CHEMISTRY 10 – Help Sheet #1
Chem 103 Review
Do the topics appropriate for your course
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Nuggets: Electronegativity, Bond Polarity, Lewis dot structures, Bond Order, Resonance, VSEPR, Bond angles, Molecular Polarity, Hybrid Orbitals, Intermolecular forces, Solubility/Miscibility

ELECTRONEGATIVITY: The degree to which an atom can attract electrons \textit{in a bond}. The greater the electronegativity (EN), the greater the "pull" an atom has on the electron pair. (The EN value of H places it approximately between B and C on the Periodic Table.)

\[ \text{Trend: Increases toward F.} \]

TYPES OF BONDS

Nonpolar Covalent: \textit{equal sharing} of e\textsuperscript{-} between two identical nonmetal elements (e.g., O\textsubscript{2} and Cl\textsubscript{2}).
\[ \Delta \text{EN} = 0 \text{ (same atoms)} \]

Polar Covalent: \textit{unequal sharing} of e\textsuperscript{-} between two nonmetals (e.g., H–F and C=O).
\[ \Delta \text{EN} \text{ small (different atoms)} \]
\[ \Delta \text{EN} = |\text{EN}_1 - \text{EN}_2| \text{ with EN}_1 \text{ and EN}_2 \text{ the EN for the 2 atoms in the bond} \]

BOND POLARITY: as the 2 atoms get farther apart on the Periodic Table \( \Rightarrow \Delta \text{EN} \uparrow \Rightarrow \text{bond polarity} \uparrow \)

As electrons are pulled to one side of the bond a slight positive charge on one side (\( \delta^+ \)) and a slight negative charge (\( \delta^- \)) on the other side of the bond with the more EN atom. The bond polarity vector points toward the more negative side of the bond (atom with greater EN); \( \delta^- \) \( \rightarrow \) \( \delta^+ \)

LEWIS DOT STRUCTURES

1. Count valence electrons
2. Draw skeleton structure
3. Build octets
4. Check number of electrons drawn; if equals the #valence e\textsuperscript{-} in Step 1 \( \Rightarrow \) done.
5. Too many e\textsuperscript{-} \( \Rightarrow \) Step 2 & add a multiple bond; Too few e\textsuperscript{-} \( \Rightarrow \) add electron pairs to the central atom.

\textit{For organic molecules, connect C atoms together and make sure the #bonds is correct as remaining atoms are attached: C = 4 bonds; N = 3 bonds usually; O = 2 bonds usually; H = 1 bond}?

BOND ORDER (BO): \textit{Single bond} \( \Rightarrow \) BO = 1; \textit{Double bond} \( \Rightarrow \) BO = 2; \textit{Triple bond} \( \Rightarrow \) BO = 3
\[ \text{Bond Order} \uparrow \Rightarrow \text{bond strength} \uparrow \text{and bond length} \downarrow \]
\[ \text{Bond Order}_{\text{resonance}} = \frac{\# \text{bonds involved in resonance}}{\# \text{locations in resonance are distributed}} \]

Resonance and Resonance Structures: In Chem 103/104 resonance structures typically have double bonds that move; \textit{common structures}:

\[ \text{Benzene, C}_6\text{H}_6, \text{BO}_{\text{C-C}} = 1.5 \]
\[ \text{Carboxylate, } -\text{CO}_2^-, \text{BO}_{\text{C-O}} = 1.5 \]
\[ \text{Amide, BO}_{\text{C-O}} = \text{BO}_{\text{C-N}} = 1.5 \]
VSEPR - Valence Shell Electron Pair Repulsion

1. Draw correct Lewis dot structure.
2. Determine #Domains; Domains = #lone pairs of e\(^{-}\) around central atom + #atoms bonded to central atom
3. Compare to table below.

<table>
<thead>
<tr>
<th>Total = A + B = domains</th>
<th># atoms bonded to central atom A</th>
<th># lone pairs on central atom B</th>
<th>Electron region geometry (ERG)</th>
<th>ERG Picture</th>
<th>Molecular geometry (MG)</th>
<th>MG Picture</th>
<th>Angles</th>
<th>Ex</th>
<th>Polarity</th>
<th>HO</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>2</td>
<td>0</td>
<td>Linear</td>
<td>Linear</td>
<td>180(^{\circ}) BeCl(_2)</td>
<td>NP(^{\text{a}})</td>
<td>sp</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>3</td>
<td>0</td>
<td>Triangular planar</td>
<td>Triangular planar</td>
<td>120(^{\circ}) BF(_3)</td>
<td>NP(^{\text{a}})</td>
<td>sp(^{2})</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>2</td>
<td>1</td>
<td>Triangular planar</td>
<td>Bent</td>
<td>&lt;120(^{\circ}) AlF(_2)</td>
<td>P</td>
<td>sp(^{2})</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>4</td>
<td>0</td>
<td>Tetrahedral</td>
<td>Tetrahedral</td>
<td>109.5(^{\circ}) CH(_4)</td>
<td>NP(^{\text{a}})</td>
<td>sp(^{3})</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>3</td>
<td>1</td>
<td>Tetrahedral</td>
<td>Triangular pyramidal</td>
<td>&lt;109.5(^{\circ}) NH(_3)</td>
<td>P</td>
<td>sp(^{3})</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>2</td>
<td>2</td>
<td>Tetrahedral</td>
<td>Bent</td>
<td>&lt;109.5(^{\circ}) H(_2)O</td>
<td>P</td>
<td>sp(^{3})</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

P = polar molecule; NP = nonpolar molecule
* = The nonpolar (NP) molecules assume the atoms around the central atom are the same; if not, the molecule will be polar

**HYBRIDIZATION**

**Sigma (\(\sigma\)) Bond:** Bond lies on the internuclear axis between the nuclei; all bonds contain a \(\sigma\) bond.
*In this course, \(\sigma\) bond are created from overlap of two hybrid orbitals or overlap of a 1s atomic and a hybrid orbital \(\sigma = HO-HO\) (e.g., sp\(^{2}\)-sp\(^{3}\)) or \(\sigma = 1s-HO\) (e.g., H–CH\(_3\)).*

**Pi (\(\pi\)) Bond:** Bond does not lie on the internuclear axis and is present with double and triple bonds.
*Pi (\(\pi\)) bonds are created from the overlap of 2 p-atomic orbitals*

\[\pi = p-p\]

Single bond = \(\sigma\) bond; Double bond = \(\sigma\) bond + \(\pi\) bond; Triple bond = \(\sigma\) bond + 2\(\pi\) bond

*For a video example of these types of questions click here.*
Example 1: a. Draw the Lewis dot structure for CF₂=CH₂. b. Identify the hybridization for each atom. c. Identify what orbitals make up each bond.

**Answer 1:**

![Lewis dot structure](image)

**INTERMOLECULAR FORCES (IMF):** Forces *between* molecules, not within a molecule.

- **H-bonding:** *Criteria: H bonded to N, O, or F;* look for –OH groups; examples: CH₃OH(l) (*alcohols*); H₂O(l); CH₃COOH(l) (*carboxylic acids*); CH₃NH₂ (*amines*); H-bonding IMF are usually stronger than Dipole-Dipole and LDF.

  ![H-bonding](image)

- **Dipole–Dipole:** *Criteria: polar molecule;* example: 

  ![Dipole–Dipole](image)

  *Polar molecules:*
  - diatomic X–Y
  - compounds with a carbonyl group:

  ![Polar molecules](image)

  - chemicals with *VSEPR Table: 2|0, 3|0, 4|0 with different* atoms around central atom
  - *VSEPR Table: 2|1, 3|1, 2|2 always polar*

  *Nonpolar molecules:*
  - diatomic X–X
  - hydrocarbons (only H and C; CₓHᵧ)
  - *VSEPR Table: 2|0, 3|0, 4|0 with same* atoms around central atom

- **London Dispersion Forces (LDF)** = Induced Dipole = Instantaneous Dipole: *Criteria: has electrons;* occurs in *all* molecules and individual atoms; caused by electrons shifting to one side of the molecule/atom creating an instantaneous dipole within the molecule; hydrocarbons have only LDFs for IMFs

  ![London Dispersion Forces](image)

  #e↑⇒ LDFs↑

**DETERMINE WHICH MOLECULE HAS LARGEST IMFs**

Determine IMF and compare: IMFs ↑ (generally) ⇒ H-Bonding > Dipole–Dipole > LDF

If two molecules have the *same IMF:*

a. **H-bonding:**
   1. more –OH or –NH groups ⇒ more H-bonding ⇒ *larger IMFs*

b. **Dipole-dipole:**
   1. one molecule is more polar when ∆EN is greater ⇒ greater permanent dipole ⇒ *larger IMFs*
   2. geometric isomers: *cis*-isomer: polar; *trans*-isomer: nonpolar (see HelpSheet #3)

c. **LDFs only:**
   1. larger molecule ⇒ more electrons ⇒ greater LDFs ⇒ *larger IMFs*
   2. hydrocarbon molecules: *more linear molecule* (i.e., less branched) ⇒ greater LDFs ⇒ *larger IMFs*

IMFs↑⇒ Boiling Point↑
SOLUBILITY: Liquids dissolving into liquids = miscibility (text, page 565-567):

In general: Like likes Like
- polar substances dissolve into polar substances (dissolve = miscible);
- nonpolar substances (e.g., hydrocarbons!) dissolve into nonpolar substances (dissolve = miscible);
- polar and nonpolar substances don’t dissolve into one another (don’t dissolve = immiscible);
- more H-bonding regions ⇒ more polar ⇒ more soluble in water

1. Answer the questions below about the structure and bonding in this molecule. (Hint: Draw in the lone pair e⁻.)
   a. What is the angle between the H-C-O labeled with an “a”?
   b. What is the electron region geometry around the carbon atom labeled “b”?
   c. What is the molecular geometry around the oxygen atom labeled “c”?
   d. What is the hybridization of the carbon atom labeled “a”?
   e. What orbitals overlap to make the bond between the carbon atom labeled “d” and the attached hydrogen atom?
   f. What orbitals overlap to make the pi (π) bond between the carbon atom labeled “a” and the oxygen atom attached to it?
   g. What is the electron region geometry around the carbon atom labeled “d”?
   h. How many lone pairs of electrons are there in the molecule?

2. Answer the following questions for the molecule to the right.
   a. Draw in the lone pairs of electrons in the structure.
   b. What is the bond order for the bond labeled “a”?
   c. What is the hybridization on the S labeled “b”?
   d. What is the angle for the C–O–H on the O labeled “c”?
   e. What orbitals are used to form the bond labeled “e”?
   f. How many σ and π bonds are in the molecule?
   g. What is the electron region geometry around the S atom labeled “b”?
   h. What orbitals are used to form the bonds labeled “a”?

3. For each set of chemicals, order them from lowest to highest boiling point.
   I. CH₄, C₃H₈, C₂H₆         II. CH₃CH₂Br, CH₃CH₂F, CH₃CH₂Cl
   III. H₂O, CH₃OH, CH₃OCH₃

4. Place the molecules in order of most to least soluble in water.
   CH₃CH₂CH₂OH, CH₃CH₂CH₂CH₂CH₂CH₃, HOCH₂CH₂OH

5. What is the electron regional geometry, hybridization, and angle around the bold underlined C atom of each molecule below? (Hint: Draw the Lewis dot structure for each molecule.)
   a. CH₄  b. CH₃CH₃  c. CH₂H₂  d. HC≡CH

6. Draw Lewis dot structures for each of the following molecules. Organic structures will usually have the following bonding: C = 4 bonds; N = 3 bonds; O = 2 bonds
   a. C₂H₆  b. C₂H₄  c. CH₃CH₂OH  d. CH₃CH₂OCH₂CH₃  e. C₆H₆ (benzene = 6-membered carbon ring)
7. I. Which molecule below has the strongest C—O bond? *(Hint: Draw the Lewis dot structure and determine BO)*  
II. Which molecule below has the longest C—O bond?  
   a. CO$_3^{2-}$  b. CH$_3$CH$_2$OH  c. CO  d. H$_2$CO (C is central atom)  e. Same strength/length.

8. Identify the polar regions in this molecule by labeling each atom of the molecule with $\delta^-$ or $\delta^+$.  

ANSWERS  
1. a. 120°  {3 domains ⇒ triangular planar ⇒ 120°}  
b. tetrahedral  {4 domains ⇒ tetrahedral}  
c. bent  {4 domains ⇒ 2 atoms + 2 lone pairs ⇒ MG = bent}  
d. sp$^2$  {3 domains ⇒ sp$^2$}  
e. $\sigma$: 1s(H)–sp$^2$(C)  {single bond ⇒ $\sigma$ bond; $\sigma$ bonds made from HO–HO or HO-1s usually; C has 3 domains ⇒ sp$^2$; $\sigma$ bond = 1s(H)–sp$^2$(C)}  
f. $\pi$: p(C)–p(O)  {all $\pi$ bonds made from overlap of p-orbitals; $\pi$ bond = p(C)–p(O)}  
g. triangular planar  {C atoms have 3 domains ⇒ triangular planar}  
h. 6  {see diagram}  

2. a.  
b. 2  {double bond with no resonance ⇒ bond order = 2}  
c. sp$^3$  {4 domains around S ⇒ sp$^3$}  
d. <109.5°  {around the vertex atom (O atom) there are 3 domains ⇒ tetrahedral ⇒ 109.5°}  
e. 1s(H)–sp$^3$(N)  {single bond ⇒ $\sigma$ bond; $\sigma$ bonds made from HO–HO or HO-1s usually; N has 4 domains ⇒ sp$^3$; $\sigma$ bond = 1s(H)–sp$^3$(N)}  
f. 22σ , 2π  {single bond = $\sigma$ bond; double bond = $\sigma + \pi$ bond; triple bond = $\sigma + 2\pi$ bonds}  
g. tetrahedral  {4 domains around S ⇒ tetrahedral ERG}  
h. $\pi$: p(C)–p(C); $\sigma$: sp$^2$(C)–sp$^2$(C)  {double bond = $\sigma + \pi$ bond; $\sigma$ bonds made from HO–HO or HO-1s usually; each C has 3 domains ⇒ sp$^2$; $\sigma$ bond = sp$^2$(C)–sp$^2$(C); all $\pi$ bonds made from overlap of p-orbitals; $\pi$ bond = p(C)–p(C)}  

3. I. CH$_4$ < C$_2$H$_6$ < C$_3$H$_8$  {as #e$^+$ ↑ ⇒ LDF↑ ⇒ IMF↑ ⇒ boiling points↑}  
II. CH$_3$CH$_2$F < CH$_3$CH$_2$Cl < CH$_3$CH$_2$Br  {as #e$^+$ ↑ ⇒ LDF↑ ⇒ IMF↑ ⇒ boiling points↑}  
III. CH$_3$OCH$_3$ < CH$_3$OH < H$_2$O  {IMFs↑ ⇒ Boiling Point↑; CH$_3$–O–CH$_3$ has dipole-dipole IMFs; CH$_3$O–H has H-bonding; H$_2$O also has H-bonding but has 2 H atoms that can H-bond; H–O–H so it has greater H-bonding while CH$_3$OH only has 1 H that can undergo H-bonding; CH$_3$OCH$_3$ has no H-bonding, is polar and therefore has dipole-dipole IMFs which are usually weaker than H-bonding and therefore has the lowest bp; CH$_3$OH is next because it has H-bonding but less H-bonding than H$_2$O; CH$_3$OCH$_3$ < CH$_3$OH < H$_2$O}
4. most soluble to least soluble in water: \( \text{HOCH}_2\text{CH}_2\text{OH} > \text{CH}_3\text{CH}_2\text{CH}_2\text{OH} > \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3 \)

Water is polar and substances that are polar will be soluble (miscible) in water. Both \( \text{CH}_3\text{CH}_2\text{CH}_2\text{OH} \) and \( \text{HOCH}_2\text{CH}_2\text{OH} \) are polar because of the polar –OH groups. Since chemical \( \text{HOCH}_2\text{CH}_2\text{OH} \) has two –OH groups it has more polar regions than \( \text{CH}_3\text{CH}_2\text{CH}_2\text{OH} \) with only one –OH group. The substance with more polar regions will be more soluble in another polar substance such as water. Chemical \( \text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3 \) is a hydrocarbon and is therefore nonpolar and least soluble in water.

5. a. tetrahedral, \( \text{sp}^3 \)
   {from LDS: \( \text{C has 4 domains} \Rightarrow \text{tetrahedral, sp}^3, 109.5^\circ \)}

b. tetrahedral, \( \text{sp}^3 \)
   {from LDS: \( \text{C has 4 domains} \Rightarrow \text{tetrahedral, sp}^3, 109.5^\circ \)}

c. triangular planar, \( \text{sp}^2 \)
   {from LDS: \( \text{C has 3 domains} \Rightarrow \text{triangular planar, sp}^2, 120^\circ \)}

d. linear, \( \text{sp} \)
   {from LDS: \( \text{H—C≡C—H} \Rightarrow \text{C has 2 domains} \Rightarrow \text{linear, sp, 180}^\circ \)}

6. a.

b.

c.

d.

e.

7. I. c {as BO\( ^\uparrow \) \Rightarrow bond strength\( ^\uparrow \); BO = 3 for structure c}

a. \[
\begin{pmatrix}
0.0 \\
\vdots \\
0.0 \\
\vdots \\
0.0 \\
\vdots \\
0.0
\end{pmatrix}^2
\] ; BO = 4/3; b. \( \text{H—C—O—H} \); BO = 1; c. \( \text{C≡O—H} \); BO = 3; d. \( \text{H—C—O—H} \); BO = 2

II. b {as BO\( ^\uparrow \) \Rightarrow bond length\( ^\downarrow \); BO = 1 for structure b}

8. The higher electronegativity of the O atoms “pull” the electrons towards those atoms and concurrently creates a positive region at the C and H atoms that are attached to the O atoms.