Chem 245
Practice Midterm Exam

Name:___________________________

I.D.______________________

Answer all FIVE of the following questions below in the space provided. Some equations, physical constants, and a Periodic Table are included at the back. All questions are of equal value.

1. (a) What are the values of the quantum numbers l and n for a 5d electron?

(b) At most, how many 4d electrons can an atom have? Of these electrons how many, at most, can have m_s = -½?

(c) A 5f electron has what value of quantum number l? What values of m_l may it have?

(d) What values of the quantum number m_l are possible for a subshell having l = 4?

(e) Calculate the effective nuclear charges experienced by a 5p and a 4d electron in a tin (Sn) atom (Z = 50).

2. (a) The second ionization of carbon ($C^+ \rightarrow C^{2+} + e^-$) and the first ionization of boron ($B \rightarrow B^+ + e^-$) can both be described by the reaction $1s^2 2s^2 2p^1 \rightarrow 1s^2 2s^2 + e^-$. Compare the two ionization energies (24.383 eV and 8.298 eV, respectively) and the effective nuclear charges Z*. Is this an adequate explanation of the difference in ionization energies? If not, suggest other factors.

(b) Select your choice by circling it:

(i) Smallest radius: Sc Ti V

(ii) Greatest volume: S^{2-} Ar Ca^{2+}

(iii) Lowest ionization energy: K Rb Cs

(iv) Highest electron affinity: Cl Br I

(v) Most energy necessary to remove an electron: Cu Cu^+ Cu^{2+}

3. The isomeric ions NSO^- (thiazate) and SNO^- (thionitrite) have been reported. Draw the most important Lewis resonance structures of *each* of these isomers, including geometries, and determine the formal charges of all atoms in the structures. Describe the hybridizations of each atom in each of the structures. Which isomer do you predict to be the most stable (lowest energy), and why?

4. The thiazyl dichloride ion, $NSCl_2^-$, is isoelectronic with thionyl dichloride, $OSCl_2$.

(a) Provide a Lewis structure, with formal charges, for each and predict the geometric **molecule** shapes.
(b) Which of these species has the smaller Cl-S-Cl bond angle? Explain.
(c) Which species would have the longer S-Cl bond length? Explain.

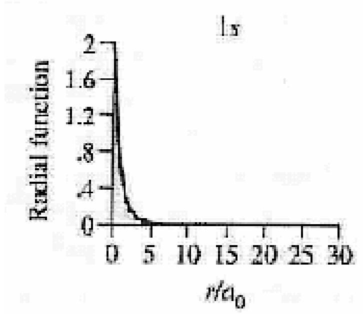
5. (a) Prepare a molecular orbital energy level diagram for SiS, silicon sulfide, showing clearly how the valence shell atomic orbitals interact to form MOs. Label the MOs with appropriate symbols and add electrons. Atomic orbital energies: Si, 3*s* = -15.89 eV, 3*p* = -7.78 eV; for S, 3*s* = -22.71 eV, 3*p* = -11.62 eV.

(b) How does your diagram illustrate the difference in electronegativity between Si and S?

(c) Predict the bond order and the number of unpaired electrons.

(d) SiS^{+} and SiS^{-} are also known. Compare the bond orders of these ions with the bond order of SiS. Which of the three species would have the shortest bond length? Why?

(e) Compare your molecular orbital description of the SiS bonding with that of the Lewis description.



$$\Delta E = hcR\left(\frac{1}{n^2} - \frac{1}{n'^2}\right)$$

$$\lambda = \frac{h}{mv} \qquad E = \frac{hc}{\lambda}$$

$$E = h\nu$$

$$E = mc^2$$

$$m = \frac{h\nu}{c^2} \rightarrow \frac{h}{c\lambda}$$

$$r_n = \frac{\varepsilon_o h^2 n^2}{\pi m_e Z e^2}$$

$$Z_{eff} = Z - S$$

$$\mu = 2\sqrt{S(S+1)}$$

$$(\Delta mv)(\Delta x) \geq \frac{h}{4\pi}$$

$$H\Psi = E\Psi$$

$$\Psi = A\sin(kx) + B\cos(kx)$$

$$\Psi = \sqrt{\frac{2}{a}}\sin\left(\frac{n\pi x}{a}\right)$$

$$E = \frac{n^2h^2}{8ma^2}$$

PERIODIC TABLE OF ELEMENTS

At. No.
X
At. Wt.

IA	IIA	IIIB	IVB	VB	VIB	VIIB		VIIIB		IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1 H 1.01																	2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.1	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.30											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc [98.91]	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57 *La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po [210.0]	85 At [210.0]	86 Rn [222.0]
87 Fr [223.0]	88 Ra [226.0]	89 †Ac [227.0]	104 Rf [261.1]	105 Db [262.1]	106 Sg [263.1]	107 Bh [262.1]	108 Hs [265.1]	109 Mt [266.1]									

58 *Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm [146.92]	62 Sm 150.36	63 Eu 151.97	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
90 †Th 232.04	91 Pa 231.04	92 U 238.03	93 Np [237.0]	94 Pu [239.0]	95 Am [241.1]	96 Cm [244.1]	97 Bk [249.1]	98 Cf [252.1]	99 Es [252.1]	100 Fm [257.1]	101 Md [258.1]	102 No [259.1]	103 Lr [262.1]

Slater's Rules

Groups: (1s) (2s,2p) (3s,3p) (3d) (4s,4p) (4d) (4f) (5s,5p), etc.

Z* = Z - S

1. For *ns* and *np* valence electrons:
   (a) Each electron in the same group contributes 0.35 to the value of S for each other electron in the group.
   EXCEPTION: a 1*s* electron contributes 0.30 to S for another 1s electron.
   (b) Each electron in *n*-1 groups contribute 0.85 to S.
   (c) Each electron in *n*-2 or lower groups contributes 1.00 to S.

2. For *nd* and *nf* valence electrons:
   (a) Each electron in the same group contributes 0.35 to the value of S for each other electron in the group.
   (b) Each electron in groups to the left contributes 1.00 to S.