Chemistry 1B, Exam III
April 22, 2009
Professor W.A. Lester, Jr.

Name KEY

TA

1.	(10)	
2.	(10)	
3.	(10)	
4.	(20)	
5.	(10)	
6.	(20)	
7.	(10)	
8.	(10)	

TOTAL (100) __________

TOTAL EXAM SCORE ( ) __________

Rules:
- Work all problems to 3 significant figures
- No lecture notes or books permitted
- No word processing calculators
- Time: 50 minutes
- Show all work to get partial credit

1. (10 points) A neon discharge, in addition to its many red photons, also emits photons of wavelength 73.6 nm. When these photons shine on a certain metal-oxide surface, photoelectrons of kinetic energy 10.1 eV are emitted. What is the work function in eV of the surface?

$$E_{TOT} = h\nu = \frac{hc}{\lambda}$$

$$= \frac{(6.626 \times 10^{-34}\ J \cdot s)(3.00 \times 10^{8}\ m/s)}{(73.6 \times 10^{-9}\ m)} \times \frac{1 eV}{1.602 \times 10^{-19}\ J}$$

$$= 16.9\ eV$$

$$E_{TOT} = \Phi + KE$$

$$\Phi = E_{TOT} - KE = 16.9\ eV - 10.1\ eV = 6.8(0)\ eV$$

2. (10 points) The spectrum of atomic hydrogen contains many lines, which appear in groups with such names as the Balmer series and Lyman series. The Humphreys series is one such group, and all of its transitions terminate at the $n = 6$ level. What is the energy in eV of a photon that results from the longest wavelength transition in this series?

longest λ → smallest ΔE

$n=7 \rightarrow n=6$ transition

$$\Delta E = E_{n=7} - E_{n=6}$$

$$= (-2.18 \times 10^{-18}\ J)\left(\frac{1}{7^2} - \frac{1}{6^2}\right) \cdot \frac{1 eV}{1.602 \times 10^{-19}\ J}$$

$$= 0.100\ eV$$

3. (10 points) Arrange the following elements in order of *increasing* ionization energy: Sr, Cs, C, N, O.

Cs , Sr , C , O , N

4. (20 points) The concept of a double bond between carbon atoms, represented by C=C, has a length near 1.34 Å. The motion of an electron in such a bond can be treated crudely as motion in a one-dimensional box.

(a) (6 points) Calculate the energy of an electron in each of its three lowest allowed states if it is confined to move in a one-dimensional box of length 1.34 Å.

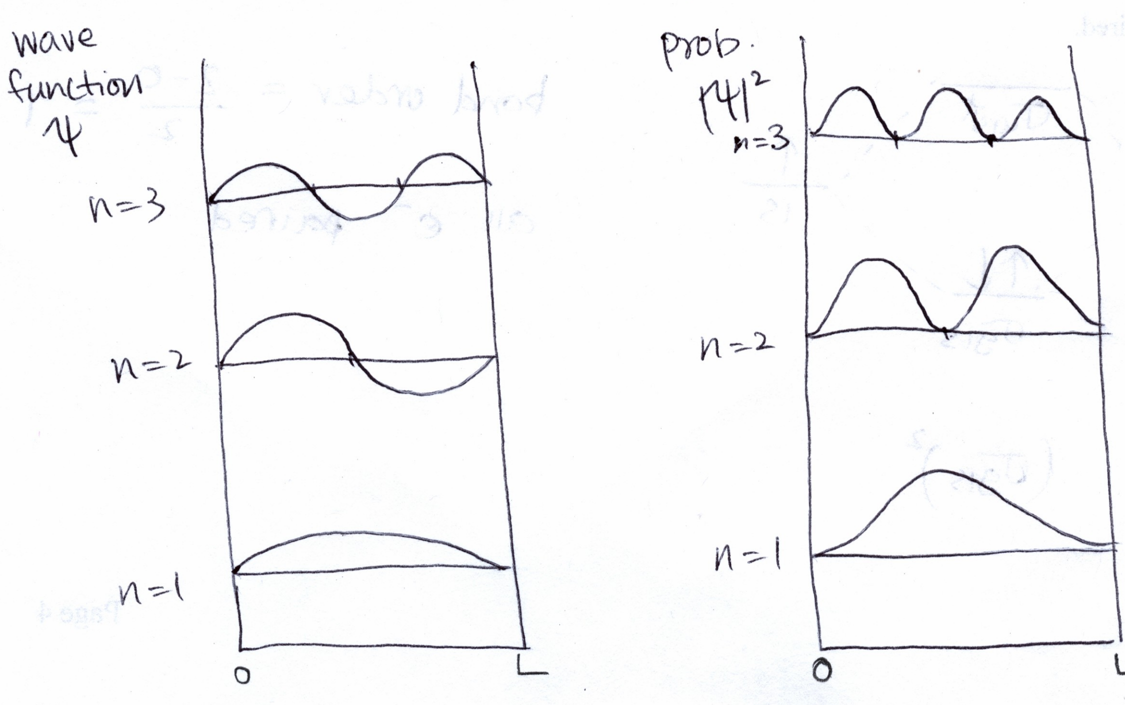
$$E_n = \frac{n^2 h^2}{8 m_e L^2} , \quad L = 1.34 \times 10^{-10} \text{ m}.$$

$$E_1 = \frac{(1)^2 (6.626 \times 10^{-34} \text{ J s})^2}{8 \cdot (9.109 \times 10^{-31} \text{ kg})(1.34 \times 10^{-10} \text{ m})^2} = 3.36 \times 10^{-18} \text{ J} \quad (21.0 \text{ eV})$$

$$E_2 = 4 E_1 = 1.34 \times 10^{-17} \text{ J} \quad (83.9 \text{ eV})$$

$$E_3 = 9 E_1 = 3.02 \times 10^{-17} \text{ J} \quad (189 \text{ eV})$$

(b) (6 points) Sketch the wave functions and the probabilities for finding the electron in the box for each of the three states. Label clearly.



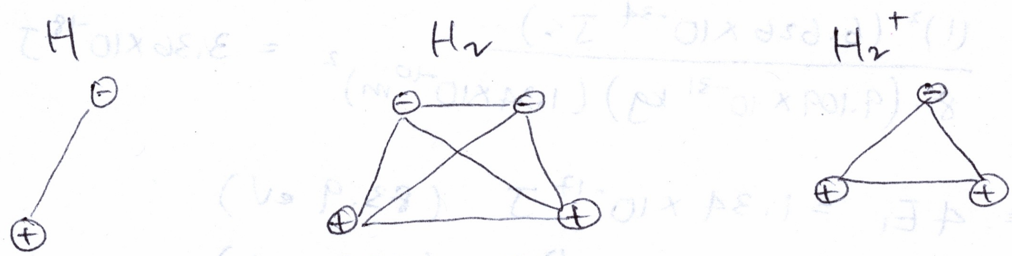
(c) (8 points) Calculate the wavelength of light necessary to excite the electron from its ground state to the first excited state.

$\Delta E = E_2 - E_1 = 1.34\times10^{-17} - 3.36\times10^{-18} = 1.00\times10^{-17}\ J$

$\Delta E = \frac{hc}{\lambda}$

$\lambda = \frac{hc}{\Delta E} = \frac{(6.626\times10^{-34})(3.00\times10^{8})}{1.00\times10^{-17}} = 1.99\times10^{-8}\ m$ (19.9 nm)

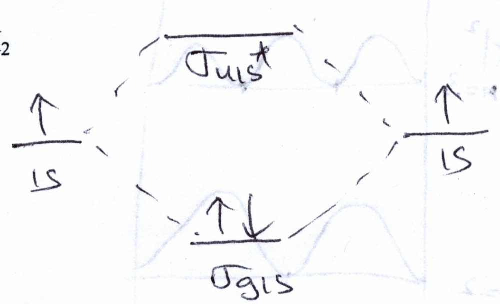
5. (10 points) Give the number of Coulomb potential energy terms in H, H_2, and H_2^+, respectively.



terms 1 6 3

6. (20 points) For each of the following write down the lowest energy electron configurations based on the H_2^+ molecular orbitals and indicate the bonding expected for each. Show whether the electrons are paired or unpaired.

(a) H_2

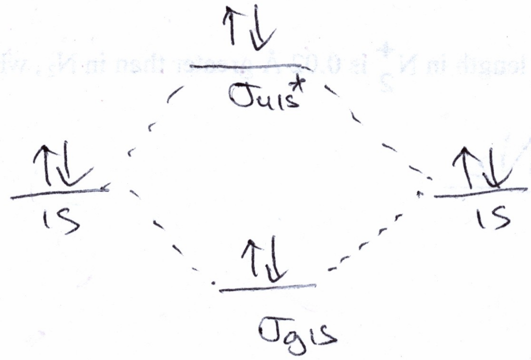


bond order $= \frac{2-0}{2} = 1$

all e⁻ paired

$(\sigma_{g1s})^2$

(b) He_2

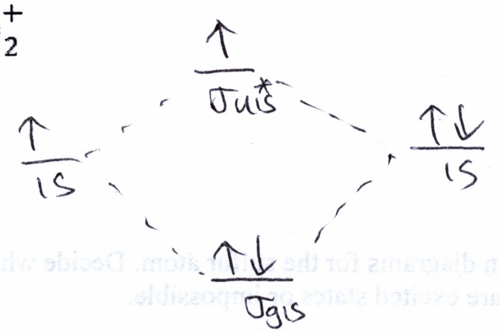


$B.O = \frac{2-2}{2} = 0$

all e^- paired

$(\sigma_{g1s})^2 (\sigma_{u1s}^*)^2$

(c) He_2^+

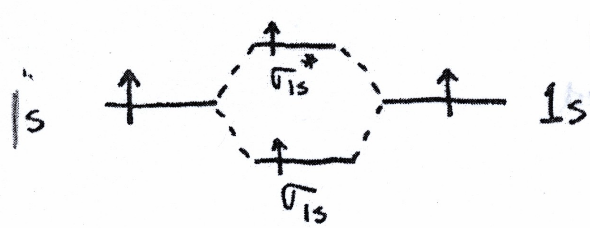


$B.O = \frac{2-1}{2} = 0.5$

1 unpaired e^-

$(\sigma_{g1s})^2 (\sigma_{u1s}^*)^1$

(d) If two H-atoms come together and they both have electrons with $m_s = +1/2$, they do not form a chemical bond. Explain this using molecular orbitals.

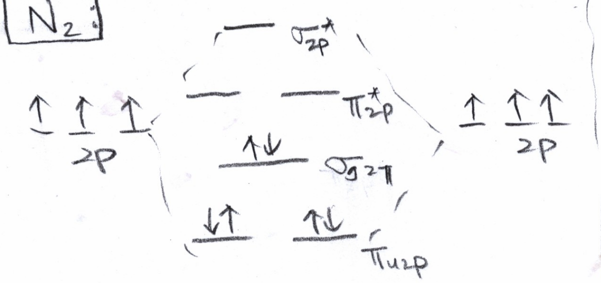


2 electrons with the same $m_s = \frac{1}{2}$ can not occupy the same orbital (Pauli exclusion), therefore they will each occupy σ_{1s} (bonding) and σ_{1s}^* (antibonding) orbitals as shown in the diagram.

$\Rightarrow B.O = \frac{1-1}{2} = 0 \Rightarrow$ no bond

7. (10 points) Explain the observations that the bond length in N_2^+ is 0.02 Å greater than in N_2, while the bond length in NO^+ is 0.09 Å less than in NO.

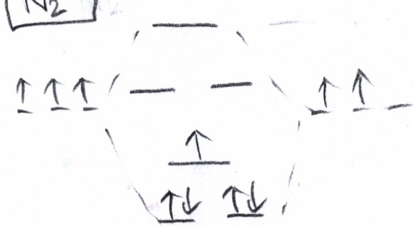
N_2:



$(\pi_{u2p})^4 (\sigma_{g2p})^2$

B.O = 3

N_2^+



$(\pi_{2p})^4 (\sigma_{2p})^1$

B.O = 2.5

∴ Larger bond order → stronger bond → shorter length

$B.O_{N_2} > B.O_{N_2^+}$

∴ N_2^+ → longer bond length

Similarly,

NO^+ $(\pi_{2p})^4 (\sigma_{2p})^2$ B.O = 3

NO : $(\pi_{2p})^4 (\sigma_{2p})^2 (\pi^*_{2p})^1$ B.O = 2.5

$B.O_{NO^+} > B.O_{NO}$

∴ NO → longer bond length

8. (10 points) The following are valence orbital occupation diagrams for the sulfur atom. Decide which of them is the ground configuration, and whether the others are excited states or impossible.

	3s	3p			4s	
a.	↑↓	↑	↓	↑↓	—	→ excited state
b.	↑↓	↑↓	↑	↑	—	→ ground state
c.	↑↓	↑↑	↑	↑	—	→ impossible
d.	↑↓	↑↓	↑↓	—	—	→ excited
e.	↑↓	↑	↑	↑	↑	→ excited