1

Chemical Bonds:

- Lewis Structure / Electron Dot Structure
  - Put molecules and ions together
  - Tell where electrons and charges located
- VSEPR: Determine shapes and bond angles
- Physical properties
  - Polarity
  - Bond length
  - Bond strength

Syllabus Learning Outcomes : 13

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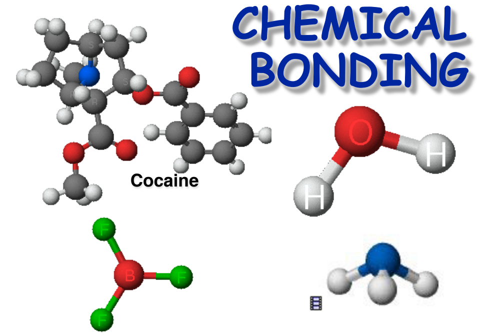
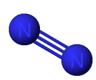
2

CHEMICAL BONDING

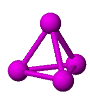


3

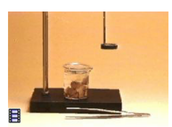
Structure & Bonding

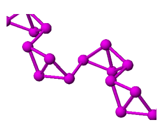


NN triple bond. Molecule is unreactive



Phosphorus is a tetrahedron of P atoms. Very reactive!

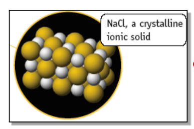




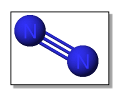
Red phosphorus, a polymer. Used in matches.

4

Bonds: Ionic and Covalent



- Ionic: transfer of e⁻



- Covalent: sharing of e⁻
- Most bonds are between

5

Ionic Compounds Have Ionic Bonds

Metal, low IE + Nonmetal, high EA

2 Na(s) + Cl_2(g) ---> 2 Na^+ + 2 Cl^-

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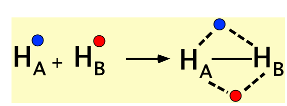
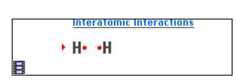
Covalent Compounds Have Covalent Bonds

Bond forms when 2 nuclei attract the same electrons.



Bond is balance of attractive and repulsive forces.



Bond Formation

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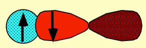
A bond can result from a "head-to-head" **overlap** of atomic orbitals on neighboring atoms.

H• + •Cl: ⟶ H:Cl:



Overlap of H (1s) and Cl (2p)

Note that each atom has a single, unpaired electron.

Chemical Bonding: Objectives

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Objectives are to understand:

1. **valence e- distribution in molecules and ions.**
2. **molecular structures**
3. **bond properties and their effect on molecular properties.**

Electron Distribution in Molecules

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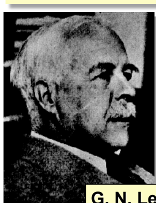


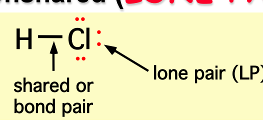
G. N. Lewis
1875 - 1946

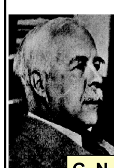
- **Electron distribution is depicted with Lewis electron dot structures**
- **Valence electrons are distributed as shared or BOND PAIRS and unshared or LONE PAIRS.**

Valence electrons in molecules can be

10

- **Shared (BOND PAIRS)**
- **Unshared (LONE PAIRS)**





G. N. Lewis
1875 - 1946

This is a LEWIS structure or ELECTRON DOT structure.

Valence electrons are the outermost s and p electrons

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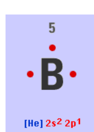
core electrons are not involved in bonding

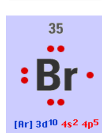
B $1s^2 2s^2 2p^1$

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



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

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



Lewis Structure Skills:

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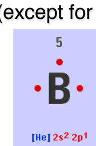
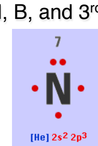
1. Count valence electrons for atoms

   Valence electrons= group number (G#)

2. Determine the number of bonds that atoms form
   - **For Groups 1A-4A, bond pairs (BPs) = G#**
   - **For Groups 5A -7A, BPs = 8 – G#**

3. Apply octet rule: BP + LP = 4 (except for H, B, and 3rd period atoms or higher)





Rules of the Game

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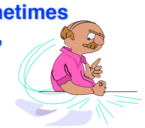
- No. of valence electrons of an atom = Group number
- For Groups 1A-4A, no. of bond pairs = group number
- For Groups 5A -7A, BP's = 8 - Grp. No.

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

BP's + LP's = 4



This observation is called the

OCTET RULE

N, O, F, Noble Gases

Octet rule helps to build molecules

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- Atoms try to get **8** valence electrons by forming bonds
- Works all the time only for F. 2nd period elements never get more than 8 e^-.
- There are many exceptions
  - H gets 2 valence e^-
  - B often gets 6 valence e^-
  - 3rd period or higher elements often get more than 8 e^- because of low energy d orbitals

Calculate the formal charge (FC) on bound atoms in molecules from the group number (G#), non-bonding electrons (NBE), and bonding electrons (BE)

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$$FC = G\# - (NBE + \tfrac{1}{2}\, BE)$$

* For this class, the preferred Lewis structure(s) have minimum sum of |FC|

Build a Lewis Structure

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Ammonia, NH_3

1. Pick central atom, not H, that has the lowest electronegativity (follows IE/EA trend).

   N is central because H cannot be

2. Count valence electrons

   H = 1 and N = 5

   Total = (3 x 1) + 5

   = 8 electrons / 4 pairs

Build a Lewis Structure

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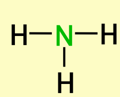
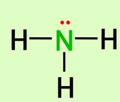
3. Form single (or sigma σ) bonds between the central atom and surrounding atoms



4. Add remaining electrons as lone pairs to complete octets for most electronegative atoms first



N has 3 bond pairs and 1 lone pair

N has 8 electrons. Each H has 2.

5. Form double bonds, calculate FC, resonance

Sulfite ion, SO_3^{2-}

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Step 1. Central atom = S

Step 2. Count valence electrons

S: 1 x 6 = 6

O: 3 x 6 = 18

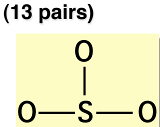
2^- charge: 2

TOTAL: 26 e^- (13 pairs)

Step 3. Form bonds - 6 e^-

20 e^-

10 pairs of e^- remain



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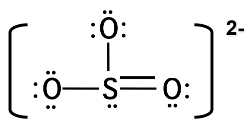
Sulfite ion, SO_3^{2-}

4) Place remaining e⁻ pairs on the most electronegative atoms first, and then on the central atom.



5) Each atom has an octet of electrons. **8**

5a) Double bond to minimize FC= G# - NBE - ½BE

5b) Three resonance structures

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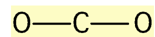
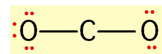
Carbon Dioxide, CO_2

1. **Central atom = ________**
2. **Valence electrons = __ or __ pairs**
3. **Form bonds.**



This leaves 6 pairs.

4. **Place lone pairs on outer atoms.**



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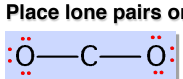
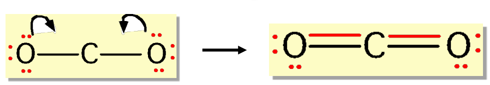
Carbon Dioxide, CO_2

4. **Place lone pairs on outer atoms.**



5. **Form DOUBLE BONDS — or pi (π) bonds — between C and O to get an octet for C.**



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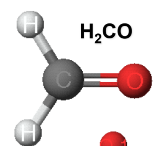
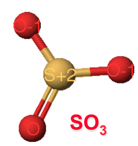
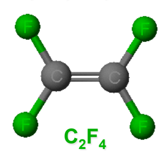
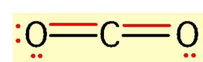
Double and even triple bonds are commonly observed for C, N, P, O, and S









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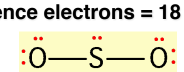
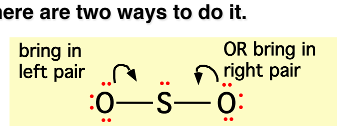
Sulfur Dioxide, SO_2

1. **Central atom = S**

2,3,4. Valence electrons = 18 or 9 pairs



5. **Form double bond so that S has an octet — there are two ways to do it.**



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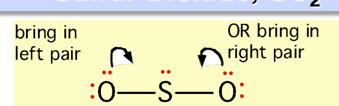
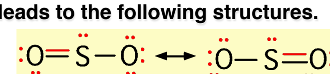
Sulfur Dioxide, SO_2



This leads to the following structures.



For this class, minimize FC by forming second double bond to 3rd period or higher central atoms. Recognize that all structures contribute

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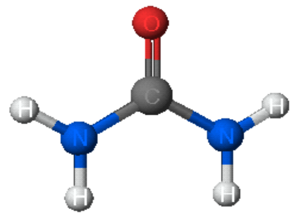
Urea, $(NH_2)_2CO$



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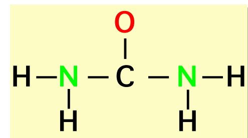
Urea, $(NH_2)_2CO$

1. Number of valence electrons = 24 e-
2. Draw sigma bonds.



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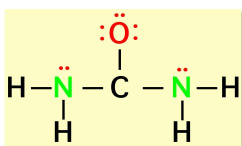
Urea, $(NH_2)_2CO$

3. Place remaining electron pairs in the molecule.



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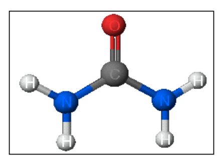
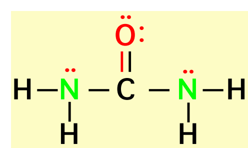
Urea, $(NH_2)_2CO$

4. Complete C atom octet with double bond.





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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 (no. of bonding electrons)
  - (no. of LP electrons)

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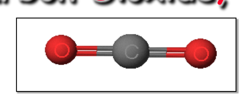
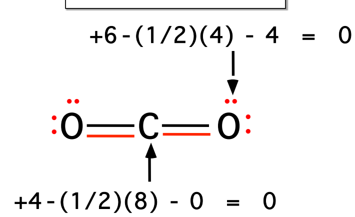
Carbon Dioxide, CO_2





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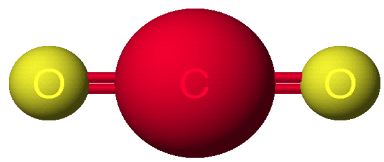
Calculated Partial Charges in CO_2



Yellow = negative & red = positive
Relative size = relative charge

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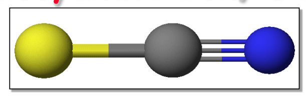
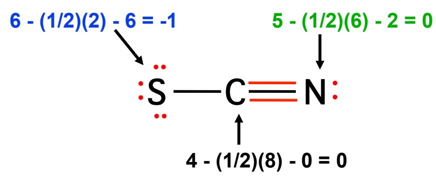
Thiocyanate Ion, SCN^-





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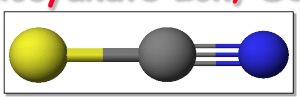
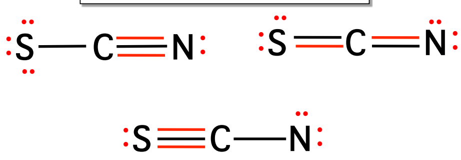
Thiocyanate Ion, SCN^-





Which is the most important resonance form?

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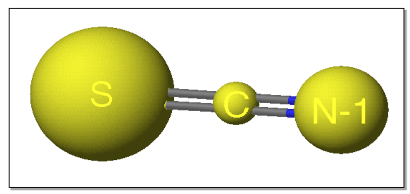
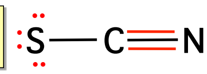
Calculated Partial Charges in SCN^-



All atoms negative, but most on the S



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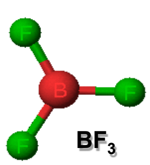
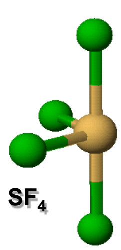
Violations of the Octet Rule

Often occurs with B and elements of 3rd and higher periods.



BF_3



SF_4

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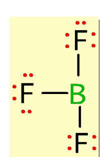
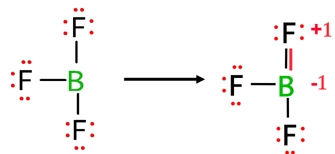
Boron Trifluoride

- Central atom = ____________
- Valence electrons = _________ or electron pairs = _________
- Assemble Lewis structure



B has only 6 e⁻ (3 pairs), so it is electron deficient.

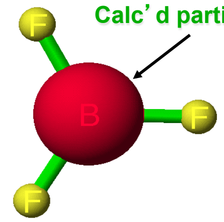
Boron Trifluoride, BF_3



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

Is There a B=F Double Bond in BF_3

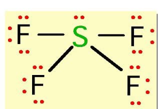
Calc'd partial charges in BF_3



F is negative and B is positive

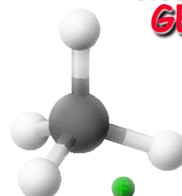
Sulfur Tetrafluoride, SF_4

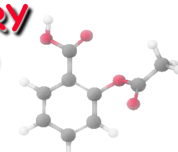
- Central atom =
- Valence electrons = ___ or ___ pairs.
- Form bonds and give e⁻ pairs.

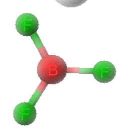


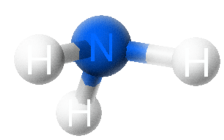
S has 10 e⁻ (5 pairs), which is common after the 2nd period.

MOLECULAR GEOMETRY (Shape)









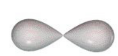
Molecular Geometry (Shape)

VSEPR

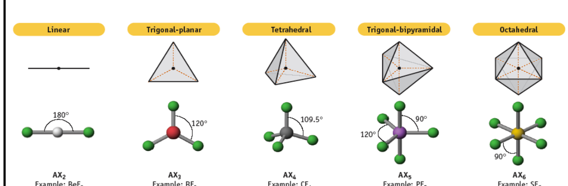
- **V**alence **S**hell **E**lectron **P**air **R**epulsion theory.

Molecule adopts the shape that minimizes the electron-pair repulsions.

VSEPR - Shape of e⁻ pairs and bonds around a central atom →



Electron Pair Geometries

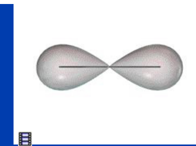


43

No. of e- Pairs Around Central Atom	Example	Geometry

44

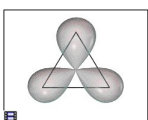
No. of e- Pairs Around Central Atom	Example	Geometry
2	F– Be– F 180 °	linear

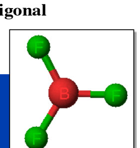


Describe Shapes

45

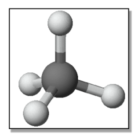
No. of e- Pairs Around Central Atom	Example	Geometry
2	F– Be– F 180 °	linear
3	F, B, F, F 120 °	planar trigonal

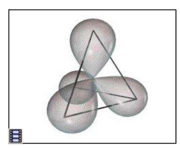




Describe Shapes

46



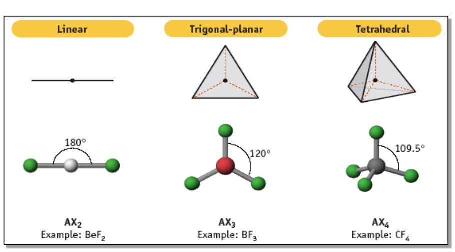


No. of e- Pairs Around Central Atom	Example	Geometry
2	F– Be– F 180 °	linear
3	F, B, F, F 120 °	planar trigonal
4	H, C, H, H, H 109 °	tetrahedral

Describe Shapes

47

Electron Pair Geometries

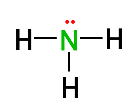


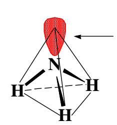
48

VSEPR Shape

Ammonia, NH_3

1. Draw Lewis structure
2. Count BP's and LP's = 4
3. The 4 electron pairs are at the corners of a tetrahedron.





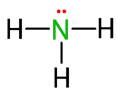
lone pair of electrons in tetrahedral position

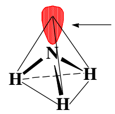
49

VSEPR Shape

Ammonia, NH_3

There are 4 electron pairs at the corners of a tetrahedron.





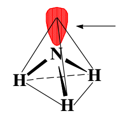
ELECTRON PAIR GEOMETRY: tetrahedral

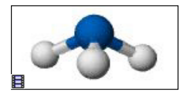
50

VSEPR Shape

Ammonia, NH_3

The electron pair geometry is tetrahedral.





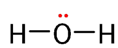
MOLECULAR GEOMETRY — the positions of the atoms — is PYRAMIDAL.

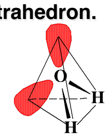
51

VSEPR Shape

Water, H_2O

1. **Draw Lewis structure**
2. **Count BP' s and LP' s = 4**
3. **The 4 electron pairs are at the corners of a tetrahedron.**



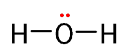


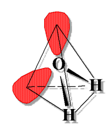
Electron pair geometry: TETRAHEDRAL

52

VSEPR Shape

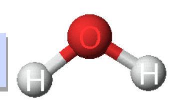
Water, H_2O





Electron pair geometry: TETRAHEDRAL

Molecular geometry: BENT.



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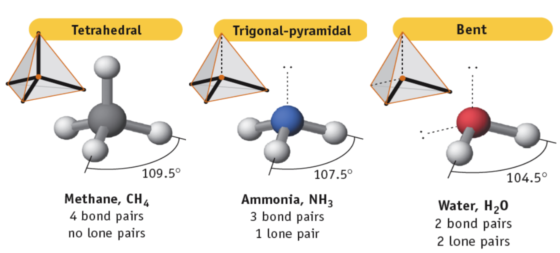
Shapes (Geometries) for Four Electron Pairs



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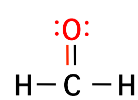
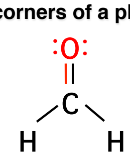
VSEPR Shape

Formaldehyde, CH_2O

1. **Draw Lewis structure**
2. **Count BP' s and LP' s at C**
3. **There are 3 electron groups around C at the corners of a planar triangle.**





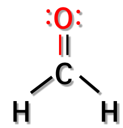
Electron pair geometry: PLANAR TRIGONAL 120° bond angles

Molecular geometry: planar trigonal

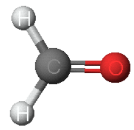
55

Structure Determination by VSEPR

Formaldehyde, CH_2O



The electron pair geometry is PLANAR TRIGONAL



The molecular geometry is also planar trigonal.

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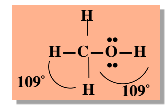
VSEPR Shape

Methanol, CH_3OH

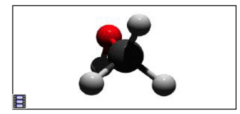


Define H-C-H and C-O-H bond angles

H-C-H = 109º

C-O-H = 109º

C and O have 4 e⁻ pairs



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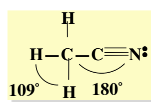
VSEPR Shape

Acetonitrile, CH_3CN



Define unique bond angles

H-C-H = 109º

C-C-N = 180º

One C is surrounded by 4 e⁻ groups and the other by 2 groups

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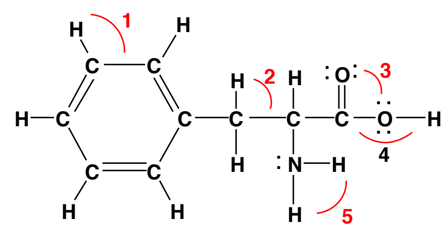
Phenylalanine, an amino acid



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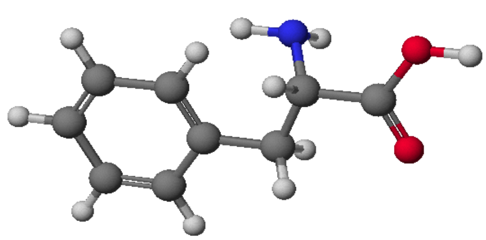
Phenylalanine



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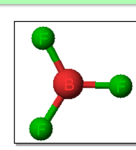
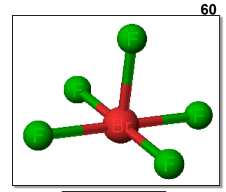
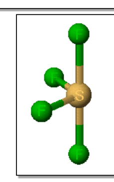
Shapes with More Than or Less Than 4 Electron Pairs







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

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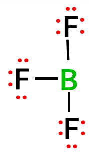
Boron Compounds

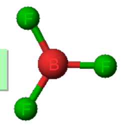
Consider boron trifluoride, BF_3



The B atom has 3 e⁻ pairs

Bond angles are 120°

Geometry: planar trigonal



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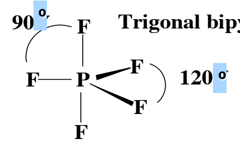
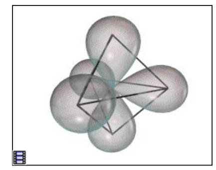
Shapes with 5 e⁻ Pairs





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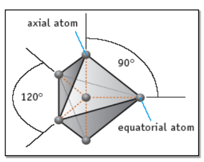
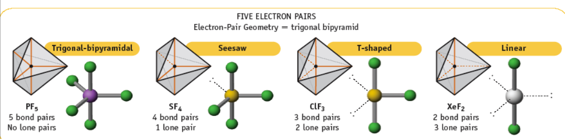
Shapes (molecular geometries) for five e⁻ pairs



Shapes based on trigonal bipyramid



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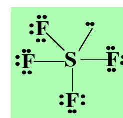
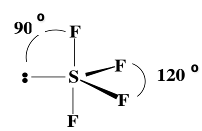
Sulfur Tetrafluoride, SF_4

- Valence electrons: 34
- Central atom: S
- Dot structure



Electron pair geometry:
trigonal bipyramid
(S has 5 e⁻ pairs)



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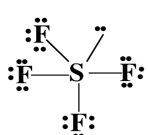
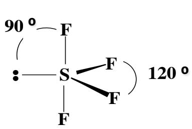
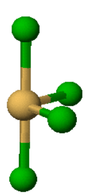
Sulfur Tetrafluoride, SF_4

Lone pair goes on the equator (where it gets more room)







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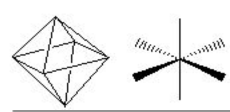
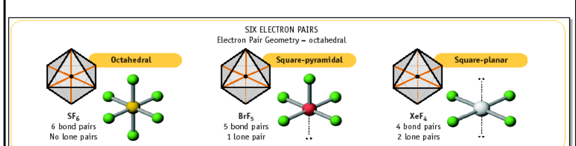
Molecular Geometries for Six Electron Pairs



All are based on the 8-sided octahedron



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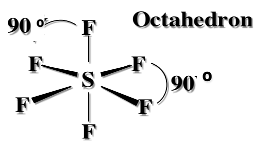
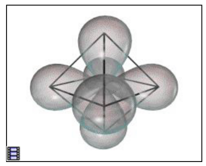
Shapes with 6 e⁻ Pairs





6 electron pairs

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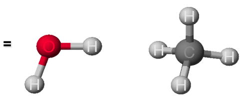
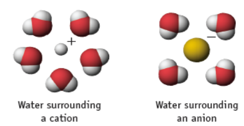
Molecular Polarity

Water
Boiling point = 100 ˚C



Methane
Boiling point = -161 ˚C



Why do ionic compounds dissolve in water?

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

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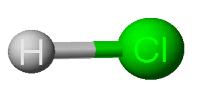
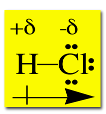
Bond Polarity



HCl is POLAR because it has a positive end and a negative end.



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

Cl has slight negative charge (-δ) and H has slight positive charge (+ δ)

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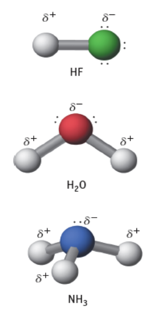

Bond Polarity

- Three molecules with polar, covalent bonds.
- Each bond has one atom with a slight negative charge (-δ) and another with a slight positive charge (+ δ)

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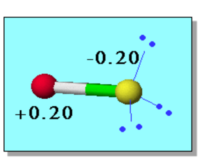
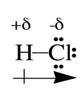
Bond Polarity



This model, calc'd using CAChe software for molecular calculations, shows that H is + (red) and Cl is - (yellow). Calc'd charge is + or - 0.20.

72



Bond Polarity

Due to the bond 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.**

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Electronegativity, χ

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

Concept proposed by Linus Pauling 1901-1994

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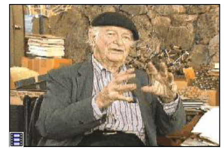
Linus Pauling, 1901-1994



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

Chemistry areas: bonding, electronegativity, protein structure

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Electronegativity

H 2.2

1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A
Li 1.0	Be 1.6											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.3											Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2
K 0.8	Ca 1.0	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 1.9	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 1.9	Te 2.1	I 2.7
Cs 0.8	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 2.4	Re 1.9	Os 2.2	Ir 2.2	Pt 2.3	Au 2.5	Hg 2.0	Tl 1.6	Pb 2.3	Bi 2.0	Po 2.0	At 2.2

<1.0 · 1.0–1.4 · 1.5–1.9 · 2.0–2.4 · 2.5–2.9 · 3.0–4.0

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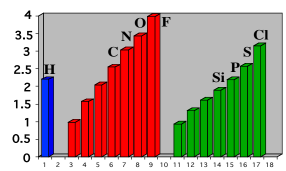

Electronegativity, χ

- 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).

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End Lecture 12

Start Lecture 13

26

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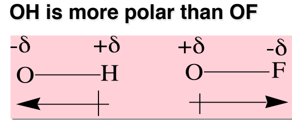
Bond Polarity

Which bond is more polar (or DIPOLAR)?

	O—H	O—F
χ	3.5 - 2.1	3.5 - 4.0
Δ	1.4	0.5

OH is more polar than OF

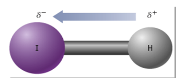


and polarity is "reversed."

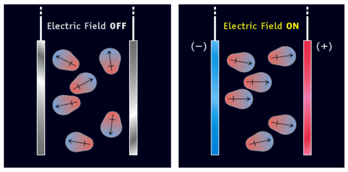
Molecular Polarity

79

Molecules—such as HI and H_2O—can be **POLAR** (or dipolar).



They have a **DIPOLE MOMENT**. The polar HCl molecule will turn to align with an electric field.



Molecular Polarity

80

The magnitude of the dipole is given in Debye units. Named for Peter Debye (1884 - 1966). Rec'd 1936 Nobel prize for work on x-ray diffraction and dipole moments.

Dipole Moments

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Table 9.8 Dipole Moments of Selected Molecules

Molecule (AB)	Moment (μ, D)	Geometry	Molecule (AB_2)	Moment (μ, D)	Geometry
HF	1.78	linear	H_2O	1.85	bent
HCl	1.07	linear	H_2S	0.95	bent
HBr	0.79	linear	SO_2	1.62	bent
HI	0.38	linear	CO_2	0	linear
H_2	0	linear			
Molecule (AB_3)	**Moment (μ, D)**	**Geometry**	**Molecule (AB_4)**	**Moment (μ, D)**	**Geometry**
NH_3	1.47	trigonal-pyramidal	CH_4	0	tetrahedral
NF_3	0.23	trigonal-pyramidal	CH_3Cl	1.92	tetrahedral
BF_3	0	trigonal-planar	CH_2Cl_2	1.60	tetrahedral
			$CHCl_3$	1.04	tetrahedral
			CCl_4	0	tetrahedral

Why are some molecules polar but others are not?

Molecular Polarity

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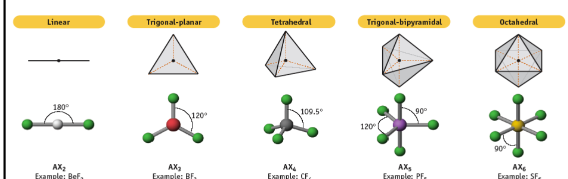
Molecules will be polar if
a) bonds are polar
AND
b) the molecule is NOT "symmetric"

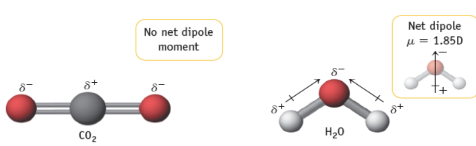


All molecules above are NOT polar

Polar or Nonpolar?

83

Compare CO_2 and H_2O. Which one is polar?



Carbon Dioxide

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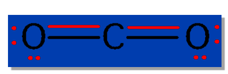
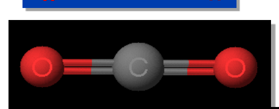
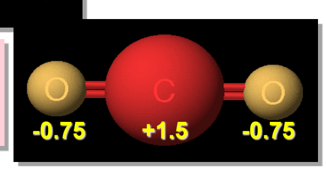


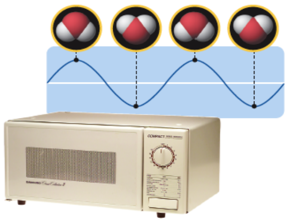
- CO_2 is NOT polar even though the CO bonds are polar.
- CO_2 is symmetrical.



Positive C atom is reason CO_2 and H_2O react to give H_2CO_3



85



Consequences of H_2O Polarity

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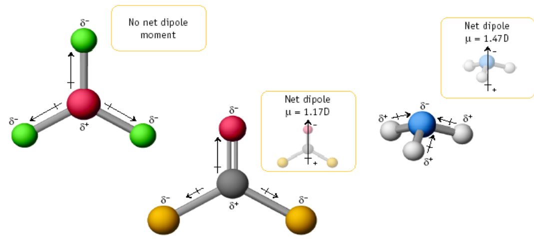
Polar or Nonpolar?

- Consider AB_3 molecules: BF_3, Cl_2CO, and NH_3.



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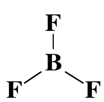
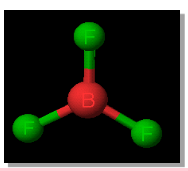
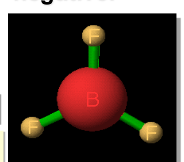
Molecular Polarity, BF_3





B atom is positive and F atoms are negative.

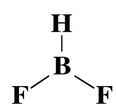


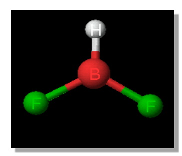
B—F bonds in BF_3 are polar.

But molecule is symmetrical and **NOT** polar

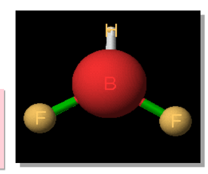
88

Molecular Polarity, HBF_2





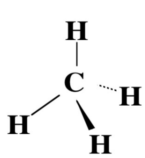
B atom is positive but H & F atoms are negative.

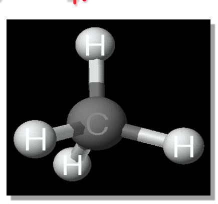


B—F and B—H bonds in HBF_2 are polar. But molecule is NOT symmetrical and is polar.

89

Is Methane, CH_4, Polar?

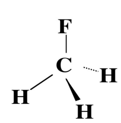


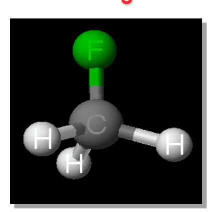


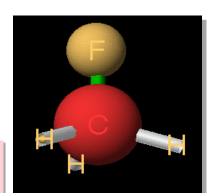
Methane is symmetrical and is NOT polar.

90

Is CH_3F Polar?



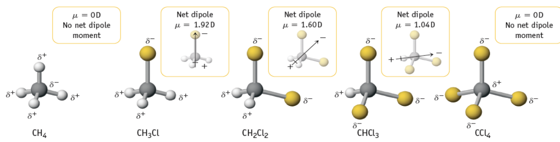




C—F bond is very polar. Molecule is not symmetrical and so is polar.

91

CH_4 ... CCl_4 Polar or Not?



- Only CH_4 and CCl_4 are NOT polar. These are the only two molecules that are "symmetrical."

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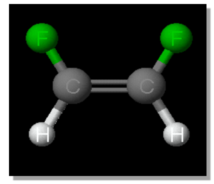
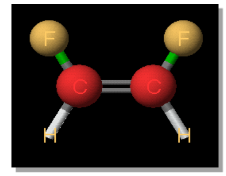
Substituted Ethylene

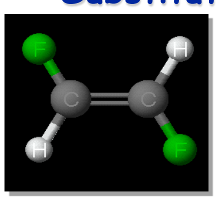


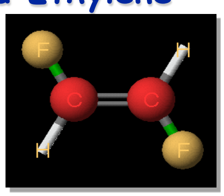


- C—F bonds are MUCH more polar than C—H bonds.
- Because both C—F bonds are on same side of molecule, molecule is **POLAR**.

93

Substituted Ethylene





- C—F bonds are MUCH more polar than C—H bonds.
- Because both C—F bonds are on opposing ends of molecule, molecule is **NOT POLAR**.

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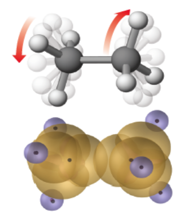
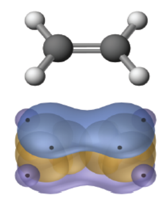
Bond Properties

- What is the effect of bonding and structure on molecular properties?



Free rotation around C–C single bond



No rotation around C=C double bond

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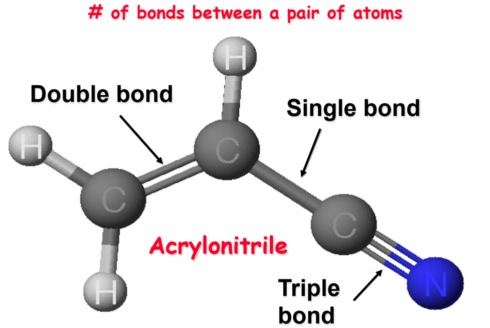
Bond Order

of bonds between a pair of atoms



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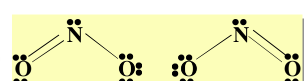
Fractional Bond Order

Occurs in molecules with resonance structures.

Consider NO_2^-

N—O bond order is 1.5



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

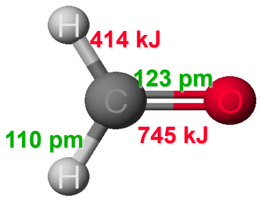
$$Bond\ order = \frac{3\ e^- \text{ pairs for N — O bonds}}{2 \text{ N — O bonds}}$$

Bond Order

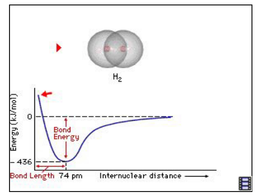
Bond order is proportional to two important bond properties:

(a) bond strength

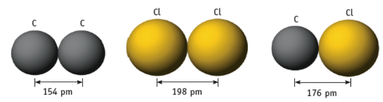
(b) bond length



Bond Length

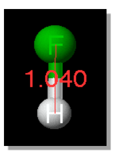


- Bond length is the distance between the nuclei of two bonded atoms.

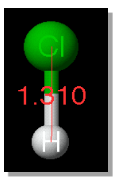


Bond Length

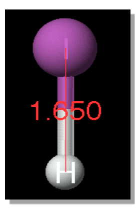
Bond length depends on size of bonded atoms.



H—F



H—Cl

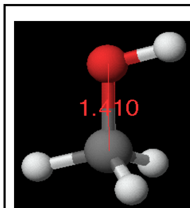


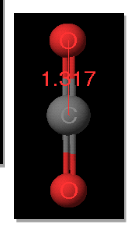
H—I

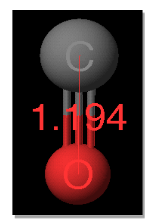
Bond distances measured in Angstrom units where 1 A = 10^{-2} pm.

Bond Length

Bond length depends on bond order.



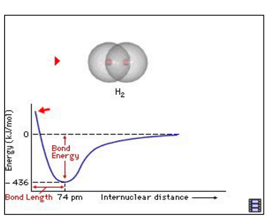




Bond distances measured in Angstrom units where 1 A = 10^{-2} pm.

Bond Strength

- —measured by the energy req'd to break a bond.



Bond Strength

energy req'd to break a bond

BOND	STRENGTH (kJ/mol)
H—H	436
C—C	346
C=C	602
C≡C	835
N≡N	945

The GREATER the number of bonds (bond order) the HIGHER the bond strength and the SHORTER the bond.

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Table 9.10 Some Average Single- and Multiple-Bond Energies (kJ/mol)*

Single Bonds

	H	C	N	O	F	Si	P	S	Cl	Br	I
H	436	413	391	463	565	328	322	347	432	366	299
C		346	305	358	485	—	—	272	339	285	213
N			163	201	283	—	—	—	192	—	—
O				146	—	452	335	—	218	201	201
F					155	565	490	284	253	249	278
Si						222	—	293	381	310	234
P							201	—	326	—	184
S								226	255	—	—
Cl									242	216	208
Br										193	175
I											151

Multiple Bonds

N=N	418	C=C	610
N≡N	945	C≡C	835
C=N	615	C=O	745
C≡N	887	C≡O	1046
O=O (in O_2)	498		

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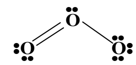
Bond Strength

Bond	Order	Length	Strength
HO—OH	1	142 pm	146 kJ/mol
O=O	2	121	498
	1.5	128	?

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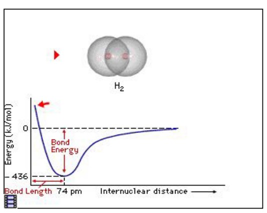
Bond Strength

- Measured as the energy req'd to break a bond.



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Bond Strength

- Measured as the energy req'd to break a bond.

BOND	STRENGTH (kJ/mol)
H—H	436
C—C	346
C=C	602
C≡C	835
N≡N	945

The GREATER the number of bonds (bond order) the HIGHER the bond strength and the SHORTER the bond.

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Using Bond Energies

Estimate the energy of the reaction

H—H(g) + Cl—Cl(g) ----> 2 H—Cl(g)

Net energy = ΔH_{rxn} =

= energy required to break bonds

- energy evolved when bonds are made

H—H = 436 kJ/mol
Cl—Cl = 242 kJ/mol
H—Cl = 432 kJ/mol

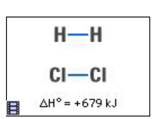

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Table 9.10 Some Average Single- and Multiple-Bond Energies (kJ/mol)*

Single Bonds

	H	C	N	O	F	Si	P	S	Cl	Br	I
H	436	413	391	463	565	328	322	347	432	366	299
C		346	305	358	485	—	—	272	339	285	213
N			163	201	283	—	—	—	192	—	—
O				146	—	452	335	—	218	201	201
F					155	565	490	284	253	249	278
Si						222	—	293	381	310	234
P							201	—	326	—	184
S								226	255	—	—
Cl									242	216	208
Br										193	175
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Multiple Bonds

N=N	418	C=C	610
N≡N	945	C≡C	835
C=N	615	C=O	745
C≡N	887	C≡O	1046
O=O (in O_2)	498		

Using Bond Energies

Estimate the energy of the reaction

H—H + Cl—Cl ----> 2 H—Cl

H—H = 436 kJ/mol
Cl—Cl = 242 kJ/mol
H—Cl = 432 kJ/mol

Sum of H-H + Cl-Cl bond energies = 436 kJ + 242 kJ = +678 kJ

2 mol H-Cl bond energies = 864 kJ

Net = ΔH = +678 kJ - 864 kJ = -186 kJ

Using Bond Energies

Estimate the energy of the reaction

2 H—O—O—H ----> O=O + 2 H—O—H

Is the reaction exo- or endothermic?

Which is larger:

A) energy req'd to break bonds

B) or energy evolved on making bonds?

Using Bond Energies

2 H—O—O—H ----> O=O + 2 H—O—H

Energy required to break bonds:

break 4 mol of O—H bonds = 4 (463 kJ)

break 2 mol O—O bonds = 2 (146 kJ)

TOTAL ENERGY to break bonds = 2144 kJ

TOTAL ENERGY evolved on making O=O bonds and 4 O-H bonds bonds = 2350 kJ

Using Bond Energies

2 H—O—O—H ----> O=O + 2 H—O—H

Net energy = +2144 kJ - 2350 kJ = - 206 kJ

The reaction is exothermic!

More energy is evolved on making bonds than is expended in breaking bonds.

End