Is That Molecule Polar? How Do We Tell?

1. Draw the Lewis structure
   A. How?
      i) Find the total number of valence electrons in the molecule.
      ii) Write the atoms in the correct orientation.
      iii) Place 2 electrons (1 bond) between each atom.
      iv) Distribute the other electrons as needed.
   B. Hints and Patterns (but there are exceptions!)
      i) Hydrogen is only bonded to one other atom (only 2 electrons needed)
      ii) Other atoms will generally (not always) have an octet (8) of electrons in a stable molecule.
      iii) A good first guess: carbon will generally have 4 bonds and 0 lone pairs (Group 4)
      nitrogen will generally have 3 bonds and 1 lone pair (Group 5)
      oxygen will generally have 2 bonds and 2 lone pairs (Group 6)
      fluorine will generally have 1 bond and 3 lone pairs (Group 7)
      iv) Rows 3 and below can exceed the octet rule if necessary.
      v) If there are “extra” electrons, they go on the central atom (if it can exceed the octet rule).

2. Determine the geometry of the molecule.
   A. Count the number of “directions”, “bonds + lone pairs”, or “effective pair”.
   B. Multiple bonds count as one “direction” or “effective pair”.

3. Determine the shape of the molecule.
   A. Only consider the atoms for the shape.
   B. The lone pairs still exist, they just do not determine the name of the shape.

4. Is the molecule symmetrical?
   A. Would the electrons be evenly distributed throughout the molecule?
   B. Polarity is the uneven distribution of electrons.
      i) Symmetrical molecules are generally non-polar.
      ii) Asymmetrical molecules are generally polar.

<u>DON’T BE FOOLED BY A FORMULA!</u>

SO_2 = polar
CO_2 = non-polar

BF_3 = non-polar
NF_3 = polar

SCl_4 = polar
CCl_4 = non-polar