Chemistry 141 Name key

Dr. Cary Willard

Exam 3a November 16, 2010

Multiple Choice (30 points)

Page 4 (12 points)

Page 5 (13 points)

Page 6 (20 points)

Page 7 (16 points)

Page 8 (24 points)

Page 9 (12 points)

Total (127 points)

Percent (100 %)

All work must be shown to receive credit. Give all answers to the correct number of significant figures

Avogadros number = 6.022 x 1023 /mol

Grossmont College

Periodic Table

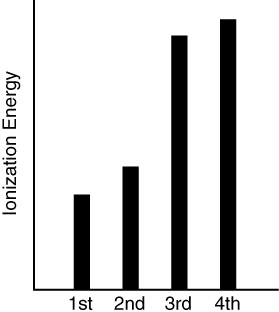
|  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| IA |  |  |  |  |  |  |  |  |  |  | |  |  |  |  |  | VIIA | NOBLE GASES |
| 1  **H**  1.008 | IIA |  |  |  |  |  |  |  |  |  | |  | IIIA | IVA | VA | VIA | 1  **H**  1.008 | 2  **He**  4.002 |
| 3  **Li**  6.941 | 4  **Be**  9.012 |  |  |  |  |  |  |  |  |  | |  | 5  **B**  10.81 | 6  **C**  12.01 | 7  **N**  14.01 | 8  **O**  16.00 | 9  **F**  19.00 | 10  **Ne**  20.18 |
| 11  **Na**  23.00 | 12  **Mg**  24.30 | IIIB | IVB | VB | VIB | VIIB | VIII VIII VIII | | | | IB | IIB | 13  **Al**  27.00 | 14  **Si**  28.09 | 15  **P**  30.97 | 16  **S**  32.06 | 17  **Cl**  35.45 | 18  **Ar**  39.95 |
| 19  **K**  39.10 | 20  **Ca**  40.08 | 21  **Sc**  44.96 | 22  **Ti**  47.90 | 23  **V**  50.94 | 24  **Cr**  52.00 | 25  **Mn**  54.94 | 26  **Fe**  55.85 | 27  **Co**  58.93 | 28  **Ni**  58.70 | | 29  **Cu**  63.55 | 30  **Zn**  65.38 | 31  **Ga**  69.72 | 32  **Ge**  72.59 | 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**  (99) | 44  **Ru**  101.1 | 45  **Rh**  102.9 | 46  **Pd**  106.4 | 47  **Ag**  107.9 | | 48  **Cd**  112.4 | 49  **In**  114.8 | 50  **Sn**  118.7 | 51  **Sb**  121.8 | 52  **Te**  127.6 | 53  **I**  126.9 | 54  **Xe**  131.3 |
| 55  **Cs**  132.9 | 56  **Ba**  137.3 | 57  **La**  138.9 | 72  **Hf**  178.5 | 73  **Ta**  180.9 | 74  **W**  183.9 | 75  **Re**  186.2 | 76  **Os**  190.2 | 77  **Ir**  192.2 | 78  **Pt**  195.1 | 79  **Au**  197.0 | | 80  **Hg**  200.6 | 81  **Tl**  204.4 | 82  **Pb**  207.2 | 83  **Bi**  209.0 | 84  **Po**  (209) | 85  **At**  (210) | 86  **Rn**  (222) |
| 87  **Fr**  (223) | 88  **Ra**  226.0 | 89  **Ac**  227.0 | 104  **Rf**  (261) | 105  **Db**  (262) | 106  **Sg**  (263) | 107  **Bh**  (262) | 108  **Hs**  (265) | 109  **Mt**  (266) | 110  **??**  (269) |  | |  |  |  |  |  |  |  |

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| 58  **Ce**  140.1 | 59  **Pr**  140.9 | 60  **Nd**  144.2 | 61  **Pm**  (147) | 62  **Sm**  150.4 | 63  **Eu**  152.0 | 64  **Gd**  157.3 | 65  **Tb**  158.9 | 66  **Dy**  162.5 | 67  **Ho**  164.9 | 68  **Er**  167.3 | 69  **Tm**  168.9 | 70  **Yb**  173.0 | 71  **Lu**  175.0 |
| 90  **Th**  232.0 | 91  **Pa**  231.0 | 92  **U**  238.0 | 93  **Np**  (237) | 94  **Pu**  (244) | 95  **Am**  (243) | 96  **Cm**  (247) | 97  **Bk**  (247) | 98  **Cf**  (251) | 99  **Es**  (252) | 100  **Fm**  (257) | 101  **Md**  (258) | 102  **No**  (259) | 103  **Lr**  (260) |

Lanthanide series

Actinide series

Part I – Multiple Choice (30 points)

1. Which of the following elements is isoelectronic with the Pb4+ ion?
   1. Rn
   2. Xe
   3. Hg
   4. Pt
2. The atomic radius of germanium is smaller than the atomic radius of potassium due to
   1. a change in the *n* quantum number.
   2. an increase in the effective nuclear charge.
   3. the fact that *p* and *d* orbitals have the same orbital penetration.
   4. a decrease in the effective nuclear charge.
3. Which arrangement is in the correct order of increasing radii?
   1. Mn > Mn2+ > Cs
   2. P < P3– < As3–
   3. Li+ > Li > Ra
   4. Cr < Cr3+ < Ca
4. Which of the following elements would you expect to have the greatest first ionization energy?
   1. C
   2. Li
   3. O
   4. Be
5. The first four ionization energies for an element are as follows. Identify the correct element from the list.
   1. Sr
   2. K
   3. S
   4. Al
6. A covalent bond results when
   1. electrons are transferred from one atom to another atom.
   2. atoms pool their electrons to form a “sea” of electrons.
   3. atoms have outer electrons with the same principal quantum number.
   4. electrons are shared between atoms.
7. Which of the following properties is typically used to predict the type of bond that forms between two elements?
   1. electronegativity
   2. atomic radius
   3. ionization energy
   4. electron affinity
8. Which of the following are listed in order of decreasing electronegativity?
   1. F, S, Na, H
   2. N, P, Si, Mg
   3. F, N, P, O
   4. F, Cl, Br, C
9. Indicate which of the following elements has the largest effective nuclear charge (*Z*eff) based on the data given.
   1. Mg (ionization energy = 738 kJ/mol)
   2. Na (ionization energy = 496 kJ/mol)
   3. Be (ionization energy = 899 kJ/mol)
   4. Li (ionization energy = 520 kJ/mol)
10. Which of the following is most likely a polar covalent bond?
    1. Na–Cl
    2. C–N
    3. H–H
    4. K–F
11. Which of the following atoms can have an expanded octet?
    1. As
    2. N
    3. C
    4. B
12. Bond length and bond strength are
    1. inversely proportional.
    2. opposite one another.
    3. directly proportional.
    4. unrelated.
13. The greatest repulsive forces in molecules are due to
    1. bonding–bonding.
    2. bonding–lone pair.
    3. unpaired–lone pair.
    4. lone pair–lone pair.
14. The H–X–H bond angle is larger in CH4 than in NH3,and ammonia has a larger angle than H2O. This trend is due to
    1. the effective nuclear charge’s decrease.
    2. an increase in the number of lone pairs, the lone pair repulsion causing the decrease in angle.
    3. the fact that carbon has a larger atom volume than nitrogen or oxygen.
    4. the change in polarity of the molecules.
15. All homonuclear diatomic molecules
    1. have polar bonds.
    2. are polar.
    3. are nonpolar.
    4. cannot vibrate.

Part 2 - Problems

1. (6 points) Write the complete electron configuration for an atom of N and of N-3.
   1. N

1s2 2s2 2p3

* 1. N-3

1s2 2s2 2p6

1. (6 points) Write the shorthand electronic configuration for an atom of Ta, of Ta+3.
   1. Ta

[Xe] 6s2 5d3 4f14

* 1. Ta+3

[Xe] 5d2 4f14

1. (5 points) Given the graphical representation of the first 21 ionization energies for zinc (in MJ/mol), explain the general trend in the curve, and any deviations from that general trend. (Larger graph on front of exam.)

The 1st 2 s electrons are very easy to remove. There is a jump for the next electron because it is a 3d electron, buried in an inner shell. There is a relatively linear increase in ionization energy for the 10 electrons. There is a significant jump in energy to remove electron 13, this is a p electron which is harder to remove. The next jump occurs at electron 19 where we again must remove the next 2 electrons from the 3 s sublevel. At electron 21, the jump is huge as we remove the next electron from the n=2 level.

1. (8 points) Two structures may be drawn for C4H5N2Br:

Structure a Structure b

* 1. Are these two resonance structures of the same molecule? Explain.

These are different molecules because they have different skeleton structures. Resonance structures must have the same skeleton structure!

* 1. How many sigma bonds are in structure a? \_\_\_\_\_\_\_\_How many pi bonds?\_\_\_\_\_\_\_\_
  2. Which bonds are longer, the CC bonds in structure a or b? Explain.

The CC bonds in b are longer because single bonds are longer than triple bonds.

* 1. Which bonds are stronger, the CN bonds in a or b? Explain.

The CN bonds in b are stronger because double bonds are stronger than single bonds.

1. (20 points) Write reasonable Lewis Electron Dot Structures for the following molecules or ions (Central atom is listed first). Tell the orbital and molecular geometry for each molecule/ion. Show formal charges for all non-zero charges. If resonance structures exist, show them.

|  |  |  |
| --- | --- | --- |
| NO2F  N is central atom | and | orbital geometry  trigonal planar  molecular geometry  trigonal planar |
| COS (C is central atom) |  | orbital geometry  linear  molecular geometry  linear |
| AlF6-3  Cryolite – used in the manufacture of aluminum |  | orbital geometry  tetrahedral  molecular geometry  linear |
| XeF2 |  | orbital geometry  trigonal bipyramidal  molecular geometry  linear |

1. (6 points) Although I3-1 is known, F3-1 is not. Using Lewis structures, explain why F3-1 does not form.

I3-1 and F3-1 both require an expanded octet on the central atom. Iodine has available d orbitals so that it can expand its octet. Fluorine’s outermost electrons are in the 2p orbitals and there are no 2d orbitals in which to place extra electrons.



1. (10 points) Answer the following questions for the structure below:



|  |  |
| --- | --- |
| What is the molecular geometry of I (arrow a)?  See saw | What is the hybridization of Kr (arrow g)?  sp3d2 |
| What is the formal charge of Rn (arrow c)?  +2 | What is the hybridization of C (arrow h)?  sp2 |
| What is the hybridization of N (arrow d)?  sp2 | What is the hybridization of Br (arrow i)?  sp3d2 |
| What is the orbital geometry of N (arrow e)?  Linear | What is the formal charge on Sb (arrow j)?  -3 |
| What is the molecular geometry of P (arrow f)?  Trigonal pyramidal | What is the formal charge on S (arrow k)?  +1 |

1. (9 points) Nitrosyl fluoride (NOF) has an atom sequence in which all atoms have formal charges of zero. Draw all possible skeleton structures and identify the structure which meets these criteria.

Good!!

bad worse 

same worse still! 

1. (5 points) Explain how sigma and pi bonds differ.

Sigma bonds have overlap between the atoms and pi bonds have overlap above and below the atoms.

1. (4 points) Describe the difference between a pure covalent bond and a polar covalent bond.

Answer: A pure covalent bond occurs when bonding electrons are shared equally (or very close to it) as in the N-N bond. A polar covalent bond is formed between 2 atoms of differing electronegativities. The bonding electrons are unequally shared between the two atoms as in the CO molecule.

1. (6 points) Look at the compound pictured below. Explain the bonding in terms of valence bond theory. That is show the atomic orbitals on the Xe atom, describe any electron promotion and hybridization necessary, and show the orbitals involved in both sigma and pi bonding as well as the orbital holding the lone pair of electrons

hybridization of xenon atom



1. (12 points) Some species with two oxygen atoms only are the oxygen molecule, O2, the peroxide ion, O2-2, the superoxide ion, O2-1, and the dioxgenyl ion, O2+1. Draw an MO diagram for each, on the following page and answer the questions. Note that each box is labeled with a particular species.
   1. Rank these species in order of increasing bond length

\_\_\_ O2-2\_\_\_>\_\_\_ O2-1\_\_\_>\_\_ O2\_\_\_\_>\_\_\_ O2+1\_\_\_

* 1. Rank these species in order of increasing bond strength

\_\_\_O2+1\_\_\_>\_\_\_O2\_\_\_>\_\_O2-1\_\_\_\_>\_\_\_O2-2\_\_\_

* 1. Give the bond order in all species

O2 2 O2-2 1

O2-1 1.5 O2+1 2.5

* 1. Identify each species as diamagnetic or paramagnetic

O2 paramagnetic O2-2 diamagnetic

O2-1 paramagnetic O2+1 paramagnetic

|  |  |
| --- | --- |
| f1q52g1  Oxygen gas, O2 | Peroxide ion, O2-2  f1q52g1 |
|  |  |
| f1q52g1  Superoxide ion, O2-1 | Dioxygenyl ion, O2+1  f1q52g1 |