Dipole Moments

Earlier: H—F: (δ+ on H, δ− on F)

F (E_n = 4.0)
H (E_n = 2.1)

$\Delta E_n = 1.9$

POLAR COVALENT BOND.

dipole moment vector

Dipole Moment (vector)

$$\mu = Q \times r$$

Q = charge separated

r = distance

Q+ ⟷ Q−

Units of charge: Coulomb (C)

Units of distance: meters (m)

Units of μ = C·m

+1C • ⟶ (1m) • −1C

$\mu = 1$ C·m

Table 10.3 Dipole Moments of Some Polar Molecules

Molecule	Geometry	Dipole Moment (D)
HF	Linear	1.92
HCl	Linear	1.08
HBr	Linear	0.78
HI	Linear	0.38
H_2O	Bent	1.87
H_2S	Bent	1.10
NH_3	Trigonal pyramidal	1.46
SO_2	Bent	1.60

Define the Debye (D)

$1\ D = 3.336 \times 10^{-30}\ C \cdot m$

Most bonds have μ on order of 0-10 D.

Since μ is a vector, we have to worry about its direction when we add μ's.

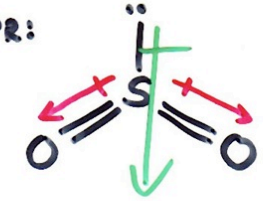
POLAR MOLECULES: $\mu_{TOTAL} \neq 0$

NON-POLAR MOLECULES: $\mu_{TOTAL} = 0$

ex: SO_2

Lewis:

VSEPR:



BOND DIPOLE MOMENTS

BENT

POLAR

$\mu_{TOT} \neq 0$

ex: CO_2

VSEPR:

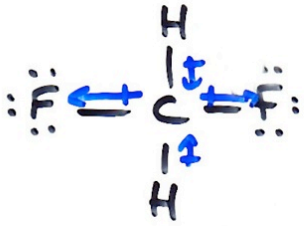
LINEAR

NON-POLAR

$\mu_{TOT} = 0$

ex: is CH_2F_2 polar?

Lewis:



C: $E_n = 2.5$
H: $E_n = 2.1$
F: $E_n = 4.0$

C–F: $\Delta E_n = 1.5$ → C–H: $\Delta E_n = 0.4$ →

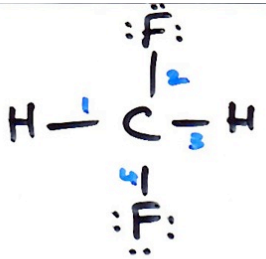
$\mu_{TOT} = 0$

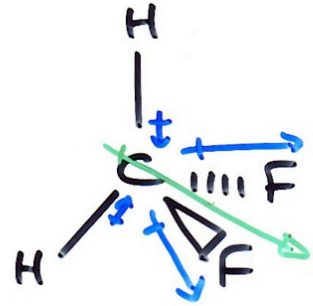
NON-POLAR.

WRONG

VSEPR! 3-D geometry!

Lewis





tetrahedral

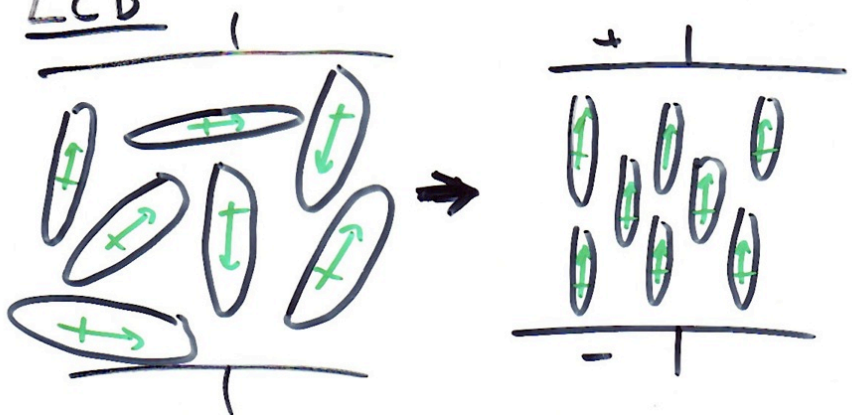
$\approx 109.5°$

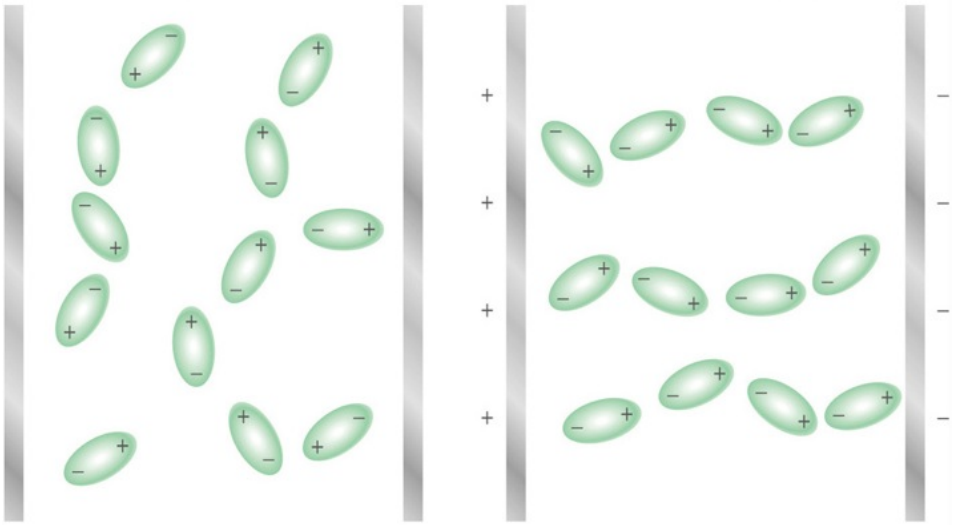
overall μ

μ_{TOT}

⇒ POLAR

LCD

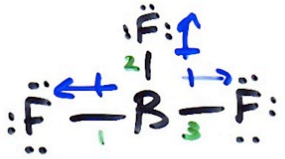




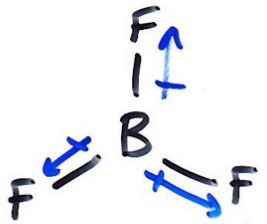
(a) (b)

ex: BF_3 POLAR ?

Lewis:



VSEPR:



$\mu_{overall} = 0$

$\Rightarrow$ NON-POLAR.

Valence-Bond Theory (VB)

Linus Pauling, 1930s

Covalent Bond = Overlap of 2 orbitals w/ 2 e⁻s.

ex: H_2

Lewis

$H\cdot \quad \cdot H \Longrightarrow H:H$ or H–H

$1s^1 \quad 1s^1 \qquad 1s^2 \quad 1s^2$

[He]

VB