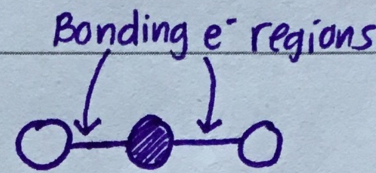
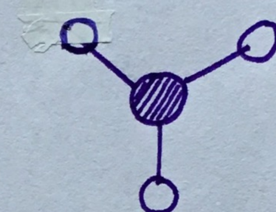
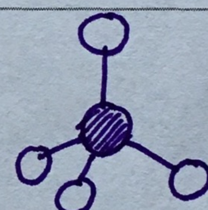
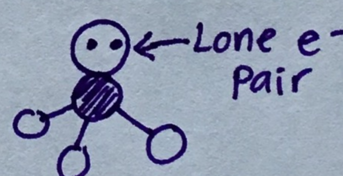
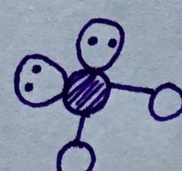
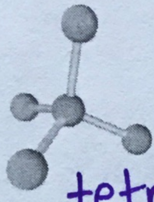
Molecular Shapes

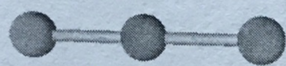
Valence Shell Electron Pair Repulsion Theory (VSPER)

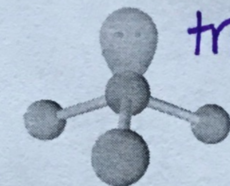
# of Total e^- Regions	# Shared e^- Regions	# Lone e^- Pairs	Shape Name	Picture	Examples
2	2	0	Linear		CO_2, Cl_2, SiO_2
3	3	0	Trigonal planar		BF_3, CH_2O
4	4	0	Tetrahedral		CH_4, CBr_4
4	3	1	Trigonal Pyramidal		NH_3, PCl_3
4	2	2	Bent		SCl_2, H_2O

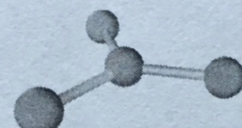
Molecular Polarity and Intermolecular Forces

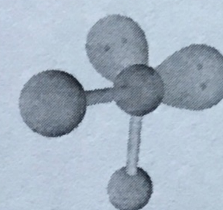
Molecular Geometry: Name each molecular shape. (linear, trigonal planar, tetrahedral, bent, pyramidal)

1. 
tetrahedral

2. 
linear

3. 
trigonal pyramidal

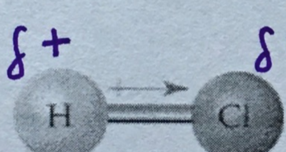
4. 
trigonal planar

5. 
bent

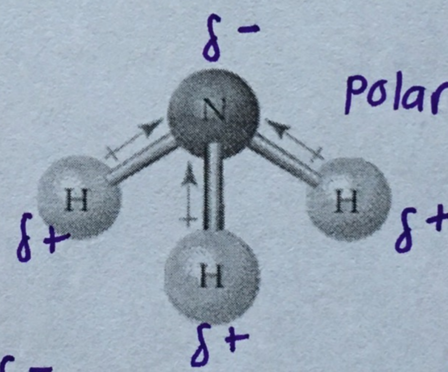
Molecular Polarity: Determine if each molecule is polar or nonpolar.

Nonpolar Molecule: equal distribution of electrical charge throughout the molecule. The charge distribution is symmetric around the central atom.

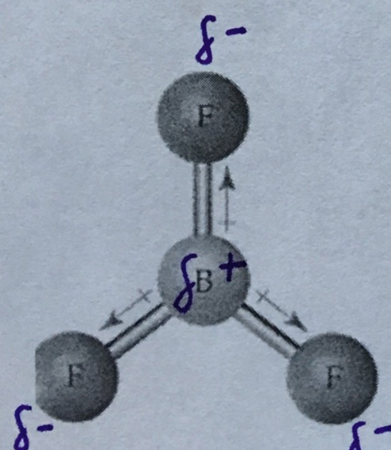
Polar Molecule: unequal distribution of electrical charge throughout the molecule. The charge distribution is asymmetric around the central atom.



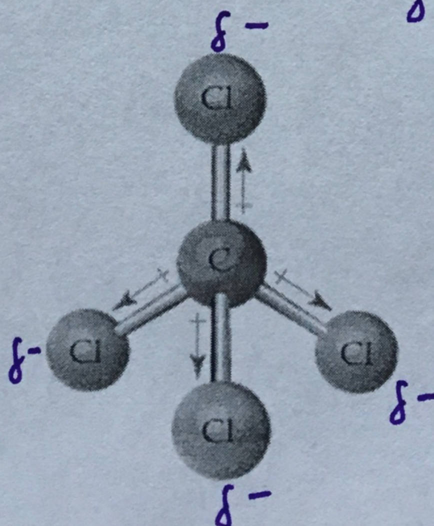
polar



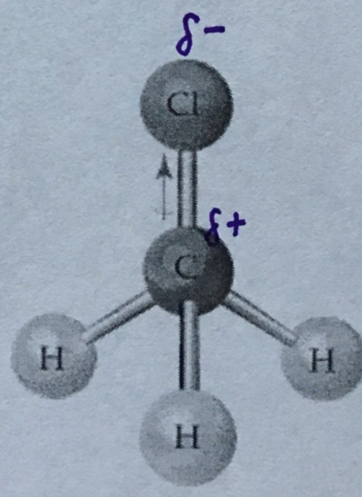
Polar



nonpolar



nonpolar



polar

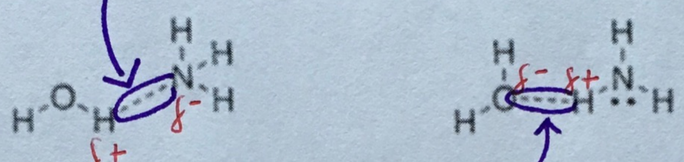
Intermolecular Forces (IMF): attractions **between** molecules (NOT the covalent bonds in the molecule!)

Stronger the attraction between molecules = harder to separate molecules from one another = more energy to separate molecules = higher melting and boiling point

3 Types of IMF's

1. Hydrogen bonding: strongest IMF; attraction between H in one molecule and N,O, or F in a different molecule

H-bond = $H^{\delta+}$ attracted to $N^{\delta-}$, $O^{\delta-}$, or $F^{\delta-}$ in a different molecule



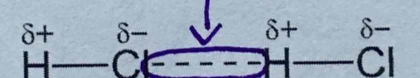
H-bond

* N, O, and F are the most electronegative elements

∴ N, O, and F have very negative δ−, making attraction to δ+ very strong

2. Dipole-dipole: attraction between polar molecules

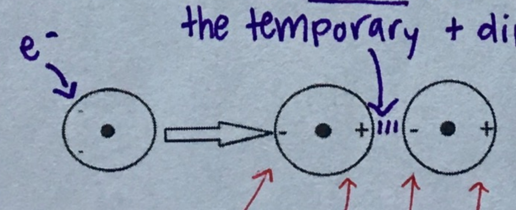
dipole-dipole attraction = δ^- dipole of one molecule attracted to δ^+ dipole of another molecule



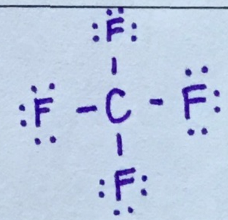
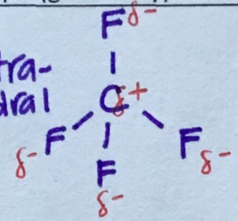
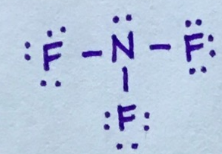
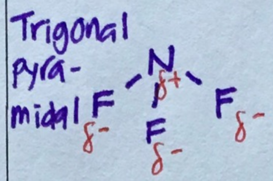
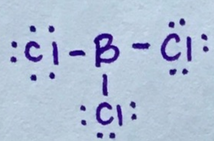
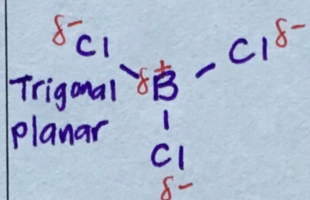
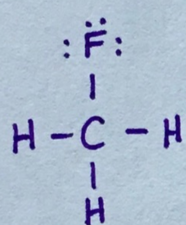
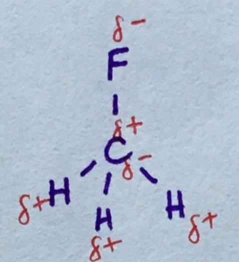
Polar molecules have permanent dipoles

3. London forces: weakest IMF; present between all molecules
*nonpolar molecules only have London forces

London Forces
the temporary + dipole of one molecule attracted to the temporary − dipole of another molecule



movement of e⁻ cause temporary positive and negative dipoles to form

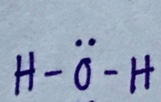
Formula	Lewis Structure	Draw and Name Molecule Shape (geometry)	Bond Polarity	Molecular Polarity	IMF
		use Lewis and e^- regions around center atom	calculate ΔEn; $\Delta En > 0.3$ = polar bond	examine molecule shape w/ δ^+ δ^-	depends on molecular polarity
CF_4		Tetrahedral 	F: 4 C: 2.5 $\Delta En = 1.5$ Polar bond	δ^- charges equally surround C atom ∴ Nonpolar Molecule	London
NF_3			F: 4 N: 3 $\Delta En = 1$ Polar bond	δ^- charges only below N atom ∴ unequal distribution ∴ Polar Molecule	London and Dipole-dipole
BCl_3			Cl: 3 B: 2 $\Delta En = 1$ polar bond	equal distribution of charge around B atom ∴ Nonpolar Molecule	London
CH_3F		 Tetrahedral	F: 4 C: 2.5 $\Delta En = 1.5$ polar bond C: 2.5 H: 2.1 $\Delta En = 0.4$ Polar bond	unequal distribution of charge around C atom (end w/ F much more negative) ∴ polar molecule	London and Dipole-dipole *Not H-bond b/c H is not bonded to N, O, or F

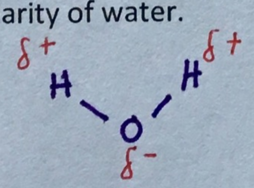
Solubility – in order for a solute to dissolve in a solvent, the solute must be attracted to the solvent (in other words, solute and solvent must have similar types of intermolecular forces)

*Polar solvents dissolve polar and ionic solutes.

*Nonpolar solvents dissolve nonpolar solutes.

1. Determine the polarity of water.





$\Delta En = 3.5 - 2.1 = 1.4$ = polar bond so add δ^+/δ^-

H_2O is a polar molecule b/c the oxygen end is δ^- and H end is δ^+
∴ unequal distribution of charge

2. Which of molecules from the table above will dissolve in water?

Since H_2O is a polar molecule w/ δ^+ and δ^- ends, H_2O attracts molecules that also have δ^+ and δ^- ends (i.e. other polar molecules).
∴ NF_3 and CH_3F dissolve in H_2O.