HAPPY CANADA DAY!

VSEPR (Valence Shell Electron Repulsion Model)

1. Electrons repel each other.
2. Electrons (bonding and lone pairs) adopt a geometry about the central atom that minimizes the electron-electrons repulsions (bonding and lone pair electrons want to be as far apart as possible).
3. There are 5 base geometries that most covalent molecules adopt: linear, trigonal planar, tetrahedral, trigonal bipyramid and octahedral
4. Bond angles determined by these geometries.
5. Shape is determined by the relative placement of only the bonded atoms around a central atom (can’t “see” lone pairs). There are many variations of different shapes for each of the 5 base geometries. Must memorize them.

Shape = geometry when central atom has no lone pairs.
Shape ≠ geometry when central atom has lone pairs.

6. A polar molecule has a partial negative charged end and a partial positive charged end separated by some distance (called a dipole moment). In general, a molecule needs polar covalent bonds (bonds with bond dipoles) in order to be a polar molecule. However, not all molecules with polar bonds are polar overall.

VSEPR Observations

- Multiple bonds count as only one set of electrons which are placed as far apart as possible from the bonding electrons and lone pair electrons.
- Lone pairs exert a stronger repulsive effect than bonding electrons; results in bond angles that are slightly smaller than predicted (can’t predict exact deviation).
- Only need one resonance structure to predict bond angles, shape and polarity.

* Courtesy of Tom Hummel’s Handouts for Chemistry 102.
Predicting Polarity of Molecules*

General Ideas:

1. From the geometry of the molecule, we can predict whether a molecule is polar or nonpolar. Polar compounds dissolve in polar solvents (H₂O). Nonpolar compounds dissolve in nonpolar solvents (C₆H₁₄).
2. Molecules with polar bonds can be polar compounds. However, not all compounds with polar bonds are polar. Polarity of molecules can only be predicted on an individual basis.

Rules:

1. Draw the Lewis Structure with the proper geometry indicated.
2. Draw in all the individual bond dipoles using arrows pointing to the more electronegative atom. Remember: B-H, C-H and P-H bonds are assumed nonpolar. All other covalent bonds between two different nonmetals are assumed polar, i.e. have a bond dipole.
3. Sum all the individual bond dipoles cancel, then the molecule is nonpolar. If they do not cancel, then the molecule is polar.

General Observations:

1. If a molecule has the form ABₓ (A = central atom, B = terminal atom) and then central atom A has no lone pairs, then the individual dipoles will always cancel giving a nonpolar molecule, e.g. CO₂, SO₃, CF₄, PCl₅, SF₆. These types of molecules are symmetric indicating that the bond dipoles are arranged so that they all cancel each other out, giving a nonpolar molecule.
2. If the central atom has lone pairs around it, then overall the molecule will usually be polar. Molecules with lone pairs around the central atom are generally nonsymmetric, meaning that the individual bond dipoles do not usually cancel each other out giving a polar molecule. Exceptions: KrF₂ (linear shape) and XeF₄ (square planar shape) type compounds are nonpolar even though the central atom has lone pairs. For KrF₂ and XeF₄ type compounds, the individual bond dipoles do cancel each other out when they are summed together (in these two cases, the bond dipoles are symmetrically arranged).
3. Another way to disrupt the symmetry of molecules is to have two or more different terminal atoms, e.g. CO₂ is nonpolar but SCO is polar.

* Courtesy of Tom Hummel’s Handouts for Chemistry 102.