1. Give the electronic configuration, valence configuration, and the number of valence electrons for the following:
   a. Bi
   b. N\(^3\)-
   c. Al\(^3\)+
   d. AlPO\(_4\)

   a) [Xe] 6s\(^2\) 4f\(^{14}\) 5d\(^{10}\) 6p\(^3\)
   b) 1s\(^2\) 2s\(^2\) 2p\(^6\)
   c) 1s\(^2\) 2s\(^2\) 2p\(^6\)

2. Draw a Lewis structure for the following giving the polarity, bond angels, hybrid orbitals, and the geometrical shape.
   a. Cl\(_3\)\(^+\)
   b. XeOF\(_2\)
   c. PF\(_4\)
   d. XeO\(_3\)F\(_2\)

3. In which of the following diatomic molecules would the bond strength be expected to weaken as an electron is remover to form the positive charged ion?
   a. H\(_2\) \(\text{YES}\)
   b. C\(_2\)\(^2-\) \(\text{YES}\)
   c. OF \(\text{NO}\)  \(\text{NO}\)

4. Use the MO theory to predict the number of bonds and the magnetic properties for the following:
   a. NO\(^-\)
   b. F\(_2\)\(^-\)
   c. B\(_2\)\(^+\)

   a) B.O. = \(\frac{8 - 5}{2} = 1.5\)
   b) B.O. = \(\frac{8 - 4}{2} = 2\)
   c) B.O. = \(\frac{3 - 2}{2} = 0.5\)
5. Describe the bonding in the O₃ molecule and the NO₂⁻ ion using the localized electron model. How would the molecular orbital model describe the pi bonding in these two species?

6. Using the p-orbitals below show how two of them combine to form: a sigma bonding molecular orbital, a sigma anti-bonding molecular orbital, a pi bonding molecular orbital, and a pi anti-bonding molecular orbital. Be sure to include phases (+, -) in your description. Be reminded that sigma bonds are head to head overlap of p-orbitals, and pi bonds are sideways overlap of p-orbitals.

7. Draw the most stable resonance form of sulfuric acid (H₂SO₄).