(9) **Today**

3.1 Lewis Structures

3.1.2 Expanded Octets

3.1.4 Failure of Lewis Structures to Predict Unusual Cases

3.2 VSEPR

Next Class (10)

3.2 VSEPR

3.3 Molecular Polarity

4.1 Symmetry elements and Operations

(11) **Second Class from Today**

4.1 Symmetry elements and Operations

4.2 Point Groups

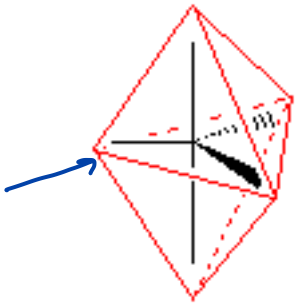
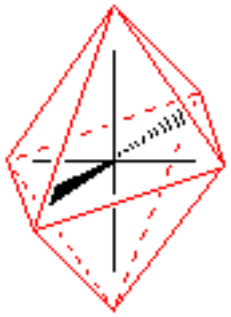
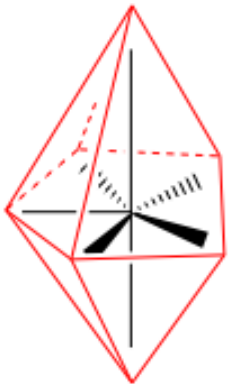
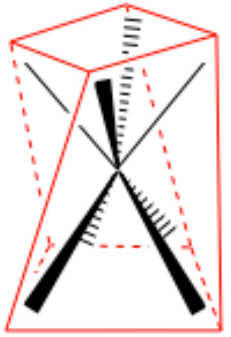
Third Class from Today (12)

4.2 Point Groups

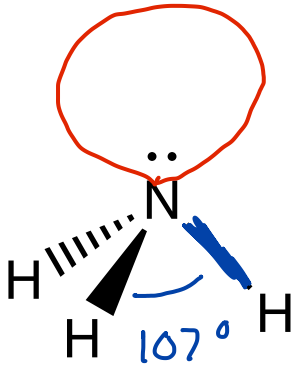
4.3 Character Tables

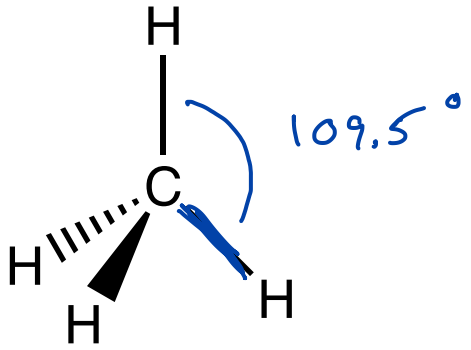
Electron pairs on the central atom of a molecule repel each other and the most stable arrangement is the one that minimizes the repulsion.

1	2	3	4
H – H	O ≡ C ≡ O (180°)	120° trigonal planar	109.5 tetrahedral
PF_5	$[PF_6]^-$	$[ZrF_7]^-$	$[XeF_8]^{2-}$

5	6	7	8
 trigonal bipyramidal	 octahedral	 pentagonal bipyramidal	 square antiprismatic

Five, seven, and eight coordinate examples from https://chem.libretexts.org/Bookshelves/Inorganic_Chemistry/Inorganic_Chemistry_(LibreTexts)/03:_Simple_Bonding_Theory/3.02:_Valence_Shell_Electron-Pair_Repulsion

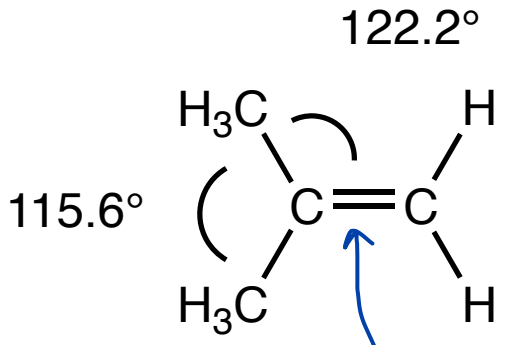




lone-pair e⁻'s don't have a second atom to help confine them

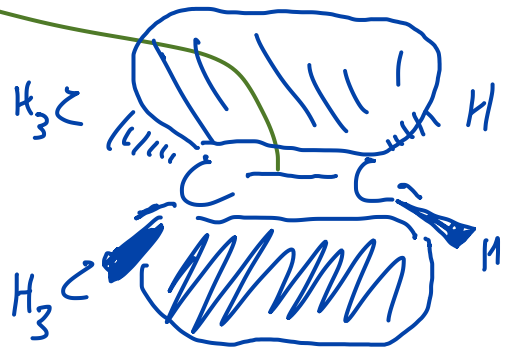
e⁻'s that are part of a bond are confined by the two atoms

lone-pair e⁻'s take up more space than e⁻'s in a σ bond (single bond)

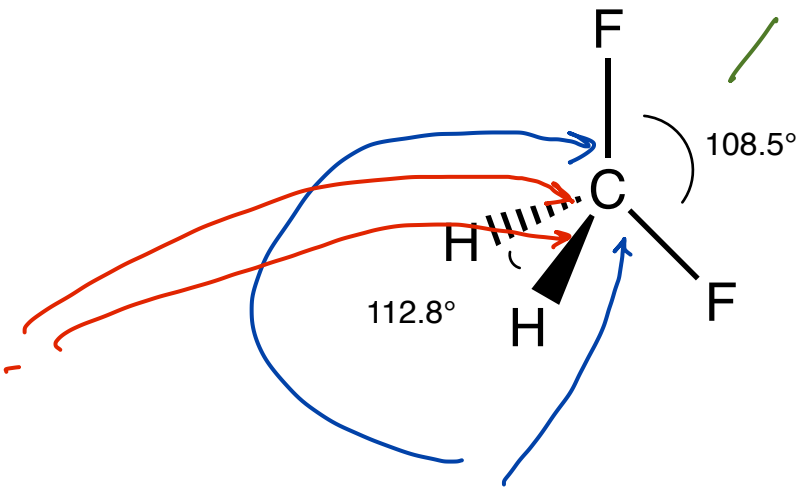


σ-bond e⁻'s confined between atoms

this π bond is taking up more space



π e⁻'s spread more than σ-bond e⁻'s and the cause other bond angles to be reduced



this bond angle is smaller even though the F atom is larger

more e⁻ density near C more repulsion larger bond angle

because the F is electronegative... They draw e⁻ density away from the C atom, lowering the e⁻ density near the C where the e⁻'s would experience the most repulsion. Less e⁻ density means less repulsion, so bond angle can be smaller

Bonds to eneg atoms take up less space as compared to bonds to more electropositive atoms

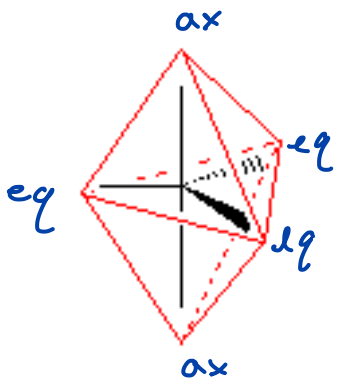
trigonal planar

the shape is based on a trigonal planar arrangement of the sets of electrons but this is not a trigonal planar molecule

V-shaped molecule

trigonal pyramidal

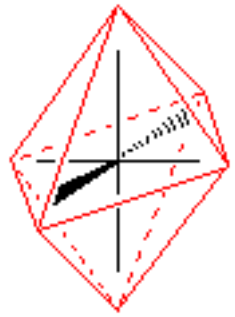
bent



ax - center - eq bond angle = 90°

eq - center - eq bond angle = 120°

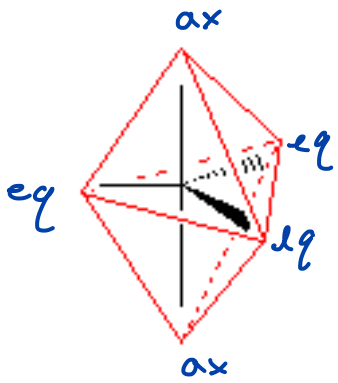
equitorial positions have a bit more space to spread out ... two bond angles are 120°, whereas in the axial positron bond angles are always 90°.



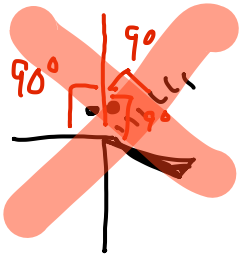
in the octahedral geometry the bond angles are 90°

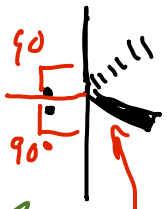
ax - center - eq bond angle = 90

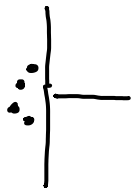
eq - center - eq bond angle = 120°

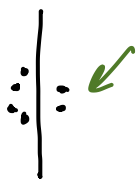


put the lonepair e⁻ in the place where they have the most space









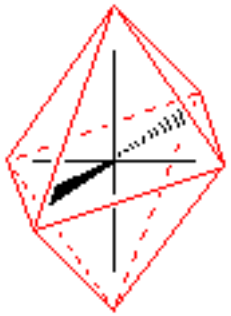
lone-pair e⁻'s in equatorial

see-saw geometry

in the plane the bond angles are 120° ... Farther away

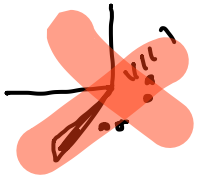
T-shaped

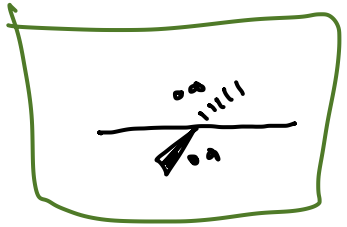
linear



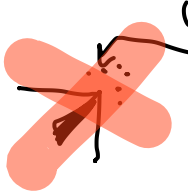
In the octahedral geometry the bond angles are 90°

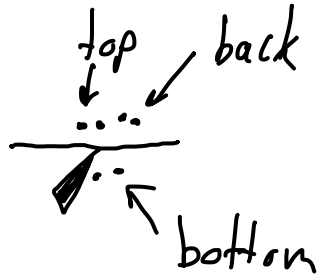
square pyramidal



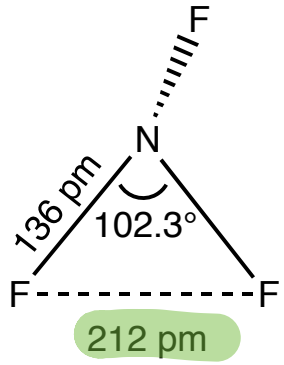


square planar

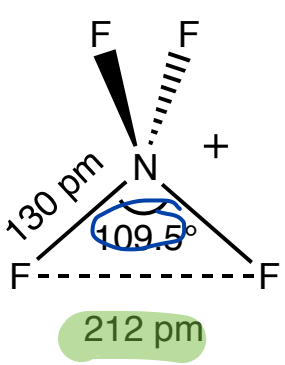




In ligand close packing, the bond length and the bond angle is determined by the packing of the pendant atoms

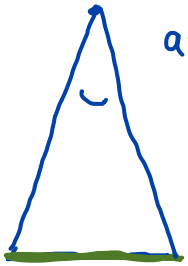


NF_3

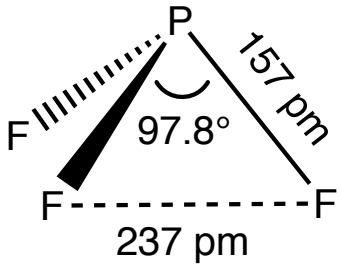


NF_4^+

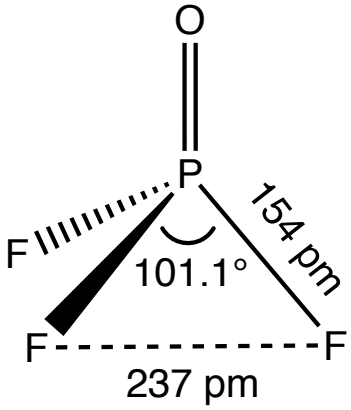
bond angle & bond length determined by the distance between the pendant atoms



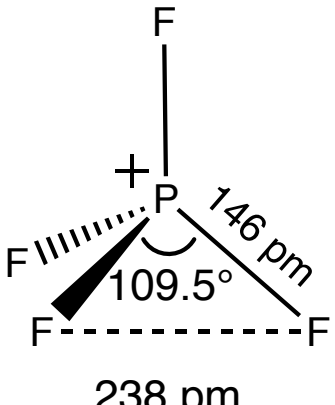
as the bond length gets longer the angle gets smaller



PF_3



OPF_3



$^+PF_4$

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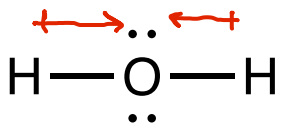
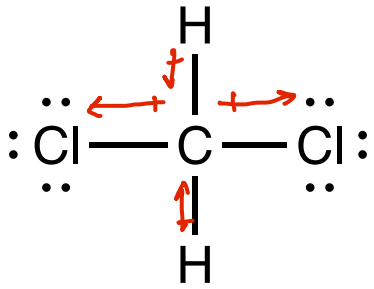
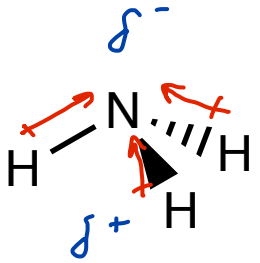


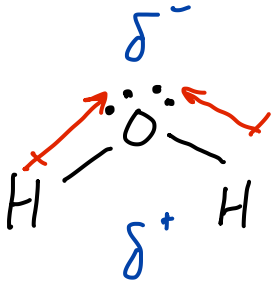
polar
1.85 D

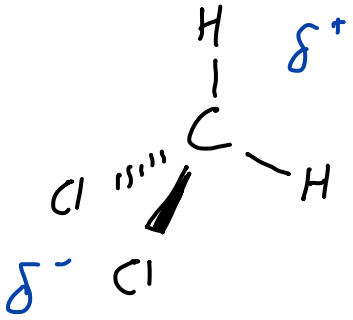


polar
1.6 D



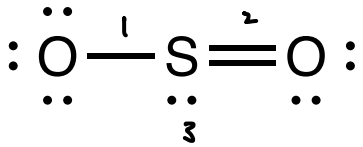
polar
1.47 D





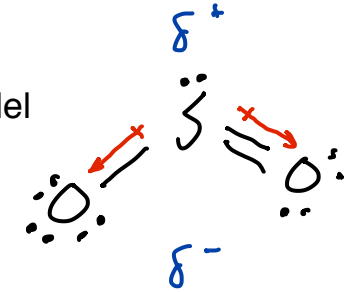
Draw Lewis Structure

Predict Shape



Find polar bonds

prediction based on Gen Chem model



Draw Lewis Structure

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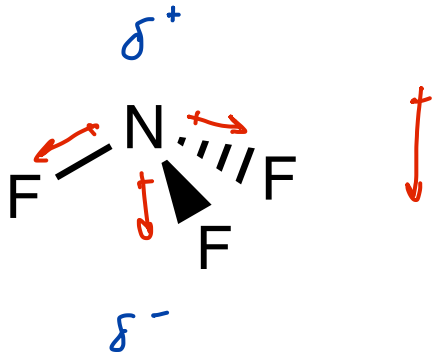
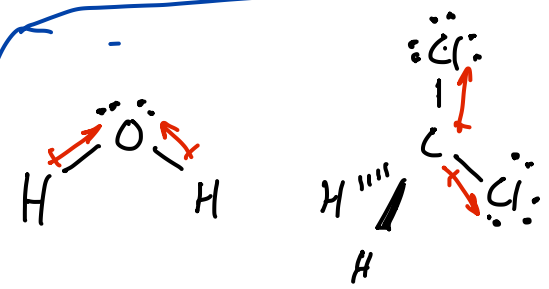
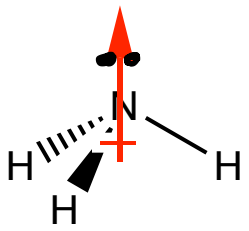


Draw Lewis Structure

Predict Shape

Find polar bonds

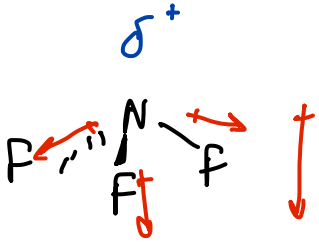


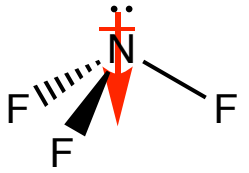


1.47 D

Does prediction based on polar bonds match the position of the lone-pair e⁻ ?

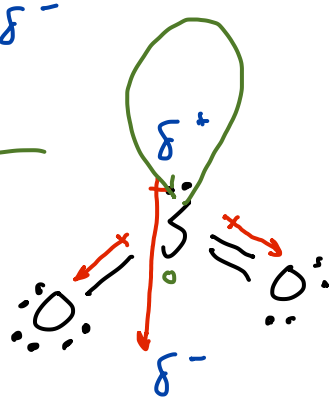
negative end of molecular dipole by lp e⁻'s + ⊕ and near area without lp e⁻? yes... then you can safely make a prediction.

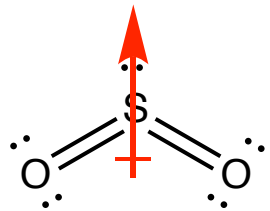




0.23 D

but there is ___ is pair of lone pair e⁻'s here





1.63 D

If not don't make the prediction