Good Morning

- Grab a whiteboard and draw the diagrams for:
- CO_2:
- H_2O:
- SCl_2:

Today

- Finish 8.2
- 8.3: Bonding Theories
- Begin 8.4: Polar Bonds and Molecules
- Lab Preview

Homework

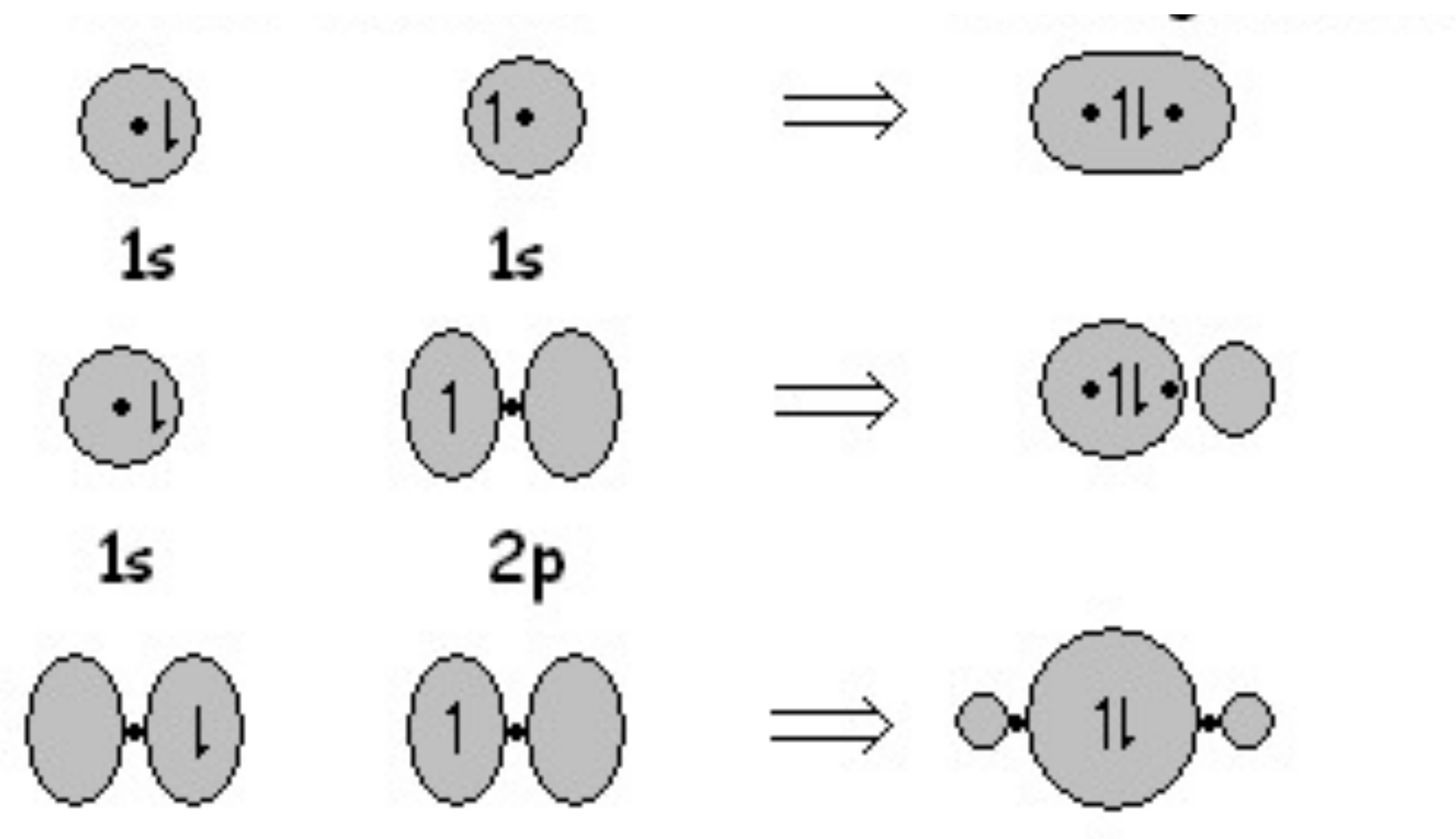
- From Monday: 1-3, 6, 13-15, 20, 21, 39, 41-46, 51.
- From Today (time permitting): 24, 27, 28, 32-35, 37, 38, 54, 57 (58-60)

8.3 Bonding Theories

- Relationships between atomic and molecular orbitals.
- The VSEPR theory for predicting the shapes of molecules.
- Orbital Hybridization.

Molecular Otbitals

- To this point we have assumed that electrons that are shared by two atoms occupy each of the atoms electron orbitals.
- It turns out that the orbitals "merge" to form new shapes.
- These are molecular orbitals.



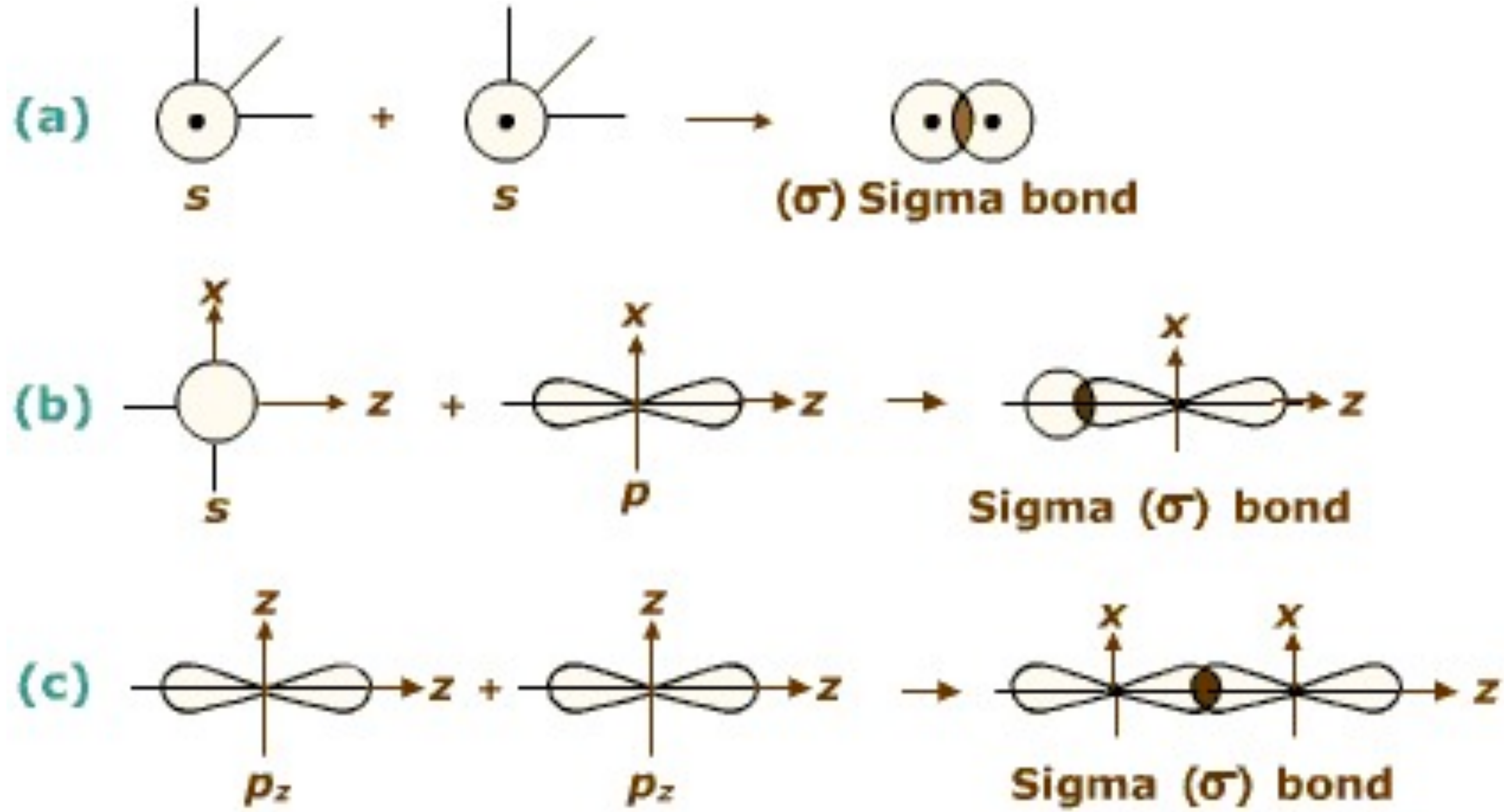
Examples

Molecular Orbitals

- These orbitals are shared by the molecule.
- Types:
  - Sigma bond (σ)
  - Pi bond (π)

Sigma Bonds σ

- Two atoms form a **symmetrical** orbital in reference to the "bond axis."
- The bond axis is a line that could be drawn along both nuclei.
- Nuclei are attracted to grater electron density.
- End to end overlap.



Formation of a sigma bond due to (a) The *s* - *s* overlap (b) The *s* - *p* overlap (c) The p_z - p_z overlap

Sigma Bonds

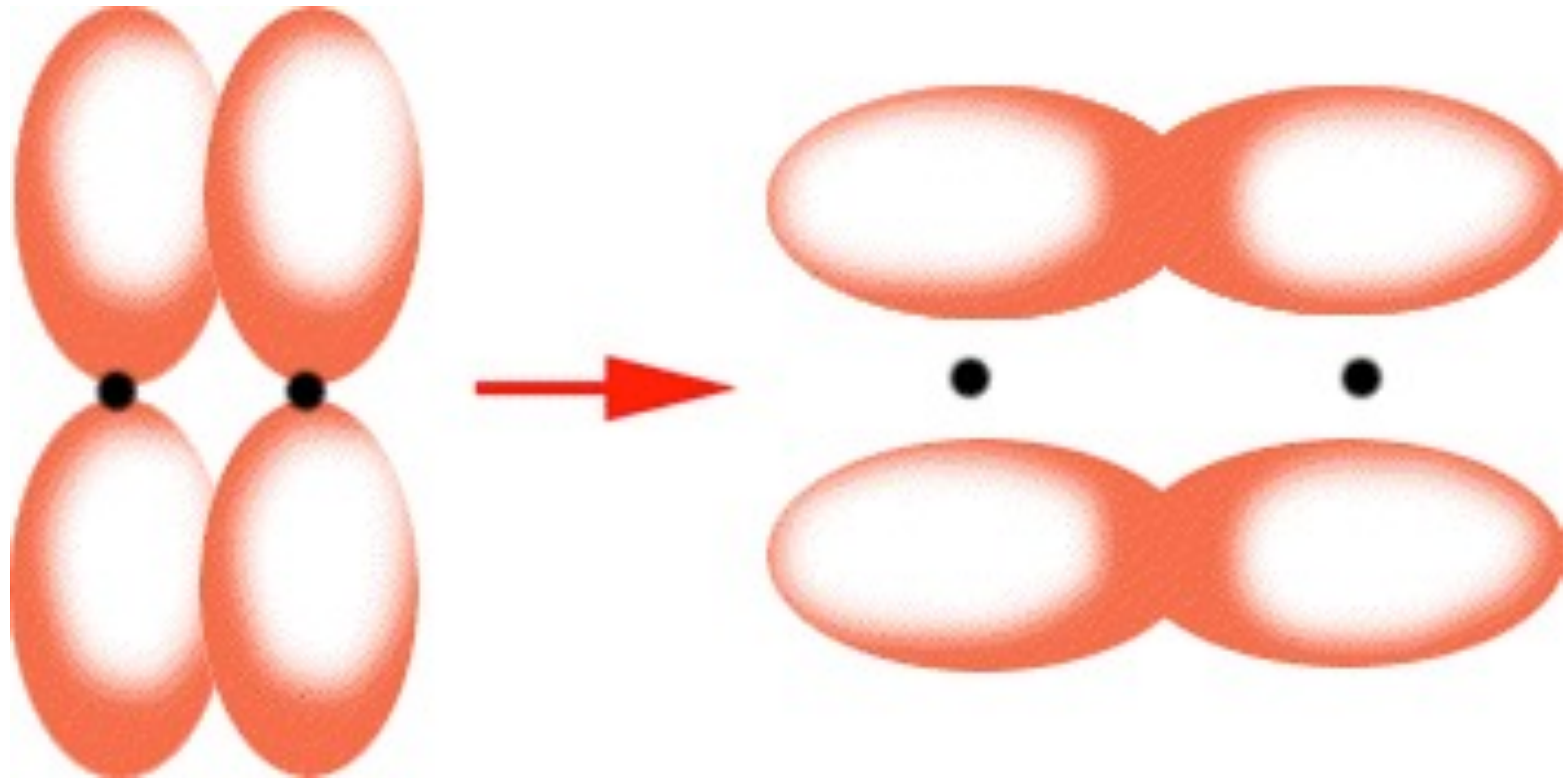
Examples

- Hydrogen gas:

- Fluorine gas:

Pi Bonds π

- Side by side overlap.
- Electrons are most likely to be found away from the bond axis.
- "Sausage-shaped"
- Less overlap than sigma bonds.
- Less attraction to nucleus due to proximity.



Example

Fact Check

- Pi or Sigma bonds are stronger?
- Why?
- Draw one of each.
- Write questions if you have them.

VSEPR Theory

- Valence
- Shell
- Electron
- Pair
- Repulsion

VSEPR Theory

- A better way to represent molecular bonding in 3-D compared to dots.
- Think of Methane CH_4
- The hydrogen atoms want to repel each other.

Formal VSEPR

- The repulsion between electron pairs causes molecular shapes to adjust so that valence electron pairs are as far apart as possible.
- Think like charges repel.

VSEPR Theory

Video Intro

How Much Space?

- The space between bonding electron pairs depends on many things.
- We will explore the reasons why different atoms make different shapes.
- Depends on the number and types of bonds.

Shape Shifting

- Shared bonding pairs repel one another.
- Unshared bonding pairs also repel pairs of electrons.
- This will lead to some things being "bent."
- All shapes are listed on p. 233.

Method

- Draw the dot diagram.
- Note where there are shared electron pairs.
- Note non-shared electron pairs.
- All electron pairs repel.

Triatomic Example

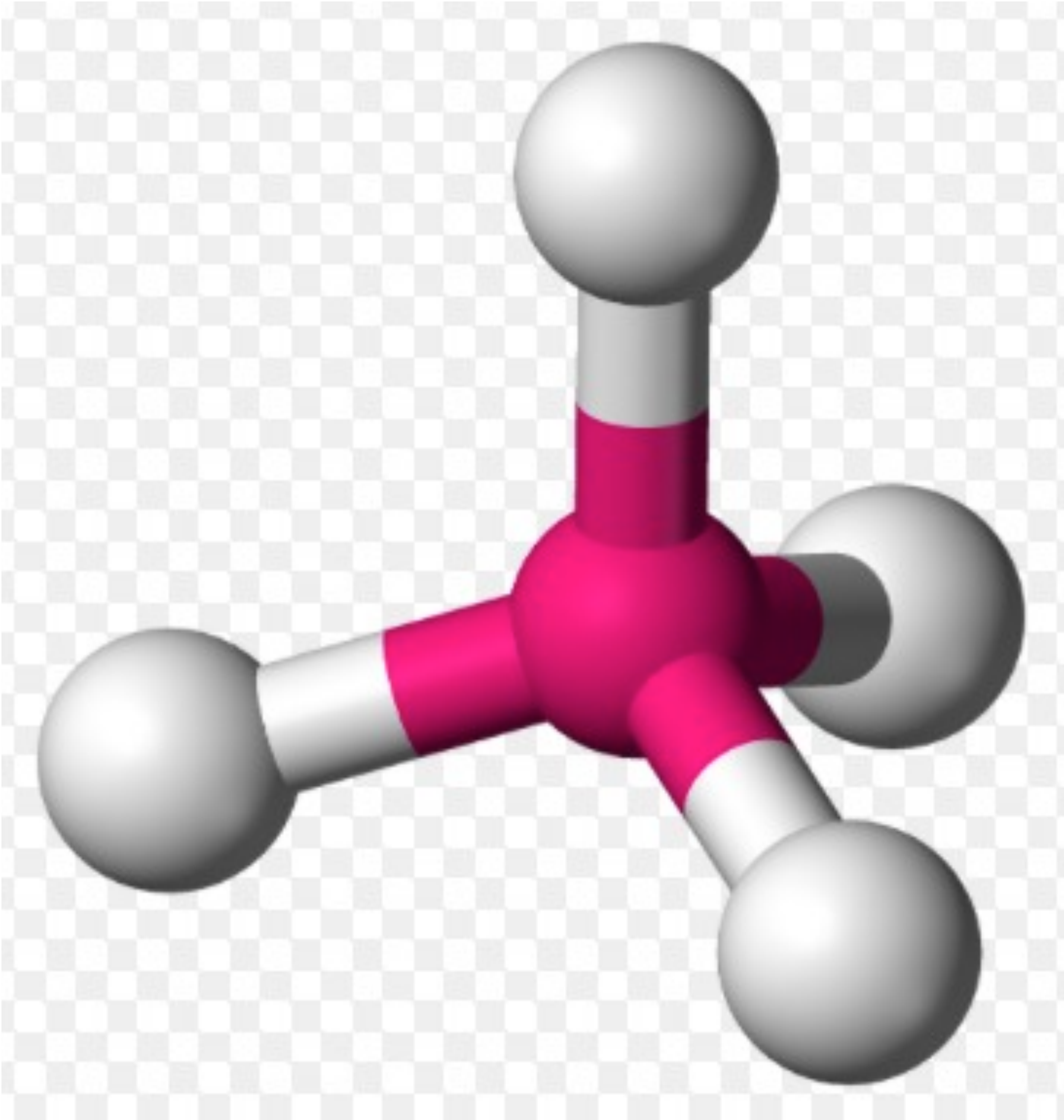
- H_20:
- CO_2:

Tetrahedral

- Tetra-Four
- Hedral-number of flat surfaces on an object.
- These molecules have a central atom and four single covalent bonds.

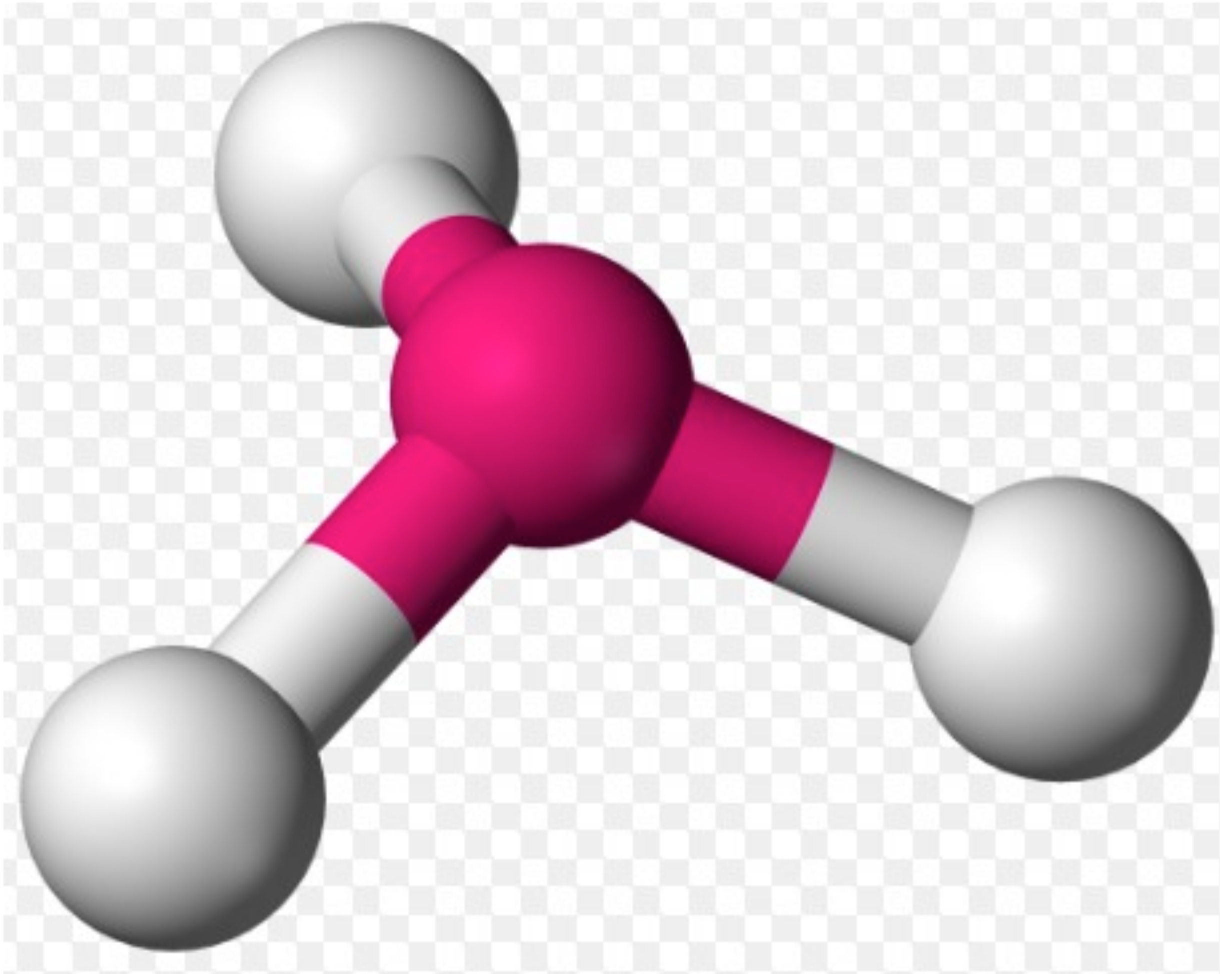
Example: Methane

- $CH4$:



Pyramidal: Ammonia

- NH_3:

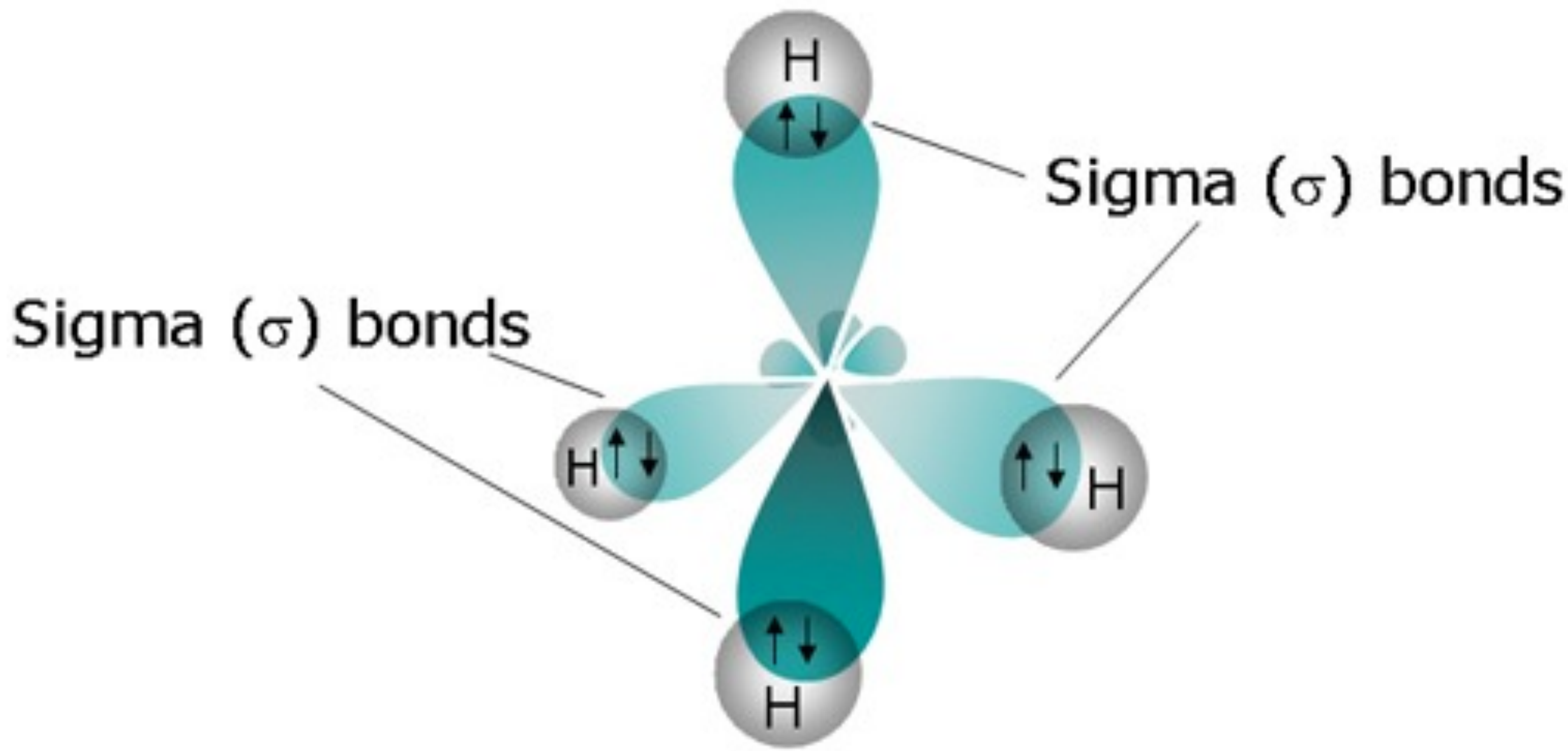


Hybrid Orbitals

- Several atomic orbitals mix together.
- The shapes change.
- They form the same **total** number of orbitals.

Example: Methane

- Carbon has 4 total valence electrons.
- 2 in the s orbital and 2 in the p orbital.
- Shared s orbitals are stronger than shared p orbitals.
- The s orbital electrons move out to the p orbital and the shape is symmetrical.



Methane

8.4 Polar Bonds and Molecules

- How does electronegativity determine the distribution of charge in a molecule?
- What happens when polar molecules are between metal plates?

8.4 Polar Bonds and Molecules

- Strength of intermolecular attractions in molecules compared to ionic compounds.
- Why do "network solids" have a high melting point?

Polar & Non-Polar Molecules: Crash Course Chemistry #23

Bond Polarity-4:45

Non-polar Covalent

- The pull of each atom on the electrons that it shares is equally distributed.
- These have a balance of charge.
- Diatomic Molecules like hydrogen and oxygen gas are examples of this.

Polar Covalent Bonds

- One atom has a higher electronegativity than another atom.
- The more electronegative atom attracts the electrons more.
- One region is more negative than another.

Difference in Electronegativity

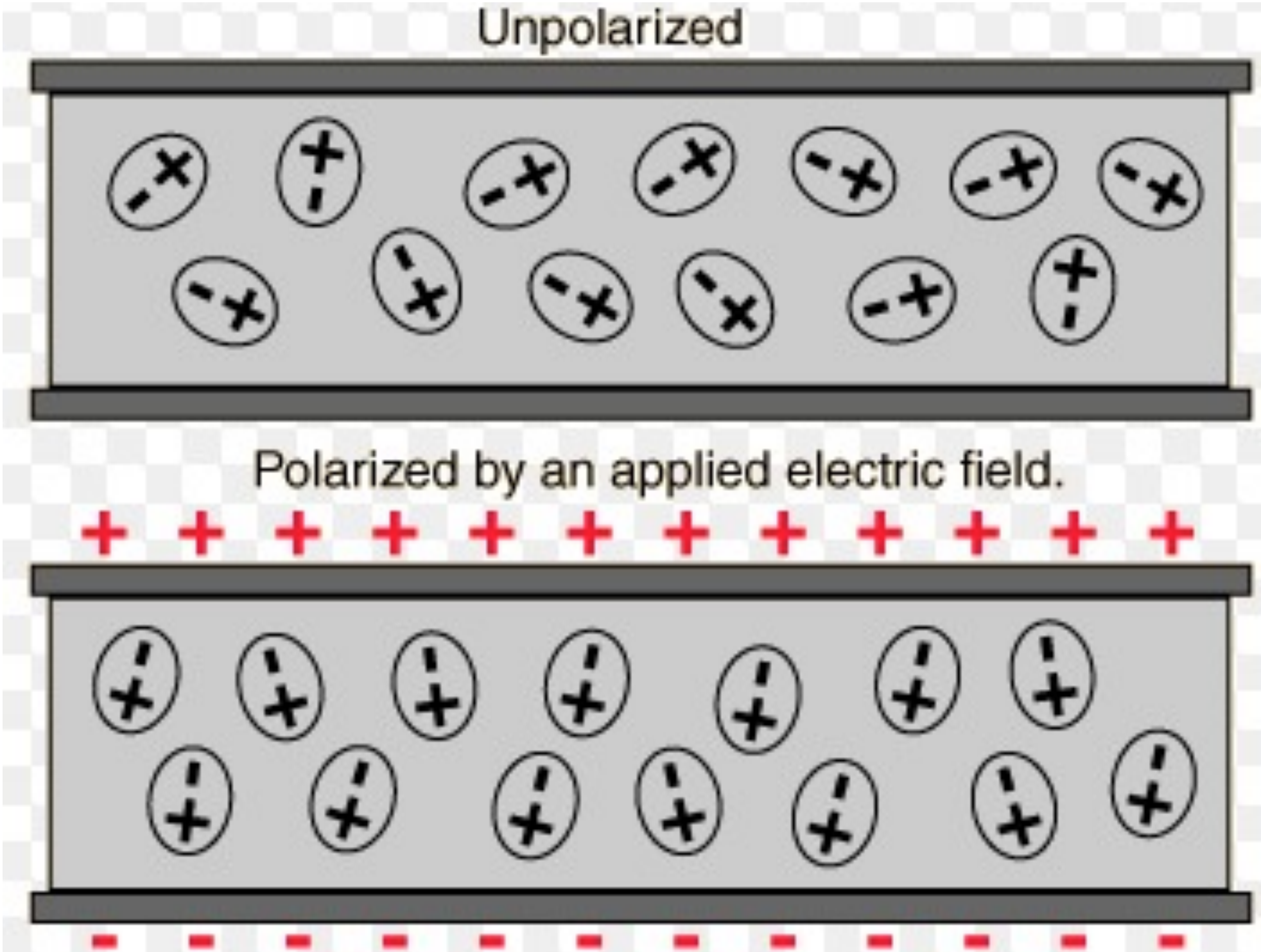
- 0.0-0.4: Non-polar covalent bond.
- 0.4-1.0: Moderately polar covalent.
- 1.0-2.0: Very polar covalent.
- More than 2.0: Ionic bond.

H 2.1																	He
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
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 1.8	Ni 1.8	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 1.1	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	Ti 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn 2.4
Fr 0.7	Ra 0.7	Ac 1.1	Unq	Unp	Unh	Uns	Uno	Une									

Ce 1.1	Pr 1.1	Nd 1.1	Pm 1.1	Sm 1.1	Eu 1.1	Gd 1.1	Tb 1.1	Dy 1.1	Ho 1.1	Er 1.1	Tm 1.1	Yb 1.1	Lu 1.2
Th 1.3	Pa 1.5	U 1.7	Np 1.3	Pu 1.3	Am 1.3	Cm 1.3	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr

PT Electronegativity

Water



Polar Molecules Between Charged Plates

Dipole Effect

- Polar molecules have charges on different sides.
- Dipole: two poles.
- These things respond to electrical charges.

Attractions Between Molecules

- Van der Waals Forces - The two weakest forces between groups of the same kind of molecule.
- 1) Dipole interactions - positive end of one molecule is attracted to the negative end of another molecule of the same kind.

Attractions Between Molecules

- 2) Dispersion Forces - electrons from one molecule affect the electrons of another molecule.
- Example: + + + + +

Hydrogen Bonds

- Hydrogen that is bonded in a molecule tends to have it's 1 electron towards the center of the molecule.
- The part facing out is more positive.
- This is attracted to negative parts of molecules.

Intermolecular Attractions

- These are "network solids"
- All of the atoms are covalently bonded to each other.
- These molecules bond to each other and for lattice structures the way that ionic compounds do, but with covalent bonds instead of electrostatic forces.

Physical Properties

- VERY high melting point.
- Not water soluble.
- Very low electrical conductivity.
- Ex: Diamond do not melt. They vaporize at 3500 degrees Celsius.