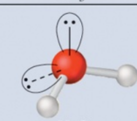
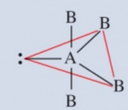
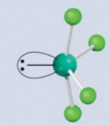
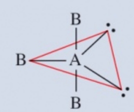
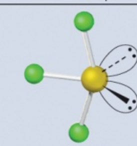
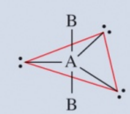
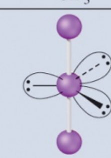
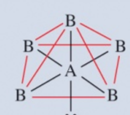
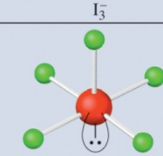
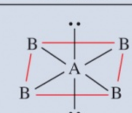
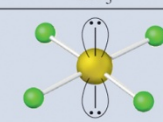


Table 10.2 Geometry of Simple Molecules and Ions in Which the Central Atom Has One or More Lone Pairs

Class of Molecule	Total Number of Electron Pairs	Number of Bonding Pairs	Number of Lone Pairs	Arrangement of Electron Pairs*	Geometry of Molecule or Ion	Examples
AB_2E	3	2	1	 Trigonal planar	Bent	 SO_2
AB_3E	4	3	1	 Tetrahedral	Trigonal pyramidal	 NH_3
AB_2E_2	4	2	2	 Tetrahedral	Bent	 H_2O
AB_4E	5	4	1	 Trigonal bipyramidal	Distorted tetrahedron (or seesaw)	 SF_4
AB_3E_2	5	3	2	 Trigonal bipyramidal	T-shaped	 ClF_3
AB_2E_3	5	2	3	 Trigonal bipyramidal	Linear	 I_3^-
AB_5E	6	5	1	 Octahedral	Square pyramidal	 BrF_5
AB_4E_2	6	4	2	 Octahedral	Square planar	 XeF_4

*The colored lines are used to show the overall shape, not bonds.

Molecules with More than 1 Central Atom

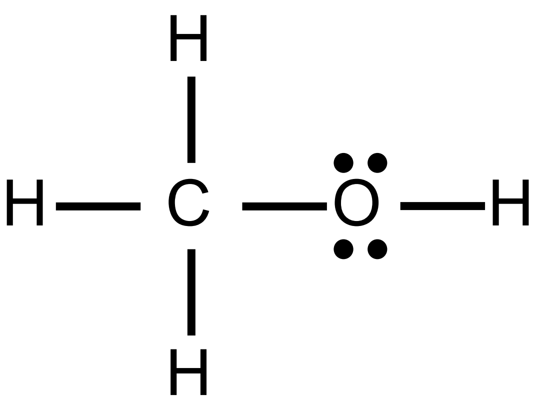
VSEPR must be done separately for each atom

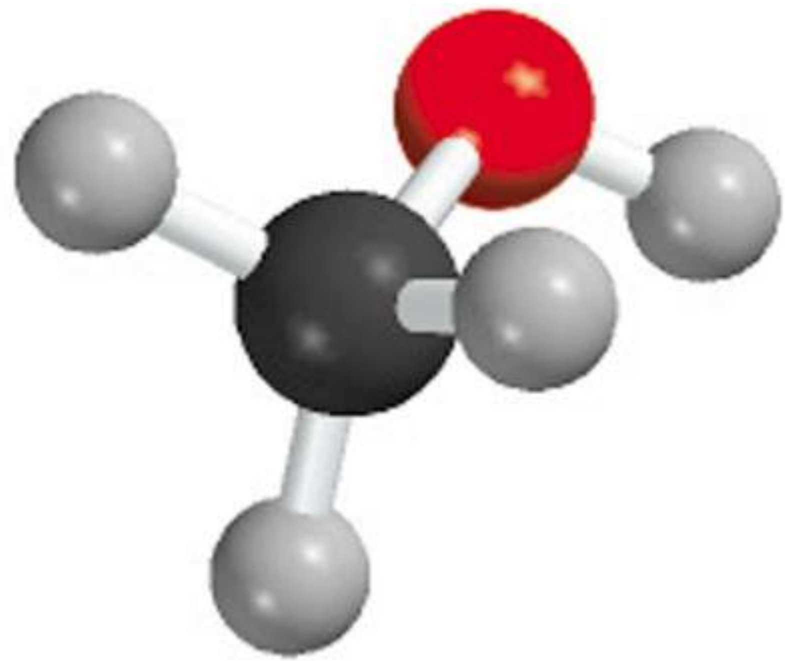
May result in a different molecular geometry around each one

Methanol CH_3OH

C: 4 spaces: tetrahedral

O: 4 spaces: bent





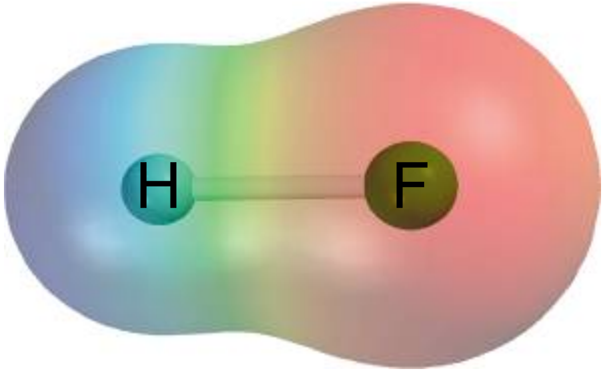
Oxoacids: Hydrogen goes on oxygens

HNO_3, H_2SO_4, etc. will also use this method

Electronegativity and Polarity

Electronegativity

The ability of an atom to attract electrons
F is the most electronegative atom
Nonmetals high electronegativies



Increasing electronegativity

Increasing electronegativity

1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
H 2.1																	
Li 1.0	**Be** 1.5											**B** 2.0	**C** 2.5	**N** 3.0	**O** 3.5	**F** 4.0	
Na 0.9	**Mg** 1.2											**Al** 1.5	**Si** 1.8	**P** 2.1	**S** 2.5	**Cl** 3.0	
K 0.8	**Ca** 1.0	**Sc** 1.3	**Ti** 1.5	**V** 1.6	**Cr** 1.6	**Mn** 1.5	**Fe** 1.8	**Co** 1.9	**Ni** 1.9	**Cu** 1.9	**Zn** 1.6	**Ga** 1.6	**Ge** 1.8	**As** 2.0	**Se** 2.4	**Br** 2.8	**Kr** 3.0
Rb 0.8	**Sr** 1.0	**Y** 1.2	**Zr** 1.4	**Nb** 1.6	**Mo** 1.8	**Tc** 1.9	**Ru** 2.2	**Rh** 2.2	**Pd** 2.2	**Ag** 1.9	**Cd** 1.7	**In** 1.7	**Sn** 1.8	**Sb** 1.9	**Te** 2.1	**I** 2.5	**Xe** 2.6
Cs 0.7	**Ba** 0.9	**La-Lu** 1.0-1.2	**Hf** 1.3	**Ta** 1.5	**W** 1.7	**Re** 1.9	**Os** 2.2	**Ir** 2.2	**Pt** 2.2	**Au** 2.4	**Hg** 1.9	**Tl** 1.8	**Pb** 1.9	**Bi** 1.9	**Po** 2.0	**At** 2.2	
Fr 0.7	**Ra** 0.9																

Predicting Polarity: NH_3

Predict molecular shape.

VSEPR

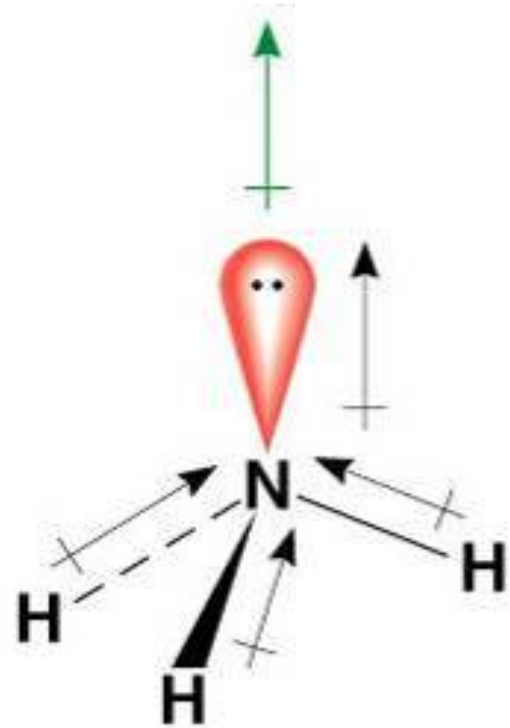
AB_3E

Tetrahedral

Predict bond dipoles.

H less electronegative than N

lone pair more electronegative than N

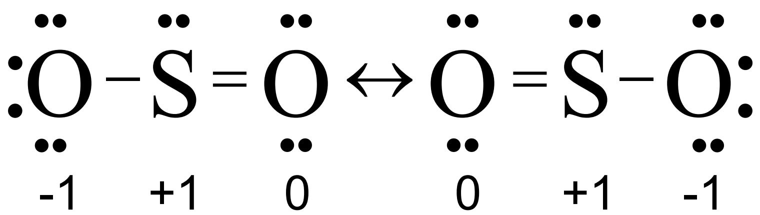


Bond dipoles cancel or combine?

Combine: Polar molecule

Resonance Theory

If a molecule or ion can be represented by 2 or more Lewis structures that differ only in electron location, the true structure is a composite of them.



Resonance Structures

Equivalent Lewis structures that can be drawn for a molecule

Formal charges will usually be present

Delocalization

Electrons are spread out over several atoms

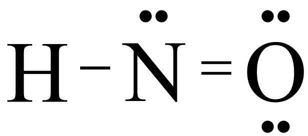
Stabilizes molecule

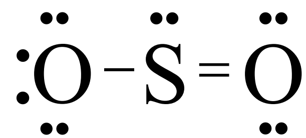
Formal Charge

Difference between the # of valence electrons in a free atom & the # of electrons assigned to that atom in a Lewis structure.

F.C. = Group # - (# of lone e- + # bonds)

Molecule is most stable if formal charge is 0 for each atom.





H: = 1-(0 +1)= 0
N: = 5-(2+3)= 0
O: = 6-(4+2)= 0

O: = 6-(6 +1)= -1
S: = 6-(2 +3)= +1
O: = 6-(4 +2)= 0

The most likely Lewis structure has the lowest formal charges

Negative formal charge must be on more electronegative atom

Sum of formal charges must =0 for molecules or = ionic charge

Recitation Questions

Answer the following questions about this compound: SO_2

1. Draw both Lewis structures for this compound. (resonance)
2. Give the VSEPR notation for the central atom in this structure.
3. Identify the molecular geometry on each central atom.
4. Show the polarity of the molecule using arrow notation.
5. Is the molecule polar or nonpolar?
6. Give the formal charge for all atoms.