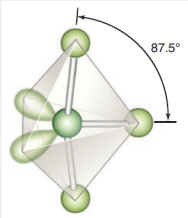
Lone pairs affect polarity

Lone pairs contribute to localize negative charge on an atom.

The effect of lone pairs alone is why the T-shaped molecule I_4 is polar, even though each bond is none polar.



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[TP] Lone pairs can either enhance or diminish the contribution of bond dipoles to the polarity of a molecule. Which molecule is more polar?

33% 1. NH_3
33% 2. NF_3
33% 3. Each of the above is nonpolar

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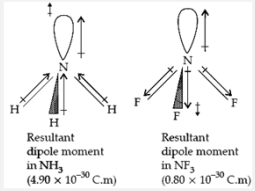
Lone pairs affect polarity

Lone pairs can either enhance or diminish the contribution of bond dipoles to the polarity of a molecule. An example of a consequence of this that NH_3 is more polar (has a larger dipole moment; see section 7-11, page 226) than NF_3.

Note that the arrows in the figure show the shift in the electron clouds due to the electronegativity differences.

The lone pairs in NH_3 augment the cloud shifts in the N-H bonds, and so enhance the polarity.

The lone pairs in NF_3 oppose the cloud shifts in the N-F bonds, and so diminish the polarity.



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Lone pairs affect polarity

In the series of T-shaped molecules ClF_3, BrF_3, and IF_3 (problem 7.36), the lone pairs oppose the shifts in bonding electron clouds and so diminish the polarity of each molecule.

However, since the effect of the lone pairs is the same in each molecule, their relative polarities mirror their relative bond polarities.

Therefore, the order of molecular polarities is

$ClF_3 < BrF_3 < IF_3$

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