following compounds. Use VSEPR theory to predict the molecular shape.

1. PH_3
2. $SnCl_4$
3. $SnCl_2$
4. O_3
5. N_3^-
6. CO_2
7. SO_2
8. HCN
9. FNO
10. H_2CO

III. Predict the orbital hybridization of the central atom in each of the molecules described in Sections I and II.

IV. Write the notation for the electronic configuration of each of the following homopolar, diatomic molecules and ions. Determine the bond order of each and determine whether the molecule or ion is paramagnetic or diamagnetic. Also draw the Lewis structure of each.

1. NO 2. CN^- 3. O_2^+

V. Predict the geometric configuration of a molecule of each of the following compounds. Describe the bonding between all atoms in the molecule. Remember that the molecular shapes will only be approximations based on the theories with which you are now familiar.

1. H_2CO, formaldehyde
2. C_6H_6, a six-carbon ring compound, benzene
3. CCl_4, carbon tetrachloride
4. HCCH, acetylene
5. SO_2, sulfur dioxide
6. FNNF, dinitrogen difluoride
7. HF, hydrogen fluoride

VI. From the following list of molecules and atoms, described by their Lewis structures, choose the ones that answer the questions. A molecule or atom may be used to answer more than one question.

EXERCISES

I. Use the concept of electron pair repulsions, VSEPR, to predict the geometric configuration of a molecule of each of the following compounds:

1. XeF_4, xenon tetrafluoride
2. AsF_5, arsenic pentafluoride
3. SnF_4, tin tetrafluoride or tin(IV) fluoride
4. IF_2^+, the iodine difluoride cation
5. ClF_3, chlorine trifluoride
6. H_3O^+, the hydronium ion
7. I_3^-, the triiodide ion
8. $SnBr_6^{2-}$, the tin(IV) hexabromide anion
9. SF_4, sulfur tetrafluoride

How to Diagram Lewis Structures

1. Find the total number of valence electrons supplied by all the atoms in the structure. The number supplied by each A family element is the same as the group number of the element:

   a. For a negative ion, increase the number by the charge of the ion.

   b. For a positive ion, decrease the number by the charge of the ion.

2. Determine the number of electrons that would be required to give two electrons to each H atom individually and eight electrons to each of the other atoms individually:

$$\textit{num. } e^- \textit{ for individual atoms} = 2(\textit{num. H atoms}) + 8(\textit{num. other atoms})$$

3. The number obtained in step 2 minus the number obtained in step 1 is the number of electrons that must be shared in the final structure:

$$\textit{num. bonding } e^- = (\textit{num. } e^- \textit{ for individual atoms}) - (\textit{total num. } e^-)$$

4. One-half the number of bonding electrons (step 3) is the number of covalent bonds in the final structure:

$$\textit{num. } e^- \textit{ pair bonds} = (\textit{num. bonding } e^-)/2$$

5. Write the symbols for the atoms present in the structure, arranging them in the way that they are found in the structure.

6. Indicate electron-pair bonds by dashes written between the symbols. Indicate one bond between each pair of symbols, and then use any remaining from the total calculated in step 4 to make multiple bonds. Note that each H atom is limited to one bond.

7. The total number of electrons (step 1) minus the number of bonding *electrons* (step 3) is the number of unshared electrons:

$$\textit{num. unshared } e^- = (\textit{total num. } e^-) - (\textit{num. bonding } e^-)$$

   Complete the electron octet of each atom (other than the H atoms) by adding dots to represent unshared electrons.

8. Indicate the formal charges of the atoms where appropriate, and evaluate the structure.