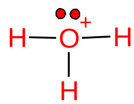
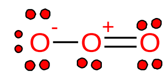
Name ______Mr. Perfect_____________________________ Date _____Sp 09__________

1. Draw Lewis dot structures for the following molecules and include any formal charges on the atoms. (20 pts)

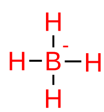
a) H_3O^+



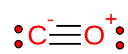
b) O_3



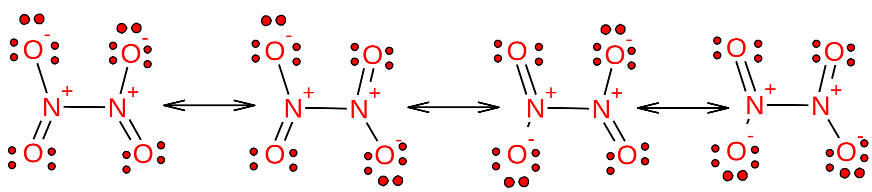
c) BH_4^-



d) CO

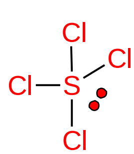


2. Draw the possible resonance structures for N_2O_4 and include any formal charges on the atoms. (10 pts)



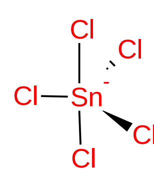
3. Using VSEPR theory, draw the following molecules in their correct molecular geometry and name the molecular geometry. (10 pts)

a) SCl_4



See-Saw

b) $SnCl_5^-$



Trigonal bipyramidal

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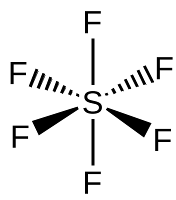
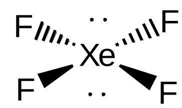
4. Give the expected bond angles for the following molecules and predict what hybrid orbitals are expected for each: (10 pts)

a) Bond angles 90°



sp^3d^2

b) Bond angles 90°



sp^3d^2

5. Write both the electron configurations and the orbital diagrams for the following elements in their ground states. Also, state if the element is paramagnetic or diamagnetic. (10 pts)

a) $_{18}$Ar

[Ar]

[Ar] diamagnetic

b) $_{12}$Mg

[Ne] $3s^2$

[Ne] [↑↓] diamagnetic
3s

c) $_{28}$Ni

[Ar] $4s^2 3d^8$

[Ar] [↑↓] [↑↓][↑↓][↑↓][↑][↑] paramagnetic
4s 3d

d) $_{13}$Al

[Ne] 3s23p1

[Ne] [↑↓] [↑][ ][ ] paramagnetic
3s 3p

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6. If a molecule is considered "unstable" with a bond order equal to zero or less, then it is assumed that the molecule does not exist. Using molecular orbital theory, determine if the following molecules are expected to exist. Draw the molecular orbital diagrams for each. (20 pts)
{The filling order for the first 2 periods are as follows: $\sigma_{1s}\sigma^*_{1s}\sigma_{2s}\sigma^*_{2s}\pi_{2p}\sigma_{2p}\pi^*_{2p}\sigma^*_{2p}$}

a) H_2^+

_____ σ_{1s}^*

1s _↑_   ___ 1s

_↑_ σ_{1s}

H atom H_2 molecule   H atom

B.O. = ½ (1-0) = ½
exists

b) He_2

_↑↓_ σ_{1s}^*

1s _↑↓_   _↑↓_ 1s

_↑↓_ σ_1

He atom He_2 molecule   He atom

B.O. = ½ (2-2) = 0
does not exist

c) B_2

____ σ_{2p}^*

____ ____ π_{2p*}

_↑_ ____ ____   _↑_ ____ ____
2p   2p

____ σ_{2p}

_↑_ _↑_ π_{2p}

_↑↓_ σ_{2s}^*

2s _↑↓_   _↑↓_ 2s

B atom   _↑↓_ σ_{2s} B atom

B_2 molecule

B.O. = ½ (4-2) = 1
exists

d) Be_2

_↑↓_ σ_{2s}^*

2s _↑↓_   _↑↓_ 2s

Be atom   _↑↓_ σ_{2s} Be atom

Be_2 molecule

B.O. = ½ (2-2) = 0
does not exist

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7. Use partial orbital diagrams to show how the valence shell atomic orbitals of the central atom of the following molecules combine to form a hybrid orbital "Valence Bond Theory". (*Hint: Use VSEPR to predict the hybridization of the central atom.*) (10 pts)

a) H_2O

O atom [↑↓] 2s [↑↓][↑][↑] 2p

↓

[↑↓][↑↓][↑][↑] sp^3

O will now accept an e for each H.

b) BF_3

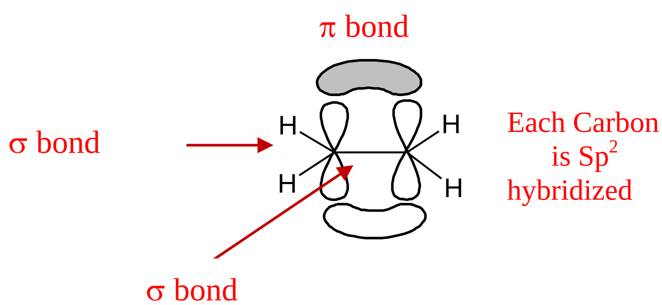
B atom [↑↓] 2s [↑][_][_] 2p

↓

[↑][↑][↑] sp^2 [_] p empty p-orbital

B will now accept an e for each H.

8. Draw C_2H_4 showing the hybrid orbitals of carbon involved in bonding. Show all of the σ bonds and π bonds. What is the hybridization of carbon in this molecule? (10 pts)



9. Extra Credit (5 pts)
Consider the following bond lengths:

C-O 143 pm C=O 123 pm C≡O 109 pm

Explain why in the CO_3^{2-} ion, all three Carbon-Oxygen bonds have identical bond lengths of 136 pm?

The bond length is an average of two single bonds plus a double bond due to the following resonance structures.