Form A/B C/D

Detailed Key

CHEMISTRY 102A/C Spring 2018
Hour Exam I Page 1

1 | 21
25 | 5

1. Measured quantities, such as length, mass, or volume, can best be described as:

All measured quantities have uncertainty associated with them; generally assumed to be ±1 in the last digit of the measurements.

a) sometimes certain b) always certain

c) always uncertain d) sometimes uncertain

2 | 22
26 | 6

2. Calculate the number of oxygen atoms in 2.0 g of perchloric acid. = $HClO_4$

$HClO_4$: 1.008 + 35.45 + 4(16.00) = 100.46 g/mol

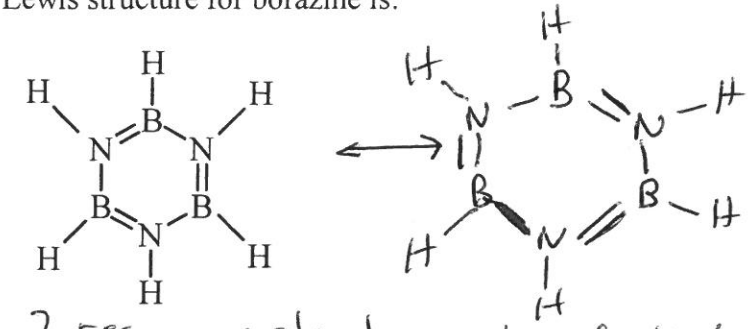
a) 4.8×10^{22} O atoms b) 6.0×10^{23} O atoms c) 1.7×10^{21} O atoms

$2.0\ g\ HClO_4 \left(\frac{1\ mol\ HClO_4}{100.46\ g}\right)\left(\frac{4\ mol\ O}{1\ mol\ HClO_4}\right)\left(\frac{6.022 \times 10^{23}\ atoms\ O}{mol\ O}\right) = 4.8 \times 10^{22}\ O\ atoms$

d) 3.4×10^{21} O atoms e) 1.2×10^{22} O atoms

3 | 23
27 | 7

3. A Lewis structure for borazine is:



$B_3N_3H_6$: 3(3) + 3(5) + 6(1) = 30 valence electrons

Bond angles about each B and N are 120°, so each B and N are sp^2 hybridized with 1 unhybridized p atomic orbital.

2 resonance structures where double bonds are rotated one position

Which of the following statements is **false** concerning borazine?

T a) Each sigma bond between boron and nitrogen is formed from overlap of an sp^2 hybrid orbital from boron with an sp^2 hybrid orbital from nitrogen.

T b) Another resonance Lewis structure can be drawn for borazine. See above

T c) All boron-nitrogen bond lengths are equivalent in borazine. Whenever resonance can be drawn, all bonds are equal.

T d) The π electrons in borazine are delocalized above and below the entire ring surface.

F e) Each boron atom and nitrogen atom in borazine has two unhybridized 2p atomic orbitals. B and N are both sp^2 hybridized. one

2 just like C_6H_6, benzene, discussed in section 9.5 of text.

4 | 24
28 | 8

4. How many of the following five processes (I-V) are examples of chemical change?

formulas change in a chemical change; in a physical change, formulas do not change, but the state of the substance changes (s, ℓ, aq, or g).

physical I. $H_2O(l) \rightarrow H_2O(g)$

physical II. $I_2(s) \rightarrow I_2(g)$

chemical III. $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(l)$

physical IV. $C_6H_{12}O_6(s) \rightarrow C_6H_{12}O_6(aq)$

chemical V. $2\ H_2O_2(aq) \rightarrow 2\ H_2O(l) + O_2(g)$

a) 1 b) 2 c) 3 d) 4

e) 5 (All are examples of chemical change.)

From Heisenberg's uncertainty principle, one cannot know the exact location of an electron. So simple circular orbits at a defined distance from the nucleus are not allowed by Heisenberg's uncertainty principle.

5|11 19|13 5. Which aspect of the Bohr model is **not** allowed by Heisenberg's uncertainty principle?

a) discrete atomic energy levels
b) simple circular orbits
c) deBroglie wavelengths
d) atomic orbitals

radius trend ←

Trend only works for neutral charged atoms. For the neutral charged atoms, the radius order is: F < Br < Ca < Sr

6|12 20|14 6. Which of the following correctly orders the radius of the atoms/ions listed from smallest to largest radius? For F^+, it has the same number of protons as F, but F^+ has one fewer electron. F^+ will be smaller than F. For Rb^-, it is isoelectronic with Sr, but Rb^- has 1 fewer proton attacking the 38 e⁻. So Rb^- is larger than Sr. Putting all this together gives:

a) $F < F^+ < Ca < Br < Rb^- < Sr$
b) $F < F^+ < Ca < Sr < Rb^- < Br$
c) $F^+ < F < Br < Ca < Rb^- < Sr$
d) $F^+ < F < Br < Ca < Sr < Rb^-$
e) $F^+ < F < Ca < Sr < Br < Rb^-$

$F^+ < F < Br < Ca < Sr < Rb^-$

7|13 21|15 7. Which of the following statements is **false** concerning bonding?

O more electronegative than C.

T a) In a C–O bond, electron density on average is greater near the O atom.
T b) A C–O bond is an example of a polar covalent bond.
F c) The bond in NaBr is formed by ~~sharing~~ transfer electrons. metal + nonmetal → ionic bond forms
T d) Elements with extremely different electronegativity values tend to form ionic bonds with each other.
T e) An N–H bond is more polar than a P–H bond.

N and H have a larger difference in electronegativity. Note that P and H have identical electronegativities

8|14 22|16 8. Consider the following **unbalanced** reaction (it is called the thermite reaction):

$9\ Fe(s) + 4\ Al_2O_3(s) \rightarrow 3\ Fe_3O_4(s) + 8\ Al(s)$

In the best balanced equation, what is the sum of the coefficients for the reactants and the products? $9+4+3+8 = 24$

a) 24 b) 6 c) 8 d) 9 e) 13

9|15 23|17 9. Which of the following is named **incorrectly**?

ok a) $Cr_2(SO_3)_3$ chromium(III) sulfite
ok b) $NaC_2H_3O_2$ sodium acetate
NO c) KNO_2 potassium nitride
ok d) $Ca(OH)_2$ calcium hydroxide
ok e) $NiCO_3$ nickel(II) carbonate

NO_2^- = nitrite; N^{3-} = nitride

Cu^+: $[Ar]4s^0 3d^{10}$ (0 unpaired e⁻)
Ni^{2+}: $[Ar]4s^0 3d^8$ ↑↓ ↑↓ ↑↓ ↑ ↑ (2 unpaired e⁻)
Zn^{2+}: $[Ar]4s^0 3d^{10}$ (0 unpaired e⁻)
Cr^{2+}: $[Ar]4s^0 3d^4$ ↑ ↑ ↑ ↑ — (4 unpaired e⁻)

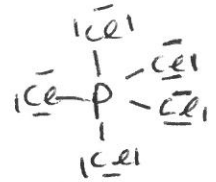
10|16 24|18 10. Consider the transition metal ions Cu^+, Ni^{2+}, Zn^{2+}, Cr^{2+}, and Ti^{2+}. How many of these five transition metal ions has/have two (2) unpaired electrons in the ground state?

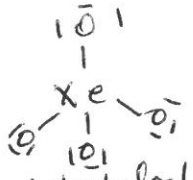
Ti^{2+}: $[Ar]4s^0 3d^2$ ↑ ↑ — — — (2 unpaired e⁻)

a) 1 b) 2 c) 3 d) 4 e) 5 (All have 2 unpaired electrons.)

Ni^{2+} and Ti^{2+} both have 2 unpaired electrons. Note that all of these are exceptions to know.

Form A|B / C|D





tetrahedral

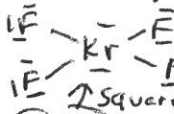
CHEMISTRY 102A/C
Hour Exam I

Spring 2018

trigonal bipyramid

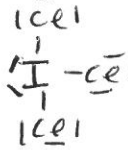
11|1 / 14|24

11. How many of the following molecules have a square planar shape?



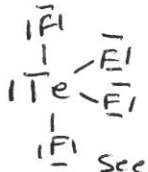
2 square planar

36e⁻ 40e⁻ 32e⁻ 34e⁻ 28e⁻
KrF_4, PCl_5, XeO_4, TeF_4, ICl_3



T-shape

a) 1 b) 2 c) 3 d) 4



see-saw

e) 5 (All have a square planar shape.)

12|2 / 15|25

12. Which ion, in each of the following three pairs of ions, has the largest ionization energy?

Both O^- and O^{2-} have 8 protons. O^- has fewer electrons, so O^- attracts e⁻ more strongly (larger IE).

O^- vs O^{2-} Mg^+ vs Mg^{2+} O^{2-} vs Mg^{2+} – isoelectronic (10 e⁻); Mg^{2+} has more protons & attracting the 10 e⁻, so Mg^{2+} has larger IE.

a) O^{2-}; Mg^+; O^{2-} b) O^-; Mg^+; O^{2-} c) O^-; Mg^+; Mg^{2+}

Mg^{2+} and Mg^+ both have 12 protons. Mg^{2+} has fewer electrons, so the electrons are attracted more strongly in Mg^{2+} which will have larger IE.

d) O^{2-}; Mg^{2+}; O^{2-} e) O^-; Mg^{2+}; Mg^{2+}

13|3 / 16|26

13. The isotope of an unknown metal, M, has a mass number of 26. The most stable ion of the isotope forms a binary compound with sulfur having the formula MS. How many neutrons does M have? For relatively light elements, # neutrons ≈ # protons

M forms +2 charged ions.

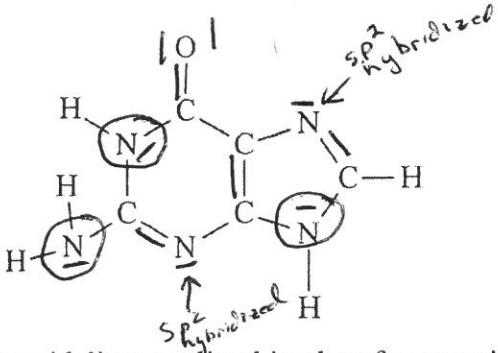
If neutrons = protons, then this isotope has 13 protons and 13 neutrons.

a) 14 b) 26 c) 12 d) 10 e) 24

Element 13 is Aluminum, but Al forms Al_2S_3 (not AlS). To fit MS data, M is Mg (12 protons), Mg forms Mg^{2+} ions.

neutrons = 26 − 12 = 14

DNA molecules are complex organic compounds found in every living cell which act as the information and control centers for the cell. Part of any DNA molecule is the organic compound guanine which has the following skeletal structure.



Organic Rules
C: 4 bonds + 0 lone pairs
N: 3 " + 1 " "
O: 2 " + 2 " "

Using the guidelines outlined in class for organic compounds, complete a Lewis structure for guanine and answer the following two questions.

14|4 / 17|27

14. How many C and N atoms are sp^3 hybridized?

The circled N atoms have tetrahedral geometry (109° bond angles), so they are sp^3 hybridized. None of the carbon atoms have tetrahedral geometry, so 0 carbons are sp^3 hybridized. All carbons have trigonal planar geometry, so they are sp^2 hybridized.

a) 1 b) 2 c) 3 d) 4 e) 5

15|5 / 18|28

15. How many double bonds are in the completed Lewis structure?

a) 1 b) 2 c) 3 d) 4 e) 5

There are 4 double bonds in the correct Lewis structure.

Form A/B e/D

$\Delta E = -R_H\left(\frac{1}{n_2^2} - \frac{1}{n_1^2}\right) = -2.178 \times 10^{-18}\,J\left(\frac{1}{2^2} - \frac{1}{6^2}\right)$

$\Delta E = -4.84 \times 10^{-19}\,J$

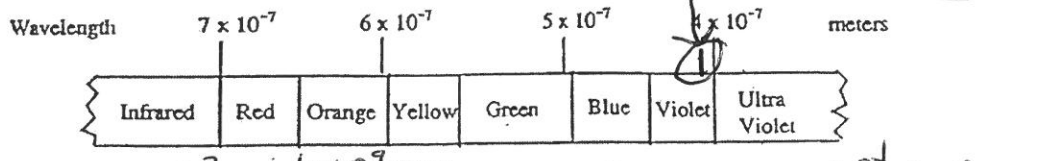
CHEMISTRY 102A/C Spring 2018
Hour Exam I Page 4

$E_{photon} = |\Delta E| = 4.84 \times 10^{-19}\,J = \frac{hc}{\lambda}$

$\lambda = \frac{hc}{E} = \frac{6.626 \times 10^{-34}\,J\,s\,(2.998 \times 10^{8}\,m/s)}{4.84 \times 10^{-19}\,J} = 4.10 \times 10^{-7}\,m$

16/25 10/1

16. Use the figure below to answer the next question:



$4.10 \times 10^{-7}\,m \left| \frac{1 \times 10^{9}\,nm}{1\,m} \right. = 410.\,nm$ ← not needed

For hydrogen, the light emitted for the n = 6 to n = 2 electronic transition is in the visible region of the electromagnetic radiation spectrum (see spectrum above). What is the color of visible light emitted for the n = 6 to n = 2 electronic transition?

From the figure, $4.10 \times 10^{-7}\,m$ light is violet.

a) red b) orange c) green d) blue (e) violet

17/26 11/2

17. Which of the following compounds will be least soluble in water? Hint: Water is a polar solvent and "like dissolves like".

Water dissolves other polar covalent compounds; it does not dissolve nonpolar compounds. Only KrF_2 is nonpolar, so it is least soluble in water. Only in KrF_2 do the effect of the individual bond dipoles cancel each other out.

(a) KrF_2 (22e) b) PF_3 (26e⁻) c) IF_5 (42e⁻) d) COS (16 e⁻) e) SO_2 (18e⁻)

F–Kr–F nonpolar

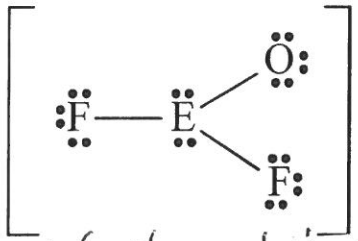
polar P, F, F, F

polar I, F, F, F, F

S=C=O polar

O, S=O + 1 other resonance structure, polar

E in the Lewis structure below is a general symbol for some element. Consider this Lewis structure for the next two questions.



EF_2O^{2-} has 28 electrons in Lewis structure.

$28 = x + 2(7) + 6 + 2$

$x = 6$ valence electrons

E is a group 6 element (has 6 valence e⁻), but it must be a row 3 element or heavier because E has more than 8 electrons around it.

18/27 12/3

18. Which of the following elements could be E?

Could be S or Se or Te. Only S is listed.

a) C b) Si (c) S d) Kr e) Ne

19/28 13/4

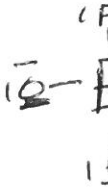
19. Which of the following statements is **false** concerning the EOF_2^{2-} ion?

I a) The predicted VSEPR shape of EOF_2^{2-} is T-shaped.

(b) The predicted VSEPR bond angles about the central E atom are ~~120°~~ 90° and 180° (no 120° in T shaped molecules).

T c) The predicted hybridization of the central E atom is dsp^3.

T d) It is impossible to draw a Lewis structure for EOF_2^{2-} which satisfies the octet rule for all atoms in EOF_2^{2-}.



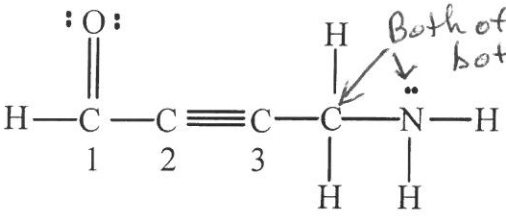
EOF_2^{2-} has a T-shaped structure. Based on trigonal bipyramidal geometry, so dsp^3 hybridized.

Form A/B C/D

single bond = σ
double bond = 1 σ + 1 π
triple bond = 1 σ + 2 π

20|17 6|9

20. Which of the following statements regarding the Lewis structure below is **false**?



Both of these are sp^3 hybridized because both central atoms exhibit tetrahedral geometry.
C_1 is sp^2 hybridized because it has trigonal planar geometry.
C_2 and C_3 are sp hybridized because they both exhibit linear geometry.

T a) An sp^2 hybrid orbital from C-1 overlaps with an sp hybrid orbital from C-2 to form the sigma bond between C-1 and C-2. Hybrids overlap to form σ bonds.
T b) This molecule has three π bonds.
T c) Two of the atoms in this compound are sp^3 hybridized.
(F) d) The π bonds between C-2 and C-3 are formed from overlap of ~~sp hybrid~~ unhybridized p atomic orbitals.
T e) There are 10 sigma bonds in this molecule. The unhybridized p atomic orbitals always form the π bonds.

21|18 7|10

21. Which of the following statements is **false** concerning ionization energy (IE)?
First IE of Ne is greater than F (from IE trend) and the second IE of Ne is greater than the first IE of Ne, so second IE of Ne has to be greater than first IE of F.

T a) The second ionization energy of Ne is greater than the first ionization energy of F.
(F) b) For an isoelectronic series, the species with the most protons should have the smallest ionization energy. species with most protons attracts the electrons most strongly; so it will have the largest IE.
T c) As the size of an atom increases, ionization energy generally decreases. trends are opposite.
T d) As the electronegativity of an atom increases, ionization energy generally increases. same trends
T e) The third ionization energy for magnesium corresponds to the enthalpy change for the reaction: $Mg^{2+}(g) \rightarrow Mg^{3+}(g) + e^- \quad \Delta H = IE_3$. ← corresponds to third electron being removed.

22|19 8|11

22. How many of the following are **correct** ground state electron configurations for the element or ion listed? Note: Element #118 is not yet discovered and Zr is element #40.

noble gas under Rn → element #118: [Rn] $7s^2 5f^{14} 6d^{10} 7p^6$
Zr: [Kr] $5s^2 4d^2$
S^{2-}: [Ne] $3s^2 3p^6$
Ge: [Ar] $4s^2 3d^{10} 4p^2$
} All are correct.

a) 0 (None are correct.) b) 1 c) 2 d) 3

(e) 4 (All are correct.)

23|20 9|12

23. Which of the following statements is **true**?

F a) Rutherford's metal foil experiment proved the existence of electrons.
F b) Dalton's atomic theory said that each element is made up of neutrons.
(T) c) Most of an atom's volume is occupied by its electrons.
F d) Most of an atom's mass comes from its electrons.
F e) It is impossible for an isotope of carbon and an isotope of nitrogen to have the same mass number.
^{14}C and ^{14}N both exist, for example.

A | B
C | D

Volume of metal = 32.37 − 30.07 = 2.30 mL
By subtraction rule, this number is known to 2nd decimal place.

24 | 6
1 | 19

24. A piece of metal has a mass of 1.0107×10^{-2} kg. When it was placed into a 100-mL graduated cylinder containing 30.07 mL of water, the volume of water increased to 32.37 mL. What is the density of the metal (in g/mL) to the correct number of significant figures?

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{1.0107 \times 10^{-2}\ \text{kg} \left(\frac{1000\ \text{g}}{1\ \text{kg}}\right)}{2.30\ \text{mL}} = 4.39\ \text{g/mL}$$

a) 4.4×10^{-3} g/mL b) 4.394×10^{-3} g/mL c) 4.4 g/mL
By multiplication rule, a 5 sig fig number divided by a 3 sig fig number gives a 3 sig fig answer.
(d) 4.39 g/mL e) 4.394 g/mL

25 | 7
2 | 20

25. Experiments show that an unknown element, X, has one unpaired electron in the ground state and has an exothermic electron affinity. Its most stable ion exhibited in ionic compounds is known to be isoelectronic with argon. Which of the following elements could be X? Cl, Na, and Ga all have 1 unpaired electron (Ca has 0 unpaired e⁻ and S has 2 unpaired e⁻). Of these only Cl forms an ion in stable ionic compounds (Cl^-) that is isoelectronic with Ar (18e⁻).

(a) Cl b) Na c) Ca d) S e) Ga

26 | 8
3 | 21

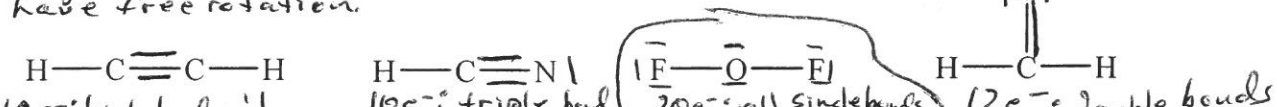
26. Which of the following does **not** correctly describe 16.0 grams of methane, CH_4?
CH_4: 12.0 + 4(1.0) = 16.0 g/mol = 1 mol CH_4 = 6.022×10^{23} molecules CH_4

T a) One (1.00) mole of methane.
T b) The amount of methane that contains 12.0 g of carbon.
T c) The amount of methane that contains 4.0 g hydrogen.
F d) The amount that contains ~~16.0~~ (6.02×10^{23}) molecules of methane.
T e) The amount that contains $4.0 \times (6.02 \times 10^{23})$ atoms of hydrogen.

↳ atoms H = $\frac{16.0\ \text{g}\ CH_4}{16.0\ \text{g}} \left|\frac{1\ \text{mol}\ CH_4}{1\ \text{mol}\ CH_4}\right| \frac{4\ \text{mol H}}{\text{mol}} \left|\frac{6.022 \times 10^{23}\ \text{atoms}}{\text{mol H}}\right. = 4(6.022 \times 10^{23})$ H atoms

27 | 9
4 | 22

27. In some bonds, the atoms can rotate freely without breaking the bond between the atoms; while in other bonds, the atoms cannot rotate freely unless the bond is broken. How many of the following four molecules have at least one bond where the atoms **cannot** freely rotate unless a bond is broken? Single bonds consist of sigma bonds, and only sigma bonds have free rotation. π bonds do not have free rotation.



10e⁻; triple bond can't rotate 10e⁻; triple bond can't rotate 20e⁻; all single bonds, so can all rotate. 12e⁻; double bonds can't rotate freely.

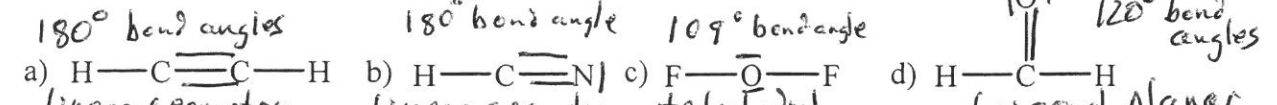
a) 0 (none) b) 1 c) 2 (d) 3

C_2H_2, HCN, and CH_2O all have double or triple bonds, so they all have at least 1 π bond. Therefore, these 3 molecules do not have free rotation about all bonds.

e) 4 (All have at least one bond where the atoms cannot freely rotate.)

28 | 10
5 | 23

28. Which of the following molecules have at least one 90° bond angle in their structure as predicted by the VSEPR model? Hint: reference the Lewis structures you drew for the previous question.



180° bond angles 180° bond angle 109° bond angle 120° bond angles

a) linear geometry b) linear geometry c) tetrahedral geometry d) trigonal planar geometry

(e) None of the above molecules (a-d) have at least one 90° bond angle.

Form A|B C|D

E_{photon} = energy difference between 2p and 2s orbital

$E = \frac{hc}{\lambda} = \frac{6.626\times10^{-34}\ J\cdot s\,(2.998\times10^{8}\ m/s)}{670.8\ nm\left(\frac{1\ m}{1\times10^{9}\ nm}\right)} = 2.961\times10^{-19}\ J$

29|29 29|29

29. When an electron in a 2p orbital of a lithium atom makes a transition to the 2s orbital, a photon of wavelength 670.8 nm is emitted. The energy difference between the 2p and 2s orbitals in lithium is:

a) 2.96×10^{-10} J (b) 2.96×10^{-19} J c) 3.38×10^{-18} J

d) 2.96×10^{-17} J e) 3.38×10^{-20} J

mass O = 100.00 − 68.42 = 31.58 g O

30|30 30|30

30. A metal M forms an oxide having the formula M_2O_3. If 100.00 g of M_2O_3 contains 68.42 g of M, what is the atomic mass of the metal, M?

mol M in 100.00 g compound = $31.58\ g\ O\left(\frac{1\ mol\ O}{16.00\ g}\right)\left(\frac{2\ mol\ M}{3\ mol\ O}\right) = 1.3158\ mol\ M$

a) 26.98 amu b) 44.96 amu (c) 52.00 amu d) 69.72 amu e) 102.9 amu

In 100.00 g compound, we have 68.42 g M which represents 1.3158 mol M.

molar mass = $\frac{68.42\ g}{1.3158\ mol}$ = 52.00 g/mol

31. Which of the following statements is **false**?

6Px, 6Py, and 6Pz all have the same energy.

T a) The three 6p atomic orbitals have the same energy but they differ in their orientation about the x, y and z axes.

T b) 3f atomic orbitals do not exist in the quantum mechanical model of the atom. 4f are first f orbitals to exist.

T c) For neutral charged atoms, the 8s atomic orbital should be lower in energy than the 7d atomic orbitals. From periodic table.

(F d)) An excited state electron configuration for an atom represents a ~~lower~~ higher energy electron configuration as compared to the ground state electron configuration.

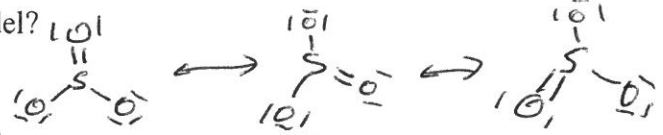
ground state = lowest energy electron configuration

32|32 32|32

32. How many of the following compounds/ions exhibit resonance and have at least one 120° bond angle as predicted by the VSEPR model?

24e⁻ 24e⁻ 24e⁻ 24e

SO_3, SeS_3, NO_3^-, TeS_3



All exhibit trigonal planar geometry so all have 120° bond angles.

a) 0 (none) b) 1 c) 2 d) 3 e) 4 (all)

All four of these substances contain 4 atoms in formula and all have 24 e⁻. All will have the same Lewis structures (like SO_3 above).

33|33 33|33

33. If the frequency of electromagnetic radiation is decreased by a factor of one-half, which of the following statements is **false**? Wavelength and frequency are inversely related.

$\lambda\nu = c$; as ν decreases by 2, λ must increase by a factor of 2.

T a) The number of cycles passing a given point per unit time halves.

(F b)) The velocity of the radiation halves. All travel at same velocity, c = speed of light

T c) The wavelength of electromagnetic radiation is doubled.

d) The photon energy of the electromagnetic radiation is halved.

$E = h\nu$; as ν decreases by 2, E_{photon} must also decrease by 2.

E and ν are directly related.

34. My answers for this Chemistry 102 exam should be graded with the answer sheet associated with:

a) Form A b) Form B c) Form C d) Form D e) Form E