Bonding Set I

1. NF₃ and PF₅ are stable molecules. Write the Lewis electron-dot formulas for these molecules. On the basis of structural and bonding considerations, account for the fact that NF₃ and PF₅ are stable molecules but NF₅ does not exist.

2. (a) Draw the Lewis electron-dot structures for CO₃²⁻, CO₂, and CO, including resonance structures where appropriate.
(b) Which of the three species has the shortest C-O bond length? Explain the reason for your answer.
(c) Predict the molecular shapes for the three species. Explain how you arrived at your predictions.

3. CF₄ XeF₄ ClF₃
   (a) Draw a Lewis electron-dot structure for each of the molecules above and identify the shape of each.
   (b) Use the valence shell electron-pair repulsion (VSEPR) model to explain the geometry of each of these molecules.

4. Answer the following questions about the structures of ions that contain only sulfur and fluorine.
   (a) The compounds SF₄ and BF₃ react to form an ionic compound according to the following equation.
      \[ \text{SF}_4 + \text{BF}_3 \rightarrow \text{SF}_3\text{BF}_4 \]
      (i) Draw a complete Lewis structure for the SF₃⁺ cation in SF₃BF₄.
      (ii) Identify the type of hybridization exhibited by sulfur in the SF₃⁺ cation.
      (iii) Identify the geometry of the SF₃⁺ cation that is consistent with the Lewis structure drawn in part (a)(i).
      (iv) Predict whether the F—S—F bond angle in the SF₃⁺ cation is larger than, equal to, or smaller than 109.50°. Justify your answer.
   (b) The compounds SF₄ and CsF react to form an ionic compound according to the following equation.
      \[ \text{SF}_4 + \text{CsF} \rightarrow \text{CsSF}_5 \]
      (i) Draw a complete Lewis structure for the SF₅⁻ anion in CsSF₅.
      (ii) Identify the type of hybridization exhibited by sulfur in the SF₅⁻ anion.
      (iii) Identify the geometry of the SF₅⁻ anion that is consistent with the Lewis structure drawn in part (b)(i).
      (iv) Identify the oxidation number of sulfur in the compound CsSF₅.
1. | NF₃ | PF₅ |
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<td><img src="image1" alt="Lewis structure of NF₃" /></td>
<td><img src="image2" alt="Lewis structure of PF₅" /></td>
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Your Lewis structures should show the sp³ hybridization in NF₃ and the sp³d hybridization in PF₅. The reason NF₃ doesn’t exist is because N has no d orbitals available for hybridization.

2. (a) ![Diagram of CO₂](image3)  ::O=\text{C}=O::  :\text{C}=\text{O}:

(b) CO has the shortest bond length because of the triple bond.
(c) CO₂⁻ is trigonal planar because of the sp² hybridization, which puts three sigma bonds in a plane with lobes of an unhybridized p orbital above and below that plane, forcing the planar shape. CO₂ has two double bonds, forcing it to be linear (because of the perpendicular pi bonds). CO has a triple bond, which is also linear for the same reason as CO₂.

3. (a) ![Diagram of CF₄](image4)  ![Diagram of XeF₄](image5)  ![Diagram of ClF₃](image6)  

(b) CF₄ has 4 shared, no unshared pairs. The maximum angular separation gives a tetrahedron.
XeF₄ has two unshared pairs on opposite sides of the Xe. This forces the remaining F atoms into a central plane.
CIF$_3$ is a modified trigonal bipyramid; the two unshared pairs go in equatorial positions leaving the T shape.

4. (a) (i) \[
\begin{array}{c}
\vdots \vdots \vdots \\
\vdots & S & \vdots \\
\vdots & \vdots & \vdots \\
\vdots \vdots \vdots 
\end{array}
\]

(ii) sp$^3$.
(iii) pyramidal
(iv) smaller than 109.5°. The lone pair of unbonded electrons occupies more space than the bonded pairs and, therefore, pushed the bonded pairs away and hence, a smaller bond angle than a perfect tetrahedron.

(b) (i) 

(ii) sp$^3$d$^2$
(iii) square pyramidal
(iv) S = +4