Announcements for Thursday, 17OCT2024

- Week 7 Homework Assignments available on eLearning
  - Graded and Timed Quiz 7 – “Bonding” due **Monday, 21OCT2024, at 6:00 PM (EDT)**
- **Exam 2 Location Requests due by Friday, 18OCT2024, 11:59 PM (EDT)**
  - for students having a class ending at 7:00 PM or later on exam day

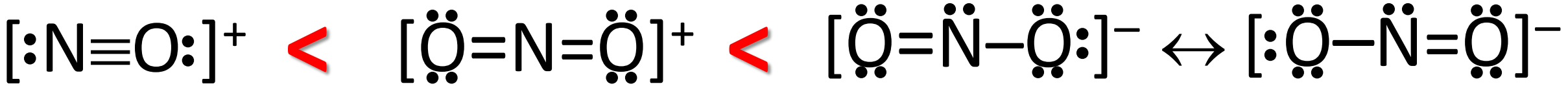
ANY GENERAL QUESTIONS? Feel free to see me after class!

Try This On Your Own

Rank the following species in order of increasing N-O bond length:

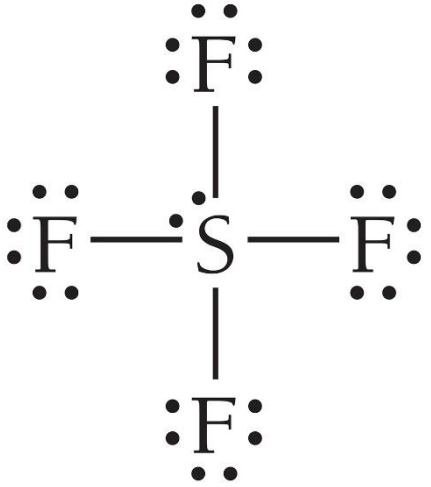
NO_2^-, NO^+, NO_2^+

Hint: Start by drawing the Lewis structures



Try This On Your Own

Predict the electron geometry, molecular geometry, and approximate bond angle(s) for the following species:

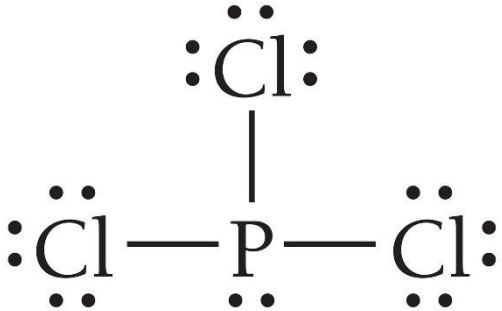


5 electron groups

e^- g.: trigonal bipyramidal

m. g.: seesaw

< 90°, < 120°, < 180°

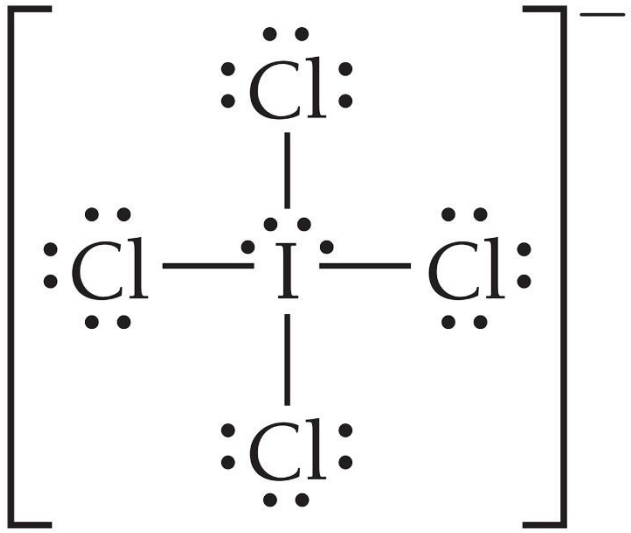


4 electron groups

e^- g.: tetrahedral

m. g.: trigonal pyramidal

< 109.5°



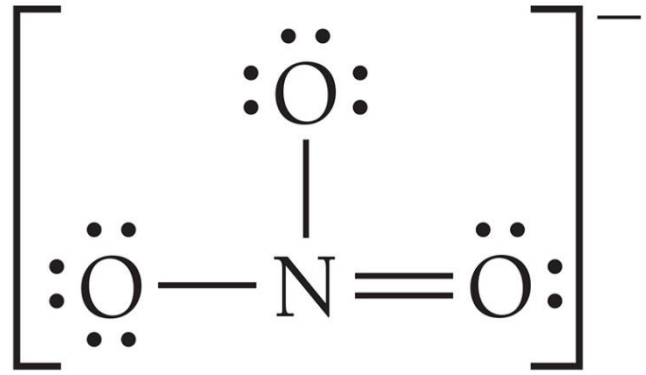
6 electron groups

e^- g.: octahedral

m. g.: square planar

= 90°, = 180°

(ideal bond angles)



3 electron groups

e^- g.: trigonal planar

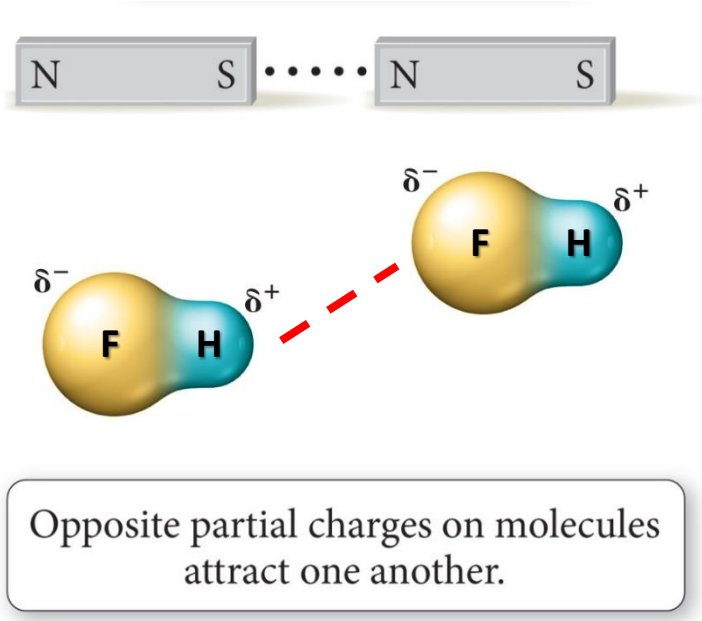
m. g.: trigonal planar

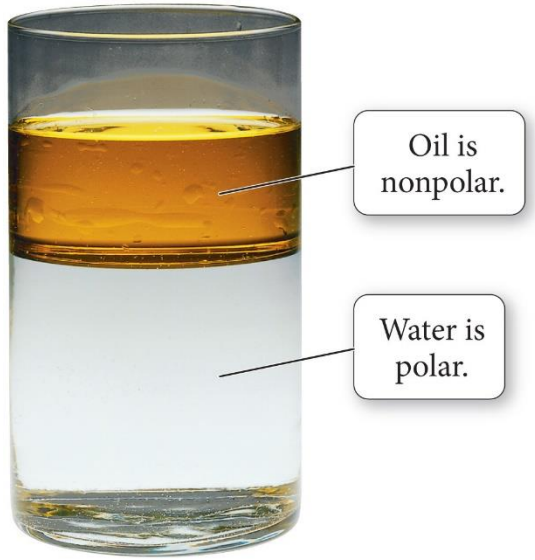
= 120°

(ideal bond angles)

Molecular Polarity

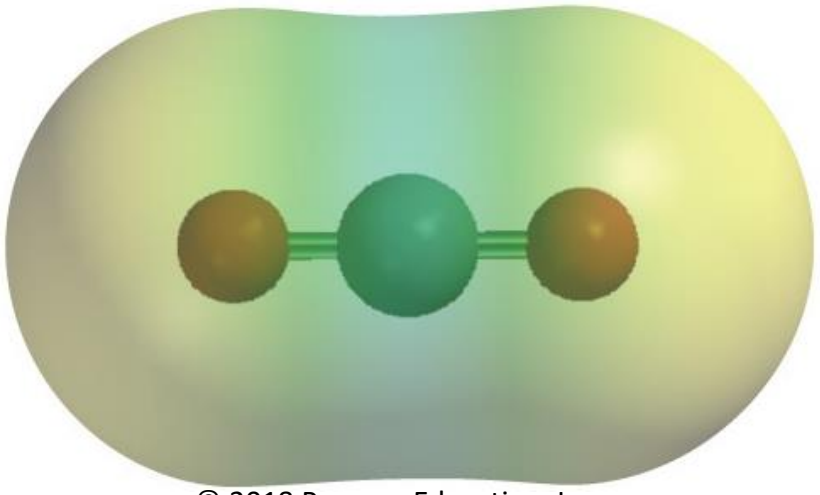
- we already learned that bonds can be polar due to uneven sharing of bonding electrons between two atoms
  - example: HF
  - the bond acts similar to a little magnet and will respond to an external electric field
- molecules also may exhibit polarity (i.e., have a net dipole moment)
- polar molecules can interact with each other much like different poles of magnets interact
- polar molecule will easily mix with other polar molecules to make homogeneous solutions
  - like dissolves like



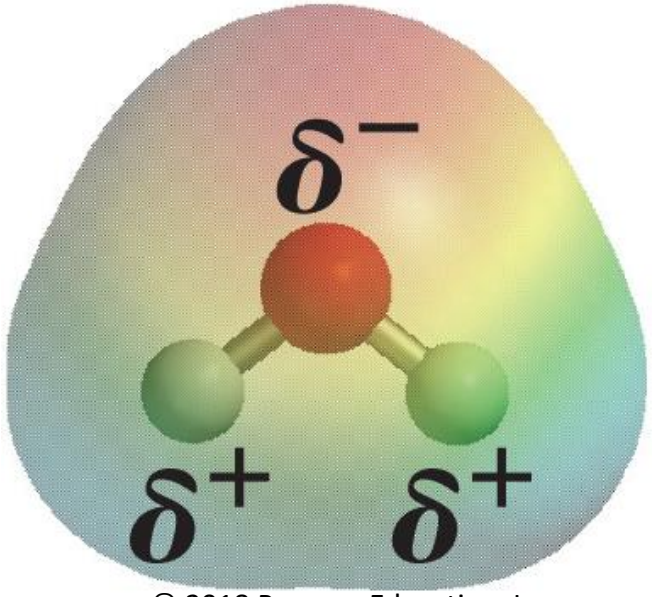


Requirements for Molecular Polarity

in molecules where there are two or more bonds, the entire molecule can be polar and may exhibit a net dipole moment provided some requirements are met



CO_2: no net dipole = nonpolar molecule

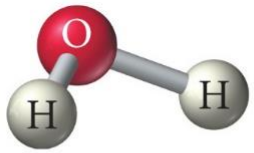


H_2O: *has* net dipole = polar molecule

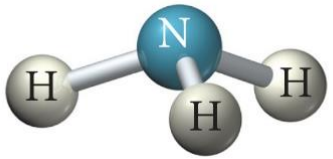
1. contains polar bonds (…we will be ignoring some exceptions to this)
2. the individual bond dipoles around the central atom do not cancel each out

Geometries That Commonly Lead to Molecular Polarity

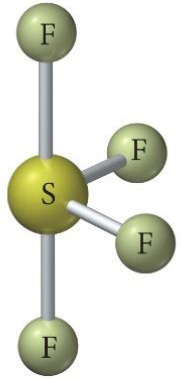
- this is assuming that ***all bonds are identical***



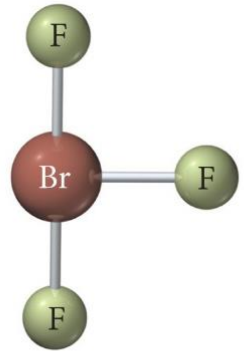
Molecular geometry: bent



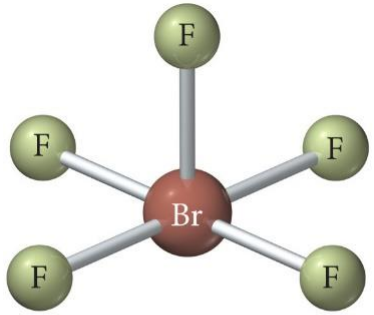
Molecular geometry: trigonal pyramidal



Molecular geometry: seesaw



Molecular geometry: T-shaped



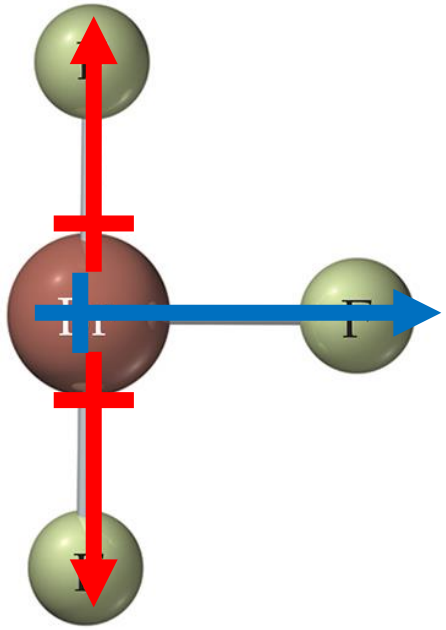
Molecular geometry: square pyramidal

- it gets more complicated when there is more than one type of bond
  - example: FCN
- bond dipoles are vector quantities
  - they have both magnitude AND direction

Try This

Determine if the following molecules are polar or nonpolar

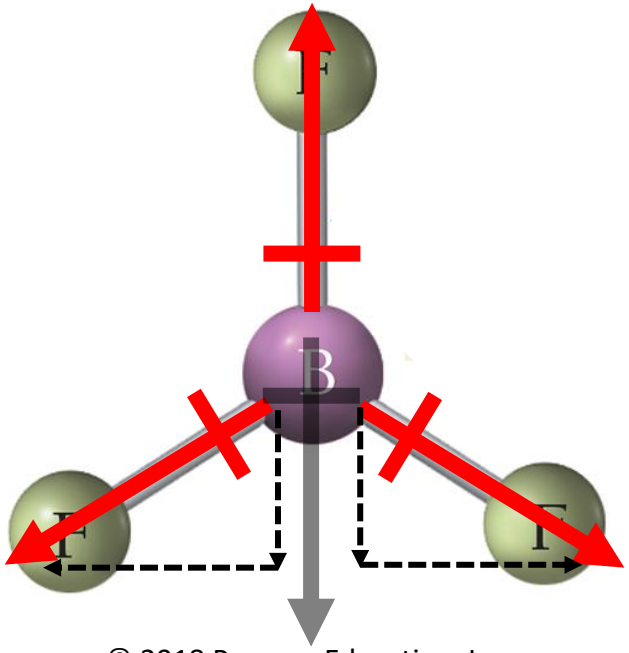
BrF_3



Molecular geometry:
T-shaped

polar

BF_3



Molecular geometry:
Trigonal planar

nonpolar

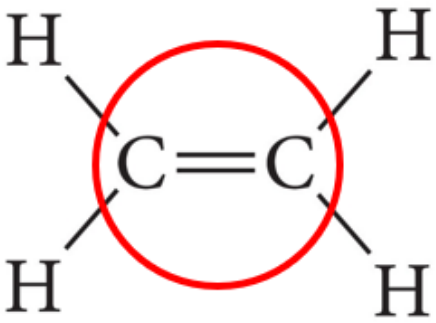
Organic Lewis Structures

alkene:	eth**ene**	$H_2C{=}CH_2$
alkyne:	eth**yne**	$H_2C{\equiv}CH_2$
alcohol:	methan**ol**	$CH_3\mathbf{OH}$
carboxylic acid:	ethan**oic acid** (acetic acid)	$CH_3\mathbf{COOH}$
amine:	methyl**amine**	$CH_3\mathbf{NH_2}$
aldehyde:	ethan**al**	$CH_3\mathbf{CHO}$

Organic Lewis Structures

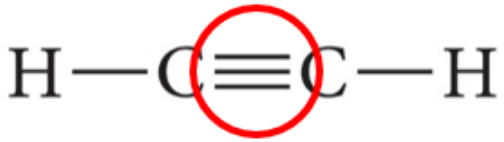
ethene (C_2H_4)

CH_2CH_2



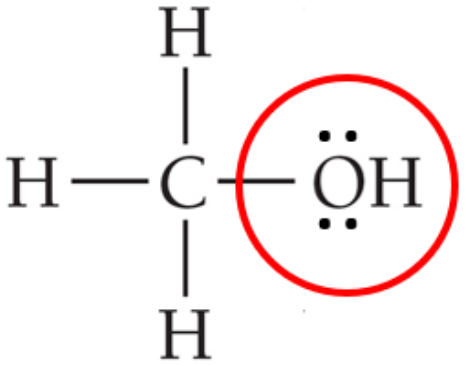
ethyne (C_2H_2)

CHCH



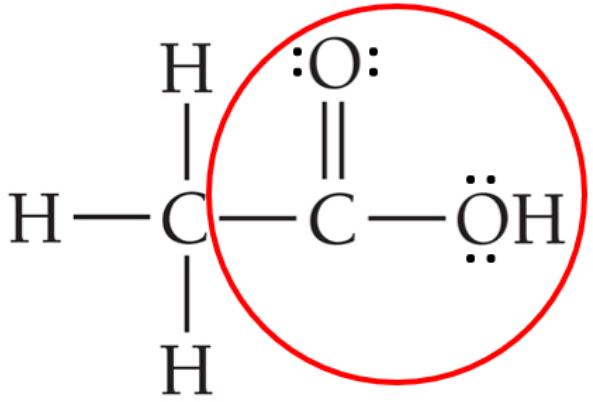
methanol

CH_3OH



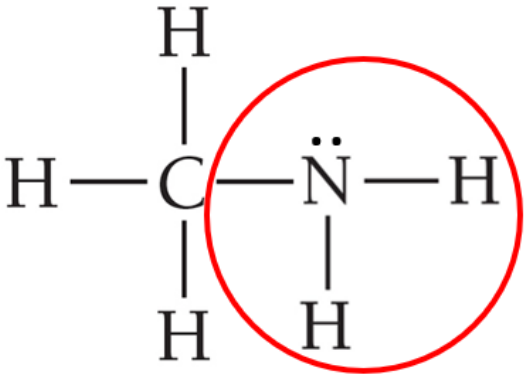
acetic acid

CH_3COOH



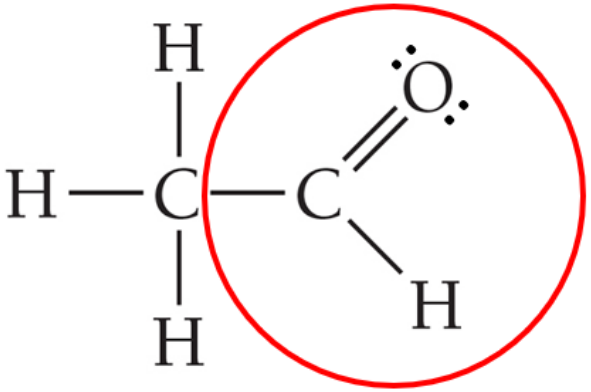
methylamine

CH_3NH_2



acetaldehyde

CH_3CHO



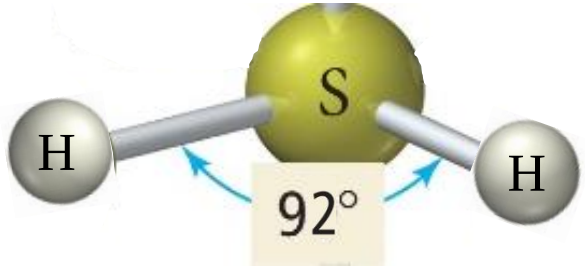
Chapter 6: Chemical Bonding II

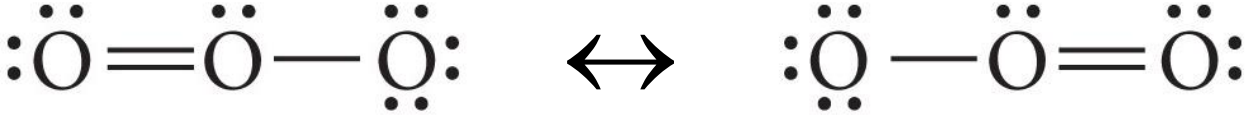
Some questions we'll try to answer

- What are some limitations to Lewis structures and VSEPR in predicting the shapes of some molecules?
- What actually happens on the atomic level when covalent bonds are formed?
- What are the different ways in which orbitals can overlap?
- How does the type of orbital overlap impact the behavior of the resulting chemical bond?
- What are hybrid orbitals and how do they fundamentally differ from unhybridized orbitals?

Problems with Lewis Theory of Bonding

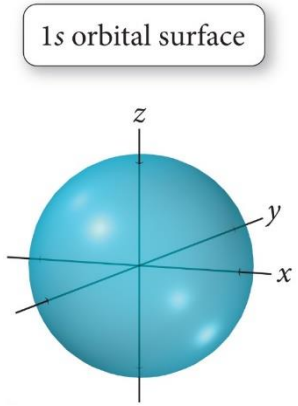
- Lewis theory (and structures) does not give good numerical predictions regarding bond lengths and energies
- Lewis theory gives good *approximations* of bond angles for a molecule but can't account for actual bond angles for many molecules
  - example H_2S is 92° not ≈ 109.5°
- molecules showing resonance are not adequately dealt with in Lewis theory
- there are other bonding theories that improve upon Lewis theory and can explain things that Lewis theory can't
  - example: how are double bonds and triple bonds made?

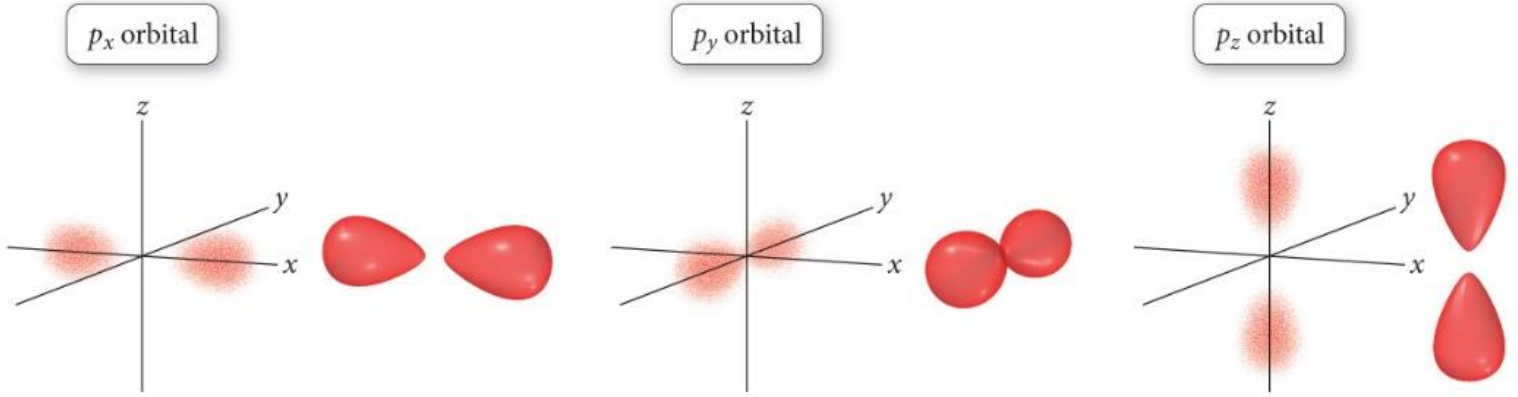


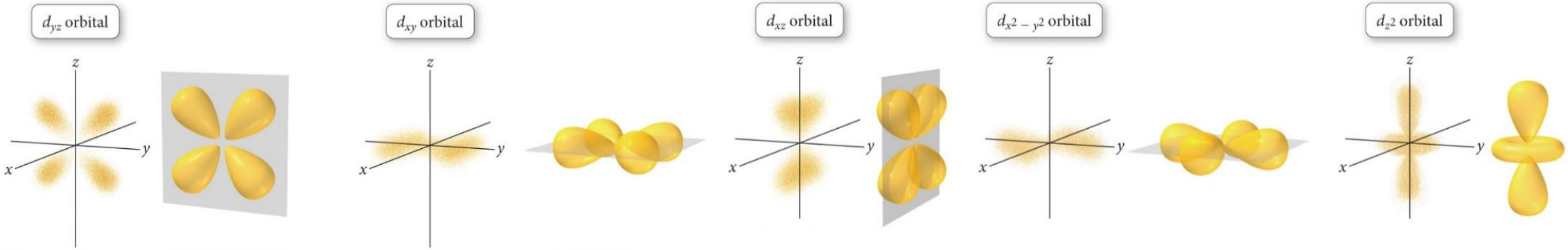


RECALL FROM CHAPTER 2

s-orbitals, p-orbitals, d-orbitals **unhybridized**

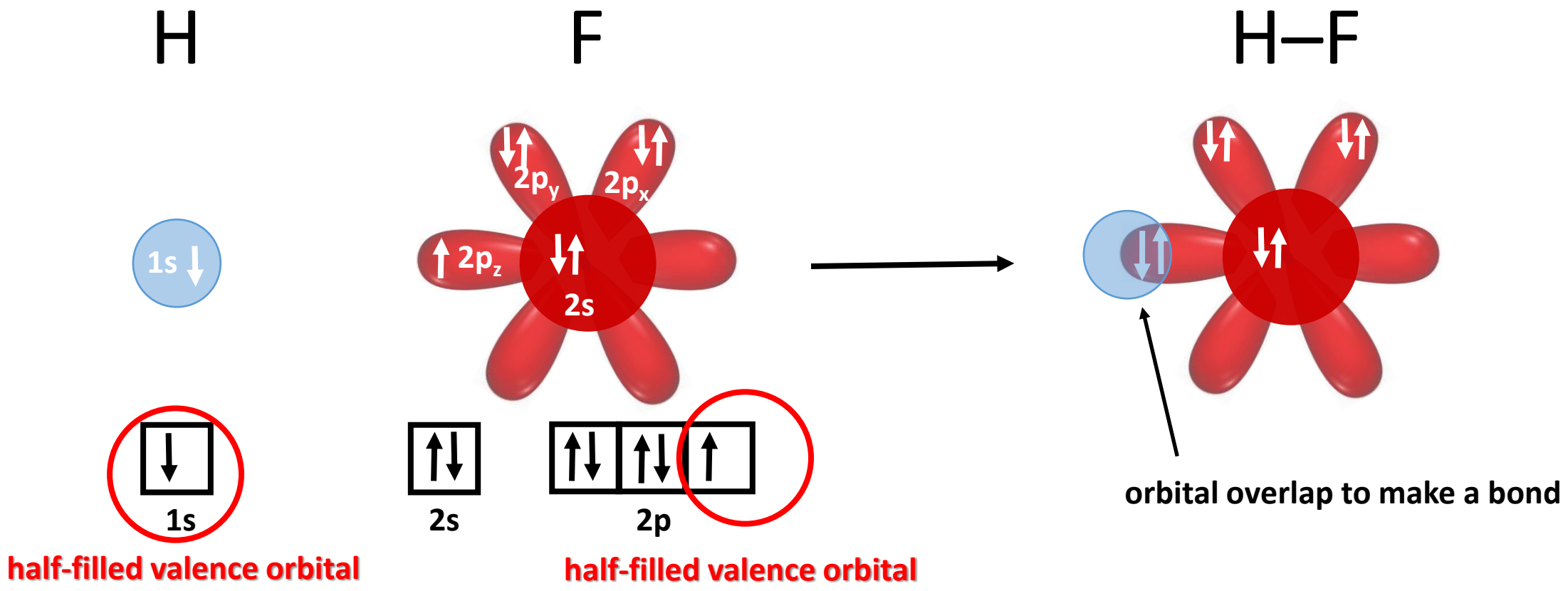






Valence Bond Theory: an Improvement to Lewis Theory

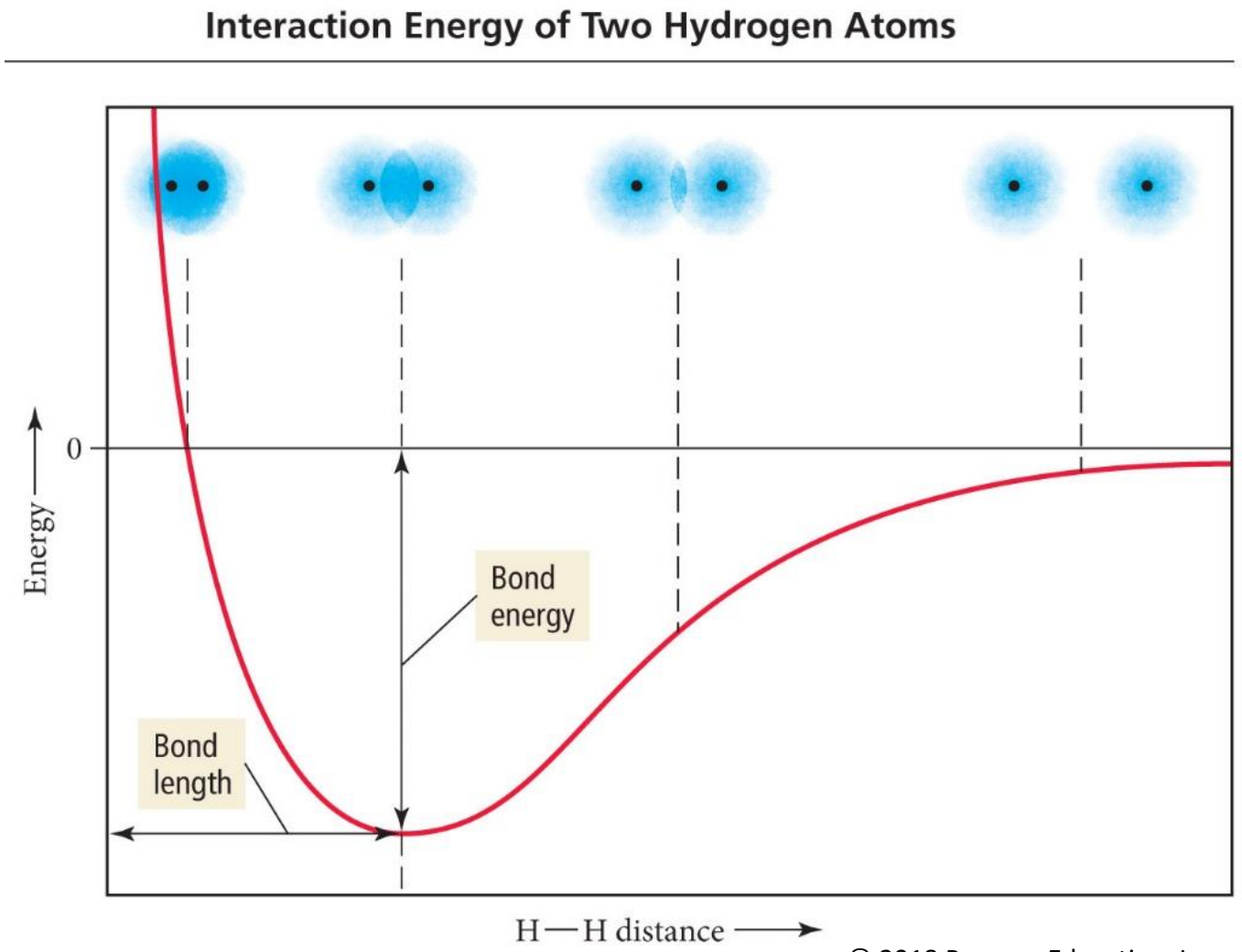
- a chemical bond is the overlap between two ***half-filled*** atomic orbitals
  - the orbitals are necessarily in the **valence** shell (i.e., outermost)
  - the electrons in the orbitals must have paired spins
  - the bonding in HF
  - coordinate covalent bonds can also occur



Valence Bond Theory: an Improvement to Lewis Theory (continued)

- the amount of orbital overlap between atoms represents an energy minimum and a balance between e^-/nucleus attraction and nucleus/nucleus repulsion
  - example: H_2

- the geometry of the overlapping orbitals determines the shape of the molecule



An Issue with Methane (CH_4) and Valence Bond Theory

valence bond theory does not adequately account for certain attributes of CH_4

- C readily makes 4 identical bonds with H
- all H–C–H bond angles are 109.5°
- carbon must **NOT** be using its regular s-orbital and p-orbitals to make bonds
- there **must** be more to the story…