MH1_Practice Problem solving strategy
Practice Mid-Summer 2020

1. **Write Name/Formula of the following compounds**:

Examples for solving problems

P_4O_{10} ___________________________________

Hg_2CO_3 ___________________________________

Naming compounds: This exam has the naming compounds for you to recognize what type of compound a given formula represents – ionic or covalent. It's important to be able to classify them correctly because they require different naming schemes to be applied. Ionic formulas start with a metal or the cationic ion Covalent formulas star with a nonmetal or metalloid element. The template I've given for **ionic compounds** is:

Cation name (space) Anion name

The template for **covalent compounds** is:

(prefix) element name (space) (prefix) element name with "-ide" ending

On the practice quiz, the **firstformula** is **P_4O_{10}**. This should be identified as covalent since P is a nonmetal. We need multiplying prefixes on both parts. It should be tetraphosphorus decoxide. Remind them that a prefix ending in "a" such as deca- has the "a" dropped when used with oxide. The **second formula** is **Hg_2CO_3**. This should be identified as ionic compound since lead is a metal. Since the anion is a polyatomic, CO_3^{2-}, that only appears once it indicates that we're dealing with mercury(I). Thus, the name of the compound should be mercury(I) or mercurous carbonate.

2. **Complete the Lewis structures:**

Examples for solving problems

Complete the Lewis structure for the simple organic molecule (the amino acidglycine) whose skeletal structure is shown below. Add multiple bonds and lone pairs as needed **BUT NO ADDITIONAL ATOMS**. Then survey the sigma, pi and lone pairs. Place the number of each in the blanks supplied.

```
       H      H      O1
       |      |      /
H  --  N  --  C1 --  C2 --  O2 --  H
              |
              H
```

In this structure there are: _____ sigma bonds, _____ pi bonds and ____ lone pairs.

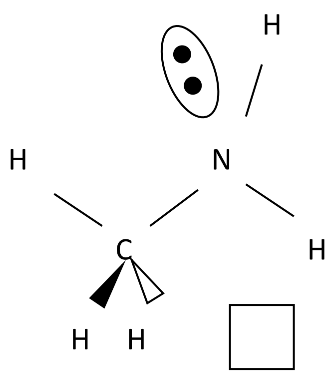
The bond angles at C_1 are _____^{o} and the bond angles at C_2 are ______^{o}.

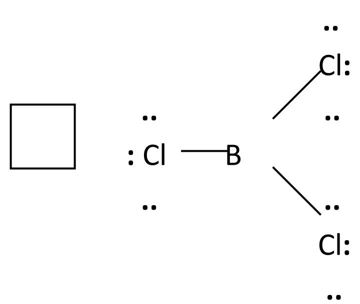
Completing a Lewis Structure and Inventorying Its Parts. This problem asks the you to recognize the typical bonding pattern in H, C, N, O, S, P and halogens for simple organic compounds. The guidelines for this are in your book and in the lecture slides. You should simply look at each atom in the order encountered and decide if it's okay or needs something more. Like in glycine (an amino acid) chemical formula NH_2-CH_2-COOH. Here, N – needs a lone pair, C_1 – okay; it has four bonds, C_2 – Needs one more bond. So does the O_1-atom above it. Make it a double bond to the upper O_1 and put two lone pairs on the O_1 and O_2 – needs two lone pairs.

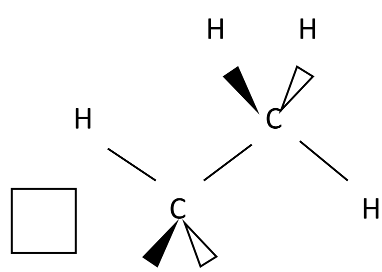
Inventorying the parts. The first bond between two atoms are all sigma bonds. There is a total of 9 of them here. Any multiple bonds that have been drawn are pi bonds. The added bond in the C=O contains one pi bond. Finally, there are 5 lone pairs – one on the N and two each on the O atoms. **Bond angles.** I've taught you two atoms around which to determine the bond angles. C_1 has 4 regions each with a bond so it's tetrahedral. C_2 has 3 regions each with a bond so its trigonal planar. These are VERY COMMON bond angles for carbon, so I thought it would be good to reinforce that here. For N, it's 4 regions with 3 bonds -- triangular pyramid. O_2 is 4 regions with 2 bonds – bent.

3. **Categorize each molecule below by writing P (polar) or N (nonpolar) in the boxes provided.**

Examples for solving problems







Polar or Nonpolar?

This exercise asks you to look at a molecular formula and decide if it's polar or nonpolar. You should first ask, "Are there any **FONCl** "phone call" atoms? If so, the next question should be, "Is the molecule totally symmetrical?" A molecule is polar if it has FON Cl atoms and is asymmetric. A molecule is nonpolar if it is symmetric, despite having some **FONCl** atoms present. I've also told you that pure hydrocarbons are nonpolar. I've included three molecules on the practice problems. The oxygen in methanol and chlorine in BCl_3 should easily be seen as asymmetric arrangements leading to polar molecules. The ethane is all "black and white", an indication of its pure hydrocarbon nature. It's nonpolar.

4. **Determine the shape and its polarity of the inorganic compound whose skeletal structure is shown below?**

Examples for solving problems

__________ e^{-}'s

Now complete the Lewis structure.

F ---Xe --- F

What shape does this species exhibit?

Regions _________ Regions with bonds _________

Shape __________________________________

If the electronegativity of fluorine is 4.0 and that for Xe is 2.6, what kind of bond forms between F and Xe? Use a Δ e. n. calculation to justify your answer.

Δ e. n. = __________________

ionic polar covalent nonpolar covalent

A General Lewis Structure

The Lewis structure for a more complicated molecule that needs the general rules for Lewis structures. Note that there are no guidelines for bonding orders in noble gasses.

Step 1 The first step in applying the general rules is to determine how many valence electrons should appear in your finished Lewis structure. If you don't do this step, you'll probably guess the wrong number! For the given molecule, the total valence electrons are found by summing the group numbers for each atom present.

Xe – Group 8A – 8 valence electrons
F – Group 7A – 7 valence electrons each for a total of 2 x 7 e^{-} = 14 e^{-}
No species charge here, it's a neutral molecule
Total e^{-}'s = 8 + 14 = 22 valence e^{-}'s

Step 2 is connecting the atoms in the skeletal structure. I've done this for them here. In another setting I might hint at this by saying that in XeF_2, xenon difluoride, Xe is the central atom.
Step 3 is to place lone pairs around non-H terminal atoms to satisfy octets. (H, if present, does not need anything more than a sigma bond for it to have $2e^{-}$, the noble gas configuration of helium.)
Now see how you're doing with respect to the total valence e^{-}'s available. You've got an octet around each F -atom for a total of 16 e^{-}'s. There are extras left, since you need to show a total of 22 e^{-}'s so go to Step 4.

Step 4 says to place the remaining e^{-}'s on the central atom as long as that's allowed. Expanded octets are available for elements of row 3 or higher (Xe is in row 5) and in Group 5A or higher for a 10-e^{-} central atom (Xe is in Group 8A). So we're okay with three lone pairs around it.

F --- Xe --- F

Finally, the students need to use the VSEPR concept to determine the overall molecular geometry. This molecule has a central atom with 5 regions of density (two single bonds and three lone pairs). Two of the regions are occupied by bonds, so it's a linear shape.

The next step is to analyze the polarity of the bonds in XeF_2. You do this with a Δ en calculation using the electronegativities that I've supplied. This appears as:

$$\underset{F}{4.0} \quad - \quad \underset{Xe}{2.6} \quad = \quad 1.4$$

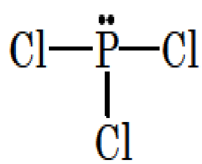
A difference of 1.4 means that the Xe-F bonds are classified as polar covalent.

5. **Predict the geometries of these species using hybridization and VSEPR method**

Examples for solving problems

PCl_3

The Lewis structure of PCl_3 is shown below. Since in the VSEPR method the number of bonding pairs and lone pairs of electrons around the *central atom* (phosphorus, in this case) is important in determining the structure, the lone pairs of electrons around the chlorine atoms have been omitted for simplicity. There are three bonds and one lone electron pair around the central atom, phosphorus, which makes this an AB_3E case. The information in Table ! shows that the structure is a trigonal pyramid like ammonia.



What would be the structure of the molecule if there were no lone pairs and only three bonds?