




NAME: Khang Tarr Lab Section: _____ Score: _____**MOLECULAR MODELS Part II**

In this experiment, you will use the Lewis structures that you draw to build models and explore the concepts of molecular shape, hybridization and bonding schemes and polarity. Show all $[6N + 2] - [VE]$ work to aid in drawing Lewis Structures.

Section III: sp^3d^2 Hybridization:

Draw the Lewis structures on the table below. Include their VSEPR Structures and hybridization of the central atom. How many holes will the central atom require to accommodate the hybridized AND unhybridized p orbitals? 6 Build the indicated models and sketch a 3-D representation of them in the table.

Molecule	Lewis Structure	Model	VSEPR structure
SF ₆	$[6N + 2] - [VE] = 42 - 48 = -6$ <p>(or $1 \text{ orbit } (2e^- \text{ at } 1 \text{ orbital } \times 6 \text{ orbit})$)</p> 		

- 1) Would you expect sulfur hexafluoride to be a polar molecule? Explain why or why not. If so, indicate its molecular dipole in the table.

$$\Delta EN_{S-F} = |2.58 - 3.98| = 1.4 \Rightarrow \text{polar bond}$$

SF₆ is a nonpolar molecule because it has a perfect symmetry structure, so the bond cancel out.

- 2) The structure of bromine pentafluoride is different from that of sulfur hexafluoride in that it has a lone pair. Because all the bond angles on an octahedral center are 90 degrees, there is no axial or equatorial position. Does it matter where you put the lone pair? Why or why not?

I do not think the position of the lone pair ~~for~~ form BrF₅ is crucial because all the ^{bond} angles on an octahedral center is 90°, and it is the same with square pyramidal center. Therefore, wherever we put the lone pair, the bond angle of the lone pair is on the center is 90°.

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