AP Chemistry
Semester 1 Final Exam Review Problems
2011-2012

Chapter 1: Matter and Measurement

Topics: Classifying Matter; Elements and Atoms; Compounds and Molecules; Physical and Chemical Properties; Physical and Chemical Changes; Units of Measurement; Precision Accuracy and Experimental Error.

1. When 10.0g of marble chips (calcium carbonate) is treated with 50.0mL of hydrochloric acid (density = 1.096g/mL), the marble dissolves, giving a solution and releasing carbon dioxide gas. The solution weighs 60.4g. How many liters of carbon dioxide gas are released if the density of the gas is 1.798g/L?

2. Some bottles of colorless liquids were being labeled when the technicians accidentally mixed them up and lost track of their contents. A 15.0mL sample withdrawn from one bottle weighed 22.3g. The technicians knew that the liquid was either acetone, benzene, chloroform, or carbon tetrachloride (which have densities of 0.792g/cm^3, 0.899g/cm^3, 1.489g/cm^3 and 1.595g/cm^3, respectively). What was the identity of the liquid?

Chapter 2: Atoms and Elements

Topics: Protons, Electrons, and Neutrons: Development of Atomic Structure; Atomic Number and Atomic Mass; Isotopes; Atomic Weight; Atoms and the Mole; The Periodic Table

3. Natural carbon, which has an atomic mass of 12.011amu, consists of carbon-12 and carbon-13 isotopes. Given that the mass of carbon-13 is 13.00335amu, what would be the average atomic mass (in amu) of a carbon sample prepared by mixing equal numbers of carbon atoms from a sample of natural carbon and a sample of pure carbon-13?

4. A monoatomic ion has a charge of 2+. The nucleus of the ion has a mass number of 62. The number of neutrons in the nucleus is 1.21 times that of the number of protons. How many electrons are in the ion? What is the name of the element?

Chapter 3: Molecules, Ions, and Their Compounds

Topics: Molecules, Compounds and Formulas; Molecular Models; Ionic Compounds: Formulas, Names and Properties; Molecular Compounds: Formulas, Names and Properties; Formulas, Compounds and the Mole; Describing Compound Formulas (Percent Composition, Empirical vs. Molecular Formula, etc.); Hydrated Compounds

5. A sample of metallic element X, weighing 3.177g combines with 0.6015L of O_2 gas (at normal pressure and 20.0^{o}C) to form the metal oxide with the formula XO. If the density of O_2 gas under these conditions is 1.330g/L, what is the mass of the oxygen? The atomic mass of oxygen is 15.9994amu. What is the atomic mass of X? What is the identity of X?

6. A sample of green crystals of nickel(II)sulfate heptahydrate was heated carefully to produce the bluish green nickel(II)sulfate hexahydrate. What are the formulas of the hydrated? If 8.753g of the heptahydrate produces 8.192g of the hexahydrate, how many grams of anhydrous nickel(II)sulfate could be obtained?

Chapter 4: Chemical Equations and Stoichiometry
Topics: Chemical Equations; Balancing Chemical Equations; Mass Relationships in Chemical Reactions: Stoichiometry; Reactions in Which One Reactant is Present in Limited Supply; Percent Yield; Chemical Equations and Chemical Analysis

7. A sample of limestone (containing calcium carbonate) weighing 438mg is treated with a solution of oxalic acid, $H_2C_2O_4$, to give solid calcium oxalate, CaC_2O_4, carbon dioxide and water. The mass of the calcium oxalate produced is 472mg. Write a balanced equation for this reaction. What is the mass percentage of calcium carbonate in this limestone?

8. Potassium superoxide, KO_2, is employed in a self-contained breathing apparatus used by emergency personnel as a source of oxygen. The reaction is
$4KO_{2(s)} + 2H_2O_{(l)} \rightarrow 4KOH_{(s)} + 3O_{2(g)}$
Say that a self-contained breathing apparatus is charged with 750.g of KO_2 and then is used to produce 195g of oxygen. Was all of the KO_2 consumed in this reaction? If the KO_2 wasn't all consumed, how much is left over and what mass of additional O_2 could theoretically be produced?

Chapter 5: Reactions in Aqueous Solution
Topics: Properties of Compounds in Aqueous Solution; Precipitation Reactions; Acids and Bases; Reactions of Acids and Bases; Gas-Forming Reactions; Classifying Reactions in Aqueous Solution; Oxidation-Reduction Reactions; Measuring Concentrations of Compounds in Solution; pH, a Concentration Scale for Acids and Bases; Stoichiometry of Reactions in Aqueous Solution

9. You order a glass of juice in a restaurant, only to discover that it is warm and too sweet. The sugar concentration of the juice is 3.47M, but you would like it reduced to a concentration of 1.78M. How many grams of ice should you add to 100.mL of juice, knowing that only a third of the ice will melt before you take the first sip? (The density of water is 1.00g/mL)

10. Mercury(II)nitrate is treated with hydrogen sulfide gas, forming a precipitate and a solution. Write the molecular equation and the net ionic equation for the reaction. An acid is formed; is it strong or weak? Name each of the products. If 81.15g of mercury(II)nitrate and 8.52g of hydrogen sulfide are mixed in 550.0g of water to form 58.16g of precipitate, what is the mass of the solution after the reaction?

Chapter 6: Principles of Reactivity: Energy and Chemical Reactions
Topics: Basic Principles of Energy (Conservation of Energy, Temperature and Heat, Heat Flow); Specific Heat Capacity and Heat Transfer; Energy and Changes of State; The First Law of Thermodynamics; Enthalpy Changes for Chemical Reactions; Calorimetry; Hess's Law; Standard Enthalpies of Formation; Product or Reactant Favored Reactions and Thermochemistry

11. Ammonia will burn in the presence of a platinum catalyst to produce nitric oxide, NO:
$4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(g)}$
What is the enthalpy of the reaction? Use the following thermochemical equations:
$N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)}$ $\Delta H = 180.6kJ$
$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$ $\Delta H = -91.8kJ$
$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$ $\Delta H = -483.7kJ$

12. When 15.3g of sodium nitrate, $NaNO_3$, was dissolved in water in a calorimeter, the temperature fell from 25.00°C to 21.56°C. If the heat capacity of the solution and the calorimeter is 1071 J/°C, what is the enthalpy change when 1mol of sodium nitrate dissolves in water? The solution process is
$NaNO_{3(s)} \rightarrow Na^+_{(aq)} + NO_3^-{}_{(aq)}$ $\Delta H = ?$

Chapter 7: Atomic Structure
Topics: Electromagnetic Radiation and It's Properties; Planc, Einstein, Energy and Photons; Atomic Line Spectra and Niels Bohr; The Wave Properties of the Electron; Quantum Mechanical View of the Atom; The Shapes of Atomic Orbitals; Atomic Orbitals and Chemistry

13. An electron microscope employs a beam of electrons to obtain an image of an object. What energy must be imparted to each electron of the beam to obtain a wavelength of 10.0pm? Obtain the energy in electron volts (eV) ($1eV = 1.602x10^{-19}J$)

14. What is the difference in energy between the two levels responsible for the violet emission line of the calcium atom at 422.7nm?

Chapter 8: Atomic Electron Configurations and Chemical Periodicity
Topics: Electron Spin (para/dia magnetism); The Pauli Exclusion Principle; Atomic Subshell Energies and Electron Assignments; Atomic Electron Configurations; Electron Configurations of Ions; Atomic Properties and Periodic Trends; Periodic Trends and Chemical Properties

15. A neutral atom has two electrons with n=1, eight electrons with n=2, eight electrons with n=3 and two electrons with n = 4. Assuming this element is in its ground state, supply the following information:
a. Atomic number b. Total number of s electrons c. Total number of p electrons
d. Total number of d electrons e. Is the element a metal, metalloid or nonmetal?

16. The configuration for an element is given here:
[Ar] (↑↓)(↑↓)(↑)(↑)(↑) (↑↓)
3d 4s
a. What is the identity of the element with this configuration?
b. Is a sample of the element paramagnetic or diamagnetic?
c. How many unpaired electrons does a 3+ ion of this element have?

Chapter 9: Bonding and Molecular Structure: Fundamental Concepts
Topics: Valence Electrons; Chemical Bond Formation; Bonding in Ionic Compounds; Covalent Bonding and Lewis Structures; Resonance; Exceptions to the Octet Rule; Molecular Shapes; Charge Distribution in Covalent Bonds and Molecules; Molecular Polarity; Bond Properties; Order, Length and Energy

17. The compound chloral hydrate, known in detective stories as knock-out drops, is composed of 14.52% C, 1.83% H, 64.30% Cl and 19.35% O by mass and has a molar mass of 165.4g/mol.
a. What is the empirical formula of this substance?
b. What is the molecular formula of this substance?
c. Draw the Lewis structure of the molecule using the fact that the Cl atoms bond to a single C atom, there is a C-C bond, and two C-O bonds in the compound.

18. Born-Haber cycles were used to obtain the first reliable values for electron affinity by considering it the unknown and using a theoretically calculated value for the lattice energy. Use a Born-Haber cycle for KF and the following values to calculate a value for the electron affinity of fluorine:

$K_{(s)} \rightarrow K_{(g)}$ $\Delta H^o = 90kJ$
$K_{(g)} \rightarrow K^+_{(g)}$ $\Delta H^o = 419kJ$
$F_{2(g)} \rightarrow 2F_{(g)}$ $\Delta H^o = 159kJ$
$K_{(s)} + ½\ F_{2(g)} \rightarrow KF_{(s)}$ $\Delta H^o = -569kJ$
$K^+_{(g)} + F^-_{(g)} \rightarrow KF_{(s)}$ $\Delta H^o = -821kJ$

Chapter 10: Bonding and Molecular Structure: Orbital Hybridization and Molecular Orbitals
Topics: Orbitals and Bonding Theories; Valence Bond Theory (Orbital Overlap, Hybridization of Atomic Orbitals, Multiple Bonds, sigma/pi bonds); Molecular Orbital Theory (Bonding, Anti-bonding and Nonbonding Molecular Orbitals)

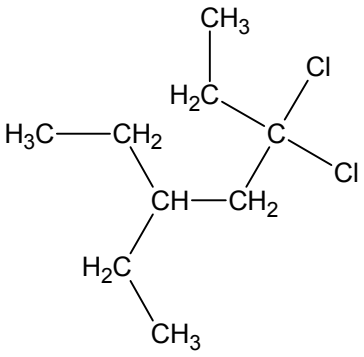
19. Draw the Lewis structure and then specify the electron-pair and molecular geometries for each of the following molecules or ions. Identify the hybridization of the central atom.
a. $XeOF_4$ b. BrF_5 c. OSF_4 d. Central Br in Br_3^-

20. Nitrogen, N_2, can ionize to form N_2^+ or add an electron to give N_2^-. Using molecular orbital theory, compare these species with regard to
a. Their magnetic character b. Net number of pi bonds c. Bond order
d. bond length e. Bond strength

Chapter 11: Carbon: More Than Just Another Element
Topics: Alkane, alkene and alkyne nomenclature; Polarity and hydrogen bonding; Isomers (structural, stereo, optical, cis/trans); Lewis structures; Functional groups; Aromatic vs aliphatic compounds; Fundamental organic reactivity; Polymerization.

21. Draw suitable structural formulas to show that there are four structural isomers of $C_3H_6Cl_2$ and then name them.

22. Give an acceptable name for the following:



23. Draw the structural formula for each of the following compounds.
a. 2-chloro-3-methylpentane
b. butanone
c. 1,1,1-chlorodifluoroethane (a refrigerant)
d. 2-methyl-1,3-butadiene (used to make elastomers)
e. p-nitrophenol
f. cyclohexene
g. isopropyl ethyl ether
h. propionaldehyde

24. Indicate the principal product(s) you would expect in
a. exposing a mixture of chlorine and propane gases to ultraviolet light.
b. heating a mixture of isopropyl alcohol and benzoic acid (using H_2SO_4 catalyst)
c. reacting 2-butene with gaseous HCl

25. For each, indicate which has:
a. the higher boiling point; C_6H_{12} or C_6H_6
b. the greater solubility in water; C_3H_7OH or $C_7H_{15}OH$

26. Draw the structural formula for:
a. 2-hydroxy-4-isopropyltoluene (thymol-flavor constituent of the herb thyme)
b. the ester methyl salicylate (oil of wintergreen) [Hint: Salicylic acid is o-hydroxybenzoic acid]

27. Compound A is an alcohol of formula $C_5H_{12}O$ that can be resolved into optical isomers. Draw three possible structures of A.