

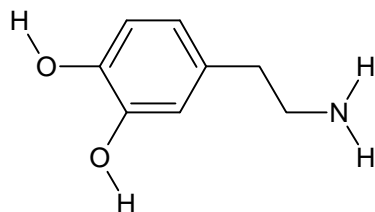


## Electronegativity (EN) and Chemical Bonds

We can indicate polar bonds on chemical compounds one of two ways:

- i) dipole moment arrows  $\begin{array}{c} | \\ \longrightarrow \end{array}$   
The arrow head points toward the element with the greater electronegativity.
- ii) partial positive ( $\delta^+$ ) and partial negative ( $\delta^-$ )  
where the Greek letter delta ( $\delta$ ) indicates partial

3. Label the polar covalent bonds in dopamine using the  $\delta^+$  and  $\delta^-$  symbols.  
Remember to add any missing lone pairs.



The difference between ionic and covalent bonds can be explained in terms of the differences in electronegativity between the atoms involved as shown in Table 1 below.

Bond	Characteristics	Example	$\Delta$ EN
Ionic	Attraction of opposite charges	$\text{Na}^+ \text{Cl}^-$	Very large
Polar Covalent	Unequal sharing of an electron pair	$\delta^+ \text{C}-\text{O} \delta^-$	Large
Nonpolar Covalent	Equal sharing of an electron pair	$\text{C}-\text{H}$ or $\text{C}-\text{C}$	Very small or none

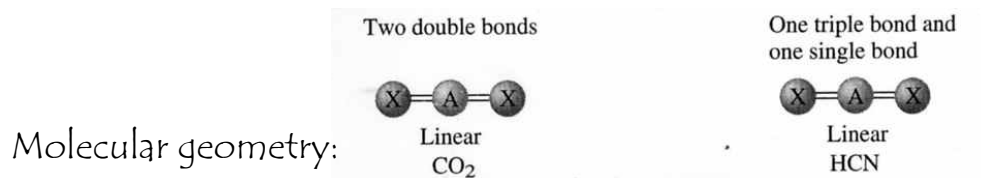
4. What types of bonds are present in sodium methoxide ( $\text{NaOCH}_3$ )?  
Draw the Lewis structure for  $\text{NaOCH}_3$  to explain your answer.

## Electron and Molecular Geometry: VSEPR = Valence Shell Electron Pair Repulsion

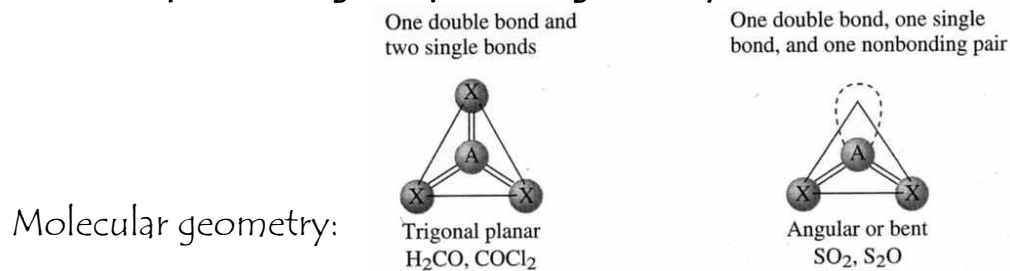
In addition to recognizing polar bonds within organic molecules, we need to visualize the 3-dimensional shape (molecular geometry). VSEPR theory predicts the molecular geometry by placing valence electrons (bonding and lone pairs) as far apart as possible around a central atom. For organic and biochemistry, we need to be familiar for the three basic arrangements of electrons around a central atom (electron geometry). Remember that when there are lone pairs around a central atom, a different molecular geometry results.

5. Draw the Lewis Structure and Perspective formula (wedges and dashes for 3-D) for each compound below.

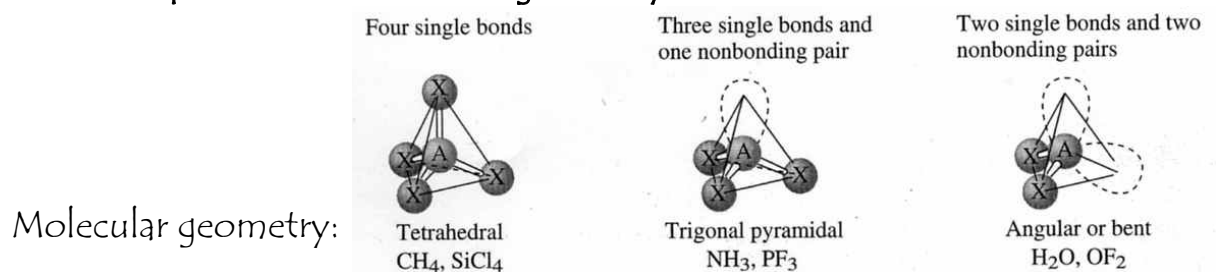
### 2 VSEPR "pairs" = Linear e<sup>-</sup> geometry



### 3 VSEPR "pairs" = Trigonal planar e<sup>-</sup> geometry



### 4 VSEPR "pairs" = Tetrahedral e<sup>-</sup> geometry



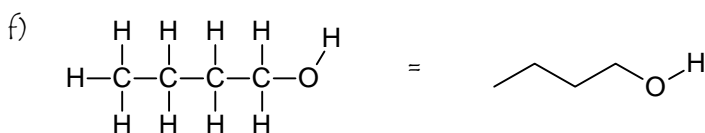
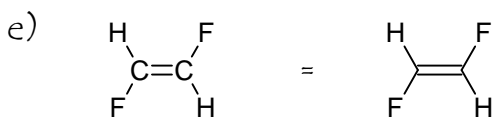
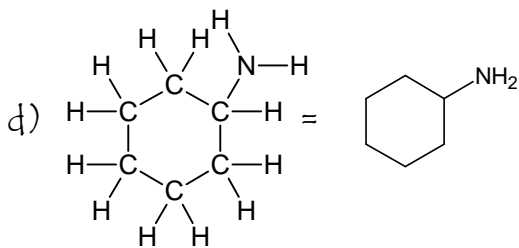
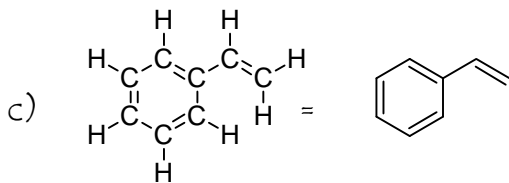
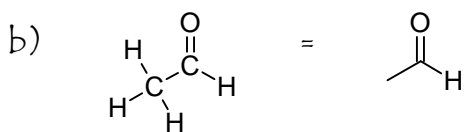
## Molecular Polarity

When we combine our understanding of bond polarity with molecular geometry, we can determine whether or not the entire molecule is polar or non-polar. In a polar molecule, one side has a partial positive charge and the other has a partial negative charge. Knowing whether or not a molecule is polar helps us to understand how it interacts with other compounds.

6. Which of the following molecules are polar? Remember to draw in the dipole moments and add any missing lone pairs. Where applicable the Lewis structure is also drawn as a simplified bond-line structure.

a)  $\text{CH}_2\text{Cl}_2$

(Hint: Draw the 3-D shape with wedges and dashes)



## Intermolecular Forces (IMFs) and Noncovalent Interactions

When neighboring molecules or ions or remote parts of the same molecule or ion interact with one another, they do so through intermolecular forces (IMFs) or noncovalent interactions, respectively. These interactions can be divided into two broad categories.

- A. Polar compounds with PERMANENT partial or full charges
  - a) H-bonding: compounds with N-H, O-H or F-H bonds
  - b) Dipole-dipole interactions: polar compounds
  - c) Ion-dipole interactions: an ion attracted to a polar compound
  - d) Salt bridge: another name for an ionic bond
  - e) Coordinate covalent: the lone pairs of a nonmetal atom associate with a metal cation
  
- B. Nonpolar compounds with TEMPORARY partial charges
  - a) London dispersion forces (also called hydrophobic interactions)

7. Indicate the type of IMF or non-covalent interaction represented in each figure.

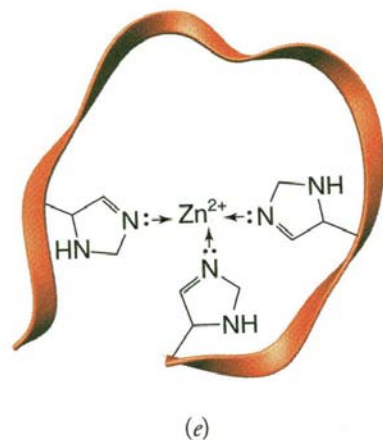
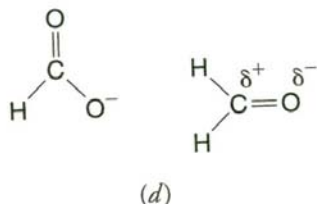
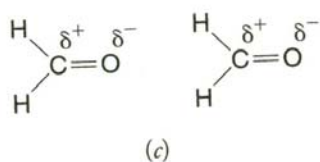
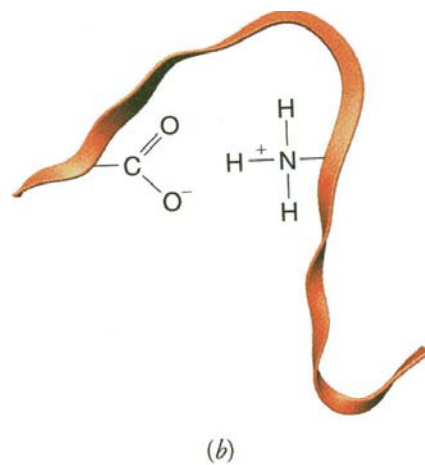
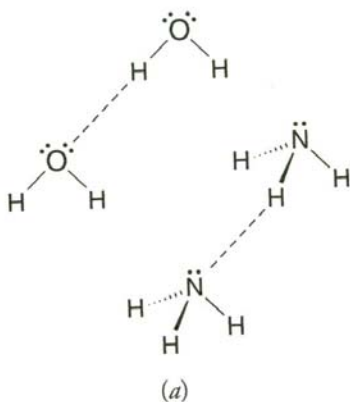
a) \_\_\_\_\_

b) \_\_\_\_\_

c) \_\_\_\_\_

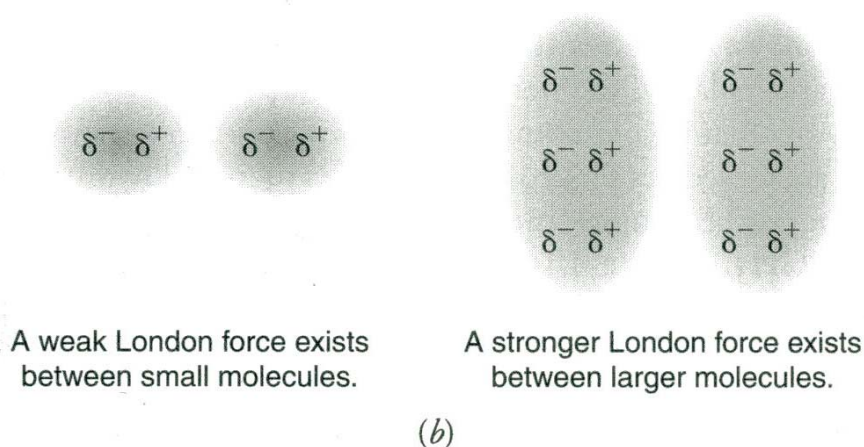
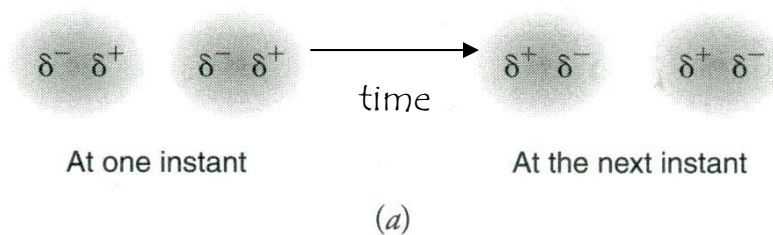
d) \_\_\_\_\_

e) \_\_\_\_\_



## London Forces (Hydrophobic Interactions)

Even though non-polar molecules do not have a permanent dipole moment, they still contain electrons which are in continuous motion. This electron movement results in temporary dipole moments which can be felt by neighboring compounds producing a temporary dipole moment in the neighbor (induced dipole). The electrons in the neighbor are attracted or repelled by these temporary partial charges. As the temporary dipole shifts, so do the dipoles on the neighbors. As the surface area of a compound increases, the chance for temporary dipole formation increases along with the ability to induce dipoles in the neighbor, therefore larger molecules with greater surface area have stronger London forces.



8. Identify the dominant IMF for each compound and  
a) for both parts (i) and (ii), which compound has the highest boiling point? Explain why.  
b) for part (ii) which compound is more soluble in water? Explain why.

