Practice Final

CHEM 200/202 **Final Exam**

May 12, 2012 **Name**______________________________________**Lab Sect. No.** ________

Please mark your answers on the scantron sheet using a #2 pencil and also mark your answers on the exam itself. You may use the blank pages at the back of the exam as scratch paper for calculations.

Mark form 'A' on your scantron.

1. The current price of zinc is \$0.002 per gram. How much zinc chlorate contains \$1.25 worth of zinc?
   (a) 2220 g
   (b) 625 g see Exam 1 #3
   (c) 2526 g
   (d) 1303 g
   (e) 1250 g

2. A detailed explanation of natural phenomena that is generally accepted and has been extensively tested is called a _____.
   (a) theory
   (b) hypothesis see Exam 1 #5
   (c) law
   (d) fact
   (e) postulate

3. What is the total molar concentration of chloride in solution when 5.00 grams of sodium chloride and 5.00 grams of barium chloride are dissolved in a beaker containing 100.0 mL of water? Assume that there is no change in volume when the salts are dissolved.
   (a) 0.856 *M*
   (b) 2.57 *M*
   (c) 1.95 *M* see Exam 1 #10
   (d) 1.34 *M*
   (e) 1.10 *M*

4. The thermite reaction is the exothermic reaction of iron(III) oxide with aluminum metal; the product of the reaction is pure iron metal and aluminum oxide. If 35.7 g of iron(III) oxide and 49.1 g of aluminum are reacted how many grams of aluminum oxide are formed and how many grams of the excess reagent (XS) remain unreacted?
   (a) Aluminum oxide: 45.6 g ; XS: 41.3 g
   (b) Aluminum oxide: 22.8 g ; XS: 37.0 g
   (c) Aluminum oxide: 35.0 g ; XS: 27.2 g
   (d) Aluminum oxide: 35.0 g ; XS: 30.6 g see Exam 1 #11, Practice Exam 1 #11,
   (e) Aluminum oxide: 22.8 g ; XS: 43.1 g & Practice Final #5

5. Calculate the volume of phosphoric acid (1.20 *M*) required to titrate 25.0 mL of a 0.900 *M* magnesium hydroxide solution
   (a) 9.38 mL
   (b) 12.5 mL
   (c) 6.25 mL see Exam 1 #17, Practice Exam 1 #17 & Practice Final #6
   (d) 18.8 mL also Practice Quiz 2 #3
   (e) 25.0 mL

6. What is the oxidation number for sulfur in sulfuric acid?
   (a) –2
   (b) 0 see Exam 1 #18 & Practice Exam 1 #18
   (c) 2 also Practice Quiz 3 #1,
   (d) 4
   (e) 6

7. What is the net ionic reaction for reaction that proceeds when silver nitrate and sodium sulfate are combined in solution? The solubility rules are provided at the end of the exam.
   (a) There is no reaction they are all spectator ions.
   (b) $2Ag^+(aq) + SO_4^{2-}(aq) \rightarrow Ag_2SO_4(s)$
   (c) $Na^+(aq) + SO_4^{2-}(aq) \rightarrow Na_2SO_4(s)$ see Exam 1 #14 & Practice Exam 1 #15
   (d) $Ag^+(aq) + NO_3^-(aq) \rightarrow AgNO_3(s)$
   (e) $AgNO_3(aq) + Na_2SO_4(aq) \rightarrow Ag_2SO_4(s) + NaNO_3(s)$

8. Select the correct statements regarding the following balanced redox reaction: (Mark all correct statements)

$$2HCl(aq) + 2FeCl_2(aq) + H_2O_2(aq) \rightarrow 2FeCl_3(aq) + 2H_2O(aq)$$

   (a) H in $HCl(aq)$ is the reducing agent.
   (b) Cl in $FeCl_2$(aq) is oxidized. see Exam 1 #19 & Practice Final #8
   (c) Fe in $FeCl_2(aq)$ is the reducing agent. and Practice Exam 1 #19
   (d) O in $H_2O_2(aq)$ is reduced. also Practice Quiz 3 #2 & Quiz
   (e) Cl in $FeCl_2(aq)$ is the oxidizing agent.

9. Which is not a basic assumption of the kinetic molecular theory?
   (a) The total kinetic energy of gas particles is constant, at a fixed temperature.
   (b) Gas particles are assumed to occupy no volume.
   (c) Individual gas particles never change in their kinetic energy.
   (d) The motion of gas particles is constant, random, and in a straight line.
   (e) Collisions between particles are elastic. see Exam 1 #24

10. A sample of ammonia gas (NH_3) occupies a volume of 3.82 L at pressure of 595 torr and at 206 K. If the temperature is increased by 62.0 °C how must the volume of the gas change in order for the pressure not to change?
   (a) Increase by 8.87 L
   (b) Increase by 4.97 L see Exam 1 #23 & practice exam 1 #27
   (c) Decrease by 1.30 L
   (d) Increase by 1.15 L
   (e) Decrease by 5.12 L

11. If it takes 30 minutes for 2.5 L of fluorine gas to diffuse through a screen at 50 °C, how much butane gas (C_4H_{10}) will diffuse through the same screen at the same temperature in 45 minutes?

$$\frac{rate_A}{rate_B} = \sqrt{\frac{M_B}{M_A}}$$

   (a) 24.2 L
   (b) 3.03 L
   (c) 4.04 L see Exam 1 #27
   (d) 0.0223 L
   (e) 1.03 L

12. How would you describe, in terms of q and ΔE, a balloon contains a gas that expands and gets colder?
   (a) $q < 0$, $\Delta E > 0$
   (b) $q < 0$, $\Delta E < 0$
   (c) $q > 0$, $\Delta E < 0$ see Exam 2 #4 & Ch. 6 HW Part 1 #8
   (d) $q > 0$, $\Delta E > 0$
   (e) $q = 0$, $\Delta E > 0$

13. Which of the following statements are true regarding the following balanced equation and enthalpy of reaction?

$$2C(g) + 2H_2(g) \rightarrow C_2H_4(g) \qquad \Delta H_{rxn} = +52.47 \text{ kJ}$$

   (a) This reaction is exothermic. see Exam 2 #2, practice exam 1 #2 & Practice Final #13
   (b) The products are more stable than the reactants.
   (c) If the reaction were done in a calorimeter, the temperature of the calorimeter would increase.
   (d) If the reaction proceeds in vessel of constant volume, the pressure will decrease.
   (e) The reaction releases heat.

14. Of the quantum numbers below, which identifies the second electron that can be added to a silicon atom?
    (a) $n = 3, l = 2, m_l = 0, m_s = -\frac{1}{2}$ see Exam 2 #20
    (b) $n = 2, l = 1, m_l = 1, m_s = -\frac{1}{2}$
    (c) $n = 3, l = 1, m_l = -1, m_s = -\frac{1}{2}$
    (d) $n = 3, l = 0, m_l = 1, m_s = -\frac{1}{2}$
    (e) $n = 3, l = 1, m_l = -1, m_s = +\frac{1}{2}$

15. Zinc sulfate ($ZnSO_{4(s)}$) can be prepared from elemental zinc, sulfur and oxygen. From the reaction below, determine the ΔH_{rxn} for the production of one mole of $ZnSO_4(s)$:

$$Zn(s) + 1/8S_8(s) \rightarrow ZnS(s) \qquad \Delta H = -183.92 \text{ kJ}$$
$$2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g) \qquad \Delta H = -927.54 \text{ kJ}$$
$$2SO_3(g) \rightarrow 2SO_2(g) + O_2(g) \qquad \Delta H = 196.04 \text{ kJ}$$
$$ZnSO_4(s) \rightarrow ZnO(s) + SO_3(g) \qquad \Delta H = 230.32 \text{ kJ}$$

    (a) –976.03 kJ
    (b) –638.70 kJ see Exam 2 #3, practice exam 2 #3, & Practice Final #14
    (c) –1077.05 kJ also Practice Quiz 4 #3
    (d) –1952.6 kJ
    (e) –684.10 kJ

16. When stoichiometric amounts of elemental sulfur and oxygen gas react in a calorimeter to produce 8.00 g of sulfur dioxide gas, the calorimeter temperature rises 9.40°C. The calorimeter contains 0.943 kg of solution with a specific heat capacity of 4.18 J/g·K. Calculate the heat of formation of sulfur dioxide, ΔH_f, in kJ/mol.
    (a) 71.2 kJ/mol
    (b) –37.0 kJ/mol see Exam 2 #10 & Practice Final #15
    (c) –314 kJ/mol also Practice Quiz 5 #3
    (d) –296 kJ/mol
    (e) 37.0 kJ/mol

17. From the list of electronic transitions for the hydrogen, given below, which is the result of the absorption of light with the largest frequency?
    (a) $n = 6$ to $n = 5$ see Exam 2 #14
    (b) $n = 4$ to $n = 3$
    (c) $n = 7$ to $n = 8$
    (d) $n = 1$ to $n = 2$
    (e) $n = \infty$ to $n = 1$

18. Which of the quantum number possibilities listed below are allowable for a ground state chlorine atom? Mark all correct answers.
    (a) $n = 1, l = 0, m_l = 0$
    (b) $l = 3, m_l = 1, m_s = -\frac{1}{2}$ see Exam 2 #24
    (c) $n = 3, l = 2, m_l = -2$
    (d) $n = 2, l = 1, m_s = -\frac{1}{2}$
    (e) $n = 2, l = 1, m_l = 1, m_s = -\frac{1}{2}$

19. Cations are typically formed from atoms with? see Exam 2 #26 and practice exam 2 #29

(a) Low first ionization energies and very negative electron affinities.
(b) High first ionization energies and slightly positive electron affinities.
(c) Low first ionization energies and moderately negative electron affinities.
(d) High first ionization energies and moderately negative electron affinities
(e) Ionization energies and electron affinities have no role in the formation of cations.

20. A beam of arsenic atoms is accelerated to such a speed that they have the same wavelength as a photon with a frequency of 4.51×10^{19} Hz. What is the average speed of the arsenic atoms?

(a) 1.33×10^{-24} m/s
(b) 8.01×10^{2} m/s
(c) 7.20×10^{13} m/s
(d) 60.0 m/s see Exam 2 #15
(e) 0.801 m/s also Practice Quiz 6 #2

21. Which one of the following properties is characteristic of typical ionic compounds? Mark all that apply.

(a) high melting point
(b) low boiling point see Exam 3 #1, Practice Exam 3 #1 & Practice Final #21
(c) brittleness
(d) good electrical conductor when solid
(e) good electrical conductor when molten

22. Choose the correct words to fill in the blanks in the following sentence. The lattice energy for ionic crystals increases as the charge on the ions ____________ and the size of the ions ________________.

(a) increases, increases
(b) increases, decreases
(c) decreases, increases
(d) decreases, decreases
(e) Choices (a) to (d) are not appropriate answers.

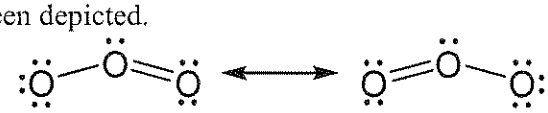
23. Of the following molecules, which has bonds with the greatest polarity?

    (a) SO_3 see Exam 3 #8, Practice Quiz 7 #5
    (b) P_2O_5
    (c) SiO_2
    (d) ClO_2
    (e) CO_2

24. Which of the following molecules has a see-saw shape?

    (a) SiH_4 see Exam 3 #17, Practice Exam 3 #13 & Practice Final #26
    (b) ICl_4^-
    (c) ICl_4^+
    (d) XeF_4
    (e) SOF_4

25. Ozone can be described by the resonance structures shown below. What does this imply? Note that formal charges have not been depicted.



    (a) The two bonds in ozone are of equal length and the electronic distribution in the O–O bonds is identical.
    (b) The single bond is longer than the double bond and the electronic distribution in the two O–O bonds is different. see Exam 3 #15 & Practice Exam 3 #12
    (c) An electron pair in the ozone molecule alternates back and forth between the two O–O bonds so that the two different bonds seem to exchange positions.
    (d) The ozone molecule revolves so that the two different bonds seem to exchange positions.
    (e) The ozone molecule is unstable.

26. Which of the following species is non-polar? Mark all correct choices.

    (a) CF_4
    (b) XeO_4
    (c) SO_2 see Exam 3 #11, Practice Exam 3 #8 & Practice Final #23
    (d) SF_4
    (e) BH_4^-

27. What is the average bond order of each Cl–O bond for the perchlorate ion, ClO_4^-?

    (a) 1 see Exam 3 #18, Practice Exam 3 #18 & Practice Final #27
    (b) 5/4
    (c) 6/4
    (d) 7/4
    (e) 2

28. What is the hybridization of the nitrogen atom in ammonia, NH_3?

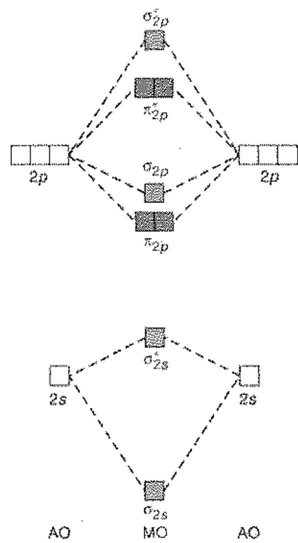
(a) *s* see Exam 3 #23, Practice Exam 3 #23 & Practice Final #28
(b) *sp*
(c) sp^2
(d) sp^3
(e) sp^4

29. Which of the following statements regarding molecular orbital (MO) theory is *false*?

(a) MO theory uses orbitals which are delocalized over the whole molecules.
(b) The number of atomic orbitals combined equals the number of molecular orbitals formed.
(c) Bonding molecular orbitals have a region of high electron density between the nuclei.
(d) Antibonding orbitals have a node between the nuclei.
(e) Antibonding molecular orbitals are lower in energy than the atomic orbitals which form it.

Error with test. Answer was announced in class. Answers above shown as intended.

30. Using the MO diagram below, determine which of the following species is diamagnetic.



(a) B_2 see Exam 3 #28-30 & Practice Exam 3 #28
(b) B_2^+
(c) C_2^+
(d) N_2^-
(e) N_2

31. In which of the following molecules are dispersion forces the strongest intermolecular force? Mark all correct answers.

(a) CH_3OCH_3 see Practice Final #31
(b) $CH_3CH_2CH_2CH_3$
(c) HCl
(d) $HOCH_2CH_2OH$
(e) CF_4

32. Liquid ammonia can be used as a refrigerant and heat transfer fluid. How much energy is needed to heat 25.0 g of $NH_3(l)$ from –75.0°C to –0.0°C?

bp = –33.4 °C mp = –77.7 °C
ΔH_{vap} = 23.5 kJ/mol ΔH_{fus} = 5.64 kJ/mol
$c(s)$ = 2.4 J/g·K $c(l)$ = 4.7 J/g·K $c(g)$ = 2.2 J/g·K

(a) 41 kJ
(b) 6800 kJ see Practice Final #32 & Practice Quiz 10 #1
(c) 12 kJ also Ch. 12 HW Part 1 #10-12
(d) 39 kJ
(e) 15 kJ

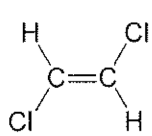
33. A liquid is considered to be boiling when ______________.

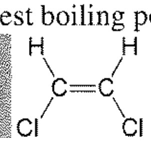
(a) it begins to vaporize see Ch. 12 Lecture Notes
(b) its vapor pressure equals one atmosphere
(c) its vapor pressure equals the pressure applied on the liquid
(d) equilibrium between the liquid and vapor phases exists
(e) it is evaporating at a rate equal to the specific vapor velocity

34. What types of forces exist between molecules of CH_2O?

(a) hydrogen bonding only see Practice Final #34 & Practice Quiz 10 #2&3
(b) hydrogen bonding and dispersion forces
(c) dipole-dipole forces only also Ch. 12 HW Part 1 #13-20
(d) dipole-dipole and dispersion forces
(e) dispersion forces only

35. Which of the following compounds is likely to have the highest boiling point?

(a) $H_2C{=}CH_2$
(b) He
(c) 

(d) 

(e) CH_4

see Practice Final #35

also Ch. 12 HW Part 2 #4

36. Nickel crystallizes in a cubic lattice which contains 4 atoms per unit cell. How many nearest neighbors surround each Ni atom?

(a) 2 see Practice Final #36 & Ch. 12 HW Part 2 #7 & 12
(b) 6
(c) 8
(d) 10
(e) 12

37. Iridium has the highest density of any element. It crystallizes in the face-centered cubic lattice with a unit cell edge length of 383.9 pm. Calculate the atomic radius of iridium.

(a) 83.1 pm see Ch. 12 HW Part 2 #16,18,19
(b) 135.7 pm
(c) 166.2 pm
(d) 192.0 pm
(e) 383.9 pm

38. Which of the following definitions is correct? see Ch. 13 Learnsmart

(a) The covalent radius is half the distance between the nuclei of two identical non-bonded atoms.
(b) The bond length is the distance between two non-bonded atoms in adjacent molecules.
(c) The van der Waals distance is the shortest distance between two covalently bonded atoms in the same molecule.
(d) The van der Waals radius is half the distance between the nuclei of two covalently bonded atoms.
(e) None of the above statements are correct.

39. Which of the following statements correctly describe the heat of solvation for a solution process? Mark all correct answers. See Ch. 13 HW Part 1 #4,8 & Ch. 13 Learnsmart

(a) The heat of solvation is the enthalpy change that occurs when a solute particle is surrounded by solvent molecules.
(b) The heat of solvation depends only on the solvent in the solution.
(c) The heat of solvation is the overall enthalpy change that occurs when a solution forms.
(d) The heat of solvation is called the heat of hydration if water is the solvent.
(e) In general, ions with greater charge densities will have more negative heats of hydration.

40. What is the strongest type of intermolecular force between $HOCH_2CH_2OH$ and CH_3OH in solution?

(a) hydrogen bond
(b) dispersion see Practice Final #40 & Ch. 13 HW Part 1 #5
(c) dipole-dipole
(d) dipole-induced dipole
(e) ion-dipole

41. The Henry's law constant (k_H) for O_2 in water at 20 °C is 1.28×10^{-3} mol/(L·atm). How many grams of O_2 will dissolve in 3.75 L of water that is in contact with pure O_2 at 1.00 atm?

(a) 0.128 g from Ch. 13 lecture question & Ch. 13 HW Part 1 #12
(b) 0.0768 g
(c) 0.0410 g
(d) 0.154 g
(e) 0.102 g

42. What is the molality of a solution prepared by dissolving 18.5 g of calcium nitrite in 83.5 g of distilled water?

(a) 0.0342 *m* see Practice Final #42 & Ch. 13 HW Part 1 #14-15 & Part 2 #1,4 #
(b) 0.0855 *m*
(c) 2.57 *m*
(d) 1.35 *m*
(e) 1.68 *m*

43. Which of the following is *not* a colligative property?

(a) density
(b) vapor pressure lowering see Ch. 13 Lecture Notes
(c) boiling point elevation
(d) freezing point depression
(e) osmotic pressure

44. From the following list of aqueous solutions and water, select the one with the highest boiling point.

(a) 1.0 *m* potassium nitrate see Ch. 13 HW Part 1 #3 & Part 2 #8,9,10
(b) 0.75 *m* sodium chloride
(c) 0.75 *m* copper(II) chloride
(d) 2.0 *m* sucrose ($C_{12}H_{22}O_{11}$)
(e) Pure water

Solubility Rules for Ionic Compounds in Water

Soluble Ionic Compounds

1. All common compounds of Group 1A(1) ions (Li^+, Na^+, K^+, etc) and ammonium ion are soluble.
2. All common nitrates (NO_3^-), acetates (CH_3COO^- or $C_2H_3O_2^-$), and most perchlorates (ClO_4^-) are soluble.
3. All common chlorides (Cl^-), bromides (Br^-) and iodides (I^-) are soluble, *except* those of Ag^+, Pb^{2+}, Cu^+, and Hg_2^{2+}.
4. All common sulfates (SO_4^{2-}) are soluble, except those of Ca^{2+}, Sr^{2+}, Ba^{2+}, Ag^+, and Pb^{2+}.

Insoluble Ionic Compounds

1. All common metal hydroxides are insoluble, *except* those of Group 1A(1) and the larger members of Group 2A(2)(beginning with Ca^{2+}).
2. All common carbonates (CO_3^{2-}) and phosphates (PO_4^{3-}) are insoluble, *except* those of Group 1A(1) and NH_4^+.
3. All common sulfides are insoluble *except* those of Group 1A(1), Group 2A(2) and NH_4^+.