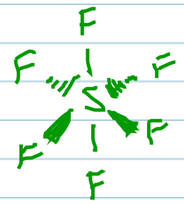
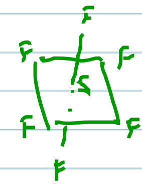
Octahedral SF_6

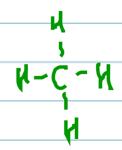
1 lone pair square based pyramid

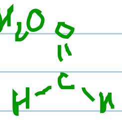
2 lp's square planar
3 lp's T-shaped
4 lp's linear

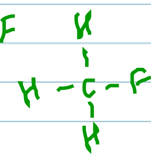
Geometry → one consequence ⇒ POLARITY

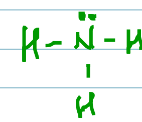
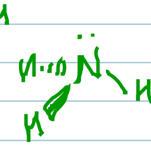
* assymetric molecules tend to be polar

1) different outside atoms
2) lone pairs on central atom
(exceptions).

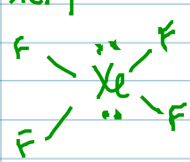
CH_4

NONPOLAR

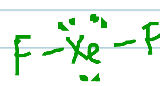
CH_2O


CH_3F

POLAR

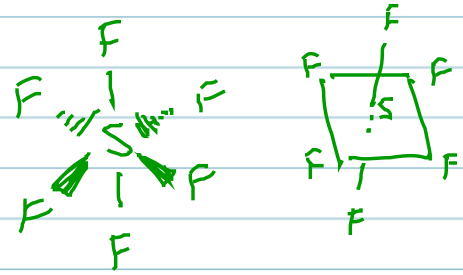
NH_3


POLAR

"lone pair exceptions"

XeF_4

square planar
NONPOLAR

XeF_2

linear
NONPOLAR

Octahedral SF_6



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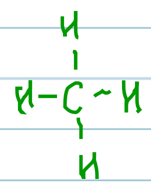
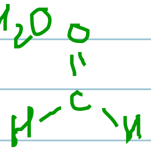
CH_4

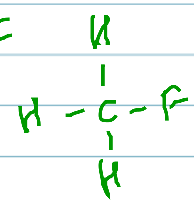


NONPOLAR

CH_2O

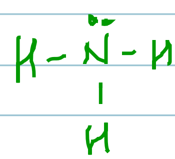


CH_3F

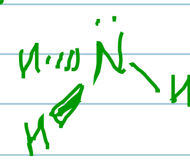


POLAR

NH_3

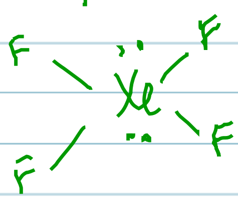


POLAR



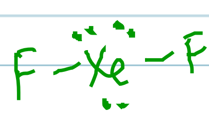
"<u>lone pair exceptions</u>"

XeF_4



square planar
NONPOLAR

XeF_2



linear
NONPOLAR

Polarity

polar molecules have a (+) and (−) end

$\delta+$ $\delta-$
H − F

dipole moment → a measure of the amount of polarity

unit "debye"
D

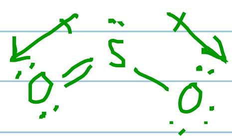
$\mu = Q \times r$

distance separating the charges

magnitude of charge

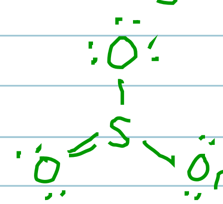
↑ μ, more polar

SO_2



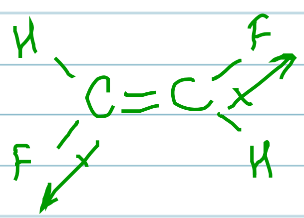
angular
POLAR
$\mu = 1.6D$

SO_3

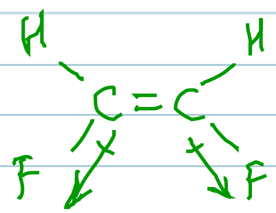


trigonal planar
NONPOLAR
$\mu = 0.0D$

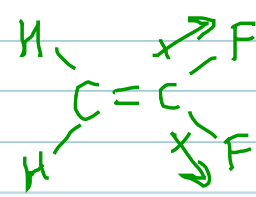
$C_2H_2F_2$



NONPOLAR



POLAR



POLAR

LEWIS THEORY

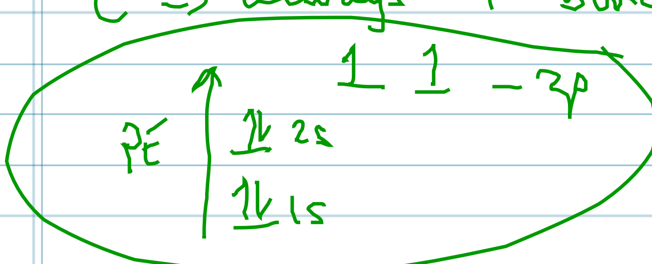
strengths: shows nature of covalent bond
↳ shared e^- pair
predicts 3-D shapes (VSEPR)

weakness: ignores ΔE altogether
does not account for Δ bond lengths

Valence Bond Theory

- a covalent bond is formed when 2 atomic orbitals from 2 different atoms overlap and an e^- pair is shared in the overlapped space
- a stable molecule forms when the PE of the 2 atoms is at a minimum
- different atoms have different distances apart to achieve minimum energy

C ⇒ always 4 bonds



predicts 2 bonds

excited state