Molecular Polarity

Molecule

Polar Molecules:

Contain δ+ and δ- ends:

Charged ends or "poles" are the result of polar bonds within the molecule and...

...the molecule's geometry.

Dipole Moment Vector

...<u>molecule</u> is said to be a "<u>polar</u>"

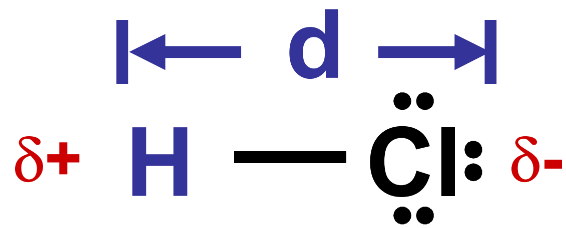
δ+ H — Cl: δ-

$E_n = 2.1$ **$E_n = 3.0$**

$\Delta E_n = 0.9 \Rightarrow$ Polar Covalent Bond

Dipole Moment: μ

$$\mu = q \times d$$



q: **Magnitude (δ: coulombs) of the charge at either end of the molecule**

...larger when more electronegative elements are involved.

d: **Distance between the charged ends of the molecule.**

Dipole Moments: $\mu = q \times d$

...differences in d

Bond Length (d)		μ
92 pm	H-F	small
127 pm	H-Cl	Larger
141 pm	H-Br	Larger
161 pm	H-I	Largest

Dipole Moments: $\mu = q \times d$

...differences in E_n

Bond Length (d)	E_n		μ
92 pm	4.0 - 2.1	$\delta+$ H–F: $\delta-$	**Largest**
127 pm	3.0 - 2.1	$\delta+$ H–Cl: $\delta-$	**Larger**
141 pm	2.8 - 2.1	$\delta+$ H–Br: $\delta-$	**Larger**
161 pm	2.5 - 2.1	$\delta+$ H–I: $\delta-$	**small**

Dipole Moments: $\mu = q \times d$

...two conflicting predictions...

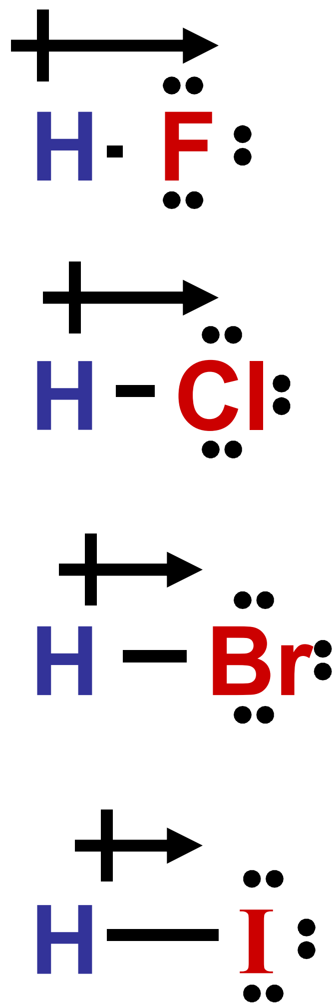
...nature tells us what's correct

Bond Length (d)	**E_n**		**μ**
92 pm	**4.0**	δ+ H - F: δ-	μ = **1.92 D**
127 pm	**3.0**	δ+ H – Cl: δ-	μ = **1.08 D**
141 pm	**2.8**	δ+ H – Br: δ-	μ = **0.78 D**
161 pm	**2.5**	δ+ H — I: δ-	μ = **0.38 D**

Dipole Moments: $\mu = \mathbf{q} \times \mathbf{d}$

H-X Series:

Electronegativity differences are more important than bond length differences.



μ = 1.92 D

μ = 1.08 D

μ = 0.78 D

μ = 0.38 D

The length of the dipole moment vector suggests the size of the molecule's dipole moment, μ.