Chem 150 – Exam 4b *Name: ____*
Fall 2011

Chemistry 150 Exam 4

Be sure to put your name on each page. This page can be removed from your exam so that you will have a Periodic Table handy throughout the exam, it does not need to be turned in. Show all your work for non-multiple choice problems which require any sort of calculation, no credit will be given for answers without work shown. If you have shown a significant amount of work or multiple drawings for a problem, draw a box around what you consider your final answer.

Avogadro's Number = 6.022×10^{23} $^{units}/_{mol}$
$32.00^oF = 0.000^oC = 273.15K$
1 foot = 12 inches
1 inch = 2.54cm (exactly)
1 pound = 453.6 g = 16 ounces
1 amu = 1.6605×10^{-24} g
Masses of subatomic particles:
Proton 1.00728amu = 1.6726×10^{-24} g
Neutron 1.00866amu = 1.6749×10^{-24} g
Electron 0.000549amu = 9.1094×10^{-28} g
Density of Water = $1.000\,^{g}/_{mL}$
R = $0.08206\ ^{L\cdot atm}/_{mol\cdot K}$
PV=nRT
1 calorie = 4.184 J = 0.001Calorie

h = 6.626×10^{-34} Jsec
$\lambda = {}^{h}/_{mv}$
1 J = 1 kg $(^{m}/_{sec})^2$
c = $\lambda\nu$ = $3.00\times10^{8}\ ^{m}/_{sec}$
E_{photon} = $h\nu$

1 **H** 1.0079																	2 **He** 4.0026
3 **Li** 6.941	4 **Be** 9.0122											5 **B** 10.811	6 **C** 12.011	7 **N** 14.007	8 **O** 15.999	9 **F** 18.998	10 **Ne** 20.180
11 **Na** 22.990	12 **Mg** 24.305											13 **Al** 26.982	14 **Si** 28.086	15 **P** 30.974	16 **S** 32.066	17 **Cl** 35.453	18 **Ar** 39.948
19 **K** 39.098	20 **Ca** 40.078	21 **Sc** 44.956	22 **Ti** 47.88	23 **V** 50.942	24 **Cr** 51.996	25 **Mn** 54.938	26 **Fe** 55.847	27 **Co** 58.933	28 **Ni** 58.69	29 **Cu** 63.546	30 **Zn** 65.39	31 **Ga** 69.723	32 **Ge** 72.61	33 **As** 74.922	34 **Se** 78.96	35 **Br** 79.904	36 **Kr** 83.80
37 **Rb** 85.468	38 **Sr** 87.62	39 **Y** 88.906	40 **Zr** 91.224	41 **Nb** 92.906	42 **Mo** 95.94	43 **Tc** (98)	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	71 **Lu** 174.97	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.23	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** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** 226.03	103 **Lr** (260)	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (263)	107 **Bh** (262)	108 **Hs** (265)	109 **Mt** (266)	110 **Ds** (269)	111 **Rg** (272)	112 **Cn** (277)	113	114	115	116	117	118

57 **La** 138.91	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	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** 168.94	70 **Yb** 173.04
89 **Ac** 227.03	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** 237.05	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (258)	101 **Md** (258)	102 **No** (259)

Score

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Multiple Choice: Circle the letter of the most correct response. (5pts. per question)

1. Which of the following is ***not*** a possible set of quantum numbers for an electron?
   a. $n = 1$, $\ell = 2$, $m_\ell = +1$, $m_s = +1/2$
   b. $n = 2$, $\ell = 0$, $m_\ell = 0$, $m_s = +1/2$
   c. $n = 3$, $\ell = 1$, $m_\ell = -1$, $m_s = -1/2$
   d. $n = 3$, $\ell = 2$, $m_\ell = +2$, $m_s = -1/2$
   e. $n = 4$, $\ell = 3$, $m_\ell = -2$, $m_s = +1/2$

2. Electronegativity
   a. Is the negative charge of an ion
   b. Is a measure of how strongly an atom attracts electrons in a covalent bond
   c. Is determined by assigning one electron to each atom of a bond
   d. Is the energy required to remove an electron from an atom in the gas phase
   e. Is the energy required to remove a *pair* of electrons from an atom

3. A covalent bond:
   a. Is always polar
   b. Forms ions in solution
   c. Always contains a metal
   d. Involves sharing electrons
   e. Always has high bond energy

4. Electronegativity ***decreases***:
   a. As the quantum number "n" decreases
   b. As atoms get smaller
   c. Top to bottom on the Periodic Table
   d. Left to right across the Periodic Table
   e. In the center of the Periodic Table

5. What orbital hybridization gives a ***square pyramid molecular shape***?
   a. sp
   b. sp^2
   c. sp^3
   d. sp^3d
   e. sp^3d^2

Trends: For each of the following, circle the correct response (1pts) and give a *brief* explanation of your choice (6pts).

6. Which atom is smaller? Explain:

   Ni (Z=28) vs. Ru(Z=44)

Ruthenium has an entire extra shell of electrons, so it should be larger

7. Which ion is smaller? Explain:

   Ti^{4+} vs. Ti^{2+} (Z=22)

Ti^{+2} has 2 electrons in the 4s subshell, while Ti^{+4} only has electrons in the n=3 subshell. Since the nuclear charge is the same (all titanium atoms or ions have 22 protons…), Ti^{+4} should be the smaller ion.
Since Ti^{+2} has more electrons but the same nuclear charge, each electron in Ti^{+2} is experiencing a little bit less of the positive charge of the nucleus, so the outermost electrons are held a little more "loosely", making Ti^{+2} a little larger

Score

8. Which bond is longer? Explain:

N-Cl vs. P-Cl

Nitrogen is smaller than phosphorus. "Bond length" is measured nucleus-to-nucleus, so chlorine should be able to get closer to the smaller "N", the N-Cl bond should be shorter.

9. Which SO bond is longer? Explain:

SO_2 vs. SO_4^{-2}

Draw Lewis structures for each. The S-O bond order for SO_2 (should be drawn with 2 double bonds, bond order 2) is higher than the S-O bond order for sulfate ion (bond order = 1.5), and higher order bonds are shorter (when the 2 elements involved in the bond are the same), so the S-O bonds in SO_2 should be a little shorter than the S-O bonds in sulfate.

10. Which element is less electronegative? Explain:

P vs. Ge

Phosphorus is smaller and closer to a full shell configuration, so it should attract electrons more strongly.
P is closer to F on the P.T.

11. Which bond is more polar? Explain:

Se-S vs. Si-As

Polarity is determined by differences in electronegativity. The difference in electronegativity between Si and As is less than the difference in electronegativity between Se and S, so the Si-As bond should be less polar than the Se-S bond.

For each of the following, write out a correct electron configuration. You may use noble gas shorthand notation for species below the 2nd row of the Periodic Table. (6pts each)

12. Germanium (At.# = 32)

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$

13. Cesium (At.# = 55)

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$

14. Phosphide ion (At.# = 15)

$1s^2 2s^2 2p^6 3s^2 3p^6$

15. Iron(III) ion (At.# = 26)

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$

Score

16. What are the 3 most likely charges (+ or -) of a tellurium ion (At.# = 52)? Explain your answers. (15pts)

The electron configuration for a Te atom is: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^4$ or $[Kr]5s^2 4d^{10} 5p^4$
Te^{-2} – Gaining 2 electrons will result in a stable "full shell" configuration, $[Kr]5s^2 4d^{10} 5p^6$
Te^{+4} – Losing 4 electrons will result in a stable "full sub-shell" configuration, $[Kr]5s^2 4d^{10}$
Te^{+6} – Losing 6 electrons will result in a stable "full shell" configuration, $[Kr]4d^{10}$
Te^{+1} – Losing 1 electrons will result in a relativelystable "half-full sub-shell" configuration, $[Kr]5s^2 4d^{10} 5p^3$

For each of the following, draw a correct Lewis Structure, determine the formal charge on each atom, name the electronic geometry, draw an appropriate VSEPR structure, name the molecular shape, and show the dipole moment of any polar molecules/ions. (15pts each)

17. RnF_4

No drawings here…
Electronic geometry = octahedral
Molecular geometry = square planar
With single bonds, the formal charge on all atoms is zero
Polar bonds (bond dipoles point from Rn to F) but a nonpolar molecule

18. PO_3^{-3}

No drawings here…
Electronic geometry = tetrahedral
Molecular geometry = trigonal pyramid
With all single bonds, the formal charge on P is zero and the formal charge on each oxygen is -1.
Polar bonds (bond dipoles point from P to O) and a polar molecule with the molecular dipole pointing through the P and between all 3 oxygens

19. $SbOBr_3$

No drawings here…
Electronic geometry = tetrahedral
Molecular geometry = tetrahedral
With single Sb-Br bonds and a double bond between Sb and O, the formal charge on all atoms is zero
Polar bonds (bond dipoles point from Sb to Br and Sb to O) and a polar molecule. The Sb=O bond is MORE polar than the Sb-Br bonds, so the molecular dipole is along the Sb=O bond pointing toward the oxygen

Score