

**1. 10 points**

The absorption of energy results in the excitation of an electron in an **ion**.

The excited electronic configuration of the ion is  $5p^1$  i.e. the electron in the ion is excited to the 5p orbital.

In the **emission** spectrum of the ion, the **longest** wavelength observed is 112.54 nm. Using Bohr's or Rydberg's equation, determine which ion does this spectrum correspond to?

Show your work and calculations, including formulae used.

(Assume all transitions are possible for the ion)

**Answer:**

$$\frac{1}{\lambda} = R \left( \frac{Z^2}{n_1^2} - \frac{Z^2}{n_2^2} \right)$$

Since it is emission spectra: The electron must go to a lower energy level. Longest wavelength will be for the smallest energy difference:

Longest wavelength  $n_{\text{final}} (n_1) = 4$ ;  $n_{\text{initial}} (n_2) = 5$

Given  $\lambda = 112.54 \text{ nm} = 112.54 \times 10^{-9} \text{ m}$

Solve for Z using either Bohr's equation or Rydberg's equation given above:

$Z = 6$ ; So the atom is C, and the ion is  $\text{C}^{5+}$

Should show the equation and their work. Can get the answer through multiple different ways.

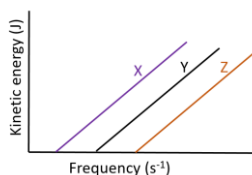
Common mistakes:

Using incorrect  $n_{\text{initial}}/n_{\text{final}}$

Using correct  $n_{\text{initial}}/n_{\text{final}}$  but incorrect calculations

All correct but final answer is the atom and not the ion

**Question 2:**

**2. 10 points**

The plot of kinetic energy of the ejected electron versus frequency, for three unknown metals (X, Y, and Z) is shown above (not drawn to scale).

$$1 \text{ eV} = 1.6022 \times 10^{-19} \text{ J}$$

**a. 8 points**

The three possible values corresponding to work functions for the metals shown, are 4.14 eV, 2.26 eV, and 2.07 eV (in no order).

Which one of the following metals (X, Y, or Z) will emit the slowest (lowest velocity) electron when excited with a photon of an ultrafast laser pulse (400 nm)?

What is the velocity ( $\text{ms}^{-1}$ ) of the ejected electron (3 significant figures)? Show your work and calculations.

X and Y would eject an electron. Y would be slowest electron since it has a higher work potential than X

$$2.07 \text{ eV} = 3.31 \times 10^{-19} \text{ J}$$

$$2.26 \text{ eV} = 3.61 \times 10^{-19} \text{ J}$$

$$4.14 \text{ eV} = 6.626 \times 10^{-19} \text{ J} - \text{no electron ejected}$$

400 nm wavelength has  $4.97 \times 10^{-19} \text{ J}$  of energy (Use,  $E = hc/\lambda$ )

$$E_{\text{photon}} = \text{KE} + \text{work function}$$

$$E_{\text{photon}} = 4.97 \times 10^{-19} \text{ J}; \text{ Work function} = 3.61 \times 10^{-19} \text{ J}$$

$$\text{So, KE} = 1.36 \times 10^{-19} \text{ J}$$

$$\text{KE} = \frac{1}{2} mu^2; \text{ Solve for } u \text{ (} m = \text{mass of electron} = 9.109 \times 10^{-31} \text{ kg)}$$

$$u = 5.46 \times 10^5 \text{ ms}^{-1} \text{ (Accept answers that are } 5.4x \text{ to } 5.50)$$

**b. 2 points**

Assume the plot shown represents the ejection of *easiest* to remove electron in metals Mg/Sr/Ca.

Which of the plots (X, Y, or Z) corresponds to Ca? Explain your reasoning.

Y corresponds to Ca – the ionization energy of Ca is in-between that of Mg and Sr.

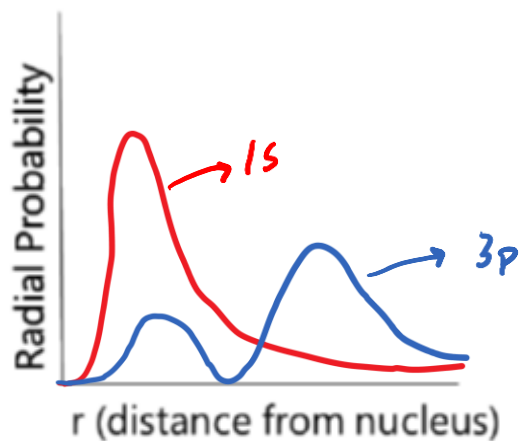
Based on the periodic table the IE decreases as we go down the group since “n” is increasing and the outer electrons experience lower  $z_{\text{eff}}$  (effective nuclear charge).

12 points

a. 6 points

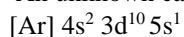
Name the orbital and draw the radial probability distribution for the following orbitals **on one plot**, for Cl (chlorine) atom (Assume ground state electronic configuration)

- I. Orbital containing the electron that requires the maximum energy to be removed. (1s)
- II. Orbital containing the electron that experiences the minimum effective nuclear charge. (3p)



b. 4 points

An unknown cation ( $A^{2+}$ ) has the following *excited* electronic configuration:



Give the ground state electronic configuration of the atom, A and the anion,  $A^{2-}$ . Show your work. (Assume all electronic transitions are possible)

$A^{2+}$ : 31 electrons; So A has 33 electrons and is As



c. 2 points

In a single atom:

How many electrons can have the following two quantum numbers?

$$n = 5; m_l = +1$$

Explain your choice, briefly.

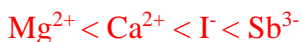
5p/5d/5f/5g all have 1 orbital where  $n = 5$  and  $m_l = +1$

8 total electrons

**Question 4****4. 10 points**

Based on what you have learned in Chem110 so far, arrange the following in increasing order, and *briefly* explain your reasoning:

A) In increasing (smallest to largest) order of size:



$\text{I}^-$ ,  $\text{Sb}^{3-}$  are isoelectronic. Higher the negative charge, larger the size due to electron-electron repulsions/lower  $z_{\text{eff}}$ . Both are isoelectronic with  $\text{Xe}/\text{Ba}^{2+}$ . Between  $\text{Ca}^{2+}$  and  $\text{Mg}^{2+}$ ; Ca and Mg are in the same group, and Mg has lower "n" (higher in the group/periodic table) and hence a smaller size. For the same reason both will be smaller than  $\text{Ba}^{2+}$  and hence smaller than  $\text{I}^-$  and  $\text{Sb}^{3-}$ .  
(Any correct chemical reasoning is acceptable)

B) In increasing (smallest to largest) order of bond angle for the central atom



Tetrahedral(109.5)

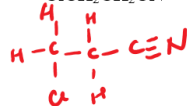
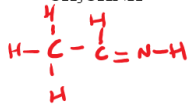
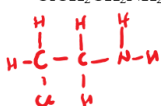
Bent (<109.5)

Linear (180)

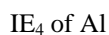
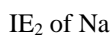
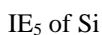


Can also show geometry

C) In increasing order (shortest to longest) of C-N bond length:



D) In increasing order (smallest to largest) of ionization energy:



All are isoelectronic cations at this point. The one with the highest charge (and highest atomic number/Z) will have the highest IE as it will be tougher to remove the electron since it has the highest  $Z_{\text{eff}}$ .



## Question 5

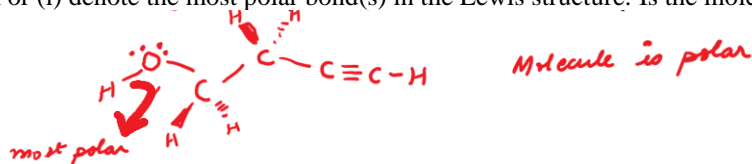
5. 10 points

a. 8 points

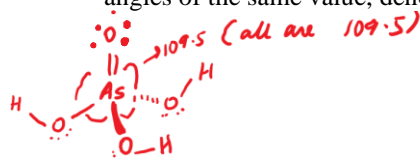
For each molecule below, draw the most stable Lewis structure showing appropriate shape and dashed/wedged bond where appropriate based on VSEPR. Show lone pairs.

(i)  $\text{HOCH}_2\text{CH}_2\text{CCH}$ 

For (i) denote the most polar bond(s) in the Lewis structure. Is the molecule polar?

(ii)  $\text{H}_3\text{AsO}_4$  (there are no O-O bonds)

For (ii) denote the *largest* bond angle in the Lewis structure. (If there are multiple bond angles of the same value, denote any one).



b. 2 points

You have discovered a new ionic compound  $\text{Y}_2\text{X}$  (Y forms 1+ cation and X forms 2- anion).

The lattice energy of  $\text{Y}_2\text{X}$  is the energy change for which process (use a chemical equation, showing reactants and products and their states, to describe the process)?



(For this question, no part marks if unbalanced equation was shown)



## Question 6

6. 6 points

Below are the subsequent ionization energies (kJ/mol) of Chromium (Cr, atomic number 24)

	IE1	IE2	IE3	IE4	IE5	IE6
Cr	652.9	1590.6	2987	4743	6702	8744

Using the ionization energy data for Chromium (Cr) given above, determine the **longest wavelength (in nm; 3 significant figures)** of a photon required to remove the first electron from the 3d orbital of the appropriate Cr cation. Show your calculations.

**Solution:**Cr electronic configuration:  $[\text{Ar}] 4s^1 3d^5$ 

IE1: removal of first 4s electron

IE2: removal of first 3d electron – This is the IE to use. IE2: 1590.6 kJ/mol

Convert to J =  $(1590.6 \text{ kJ/mol} \times 1000 \text{ J/kJ}) (1 \text{ mol}/6.023 \times 10^{23}) = 2.641 \times 10^{-18} \text{ J}$ 

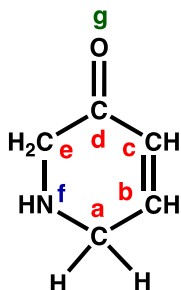
We know energy required, and can now calculate the wavelength associated with this energy

 $E = hc/\lambda$ ; solve for  $\lambda$ Wavelength =  $7.52 \times 10^{-8} \text{ m} = 75.2 \text{ nm}$  (75.0 to 75.5 accepted)



## 7. 15 points

Consider the molecule given below and note that the lone pairs are not shown in the structure.



Using chemical bonding theories that we have learnt in class (valence bond theory, hybridization), answer the following questions:

- i) **(5 points)** Consider the carbon atoms marked in red (**a**, **b**, **c**, **d**, and **e**) in the molecule, one at a time, and determine hybridization at each of the carbon centers.  
**a:  $sp^3$ ; b:  $sp^2$ ; c:  $sp^2$ ; d:  $sp^2$ ; e:  $sp^3$**
- ii) **(2 points)** Now consider the nitrogen atom marked in blue (**f**), and oxygen marked in green (**g**) and determine each of their hybridization.  
**N<sub>f</sub>:  $sp^3$ ; O<sub>g</sub>:  $sp^2$**
- iii) **(4 points)** Describe all the bonds formed by the carbon atoms marked in red as **a** and **b**. Indicate sigma/pi bonds.  
**C<sub>a</sub>: two  $sp^3$ -s  $\sigma$  bonds with H, one  $sp^3$ - $sp^2$   $\sigma$  bond with C<sub>b</sub>, and one  $sp^3$ - $sp^3$   $\sigma$  bond with N<sub>f</sub>.  
 C<sub>b</sub>: one  $sp^2$ -s  $\sigma$  bond with H, one  $sp^2$ - $sp^3$   $\sigma$  bond with C<sub>a</sub>, and one  $sp^2$ - $sp^2$   $\sigma$  bond with C<sub>c</sub>, and one p-p pi bond with C<sub>c</sub>.**
- iv) **(1 point)** How many total  $sp^3$ -s sigma bonds are in the molecule?  
**5 (no partial points)**
- v) **(2 points)** Determine electronic and molecular geometry at carbon **d** and nitrogen **f** centres.  
**C<sub>d</sub>: Electronic, trigonal planar. Molecular, trigonal planar  
 N<sub>f</sub>: Electronic, tetrahedral planar. Molecular trigonal pyramidal**
- vi) **(1 point)** Indicate the value from the choices provided below that best describes the C-N-C bond angle at nitrogen marked **f**:  
**>90° to <109.5°      109.5°      >109.5      120°      >120° to <180**

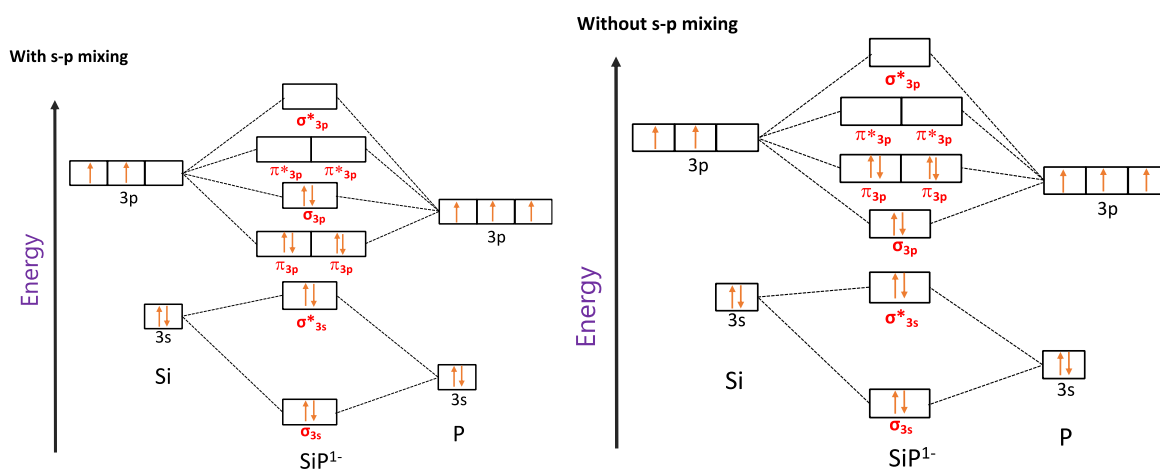
**8. 12 points** Using the valence shell orbitals only: (i) **(6 points)** Draw a **complete** molecular orbital diagram for  $\text{SiP}^{1-}$ , and then use it to answer the questions (ii to v) given below. Assume that 3s and 3p atomic orbitals will form molecular orbitals similar to those formed by the combination of 2s and 2p atomic orbitals that we learnt in class. Your molecular orbital diagram should clearly show energy arrow, atomic orbitals involved, the molecular orbitals that are formed etc. Make sure to label all the atomic and molecular orbitals.

MO Diagram of  $\text{SiP}^{1-}$ :

$\text{Si } 3s^2 3p^2$ ;  $\text{P } 3s^2 3p^3$

Total number of electrons in  $\text{SiP}^{1-} = (4+5) + 1 = 1$

Accept either of the two MO diagrams shown below: with or without sp mixing:



(ii) **(1 point)** Which one of the following will have the highest bond order:

- A)  $\text{SiP}^{1-}$       B) SiP      C)  $\text{SiP}^{2+}$

(They don't need to explain) Bond order in  $\text{SiP}^{1-}$  is  $\frac{1}{2}(8-2) = 3$ ; Bond order in SiP =  $\frac{1}{2}(7-2) = 2.5$ ; Bond order in  $\text{SiP}^{2+} = \frac{1}{2}(5-2) = 1.5$

(iii) **(1 point)** Which one of the following will have the longest bond length:

- A)  $\text{SiP}^{1-}$       B) SiP      C)  $\text{SiP}^{2+}$

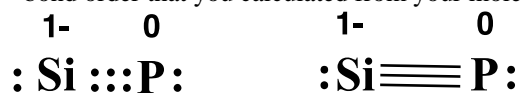
(They don't need to explain) Based on the bond orders given above,  $\text{SiP}^{2+}$  has the smallest bond order, and this will have the longest bond length.

(iv) **(1 point)** Which one of the following will be diamagnetic:

- A)  $\text{SiP}^{1-}$       B) SiP      C)  $\text{SiP}^{2+}$

(They don't need to explain) **A)  $\text{SiP}^{1-}$**  will be diamagnetic, as it has paired electrons.

(v) **(3 points)** Draw Lewis dot structure of  $\text{SiP}^{1-}$ . Does the bond order from the Lewis dot structure agree with the bond order that you calculated from your molecular orbital diagram?



It shows three bonds between SiP and it agrees with the bond order calculated for  $\text{SiP}^{1-}$  in i) above.