1. Which of the following statements are true?
I. Bond polarity can be calculated using the differences in electronegativity values.
II. Bond polarity describes to what degree electrons are shared between two atoms (i.e. Is one atom "hogging" the electrons and how much?).
III. Homonuclear bonds are always polar.
IV. F—Br is a more polar bond than F—I
V. Li—B is less polar than O—H
VI. O—O is more polar than H—H
    1. I, II, and VI
    2. I only
    3. I and II
    4. I, II,V correct
    5. II, V, VI
    6. II, IV, V
    7. II and V
Explanation:
I is true – you will need to be able to do these calculations on the test because they will allow you to rank the polarities of various bonds.
II is true and describes bond polarity conceptually.
III is false because homonuclear bonds, bonds between two atoms of the same element (i.e. H—H or O—O) are always NON-POLAR because the difference in EN is always zero.
IV is false because the EN difference between F and I is larger than the EN difference between F and Br, implying a more polar bond. We know this because EN decreases down the periodic table.
V is true. You must have memorized the EN values of all elements in the second row of the periodic table. You will need them on the test. In this case we see Li—B has a difference in EN of 1.2 while O—H has an EN difference of 1.3. Remember than a larger EN difference is a more polar bond.
VI is false. These are both homonuclear bonds with an EN difference of ZERO meaning both bonds are completely nonpolar.

2. Which of the following molecules is non-polar?
    1. CO_2 correct
    2. H_2O
    3. CH_3F
    4. O_3
    5. NO
Explanation: The only symmetrical molecules is carbon dioxide and it therefore cannot possibly be polar. You might be tempted to pick ozone (O_3) because it has nonpolar bonds but if you draw the lewis structure, you'll see that resonance puts a partial negative charge on the peripheral oxygens and that the structure is angular (see Worksheet 4 2007) – so the molecule ends up being polar!

3. Which of the following statements are correct?
I. If a central atom is bonded to five other atoms, it is hypovalent and trigonal bipyramidal.
II. If a central atom is bonded to three identical atoms, it must be nonpolar.
III. A molecule of generalized formula AB_2U_2 always has an angular shape.

    1. II only
    2. I only
    3. III only   correct
    4. I and II
    5. I and III

6. II and III
7. I, II, and III

Explanation:

I is false because having five bonds would mean TEN electrons on one atom which has more than the octet rule requires (namely eight) so it is HYPERvalent not hypovalent. The shape is correct.

II is false because it could be like ammonia, NH_3, and be trigonal pyramidal and polar.

III is true – look it up. You should be able to name the shape of any molecule, generalized or specific. Here we have used A to mean the central atom, B to represent a bond to another atom, and U to represent a lone pair.

4. What bond angles exist in PCl_5?
1. 109.5°, 180°
2. 90°, 60°
3. 120°
4. 30°, 109.5°
5. 90°, 120°, 180°   correct

Explanation: The shape is trigonal bipyramidal so the angles are 90, 120 and 180. Look it up if you don't believe me. You should know the bond angles between every three atoms in all of our structures.

5. Which of the following statements are true regarding Valence Bond Theory?

I. The number of orbitals is conserved when hybridizing.

II. An sp orbital has twenty five percent more s character than an sp^3 orbital.

III. Valence bond theory is lame.
1. I only
2. II only
3. III only
4. I and II   correct
5. I and III
6. II and III
7. I, II and III

Explanation:

I is true – if you hybridize three orbitals together you get three out. For instance, one s + one p + one p = three sp^2.

II is true – the percent character of a hybrid orbital is just what proportion of the original orbitals is that type. So for sp 1 s / 2 orbitals = 50% s character. For sp^3, 1 s / 4 orbitals = 25% s character. This describes how much the orbital is like an s orbital. So you would expect the sp orbital to be more spherical and closer to the nucleus, while the sp^3 orbital would be more dumb-bell like and farther from the nucleus.

III is false – valence bond theory is totally awesome, duh. Are you even TRYING to learn?

6. Which of the following molecules is matched with its correct electronic geometry?
1. BH_3 , trigonal planar   correct
2. H_2O, angular
3. NH_3 , trigonal planar
4. XeF_4 , tetrahedral
5. CH_4, square planar

Explanation: BH_3 is the only correctly assigned one. Angular and square planar are both MOLECULAR geometries so they are obviously wrong without even having to draw them. Obviously… And NH_3 has a lone pair so it would be tetrahedral. XeF_4 has TWO lone pairs so it would be octahedral. Make sure you know the difference between electronic and molecular geometries and can assign them readily.

7. What molecular geometries can you find in this molecule?

H

$$H_3C-C\equiv C\text{-}\overset{..}{\underset{\cdot,}{O}}-\overset{|}{\underset{|}{C}}-NH_2$$
$$H$$
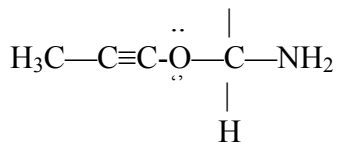
    I.   tetrahedral
    II.   linear
    III.   angular
    IV.   trigonal bipyramidal
    V.   trigonal planar
    VI.   trigonal pyramidal
    VII. Waldo
  1. I, II, III, VI correct
  2. I, II, V
  3. I II, VI
  4. II and V
  5. II, III and VI
  6. VII only

Explanation: From left to right, the geometries around the central atoms are tetrahedral, linear, angular, and trigonal pyramidal. If you picked VII. Waldo you are awesome at life but you fail at chemistry.

8. How many sigma (σ) and pi (π) bonds are in the Lewis structure for $HCOCHCHCH_3$?
  1. 7 σ and 1 π
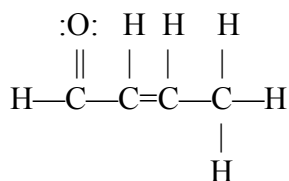 **2. 10 σ and 2 π correct**
  3. 12 σ and 0 π
  4. 8 σ and 2 π
  5. 7 σ and 2 π

Explanation: A single bond is one sigma bond, a double bond is one sigma and one pi bond, and a triple bond is one sigma and two pi bonds. The molecule looks like this so ….just count them?

$$:O:\quad H\ \ H\ \ \ H$$
$$\|\quad |\ \ |\quad\ |$$
$$H-C-C=C-C-H$$
$$|$$
$$H$$

9. In carbon dioxide, the carbon must hybridize its _____ atomic orbitals to make _____ to bond to the oxygens.
  1. 2s and 2p; an sp orbital
  2. 1s and 1p; two sp orbitals
  3. 2s and 2p; two sp orbitals correct
  4. 1s, 2s, and 2p; three sp orbitals
  5. 1s, 2s and 2p; two sp orbitals

Explanation: Bonding involves the valence electrons. For carbon, this means its 2s and 2p electrons. It would not use its 1s electrons because they are not valence. It would not use 1p electrons because there is no such thing. Furthermore, when you hybridize n atomic orbitals you get n molecular orbitals back. So combining two atomic orbitals, the 2s and the 2p, yields two molecular orbitals, two sp orbitals.

10. Which of the following statements concerning molecular orbital theory is/are true?
  I. MO theory explains resonance but cannot predict paramagnetism. VB theory is superior because it can do both.
  II. MO theory describes the nature of bonds and why they form.

III. To fill up a MO electronic diagram, you use the same rules we used for an AO electronic diagram.
IV. He$_2$ does not exist because its four electrons all raise its energy relative to unbonded He atoms.
1. I only
2. II only
3. III only
4. I and II
5. I and III
6. I and IV
7. **II and III**   **correct**
8. II and IV
9. III and IV
10. I, II, and III
11. I, II, and IV
12. I, III, and IV
13. II, III and IV

Explanation:
I is false – MO theory DOES predict paramagnetism. VB does not.
II is true.
III is true – you can still use aufbau, hund, and pauli, the only difference is that we fill up MOs instead of AOs.
IV is false – two electrons in the bonding orbital are stabilizing. Only the two antibonding electrons are destabilizing. (However, He$_2$ DOES not exist due to this destabilization)

11. List, in order of increasing energy, the names of the molecular orbitals for N$_2$.
1. $\sigma_{1s} < \sigma^*_{1s} < \sigma_{2s} < \sigma^*_{2s} < \pi_{2py} = \pi_{2pz} < \sigma_{2p} < \pi^*_{2py} = \pi^*_{2pz} < \sigma^*_{2p}$ **correct**
2. $\sigma^*_{1s} < \sigma_{1s} < \sigma^*_{2s} < \sigma_{2s} < \pi^*_{2py} = \pi^*_{2pz} < \sigma^*_{2p} < \pi_{2py} = \pi_{2pz} < \sigma_{2p}$
3. $\sigma_{1s} < \sigma_{2s} < \pi_{2py} = \pi_{2pz} < \sigma_{2p}$
4. $\sigma_{1s} < \sigma^*_{1s} < \sigma_{2s} < \sigma^*_{2s} < \pi_{2p} < \pi_{2pz} < \sigma_{2p} < \pi^*_{2py} < \pi^*_{2pz} < \sigma^*_{2p}$
5. $\sigma_{1s} < \sigma^*_{1s} < \sigma_{2s} < \sigma^*_{2s} < \sigma_{2p} < \pi_{2py} = \pi_{2pz} < \pi^*_{2py} = \pi^*_{2pz} < \sigma^*_{2p}$

Explanation: See p111 of your course manual. We won't ask you to reproduce that chart on the test, but you will need to know it to answer some questions (excluding bond lengths and bond energies, don't worry about that). It's not really that much to memorize – just learn the names and ordering of the MOs and then you should be able to fill them out with electrons using aufbau, hund and pauli. That in turn will allow you to predict the bond order and paramagneticity. Try doing it for one of these molecules without looking – can you reproduce all the information? If not…learn it more betterer [sic].
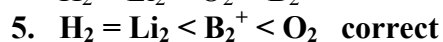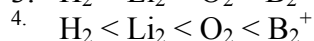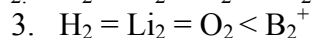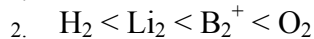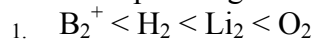
12. Calculating bond order from MO
What is the bond order in carbon monoxide, CO?
1. 0.5
2. 1.0
3. 1.5
4. 2.0
5. 2.5
6. **3.0**   **correct**
7. 3.5
8. 4

Explanation: Bond order = (# antibonding electrons - # bonding electrons)/2 = (10-4)/2 = 3. To figure out the number of antibonding and of bonding electrons, you will need to draw the MO diagram. Because CO has ten electrons, its diagram is the same as that of N$_2$ given on p111 of your course manual. All electrons in orbitals with a * are antibonding. Unstarred orbitals are bonding and contain bonding electrons. On the test, if you forget how to draw the MO diagram (Gasp! Don't you love chemistry??), you could just draw the lewis structure and see that CO has a triple bond and figure out that the bond order is 3. This method doesn't always work but it's better than nothing if you forget how to do the REAL method on the test. Make sure you can do
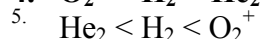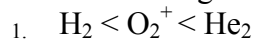
bond order calculations for homonuclear And heteronuclear molecules and molecular IONS (i.e. N_2^{2+}).

13. Rank the paramagnetism of the following compounds from least to greatest: B_2^+, H_2, O_2, Li_2.
    1. $B_2^+ < H_2 < Li_2 < O_2$
    2. $H_2 < Li_2 < B_2^+ < O_2$
    3. $H_2 = Li_2 = O_2 < B_2^+$
    4. $H_2 < Li_2 < O_2 < B_2^+$
   **5. $H_2 = Li_2 < B_2^+ < O_2$ correct**

Explanation: Draw the MO diagrams (see p111 of your course manual). Both dihydrogen and dilithium are NOT paramagnetic – they have NO unpaired electrons. Ionized diboron has only ONE unpaired electron. Dioxygen has TWO unpaired electrons so it is more paramagnetic than the ionized diboron.

14. Rank bond length of the following species from shortest to longest: H_2, He_2, O_2^+
    1. $H_2 < O_2^+ < He_2$
    2. $H_2 < He_2 < O_2^+$
    3. $O_2^+ < He_2 < H_2$
    4. $O_2^+ < H_2 < He_2$ correct
    5. $He_2 < H_2 < O_2^+$

Explanation: To rank bond length, calculate the bond orders. High bond order = strong bond = short bond. Bond orders of H_2, He_2, and O_2^+ respectively are 1, 0, and 2.5.

15. Choose the species below that has the most delocalized electrons.
    1. HCOO⁻   correct
    2. O_2
    3. C_2H_2
    4. NaCl
    5. XeF_4
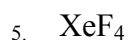
Explanation: Formate, HCOO⁻, is the only molecule with resonance and therefore the only molecule with delocalization.

16. Which of the following statements about gas laws is/are true?
    1. There exists an inverse proportionality between the pressure and volume of a gaseous system.
    2. There exists an inverse proportionality between number of moles and volume of a gaseous system.
    3. The ideal gas constant is larger for heavier gases.
    1. I   Correct
    2. II only
    3. III only
    4. I, II only
    5. I, III only

Explanation: The pressure and volume of a gaseous system is inversely related (Boyle's Law), thus I is correct. The number of moles and volume of a gaseous system are directly related (Avogadro's Law), thus II is incorrect. The ideal gas constant is R, which is the same for all gases, thus III is incorrect.

17. If we increase the volume of a gaseous system by a factor of 5 and raise the pressure by a factor of 10, then the temperature of the system will:
    1.   Increase by a factor of 50   Correct
    2.   Decrease by a factor of 50
    3.   Increase by a factor of 2
    4.   Decrease by a factor of 2
    5.   Not change.

Explanation: PV=nRT. Increasing the volume by a factor of 5 would increase the temperature of the system by a factor of 5. Increasing the pressure by a factor of 10 would further increase the temperature by a factor of 10.

Thus the total temperature would increase by a factor of 50.

18. What is the density of O_2 gas maintained at a pressure of 2 atm and temperature of 600 K?
1. 2.5 g/L
2. 1 g/L
3. 2 g/L
4. 1.3 g/L   Correct
5. 3.7 g/L
Explanation: PV= nRT. Density = mass/volume = n x MW / V = PV/(RT) x MW/ V = P x MW / RT = 2 atm x 32 g/mol / (0.0821 L·atm·mol^{-1}·K^{-1} x 600 K) = 1.3 g/L

19. The reversal of the Haber Process for ammonia synthesis is the reaction below:
2NH3(g)→N2(g) + 3H2(g)
Assuming that it goes to completion starting with 10 moles of ammonia in a 50 L container at standard temperature, what is the final pressure?
1. 10.7 atm
2. 8.5 atm
3. 9.0 atm   Correct
4. 20.5 atm
5. 25.6 atm
Explanation: 10 moles of ammonia would produce 5 moles of nitrogen and 15 moles of hydrogen (n_{tot} = 20 mol). P = (nRT)/V = (20 mol x 0.0821 L·atm·mol^{-1}·K^{-1} x 275 K)/50 L = 9.0 atm.

20. At any given temperature, how much more quickly will N_2 diffuse than Cl_2?
1. 17 times more quickly
2. 0.5 times more quickly
3. 1.6 times more quickly   correct
4. .03 times more quickly
5. Cl_2 will defuse faster than N_2
Explanation: $m_1v_1^2 = m_2v_2^2$
$v_1/v_2 = (m_2/m_1)^{1/2} = (70.91/28)^{1/2} = 1.6$

21. Rank the following gases in terms of increasing non-ideality: O_2, HF, H_2O, Xe.
1. Xe < HF < H_2O< O_2
2. HF < H_2O< O_2 < Xe
3. O_2 < Xe < HF < H_2Ocorrect
4. H_2O< HF< Xe< O_2
5. H_2O< Xe< O_2 <HF
Explanation: O_2 and Xe are both non-polar and thus closest to ideal, but Xe is substantially larger, and so is more non-ideal. HF is polar, and therefore more non-ideal than the non-polar gases. Lastly, H_2O has H-bonding and is highly non-ideal.

22. Which of the following contribute to the non-ideality of gases:
        I.   The presence of intermolecular forces
        II.   Pressure of the gas
        III.   Temperature of the gas
        IV.   Bond length of the gas
1. I and III only
2. I only
3. II and III only
4. I, II, and III   correct
5. All of the above

Explanation: Weak IMF, low pressure, and high temperature all favor ideal behavior, and thus affect whether a gas behaves ideally or not. The bond length of the gas is irrelevant.

23. Which of the following statements regarding intermolecular forces (IMF) is/are true?
    I.   All intermolecular forces are characterized by bonds inside molecules
    II. Intermolecular forces are responsible for solution properties such as surface tension
    III. All molecules have London dispersion forces
    1. III only
    2. I only
    3. II only
    4. II and III   correct
    5. I and III
    6. All of the above
Explanation: Statement I is false: all IMF are characterized by bonds between, not inside, the molecules. Statement II is true: the stronger the IMF, the greater the surface tension. Statement III is true: all molecules have London dispersion forces

24. Which of the following compounds is NOT correctly paired with the strongest IMF that it exhibits?
    I.   HBr, hydrogen bonding   correct
    II. NH_3, hydrogen bonding
    III. CaO, ion-ion
    IV. C_6H_6 (benzene), van der Waals forces
    V. SiH_4, London forces
Explanation: HBr exhibits dipole-dipole interactions.

25. Which of the following compounds is not correctly paired with the strongest IMF that it exhibits?
    1. H_2O, hydrogen bonding
    2. CH_4, London Dispersion Forces
    3.   HF, hydrogen bonding
    4.   HI, dipole-dipole
    5.   NaCl, dipole-dipole Correct
Explanation: NaCl exhibits ion-ion interactions

26. Capillary action is best described as:
    1. The tendency of molecules in the liquid phase to resist motion
    2. The tendency of a liquid to climb the walls of a burette Correct
    3. The phenomenon of an insect walking on water
    4. The likelihood that a molecule in the liquid phase will enter the gas phase
    5. A liquid's resistance to flow
Explanation: Capillary action is the tendency for the liquid to climb up the walls of a capillary, forming the meniscus that you see in a burette or straw.

27. Rank the following compounds in terms of increasing vapor pressure: KCl, CH_3OH, H_2O, $CaCl_2$, CH_4.
    1. $CaCl_2 < KCl < H_2O < CH_3OH < CH_4$ correct
    2. $CH_4 < KCl < CaCl_2 < CH_3OH < H_2O$
    3. $CaCl_2 < H_2O < CH_3OH < KCl < CH_4$
    4. $CaCl_2 < KCl < CH_3OH < H_2O < CH_4$
    5. $CH_4 < CaCl_2 < CH_3OH < KCl < H_2O$
    6. $CH_4 < CH_3OH < H_2O < KCl < CaCl_2$
Explanation: CH_4 has only dispersion forces. H_2O and CH_3OH both have H-bonding, but the former has twice the bonding potential. KCl and $CaCl_2$ are both salts (ion-ion interactions), but the latter has higher charge density. Vapor pressure is inversely proportional to the strength of IMF.

28. Ranking physical properties by IMF

Rank the following liquids in terms of decreasing boiling point: $C_{10}H_{22}$, CH_4, C_5H_{12}, C_2H_6, C_7H_{16}.

1. $C_{10}H_{22} > C_2H_6 > CH_4 < C_7H_{16} > C_5H_{12}$
2. **$C_{10}H_{22} > C_7H_{16} > C_5H_{12} > C_2H_6 > CH_4$**   **correct**
3. $CH_4 > C_2H_6 > C_5H_{12} > C_7H_{16} > C_{10}H_{22}$
4. $CH_4 > C_7H_{16} > C_{10}H_{22} > C_2H_6 > C_5H_{12}$
5. $CH_4 > C_{10}H_{22} > C_2H_6 > C_7H_{16} > C_5H_{12}$

Explanation: All of these molecules are non-polar, and boiling point is directly proportional to IMF, so one simply has to rank them from largest to smallest.

29. Rank the following species in terms of decreasing evaporation rate: NaCl, H_2, HF, HI, BaS.

1. $BaS > NaCl > HF > HI > H_2$
2. $HF > NaCl > BaS > H_2 > HI$
3. **$H_2 > HI > HF > NaCl > BaS$**   **correct**
4. $HF > HI > H_2 > BaS > NaCl$
5. $H_2 > BaS > NaCl > HF > HI$

Explanation: Evaporation rate is inversely proportional to IMF, so ranking by decreasing evaporation rate requires ranking by increasing IMF. Hydrogen gas is non-polar, and thus has very low IMF. Hydroiodic acid and hydrofluoric acid are both polar, but HF has a larger difference in electronegativity and this stronger IMF. Lastly, sodium chloride and barium sulfite have ion-ion interactions, but the later has a higher charge density.

30. Glucose $(C_6H_{12}O_6)$ is an example of what class of solids?

1. metallic solid
2. covalent network
3. **molecular covalent   correct**
4. ionic solid
5. None of the above

Explanation: Glucose molecules interact with each other mainly via dipole-dipole interactions and thus consist of molecular covalent bonds.