1. Classify the bonds in the following compounds as ionic, polar covalent, or non-polar covalent: NH_3, LiF, H_2, respectively.

A.   **Polar covalent, ionic, non-polar covalent   Correct**
B.   Ionic, polar covalent, non-polar covalent
C.   Polar covalent, non-polar covalent, ionic
D.   Ionic, non-polar covalent, polar covalent
E.   Non-polar covalent, ionic, polar covalent

Explanation: Based on the simple rule that $\Delta EN > \sim 1.5$ is ionic, $\sim 1.5 > \Delta EN > 0$ is polar covalent, and $\Delta EN = 0$ is non-polar covalent: The N-H bonds in NH_3 are polar covalent ($\Delta EN = 3.0-2.2 = 0.8$); LiF is ionic ($\Delta EN = 4.0-1.0 = 3.0$); and the H-H bond in H_2 is non-polar covalent ($\Delta EN = 0$).

2. In the Lewis structure for acetone, CH_3COCH_3, all of the following bond angles, hybridizations, or electronic geometries are described by some part of the molecule EXCEPT:

A. 120°
B. sp^3
C.   Tetrahedral
D.   **90° Correct**
E.   Trigonal planar

Explanation: In the Lewis structure for acetone, the central carbon has three electron-dense regions (one double bond to O and two single bonds to C) and thus has a trigonal planar configuration with 120° bond angels. The other two carbons are bonded to 4 atoms each (3 H's and 1 C). Thus their configuration is tetrahedral, with sp^3 hybridization.

3. In the Lewis structure for methylamine, CH_3NH_2, all the following bond angles, hybridizations, or electronic geometries are described by some part of the molecule EXCEPT:

A. sp^3
B.   Tetrahedral
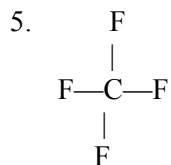C.   **120° Correct**
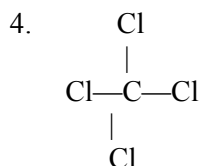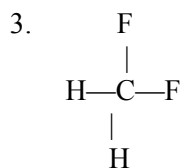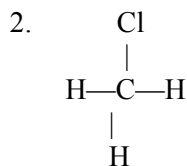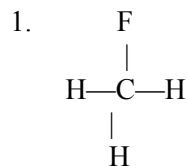D. 109.5^o
E.   None of the above

Explanation: In the Lewis structure for methylamine, the carbon has four electron-dense regions (bonded to 3 H's and 1 N) and thus has a tetrahedral configuration with 109.5^o bond angels. The nitrogen also has 4 electron-dense regions (bonded to 2 H's, 1 C, and has a pair of lone electrons) and thus has a tetrahedral configuration with 109.5^o bond angles.

4. Which of the following best describes the molecular geometry in ozone, O_3?

A.   **Angular   Correct**
B.   Linear
C. 180°
D.   Pyramidal
E.   Both B and C

Explanation: The central oxygen in ozone has two bonds (one single, one double) and a lone pair of electrons and is thus described by AB2U. This gives it the angular geometry.

5. Rank the polarity of the following Lewis structures:

1.
```
      F
      |
  H—C—H
      |
      H
```

2.
```
      Cl
      |
  H—C—H
      |
      H
```

3.
```
      F
      |
  H—C—F
      |
      H
```

4.
```
      Cl
      |
  Cl—C—Cl
      |
      Cl
```

5.
```
      F
      |
  F—C—F
      |
      F
```

1. **IV = V < II < I < III correct**
2. II < I < III < IV < V assumes more polar bonds is more polar molecule
3. IV < V < II < I < III ignores that IV and V are same, zero
4. I = II < III < IV = V ignores magnitude of EN diff
5. V < IV < III < II < I doesn't understand that III is stronger than I and II
6. IV = V < I = II < III ignores

Explanation: Polarity is a measure of the strength of the molecular dipole moment. Molecular dipole moment, in turn, is the vector sum of the electronic dipoles. The direction of the electronic dipole vectors is set by the geometry of the bond in 3D space and its length is proportional to the EN difference between the two atoms involved in the bond. So for tetrachloromethane and tetrafluoromethane, the molecular dipole is zero because the sum of the electronic dipoles is zero (the symmetry causes them to cancel out). So these are the lowest polarity and they are equal. We know that chloromethane will be less polar than fluoromethane because the EN difference between C and F is larger than between C and Cl. However, both are polar so they are higher than tetrachloro- and tetrafluoromethane. Finally, difluoromethane is the most polar because the sum of the two electronic dipoles between C and F will sum to be larger than the single dipoles in the previous two molecules. Note that the symmetry of these two dipoles will not result in their full cancellation, but rather an additive effect.

6. How many sigma and pi bonds do the following molecules have?

I. HC=CH

II. $H_2C=CH_2$

III. H_3C-CH_3

IV. $H_2C=C=CH_2$

1. I. 3,0; II. 2,0; III. 1, 0; IV. 4,0
2. I. 1,2; II. 1,1; III. 1, 0; IV. 2,2
3. I. 4,1; II. 5,1; III. 7, 0; IV. 6,2
4. I. 2,3; II. 4,2; III. 7, 0; IV. 4,4
5. **I. 3,2; II. 5,1; III. 7, 0; IV. 6,2   correct**

Explanation: For molecule I, there are two C—H sigma bonds, one C—C sigma bond, and one carbon to carbon pi bond. For molecule II, there are four C—H sigma bonds, one C—C sigma bond, and one carbon to carbon pi bond. For molecule III, there are six C—H sigma bonds and one C—C sigma bond. For molecule IV, there are four C—H sigma bonds, two C—C sigma bonds, and two carbon to carbon pi bonds.

7. Which of the following are true about hybrid orbitals?

I. They are used because atomic orbitals were not a good model for molecular bonding.
II. They are the result of mixing atomic orbitals of various types.
III. They always include at least one of each orbital type (s,p, and d).
IV. The use of a hybridized orbital model better predicts molecular orbital energy, bonding patterns, as well as molecular shape.
V. Unlike our previous models, which were only approximations of reality, hybridization really does reflect the absolute, empirical truth.

1. I, IV, V
2. I only
3. II only
4. I, II
5. I, III
6. **I, II, IV   correct**
7. I, II, V

Explanation: Statement I is true. Atomic orbitals lead to several incorrect predictions – in your course notes, we mention that the energies of the molecular orbitals would be incorrectly predicted. Statement II is true. All hybrid orbitals are made up of a mix of s and p orbitals (some also have some d orbitals mixed in). Statement III is false. Not all hybrid orbitals include d orbitals mixed in. Statement IV is true. Hybridization correctly predicts, for instance, that all the C—H bonds in methane are the same energy. It also predicts that methane, not carbene, is formed. It also correctly predicts bond angles. Atomic orbital models predict all of these incorrectly. Statement V is false. Hybridization is just another approximation we make. It works better than atomic orbitals for the types of thing we are talking about, but it is just another approximation. But don't feel bad, it's easier to learn than the truth, just ask a quantum physicist.

8. Here's a wacky molecule you might learn about later in organic chemistry or biochemistry, called isoprene. It's a building block for rubbers:

```
      CH₃
       |
       C     CH₂
      // \ //
 H₂C     C
          |
          H
```
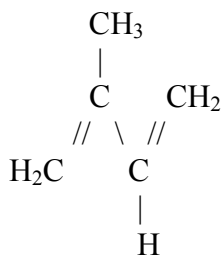
How many of the following orbital types do you spy involved with sigma and pi bonding: s, p, sp, sp^2, sp^3 ?

1. Eight s, two sp, four sp^2, one sp^3 for sigma and zero p for pi.
2. **Eight s, zero sp, twelve sp^2, four sp^3 for sigma and four p for pi.   correct**
3. Eight s, zero sp, four sp^2, one sp^3 for sigma and two p for pi.
4. Four s, zero sp, twelve sp^2, four sp^3 for sigma and four p for pi.
5. Eight s, zero sp, twelve p, four sp^3 for sigma and four sp^2 for pi.

Explanation: The eight hydrogen each bond using their 1s orbitals. There are no atoms with only two regions of electron density so there are no sp orbitals. Every carbon involved with a double bond is using three sp^2 for a total of twelve and one p orbital for the pi bonds for a total of four. The topmost carbon has four regions of electron density so it must use four sp^3 orbitals.