

# Chemistry 455A

Review Quiz  
April 6, 2007

Name \_\_\_\_\_ KEY \_\_\_\_\_

Possibly Helpful Information:

$\omega = ck$	$\hbar = 1 \cdot 10^{-34} J - \text{sec}$
$\lambda \nu = c$	$m_e = 1 \cdot 10^{-30} Kg$
$E = \hbar \omega$	$c = 3 \cdot 10^8 m / \text{sec}$
$E = \frac{p^2}{2m}$	$N_A = 6 \cdot 10^{23} \text{ molecules/mole}$
$p = \hbar k$	$V = \frac{q_1 q_2}{(4\pi\epsilon_o) r}$
$ka = \pi n$	$E_n = -22 \cdot 10^{-19} \left( \frac{Z}{n} \right)^2 J$

Each problem is worth 10 points.

Q1) (Z12:27) The work function of an element is the energy required to remove an electron from the surface of a solid. The work function for lithium is 300 kJ/mole (that is, it takes 300 kJ of energy to remove one mole of electrons from one mole of Li atoms on the surface of Li metal). What is the maximum wavelength of light that can remove an electron from an atom in lithium metal?

$$\lambda \nu = c$$

$$\Delta E = E_{\text{photon}} = \hbar \omega = h\nu = \frac{hc}{\lambda}$$

$$\lambda = \frac{hc}{\Delta E} = \frac{2\pi\hbar c}{\Delta E} = \frac{2\pi \cdot 10^{-34} \cdot \text{J} \cdot \text{sec} \cdot 3 \cdot 10^8 \text{ m/sec}}{\frac{3 \cdot 10^5 \text{ J/mole}}{6 \cdot 10^{23} \text{ molecules/mole}}} = 12\pi \cdot 10^{-34+23-5+8} \text{ m} = 38 \cdot 10^{-8} \text{ m} = 380 \text{ nm}$$

This wavelength is less than 400 nm, which is blue light, so it is UV.

Q2) (Z12:32) What is the maximum wavelength of light capable of removing an electron from a hydrogen atom in the energy state characterized by  $n=3$ ?

$$E_n = -22 \cdot 10^{-19} \left( \frac{Z}{n} \right)^2 \text{ J}$$

$$\Delta E = E_{\text{photon}} = \frac{hc}{\lambda}$$

$$\Delta E = E_{n=\infty} - E_{n=3} = 22 \cdot 10^{-19} \left( \frac{1}{3} \right)^2 \text{ J}$$

$$\lambda = \frac{hc}{\Delta E} = \frac{2\pi \cdot 10^{-34} \cdot 3 \cdot 10^8 \text{ Jm}}{22 \cdot 10^{-19} \left( \frac{1}{3} \right)^2 \text{ J}} = \frac{27\pi}{11} \cdot 10^{-34+8+19} \text{ m} = 7.7 \cdot 10^{-7} \text{ m} = 0.77 \mu$$

This wavelength is very close to 750 nm or 0.75 microns, which is red light, so it is just below red light, in the very near IR.

Q3) (Z12:72) Using only the periodic table (on the wall) write the expected ground-state electron configuration for the following:

a) The third element of Group 5A (As)

b)



As a sum check:  $18+15=33$ , and that accounts for all As electrons.

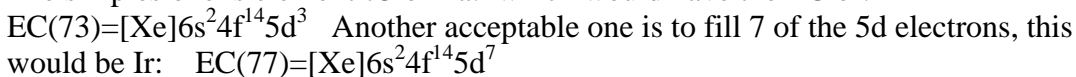
c) Element number 80



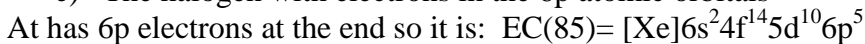
As a sum check:  $54+26=80$ , and that accounts for all Hg electrons.

d) An element with three unpaired 5d electrons

The simplest one is element 73 or Ta: which would have the EC of:

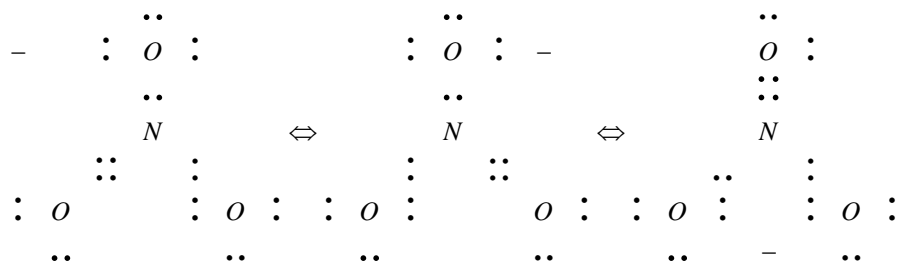
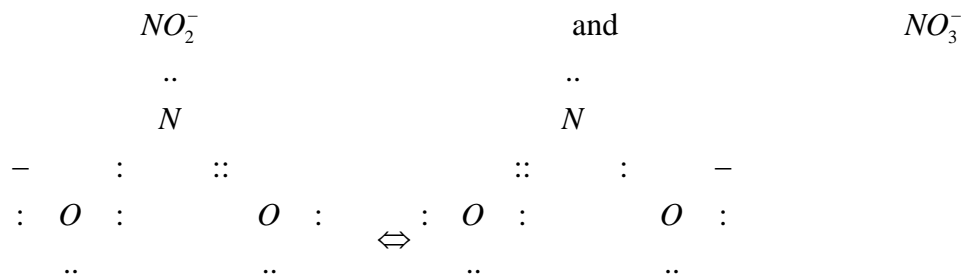


e) The halogen with electrons in the 6p atomic orbitals



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Q4) (Z13:48) Draw Lewis structures for the following two ions. Show all resonance structures, where applicable.

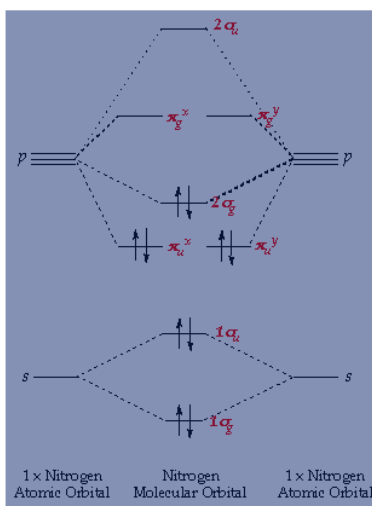


Q5) (Z14:3) Consider the MO diagram of the  $B_2$  diatomic molecule: Because  $B_2$  is known to be paramagnetic, the  $\pi_{2p}$  and  $\sigma_{2p}$  MOs must be switched from the standard order. Show an MO diagram for  $B_2$  with the electrons in the levels (label the MO levels with  $\sigma$  and  $\pi$  and s and p notation), taking into account the paramagnetic nature of this diatomic molecule:

See figure 14.40 in Zumdahl for  $B_2$ . It is much like the  $N_2$  figure shown here, with 4 fewer electrons in the diagram. For  $B_2$  there would be 2 unpaired electrons in the  $\pi_{2p}$  Orbitals. Otherwise this is the same diagram.

**Nitrogen:**

This molecule has ten electrons. The atomic orbitals combine to produce the following molecular orbital diagram



Here the  $2\pi_g$  orbital is occupied by two electrons to give a total bond order of three. This corresponds well with the Lewis structure ( $N \equiv N$ ), although the orbital approach tells us that there is one  $\sigma$  and two  $\pi$

This diagram obtained from: [www.ch.ic.ac.uk/.../course/mo\\_theory/o2ed.gif](http://www.ch.ic.ac.uk/.../course/mo_theory/o2ed.gif)

Explain how your diagram predicts that  $B_2$  is paramagnetic.

The diagram shows that the filling of the last two valence electrons goes into two distinct orbitals of the same energy (doubly degenerate), and hence they can remain unpaired. Molecules (or Atoms) with unpaired electrons will be paramagnetic.

Why might one expect the  $\sigma_{2p}$  to be lower in energy than the  $\pi_{2p}$  MO?

Ordinarily, before filling with electrons, the  $\sigma_{2p}$  orbital would be lower than the  $\pi_{2p}$  because the  $\sigma_{2p}$  puts the electrons directly between the two nuclei, and looks like a good bonding position. But as electrons are added, that region of space fills with electrons from the sigma-s  $\sigma_{2s}$  and  $\sigma^*_{2s}$  orbitals and so it is not as favorable a region of space anymore for the electrons to go, so the energy of the sigma  $\sigma_{2p}$  orbital goes up relative to the  $\pi_{2p}$ .