Chem 309 Name Key

Chapter 3 Supplemental Homework
Polarity and IMFs

Bond Polarity

When we talk about a covalent bond as a shared pair of valence electrons, we have learned that these bonding electrons are NOT always shared equally.

Electronegativity: the ability of an atom to attract bonding electrons
EN

1. On the periodic table below, indicate the element with the greatest electronegativity and the element with the least electronegativity, then draw arrows to show the overall trend for increasing electronegativity.

EN ↑ greatest EN

EN ↑

1 Group IA	2 Group IIA	3 Group IIIB	4 Group IVB	5 Group VB	6 Group VIB	7 Group VIIB	8 Group	9 Group VIIIB	10 Group	11 Group IB	12 Group IIB	13 Group IIIA	14 Group IVA	15 Group VA	16 Group VIA	17 Group VIIA	18 Group VIIIA
1 H 1.01																	2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.30											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.84	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.64	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57 La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (269)	109 Mt (268)	110 – (271)	111 – (272)	112 – (277)		114 – (289)		116 – (289)		118 – (293)

least EN

2. Write a complete sentence summarizing the trend in electronegativity.
Electronegativity increases as we move up a to the right.

We can think of electronegativity as "electron greed". When the electrons in covalent bonds are shared by atoms with different electronegativities, a polar covalent bond is created. In a polar covalent bond, the unequal sharing of electrons gives the bond a partially positive end or pole ($\delta+$) and a partially negative end or pole ($\delta-$).

Even though carbon and hydrogen have slightly different electronegativities, we will consider ALL C–H bonds to be nonpolar.

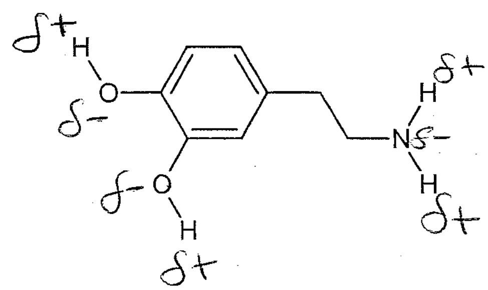
Electronegativity (EN) and Chemical Bonds

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

i) dipole moment arrows $+\!\!\!\longrightarrow$
The arrow head points toward the element with the greater electronegativity.

ii) partial positive (δ+) and partial negative (δ–)
where the Greak letter delta (δ) indicates partial

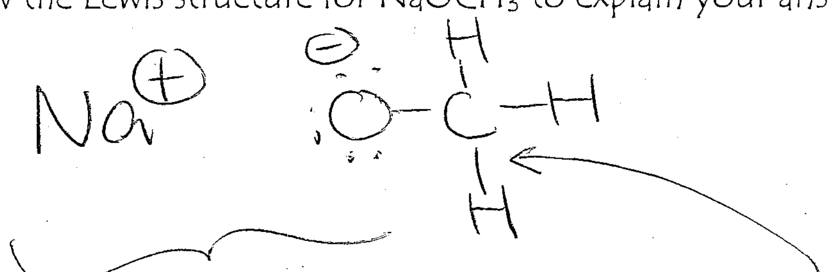
3. Label the polar covalent bonds in dopamine using the δ+ and δ– symbols.
Remember to add any missing lone pairs.



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

4. What types of bonds are present in sodium methoxide ($NaOCH_3$)?
Draw the Lewis structure for $NaOCH_3$ to explain your answer.



ionic bonds & covalent bonds

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

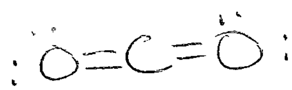
In addition to recognizing polar bonds within organic molecules, we need to visualize the 3-dimensional shape (molecular geometry). VSEPR theory predicts the molecular geometry by placing valence electrons (bonding and lone pairs) as far apart as possible around a central atom. For organic and biochemistry, 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.

2 VSEPR "pairs" = Linear e^- geometry

Two double bonds

Molecular geometry: Linear

CO_2



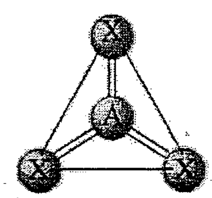
One triple bond and one single bond

Linear

HCN

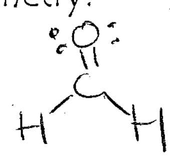
3 VSEPR "pairs" = Trigonal planar e^- geometry

One double bond and two single bonds

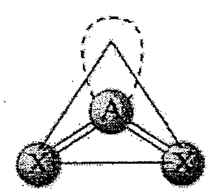


Molecular geometry: Trigonal planar

H_2CO, $COCl_2$

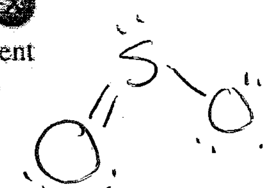


One double bond, one single bond, and one nonbonding pair



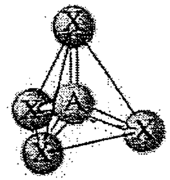
Angular or bent

SO_2, S_2O



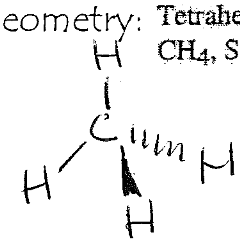
4 VSEPR "pairs" = Tetrahedral e^- geometry

Four single bonds

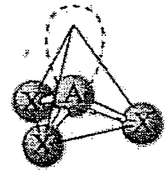


Molecular geometry: Tetrahedral

CH_4, $SiCl_4$

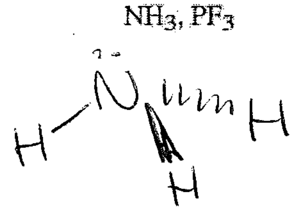


Three single bonds and one nonbonding pair

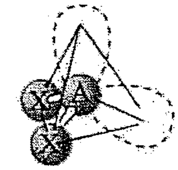


Trigonal pyramidal

NH_3, PF_3

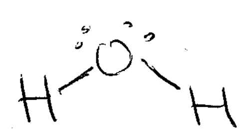


Two single bonds and two nonbonding pairs



Angular or bent

H_2O, OF_2



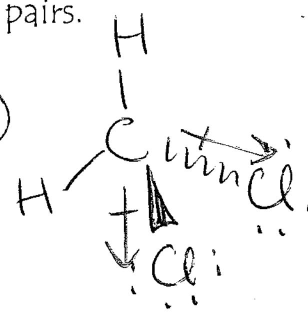
Molecular Polarity

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.

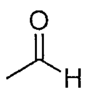
5. Which of the following molecules are polar? Remember to draw in the dipole moments and add any missing lone pairs.

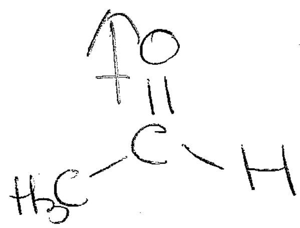
a) CH_2Cl_2
(Hint: Draw the 3-D shape with wedges and dashes)

polar

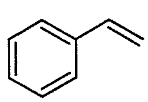


or equivalent

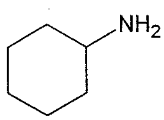
b) 

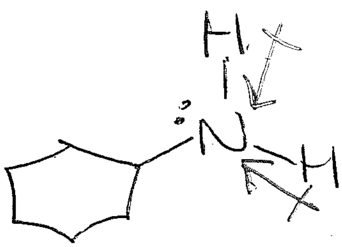


polar

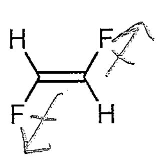
c) 

all C's & H's ∴ non-polar

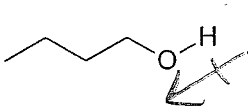
d) 



polar

e) 

non-polar

f) 

polar

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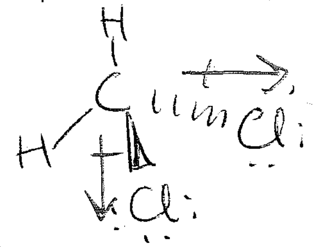
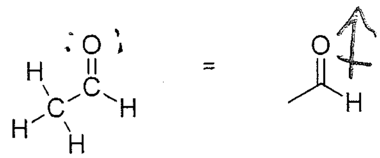
5. Which of the following molecules are polar? Remember to draw in the dipole moments and add any missing lone pairs. Where applicable the Lewis structure is also drawn as a simplified bond-line structure.

a) CH_2Cl_2
(Hint: Draw the 3-D shape with wedges and dashes)



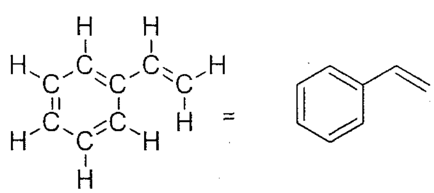
polar

b)



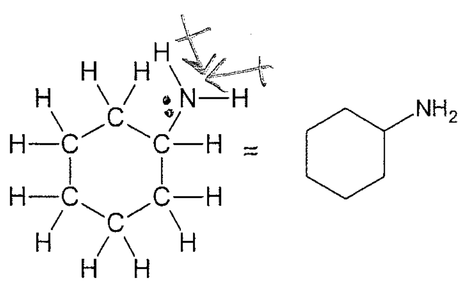
polar

c)



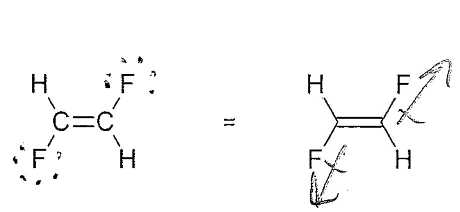
non-polar (all C's & H's hydrocarbon)

d)



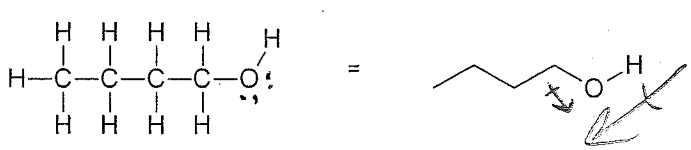
polar

e)



symmetrical pull
⇓
non-polar

f)



polar

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.

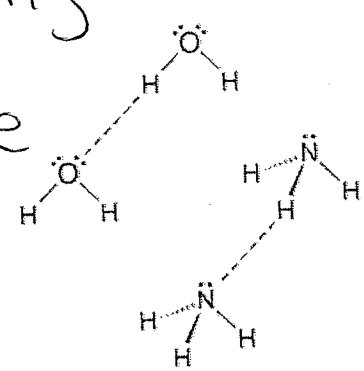
A. Polar compounds with PERMANENT partial or full charges
   a) H-bonding: compounds with N-H, O-H or F-H bonds
   b) Dipole-dipole interactions: polar compounds
   c) Ion-dipole interactions: an ion attracted to a polar compound
   d) Salt bridge: another name for an ionic bond
   e) Coordinate covalent: 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)

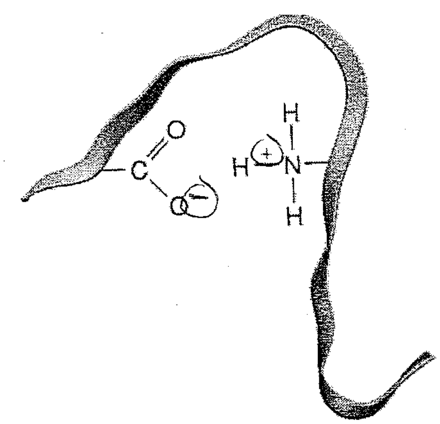
6. Indicate the type of IMF or non-covalent interaction represented in each figure.

a) H-bonding

b) salt bridge



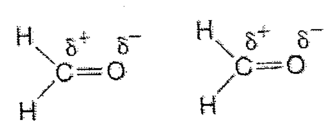
(a)



(b)

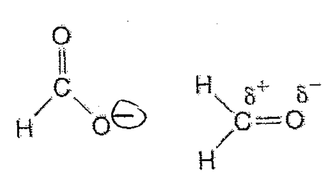
c) dipole-dipole

d) ion-dipole

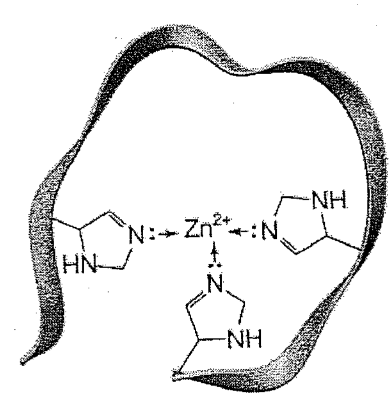


(c)

e) coordinate covalent



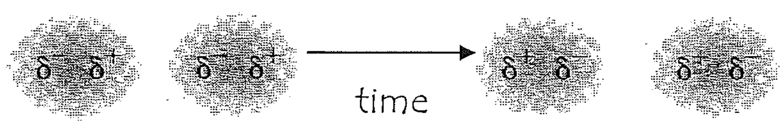
(d)



(e)

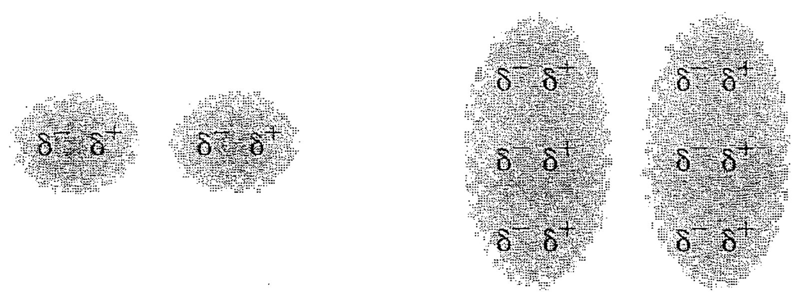
London Forces (Hydrophobic Interactions)

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.



At one instant   At the next instant

(*a*)



A weak London force exists between small molecules.   A stronger London force exists between larger molecules.

(*b*)

7. 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.

higher bp b/c more surface area

i) $CH_3CH_2CH_3$ vs $CH_3CH_2CH_2CH_2CH_2CH_3$

London   London

higher bp b/c stronger IMF

ii) $CH_3CH_2CH_2OH$ vs $CH_3CH_2CH_3$

H-bond   London

more H_2O soluble b/c it can H-bond w/ H_2O