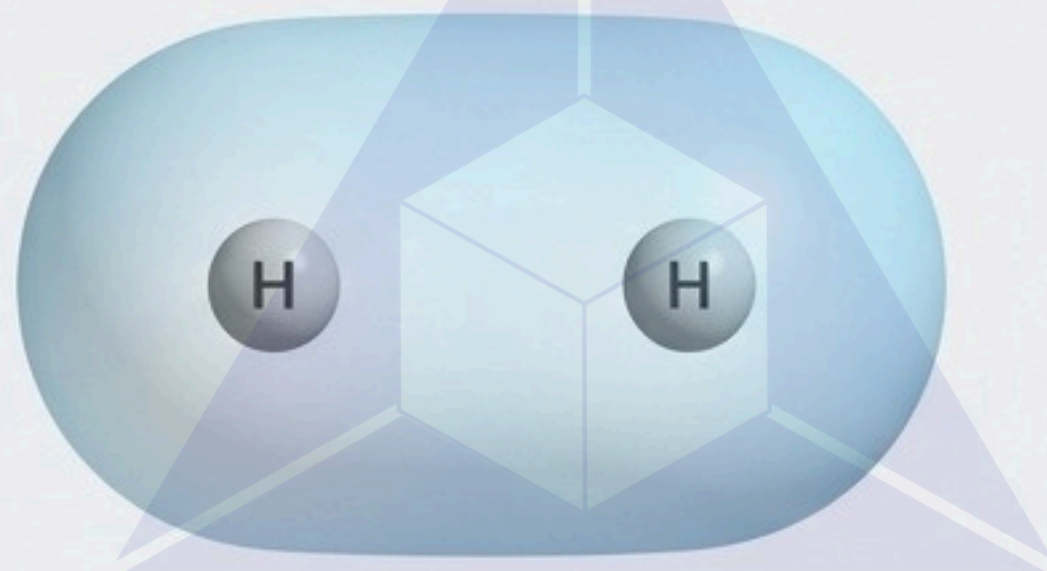


# Not All Sharing is Equal

The difference between Nonpolar and Polar covalent bonds.

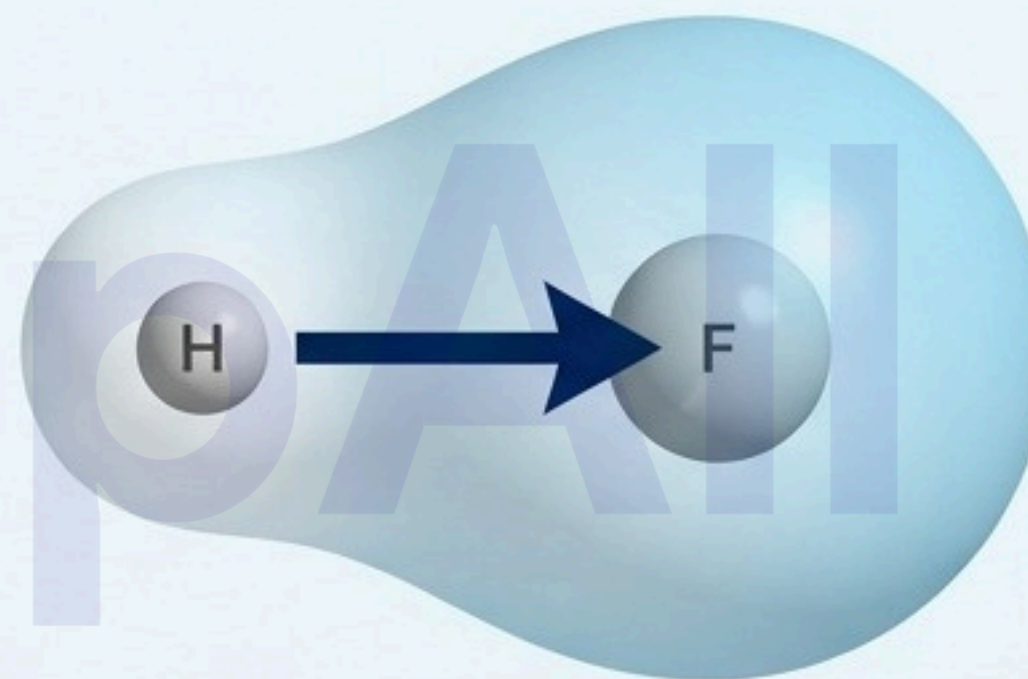
## Nonpolar Covalent Bond



### Symmetrical Sharing

When two identical atoms (like  $\text{H}_2$ ,  $\text{O}_2$ , or  $\text{F}_2$ ) bond, the shared electron pair sits exactly in the middle. The pull is equal.

## Polar Covalent Bond



### Asymmetrical Sharing

When atoms differ in their electronegativity, the shared electrons are pulled off-center, creating a polar shift in the molecule's charge.

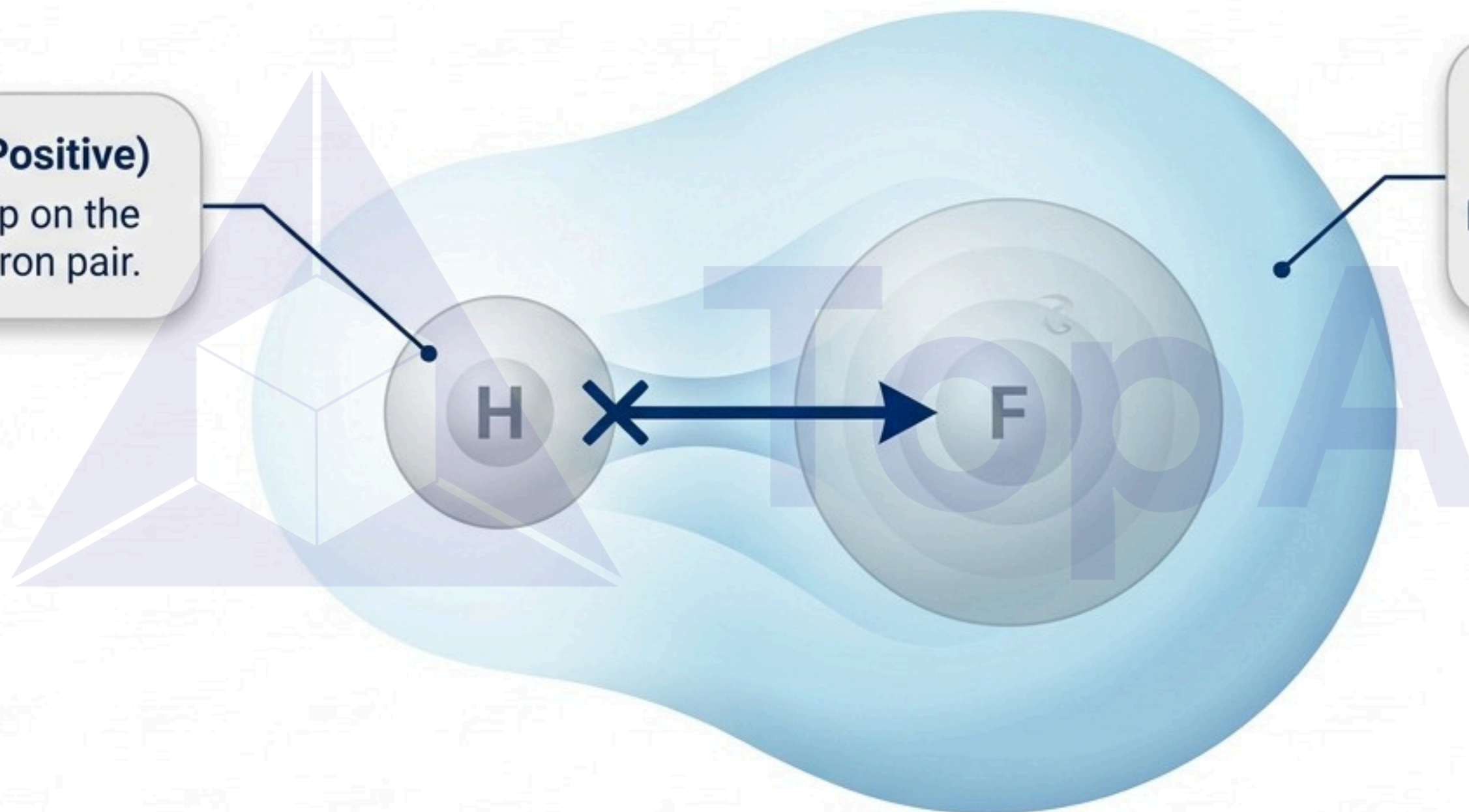
# The Anatomy of a Polar Shift

Electronegativity drives a localized tug-of-war for electron density.



**$\delta+$  (Partial Positive)**

Loses its grip on the shared electron pair.



**$\delta-$  (Partial Negative)**

Higher electronegativity pulls the electron density heavily toward itself.

This localized displacement of electrons creates a permanent electrical imbalance across the bond.

# Quantifying the Shift: The Dipole Moment

How scientists measure the exact strength of molecular polarity.

$$\mu = Q \times r$$

$\mu$  (Dipole Moment):

The resulting strength of the polarity.

$Q$  (Charge):

The magnitude of the separated partial charge.

$r$  (Distance):

The distance separating the centers of positive and negative charge.

**Unit of Measurement:**  
Debye units (D)

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

*Named after Dutch chemist Peter Debye, honoring his Nobel-winning work on X-ray diffraction and dipole moments.*



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$$E=mc^2$$



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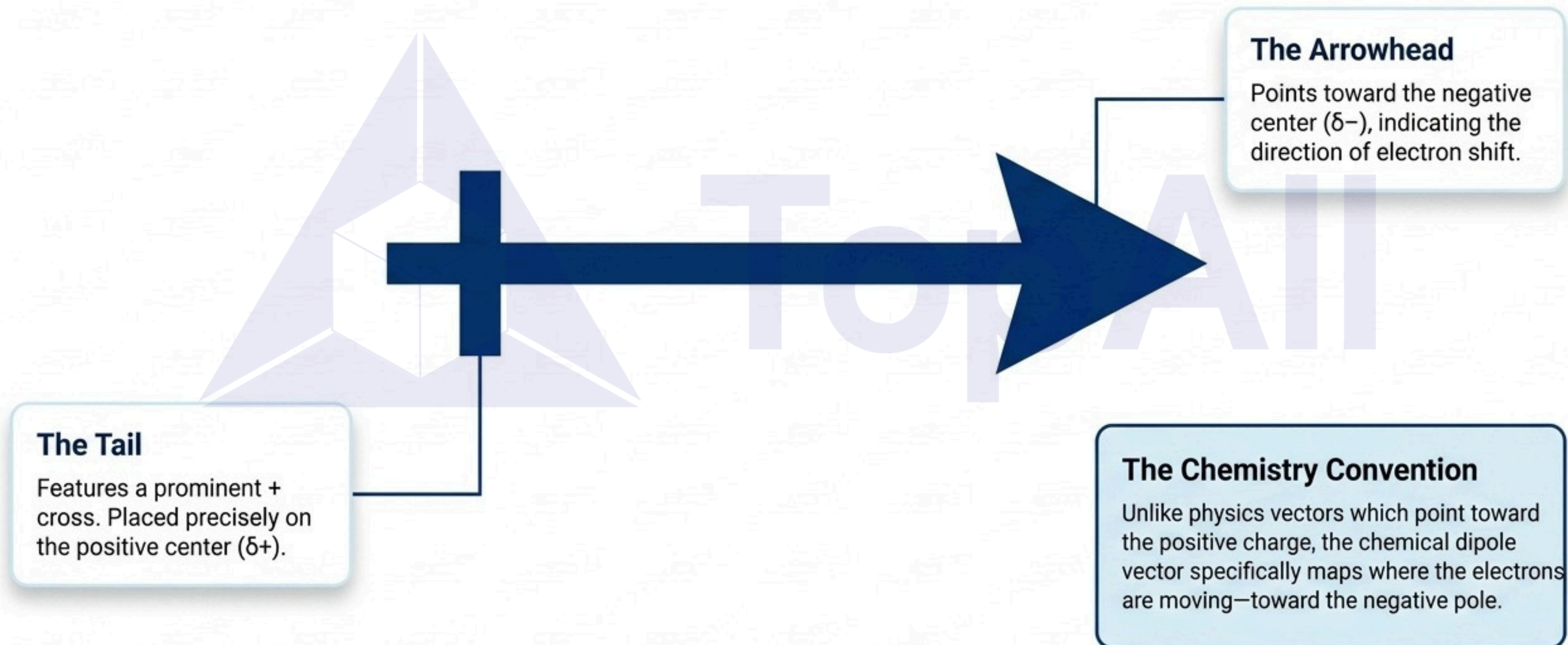
### Key Highlights

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# Dipoles as Vectors

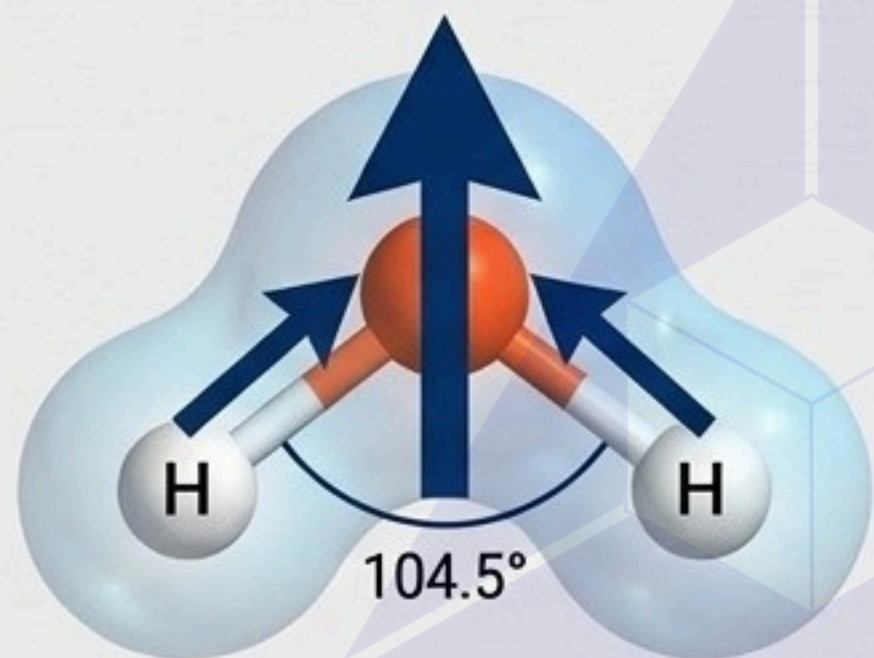
Mapping the magnitude and direction of the electron shift.



# Geometry Dictates Destiny

Individual polar bonds do not guarantee a polar molecule. Symmetry can cancel them out.

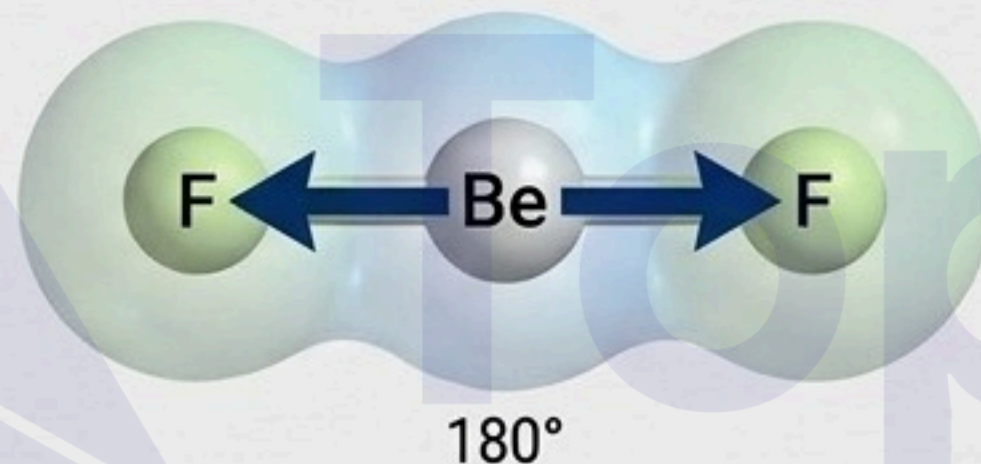
**H<sub>2</sub>O (Net Dipole)**



$$\mu = 1.85 \text{ D } (6.17 \times 10^{-30} \text{ C m})$$

Asymmetrical bent shape.  
Vectors add together.  
Highly Polar.

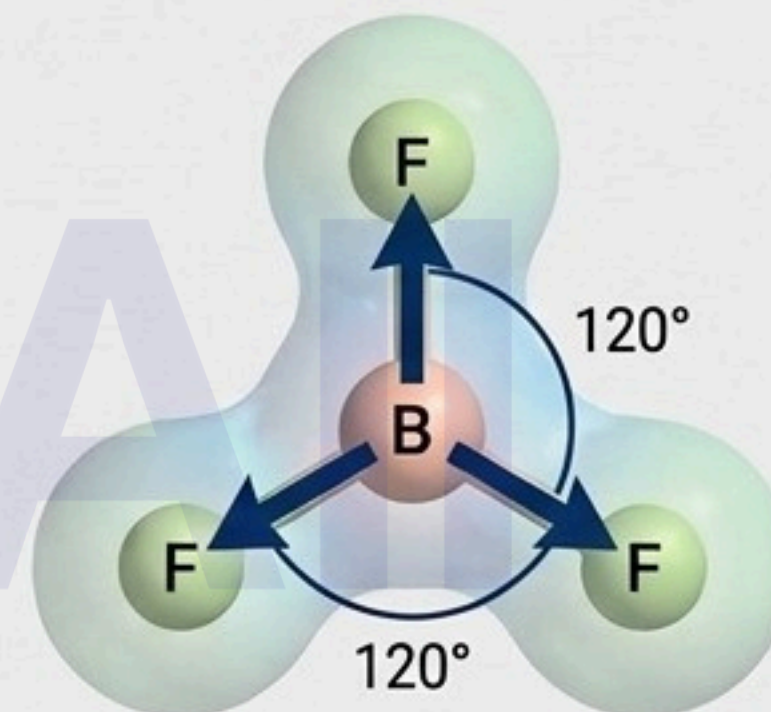
**BeF<sub>2</sub> (Net Zero)**



$$\mu = 0 \text{ D}$$

Perfect linear symmetry. Equal  
vectors perfectly cancel each other.  
Nonpolar.

**BF<sub>3</sub> (Net Zero)**



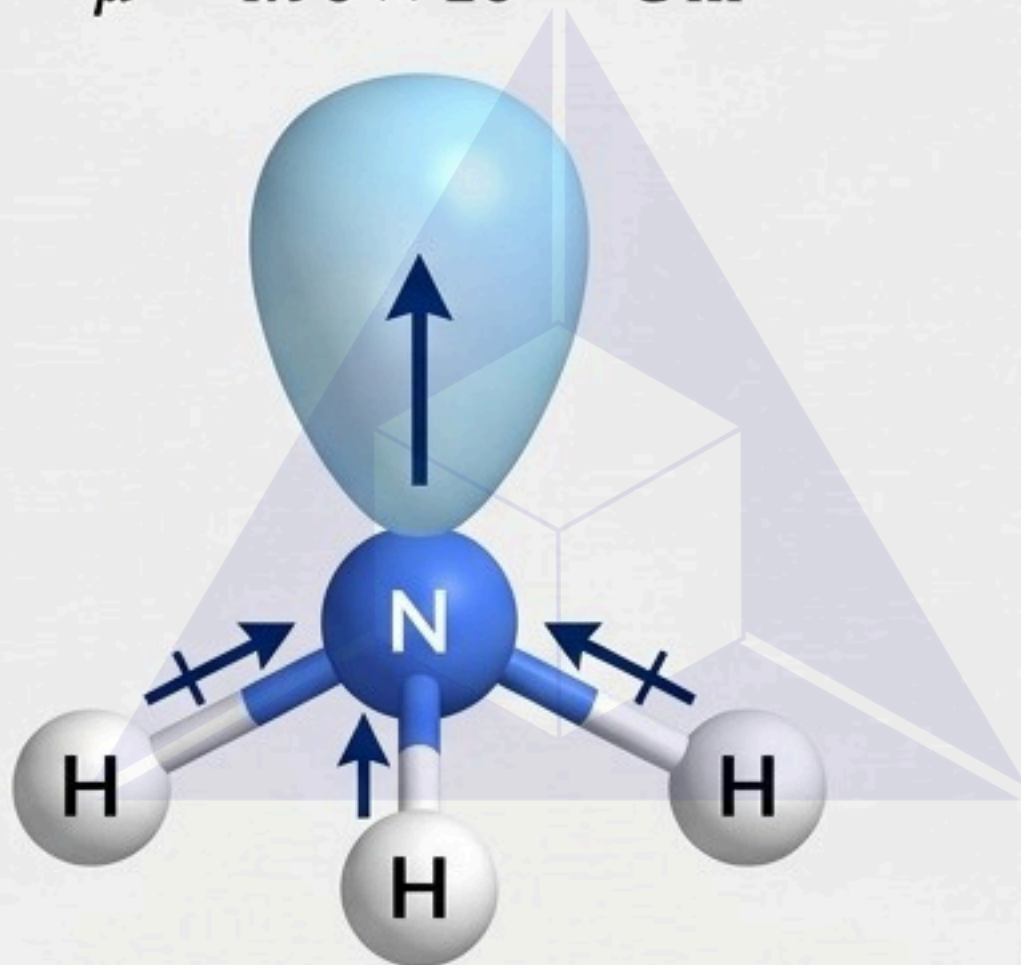
$$\mu = 0 \text{ D}$$

Radial symmetry. The resultant of  
any two vectors is equal and  
opposite to the third. Nonpolar.

# The Lone Pair Effect: $\text{NH}_3$ vs $\text{NF}_3$

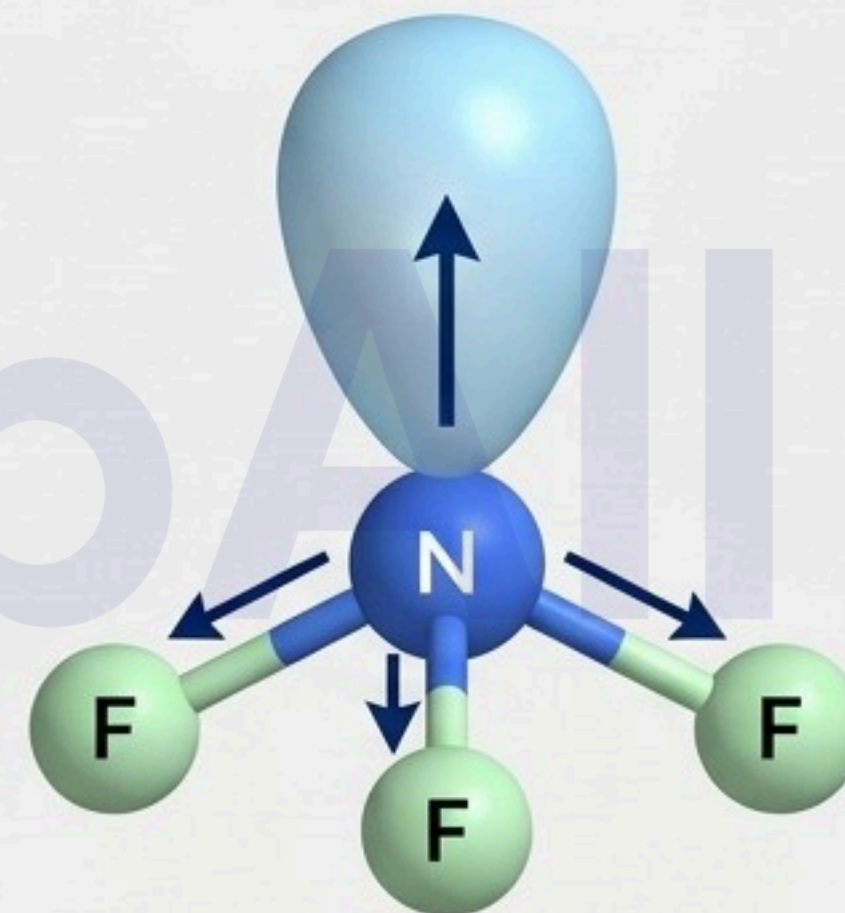
Why do two identically shaped pyramidal molecules have vastly different polarities?

Ammonia ( $\text{NH}_3$ )  
 $\mu = 4.90 \times 10^{-30} \text{ C m}$



**Vector Addition.** The N-H bond dipoles and the lone pair orbital dipole push in the same direction, resulting in a massive net dipole moment.

Nitrogen Trifluoride ( $\text{NF}_3$ )  
 $\mu = 0.80 \times 10^{-30} \text{ C m}$



**Vector Subtraction.** Fluorine is more electronegative than Nitrogen. The N-F bonds pull down, heavily canceling out the upward lone pair dipole, resulting in a tiny net dipole moment.

# Revision Frame: Bond Polarity

The core mechanics of molecular dipole moments.

## 1. Core Concept

Unequal Sharing. Polarity arises when a difference in electronegativity creates an asymmetrical electron cloud, resulting in partial charges ( $\delta+$  and  $\delta-$ ).

## 2. The Math

$$\mu = Q \times r$$

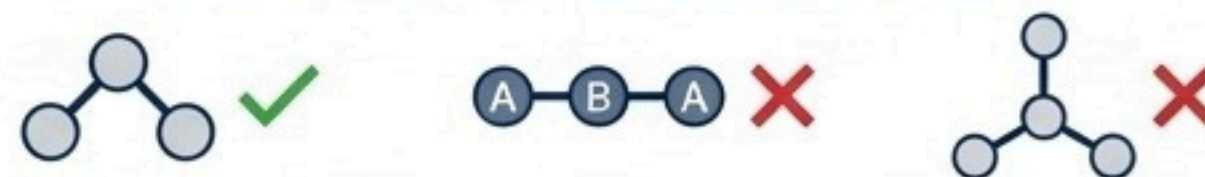
Dipole Moment ( $\mu$ ) equals Charge magnitude ( $Q$ ) multiplied by separation distance ( $r$ ). Measured in Debye (D).

## 3. The Vector Rule



Dipoles are directional vectors. The cross sits on the positive center; the arrow points toward the negative electron density.

## 4. The Geometry Rule



Polar bonds only create a polar molecule if the 3D geometry is asymmetrical. Symmetrical shapes (like  $\text{BeF}_2$  or  $\text{BF}_3$ ) perfectly cancel out individual bond vectors resulting in a Net Zero dipole ( $\mu = 0$ ).



# You've Learned the Theory.

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