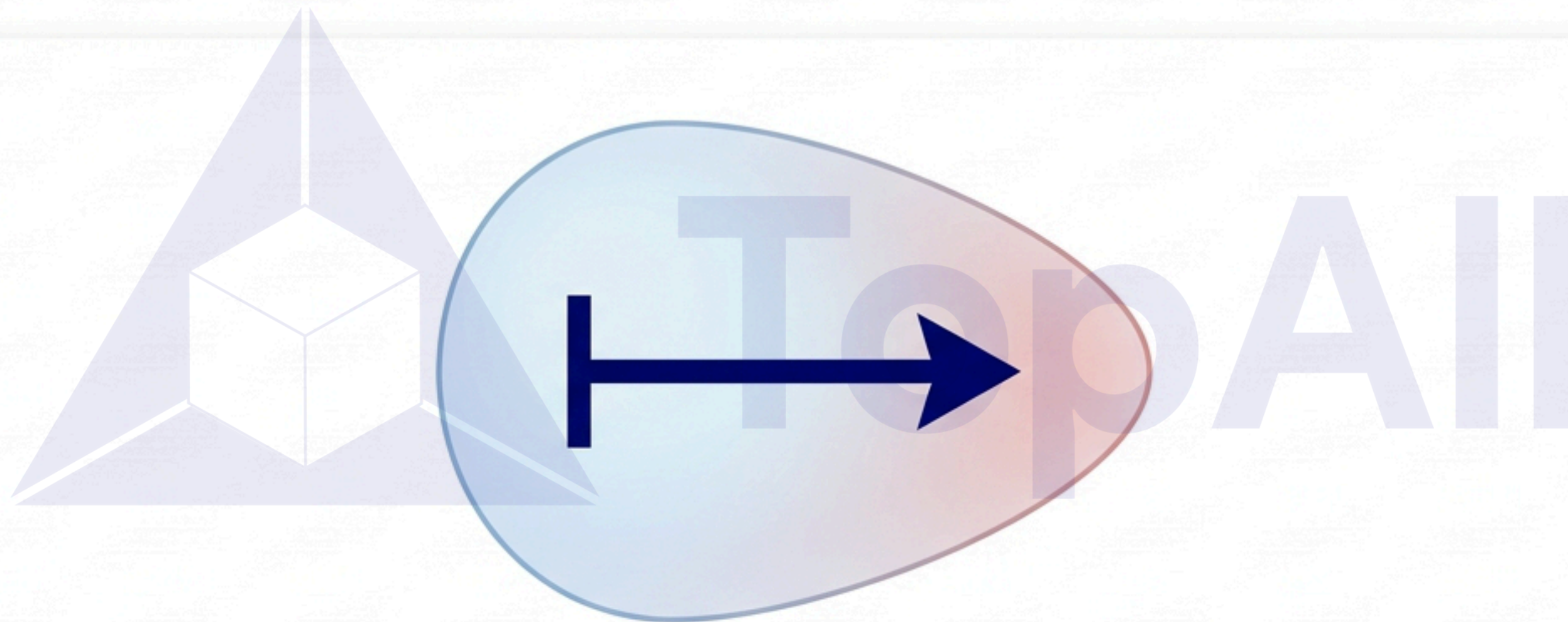


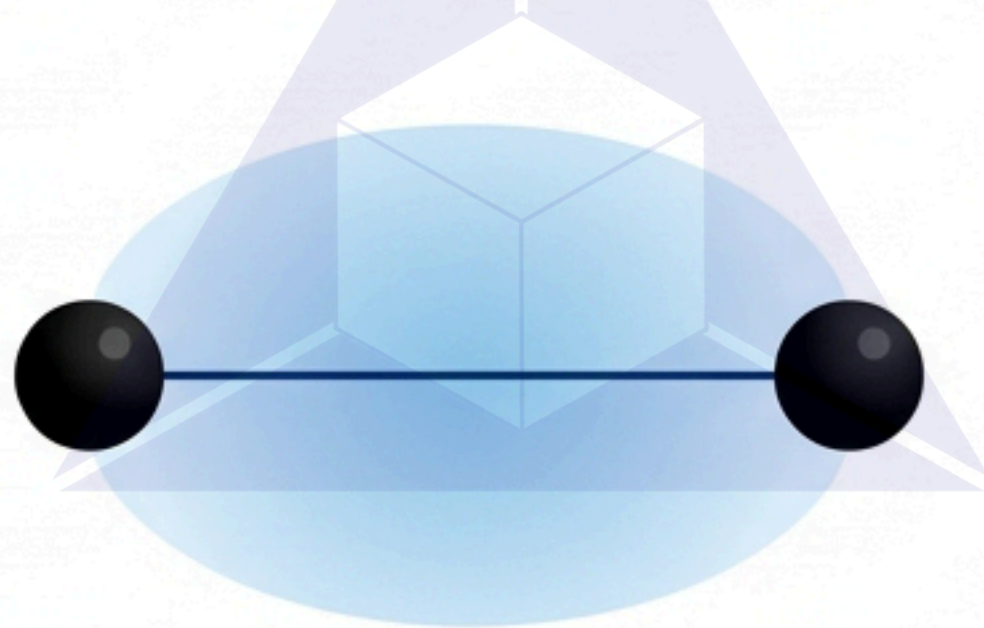
# Polarity of Bonds



# Nonpolar vs. Polar Covalent Bonds

## Ideal Sharing

[Homodiatomic] e.g.,  $\text{H}_2$ ,  $\text{O}_2$ ,  $\text{Cl}_2$



The shared electron pair is equally attracted by both atoms.

## Unequal Sharing

[Heterodiatomic] e.g.,  $\text{HF}$

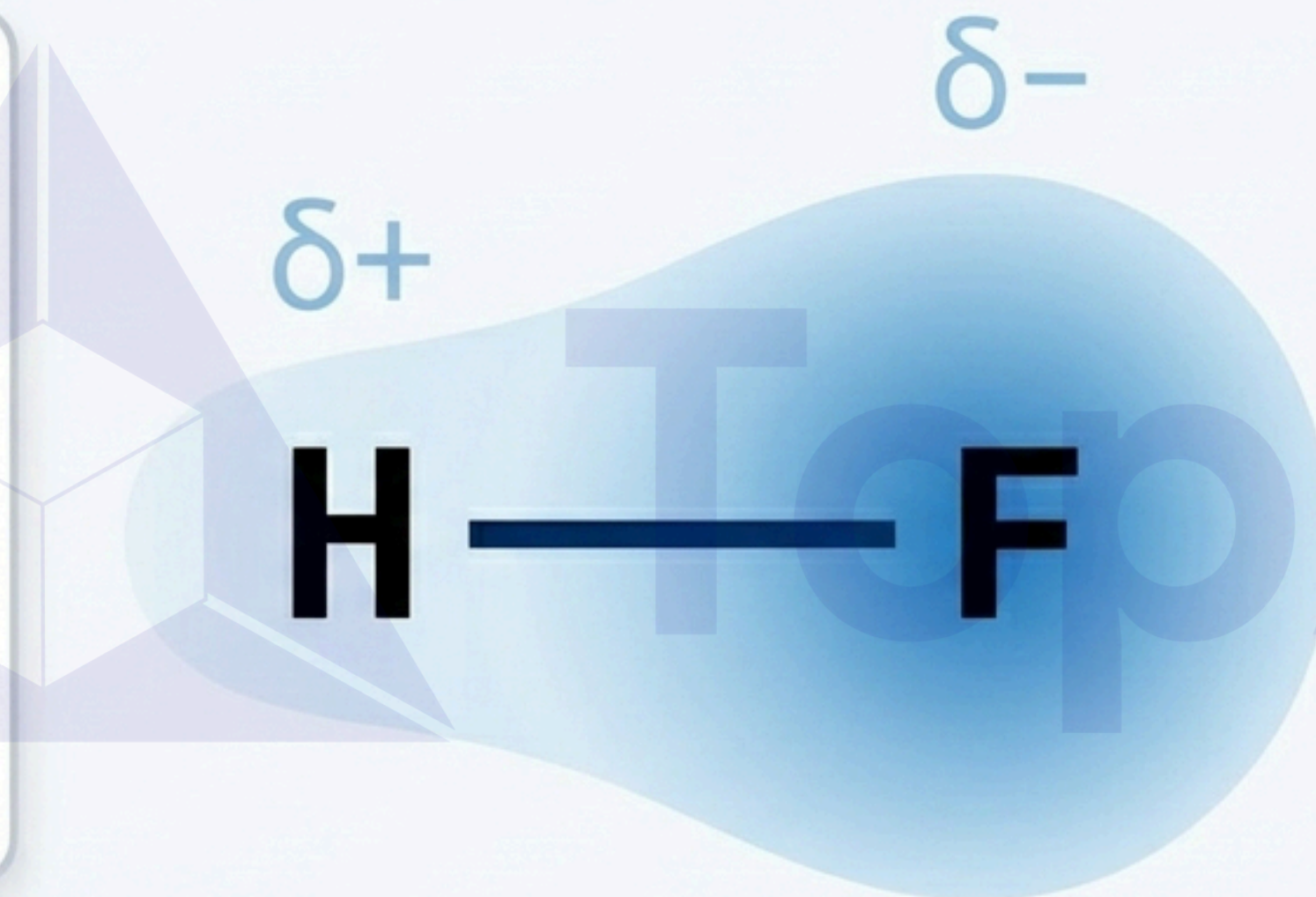


The shared electron pair gets displaced toward the more electronegative atom.

# The Hydrogen Fluoride (HF) Shift

## The Cause

Fluorine's highly electronegative nature acts as an electron magnet, pulling the shared pair away from Hydrogen.



## The Effect

Creates distinct partial charges.  $\delta+$  (delta positive) on Hydrogen and  $\delta-$  (delta negative) on Fluorine.

Polarization turns a neutral bond into a miniature magnet with two distinct poles.

# Calculating the Dipole Moment

$$\mu = Q \times r$$

$\mu$

Dipole moment  
(Vector quantity)

$Q$

Magnitude of the  
separated charge

$r$

Distance of separation  
between charge centers

Measured in Debye units (D)

$1 \text{ D} = 3.33564 \times 10^{-30} \text{ C m (Coulomb meters)}$



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## Key Highlights

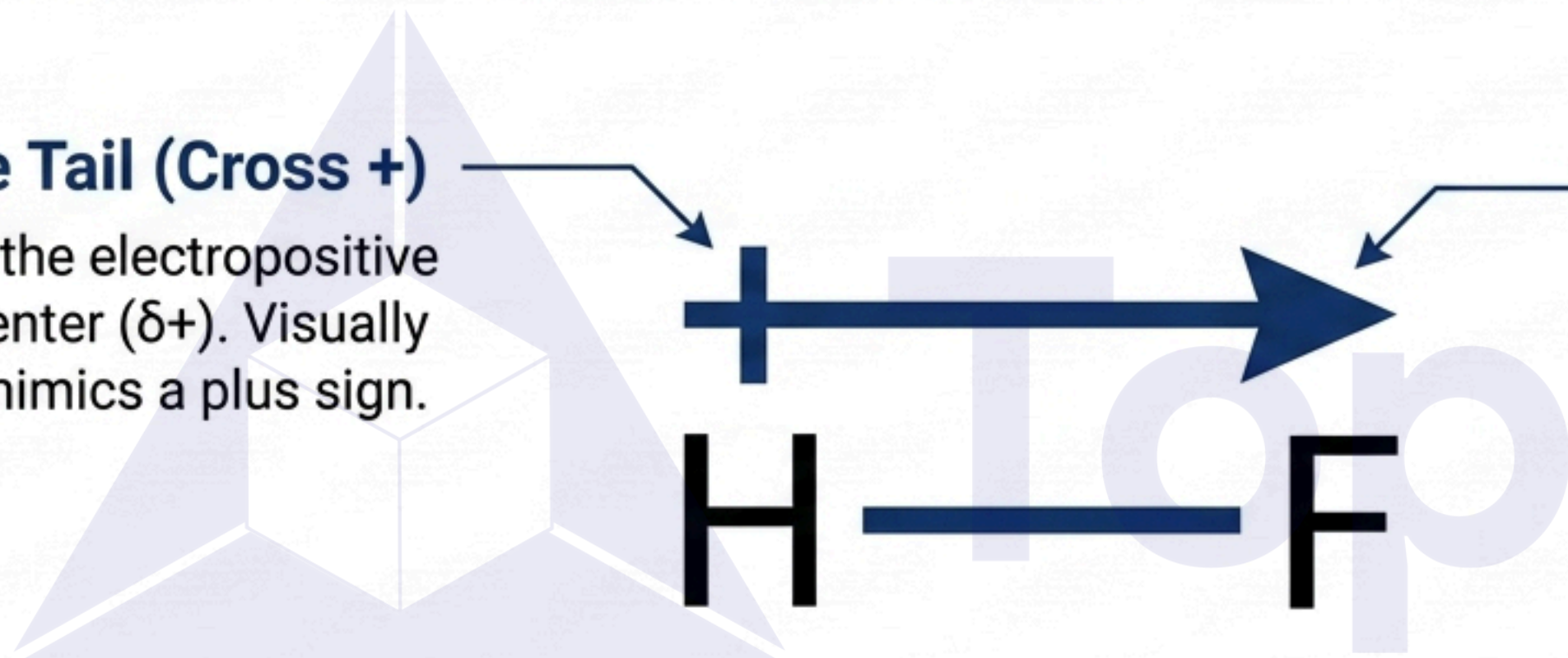
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# Anatomy of a Dipole Vector

## The Tail (Cross +)

Placed on the electropositive center ( $\delta+$ ). Visually mimics a plus sign.



## The Head (Arrow)

Points toward the electronegative center ( $\delta-$ ). Symbolizes the direction of electron density shift.

**Exam Smart Note:** In chemistry, the crossed arrow points toward the negative center. This is exactly opposite to the conventional physics dipole vector!

# When Geometry Cancels Polarity

Linear

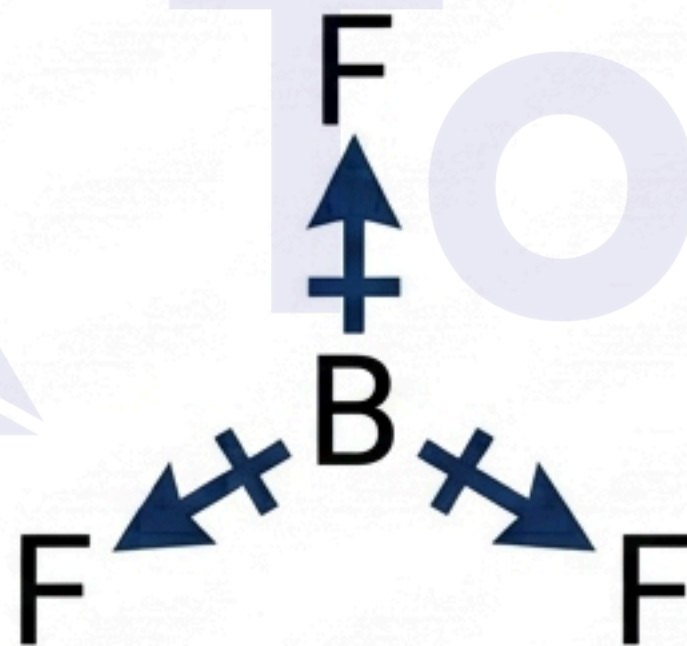
BeF<sub>2</sub>



Vectors cancel.  
Net  $\mu = 0$

Trigonal Planar

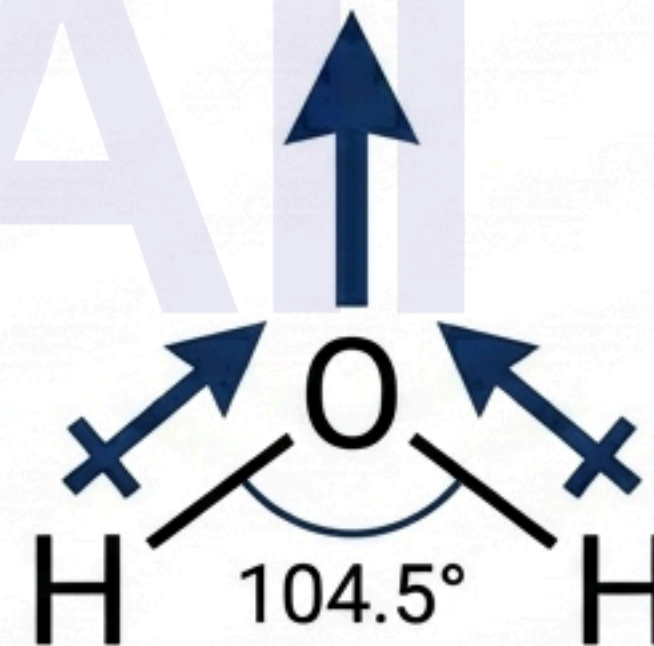
BF<sub>3</sub>



Vectors cancel.  
Net  $\mu = 0$

Bent

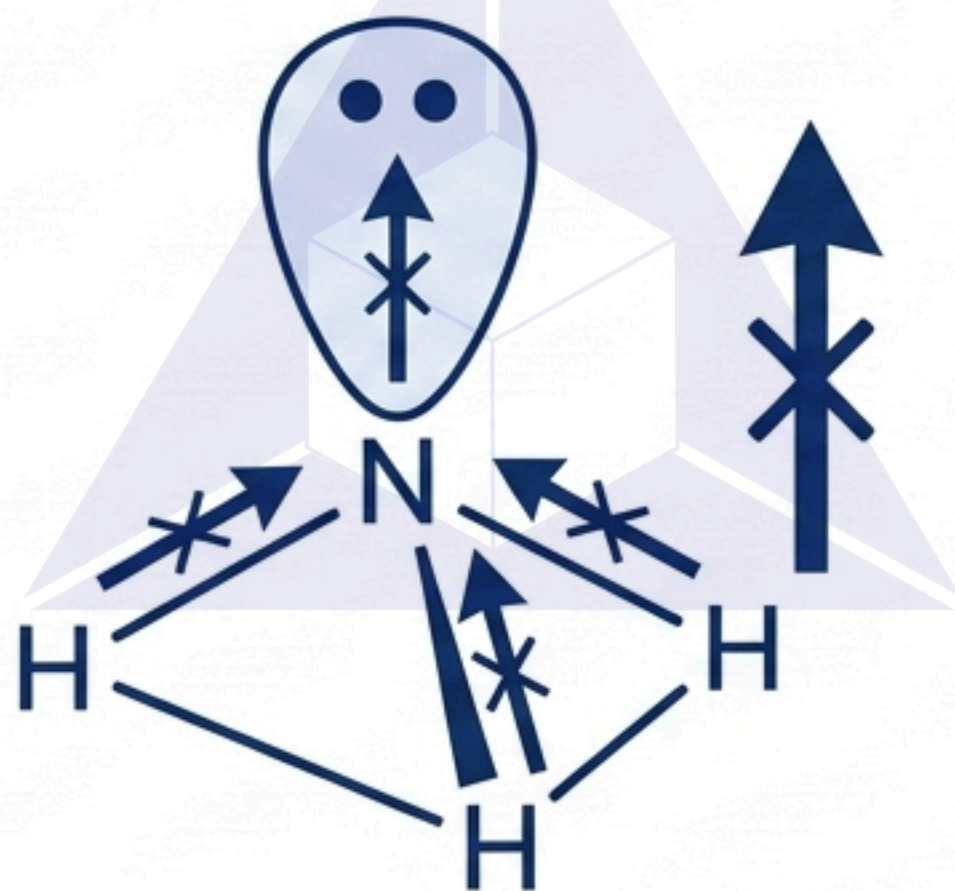
H<sub>2</sub>O



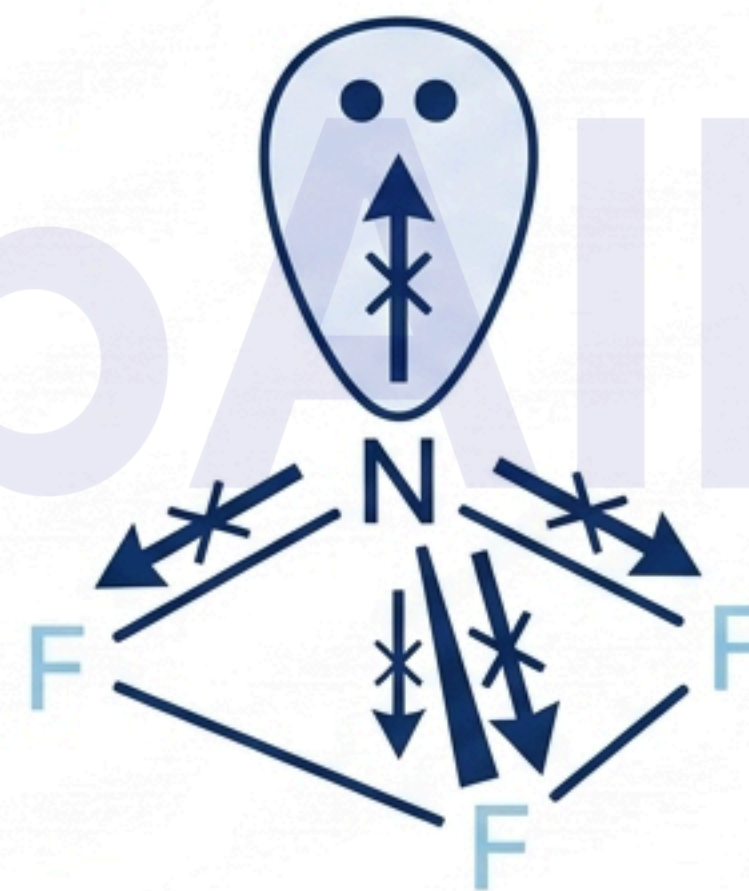
Vectors compound.  
Net  $\mu = 1.85 \text{ D}$

# The $\text{NH}_3$ vs. $\text{NF}_3$ Paradox

Both are pyramidal with a lone pair. F is *more* electronegative than N. Yet,  $\text{NH}_3$  has a *higher* dipole moment ( $\mu = 1.47 \text{ D}$ ) than  $\text{NF}_3$  ( $\mu = 0.23 \text{ D}$ ). Why?

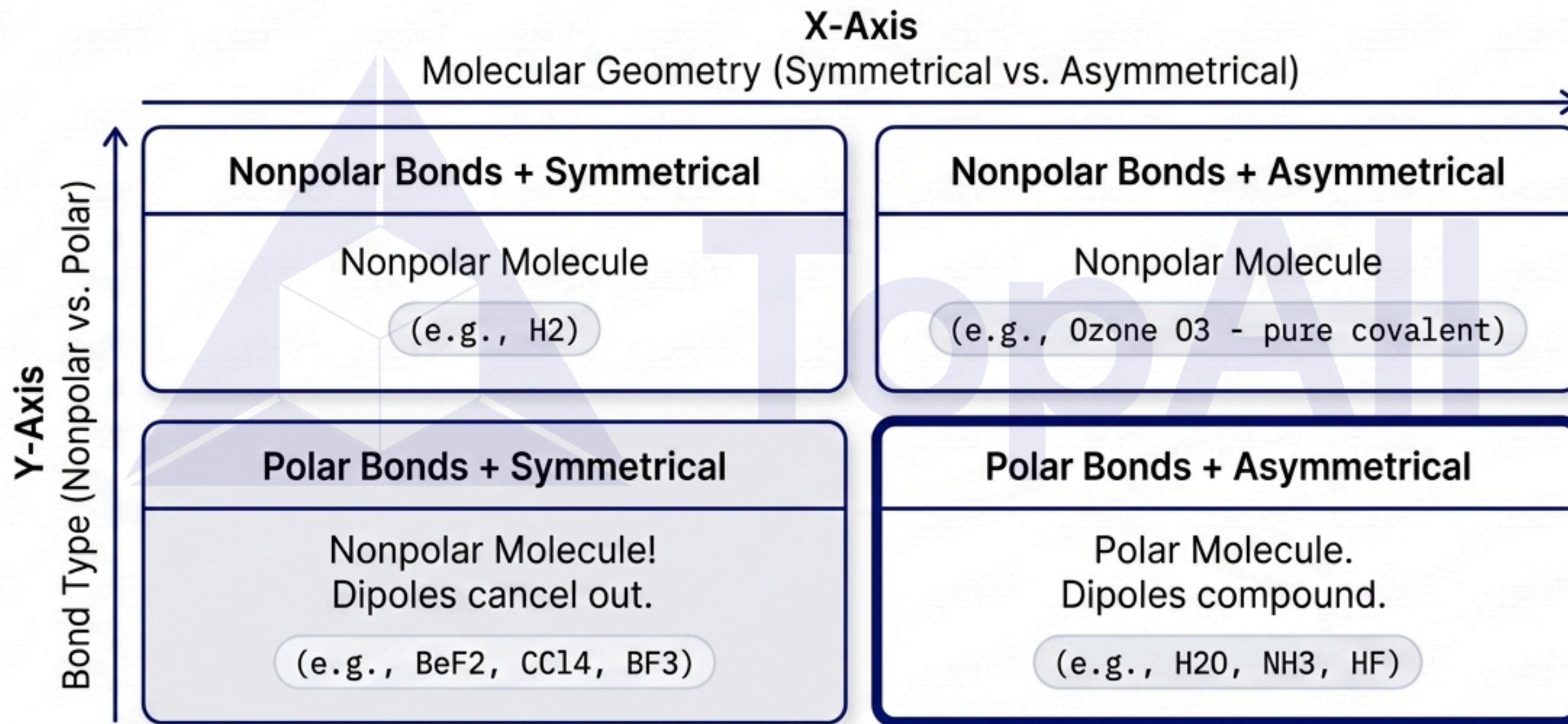
 $\text{NH}_3$  $\text{NH}_3$  $\mu = 1.47 \text{ D}$ 

Bond dipoles and orbital dipole are in the same direction. They compound.

 $\text{NF}_3$  $\text{NF}_3$  $\mu = 0.23 \text{ D}$ 

Bond dipoles and orbital dipole are in opposite directions. They subtract.

# The Polarity Decision Matrix



**Molecular Polarity = Bond Electronegativity + 3D Asymmetry**



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