

Name _____

Chem 1311

Final Exam

14 December 1999

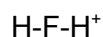
In taking this examination, you are expected to adhere to the GT academic honor code. At a minimum this requires that you utilize only the materials supplied to you (except for a non-programable calculator), and that you do not give help to, or accept help from, others.

Calculators cannot be shared. Numerical problems must show all equations used to arrive at an answer; zero credit will be given for just an answer.

1. Dissect the following molecules or ions into the acid and base that could have been used in their formation.

Acid

Base



2. Draw all diastereoisomers of the complex $[\text{Co}(\text{H}_2\text{NCH}_2\text{CH}_2\text{S})(\text{NH}_3)_2(\text{Br})_2]$. Indicate those isomers that are optically active (have nonsuperimposable mirror image forms) by drawing a box around the structure.

3. Addition of 0.001 mol of HCl to 1.0 L of neutral water results in a much greater change in pH than the addition of the same amount of HCl to 1.0 L of a solution that is 0.1 M in acetic acid and 0.15 M in sodium acetate. The pK_a of acetic acid is 4.74.

a) Calculate the change in pH for both solutions.

b) Why is the change in pH so much different in the two solutions.

4. a) Draw structures of the following molecules and ions. Your structures should: i) depict the geometry of the species, ii) show the location of all non-bonding electron pairs, if any, on all terminal atoms and, iv) show formal charges on individual atoms consistent with the number of valence shell electrons assigned to each and consistent with the overall charge on the species.
b) For each of the above species indicate what hybrid orbital set would be utilized for bonding by the central atom.

Hybrid orbital set

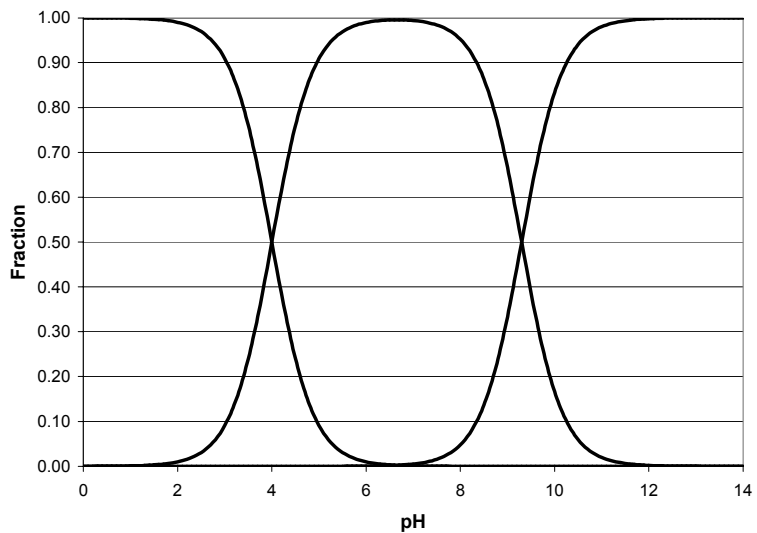


5. Draw the Lewis structure of the molecule that you predict would be formed from the reaction of NO and NO₂.

6. Shown to the right is a plot of the distribution of the species H₂A, HA⁻ and A²⁻ as a function of pH.

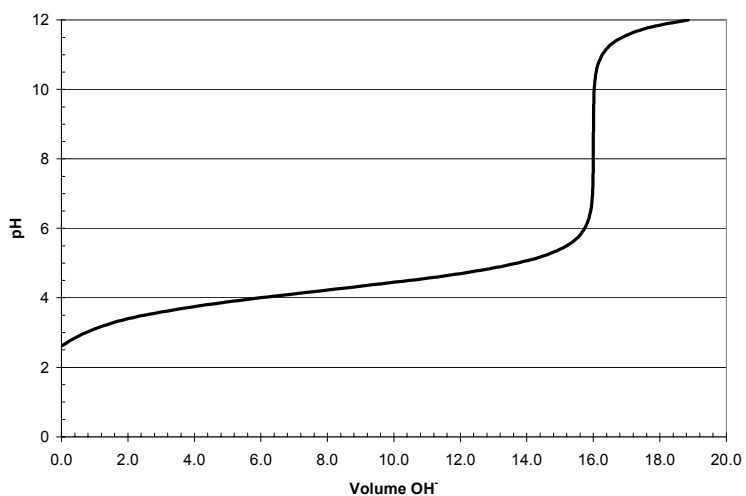
a) What are the K_a values for the first and second ionization equilibria?

b) What is the predominant species in solution at pH = 8?



7. Shown to the right is the titration curve for 24 mL of a 0.1 M solution of a weak acid.

a) What is the K_a of the acid?



b) What was the concentration of the base used for the titration?

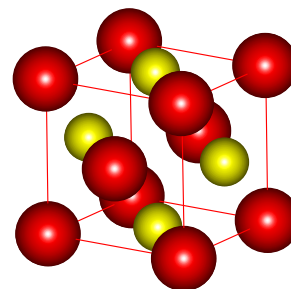
8. Draw representations of the following orbitals. Show the relative phases (signs of the wavefunction) in each lobe of an orbital, and identify nuclei and show relative sizes of portions of orbitals where appropriate.

$d_{x^2-y^2}$ (on an appropriate coordinate system)

the sigma bonding orbital in HF

the pi antibonding orbital for CN^-

9. The unit cell of a copper-gold intermetallic compound is shown to the right. The large, dark spheres are the gold atoms.
a) What is the empirical formula of the compound?

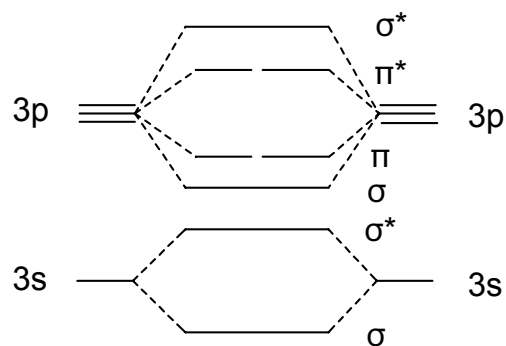


b) Assume that the unit cell is cubic with a cell dimension of 3.96 \AA . What is the density in g cm^{-3} of the compound?

10. The blue gemstone ultramarine, which is paramagnetic, has an empirical formula of $\text{Na}_6\text{Al}_6\text{Si}_6\text{O}_{14} \cdot \text{Na}_2\text{S}_4$. However, the sulfur species present, which is responsible for the color and paramagnetism, is actually the diatomic species S_2^{1-} .

a) Using the diagram to the right, show the electron configuration of S_2^{1-} .

b) Is the bond distance of the S_2^{1-} ion expected to be less than or greater than the bond distance of S_2 ? Why?



c) Should the bond distance of S_2^{1-} be less than or greater than that of O_2^{1-} ? Why?

11. The hydrogen molecule can be protonated to give H_3^+ . Assume that the ion is linear.

a) Draw the three molecular orbitals that are formed from the three hydrogen 1s orbitals.

b) Indicate the relative energy of the three molecular orbitals and the number of electrons that occupy each one.

c) Explain why this ion is considered to have electron deficient bonding.

12. Circle the answer corresponding to the correct answer for each of the following.

Which of the following molecules does not have a dipole moment?

- a) ICl_3 b) PCl_3 c) SCl_2 d) all have dipole moments

The number of unpaired electrons in the ground state of the free ion Fe^{3+} is

- a) 3 b) 4 c) 5 d) none of these

The number of unpaired electrons in the ground state of the complex ion $[\text{FeF}_6]^{4-}$ (fluoride is a weak field ligand) is

- a) 2 b) 3 c) 4 d) none of these

Which species requires the most energy for ionization?

- a) Cl^{1-} b) Ar c) K^{1+} d) all are the same

Which molecule should have bonds that are the most polar?

- a) OCl_2 b) O_3 c) SO_2 d) CO_2

Which of the following salts would be expected to have the highest lattice energy?

- a) CoF_2 b) CoF_3 c) CoCl_2 d) CoCl_3

Which of the following substances is most likely to be a solid at room temperature?

- a) SO_2 b) ClO_2 c) GeO_2 d) NO_2

Which of the following molecules will have an ionization energy that is less than that of the atom of the element of which it is composed?

- a) N_2 b) C_2 c) B_2 d) none of these

Which should be the most acidic species

- a) $\text{Al}(\text{OH}_2)_6^{3+}$ b) $\text{Ni}(\text{OH}_2)_6^{2+}$ c) $\text{Mg}(\text{OH}_2)_6^{2+}$ d) none of these is acidic

Which of the following Ti(III) complexes will have the longest wavelength absorption?

- a) $[\text{TiF}_6]^{3-}$ b) $[\text{Ti}(\text{CN})_6]^{3-}$
c) $[\text{Ti}(\text{OH}_2)_6]^{3+}$ d) all are d^1 systems and therefore absorb at the same energy

Which of the following has the largest proton affinity?

- a) F^- b) Cl^- c) Br^- d) I^-

13. Answer six (**ONLY**) of the following. Letter each answer to match the question and continue on the next page as necessary. Use balanced reactions and diagrams to illustrate your answer whenever possible. Keep your rhetoric to a minimum (40 words or less per answer).

- a) MgO has a larger lattice energy than LiF (both have the NaCl structure) although the sum of the cation and anion radii is about the same for each salt.
- b) Both $[\text{Co}(\text{NH}_3)_6]^{3+}$ and $[\text{Pt}(\text{NH}_3)_6]^{4+}$ are low-spin d^6 species, but the former is colored (orange) whereas the latter is not.
- c) Doping silicon with a small amount of arsenic greatly increases its conductivity.
- d) The pK_a of NH_3 is ca. 32; however, the pK_a of $[\text{Co}(\text{NH}_3)_6]^{3+}$ is approximately 8.
- e) Addition of trimethylamine to a solution of iodine causes the color to change from violet to yellow and the absorption maximum to shift to shorter wavelengths (higher energy).
- f) Addition of SO_3 to water results in the formation of an acidic solution.
- g) The equilibrium constant for the reaction $[\text{Cu}(\text{OH}_2)_4]^{2+} + \text{en} \rightleftharpoons [\text{Cu}(\text{OH}_2)_2(\text{en})]^{2+} + 2 \text{H}_2\text{O}$ is ca. 10^3 greater than that for the reaction $[\text{Cu}(\text{OH}_2)_4]^{2+} + 2 \text{NH}_3 \rightleftharpoons [\text{Cu}(\text{OH}_2)_2(\text{NH}_3)_2]^{2+} + 2 \text{H}_2\text{O}$. (en is ethylenediamine)
- h) In an inert solvent NaNH_2 is a much stronger base than NaOH ; however, in water they have equal basicity.