**CHE1031 Mock Final Exam Key**

Be sure to use all the tables and references necessary; tables will be provided for you during the actual final exam. To thoroughly prepare for the final you should also check terms, definitions & concepts presented in each exam section.

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| ***EXAM 1: Lectures 1 & 2*** |

**Lecture 1: Fundamentals to measurement**

**1.1: Science & chemistry**

**What is matter?**

1. What two forms of matter are considered “pure substances”?

2. A separation process that depends on the differing abilities of substances to form

gases is called \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

a. filtration

b. solvation

c. distillation

d. chromatography

e. all of the above

3. A martini, no olive, is an example of a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ mixture

4. Of the following, only \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ is a chemical reaction.

a. melting of lead

b. dissolving sugar in water

c. tarnishing of silver

d. crushing of stone

e. dropping a penny into water

**1.2: Measurement, metric units & prefixes**

5. How many femtometers to the centimeter?

6. Define the term ‘least count’ used in describing an instrument’s precision.

7. In which of the following numbers are all of the zeros significant?

a. 100.090090 b. 0.05843c. 0.1000d. 00.0030020

8. Compute the answer with the proper number of significant figures? (1815 - 1806)x(9.11x7.92) =

9. Represent this number using scientific notation: 0.0000014368.

**Dimensional analysis (or conversions)**

10. Calculate the volume (cm3 or mL) of a 63.4 g piece of metal that has a density of 12.86 g/cm3.

**Lecture 2: Atoms, Molecules & Formulas**

**2.1: Dalton’s atomic theory**

11. Which is NOT a postulate of Dalton’s atomic theory?

a. Each element is composed of tiny indivisible particles called atoms.

b. All atoms of a given element are identical to each other, and different from those of other elements.

c. During a chemical reaction atoms are changed into different atoms.

d. Compounds are formed when atoms of different elements combine.

e. Atoms of an element are not changed into different atoms by chemical

reactions.

**2.2: Molecules & formulas**

12. What Law defines or describes a molecule? State that Law.

13. Which pair of substances illustrate the Law of Multiple Proportions?

a. SO2, H2SO4

b. CO, CO2

c. H2O, O2

d. CH4, C6H12O6

e. NaCl, KCl

**2.3: Discovery of atomic structure**

14. Describe Rutherford’s gold foil experiment and what it proved & disproved.

15. In the Modern Nuclear Model of the Atom, developed by Rutherford,:

a. the heavy subatomic particles, protons & neutrons, reside in the nucleus.

b. the three principle subatomic particles (protons, neutrons & electrons) all have essentially the same mass

c. the light subatomic particles, protons & neutrons, reside in the nucleus

d. mass is spread essentially uniformly throughout the atom

e. the three principle subatomic particles (protons, neutrons & electrons) all have essentially the same mass & mass is spread essentially uniformly throughout the atom

**2.4: Sub-atomic particles & the periodic table**

16. How many protons, neutrons & electrons in an atom of platinum, Pt?

**2.5: Atomic weights**

17. The element X has three naturally occurring isotopes as shown in the table below. What is the average atomic mass of element X?

isotope abundance mass 221X 74.22 220.9 220X 12.78 220.0 218X 13.00 218.1

**Periodic tables**

**Ions & ionic compounds**

18. Aluminum reacts with a non-metal to form an ionic compound with the general formula AlX. Element X is a diatomic gas at room temperature. What is element X?a. oxygenb. fluorinec. chlorined. nitrogen

e. sulfur

**Naming chemical compounds**

19. The correct name for K2S is:

a. potassium sulfateb. potassium disulfidec. potassium bisulfided. potassium sulfide

e. dipotassium sulfate

20. Name this acid: H2SO3

21. Write the formula for perchloric acid.

***EXAM 2: Lectures 3, 4 & 5***

**Lecture 3: Moles, formulas, equations, stoichiometry & limiting reactants**

**3.1: Moles & molecular weight**

22. How many moles of sodium carbonate contain 1.773x1017 carbon atoms?

23. How many atoms of hydrogen in 2.0 moles of magnesium hydroxide?

24. Calculate the molecular weight of sodium carbonate.

**3.2: Molar conversions: percent composition, empirical vs. molecular formula**

25. Calculate the percent composition of C12H22O11.

26. What is the empirical formula of a compound that is 64.8% C, 13.6% H and 21.6% O by mass?

**3.3: Stoichiometry: balancing chemical equations**

27. Balance the chemical equation: Al(NO3)3 + Na2S -->

**3.4: Patterns of chemical reactivity**

28. Label each reaction by type: a. S + 3F2 --> SF6 b. CH4 + O2 --> CO2 + H2O c. PbCO3 --> PbO + CO2

d. 2NaCl + CaO --> Na2O + CaCl2

**3.5: Stoichiometry & conversions**

29. If hydrogen and nitrogen gases react by this Haber equation: N2 + 3H2 --> 2NH3,  
 How many g of N2 are required to completely react with 9.3 g of H2?

30. Lithium and nitrogen react (combine) to form lithium nitride: 6Li + N2 --> 2Li3N

What is the theoretical yield of lithium nitride when 3.50 g of each reactant are combined?

**Lecture 4: Aqueous solution chemistry**

**4.1: Solutes & solvents**

31. Which are weak electrolytes, and therefore dissociate incompletely in water?

a. HCl b. CH3CO2H

c. NH3 d. KCl

**4.2: Solution concentration & stoichiometry**

32. What mass of potassium chloride (g) is contained in 430 mL of an 0.430 M solution?

33. What volume (mL) of a concentrated solution of 6.00 M sodium hydroxide must be diluted to 200 mL to make a 1.50 M solution of NaOH?

**4.3: Precipitation reactions**

34. A chemist (weary from a night of intense studying) reacts aqueous solutions of CoBr2 and AgNO3.

a. Write the balanced chemical equation.b. Write the complete ionic equation.c. Write the net ionic equation.d. Identify the spectator ions.e. Identify any insoluble products.

**4.4: Acids, bases & neutralization reactions**

35. All strong acids are strong \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

36. Which is a weak acid?

a. HNO3

b. HCl

c. HI

d. HF

e. HClO4

37. What base does not contain a hydroxide ion?

38. In the neutralization reaction of HCl and Na2S what gas is formed? Write a balanced equation for the reaction, and identify the gaseous product.

39. Write balanced, complete & net ionic equations for the reaction of phosphoric acid with barium hydroxide.

40. A chemist titrates a volume of 31.5 mL of a solution of HNO3 of unknown concentration with NaOH. It takes 23.9 mL of a 0.0134 M NaOH solution to reach the endpoint of titration. What was the molar concentration (M) of the HNO3 acid?

**4.5Redox reactions**

41. Define oxidation and reduction by telling me in which reaction electrons are lost

and in which electrons are gained?

42. Which is an oxidation-reduction reaction? a. Cu + 2AgNO3 --> 2Ag + Cu(NO3)2 b. HCl + NaOH --> H2O + NaCl c. AgNO3 + HCl --> AgCl + HNO3 d. Ba(C2H3O2)2 + Na2SO4 --> BaSO4 + 2NaC2H3O2 e. H2CO3 + Ca(NO3)2 --> 2HNO3 + CaCO3

43. Oxidizing agents \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ electrons.

44. In which species does sulfur have the highest oxidation number?

a. S8

b. H2S

c. SO2

d. H2(SO3)

e. K2(SO4)

45. Which of the following pairs will NOT be spontaneous (will not occur)?

a. Mg + HCl

b. Ag + H(NO3)

c. Ni + H2(SO4)

d. Al + HBr

e. Zn + HI

**Lecture 5: Electrochemistry**

**5.1: Redox agents & half-equations**

46. What is the reducing agent In the reaction shown here:

Pb + PbO2 + 2H2(SO4) 🡪 2Pb(SO4) + 2H2O

47. The reaction that reduces dichromiun ion to chromium metal involves \_\_\_\_\_\_\_\_

electrons.

48. Use half equations to determine which is oxidized and which is reduced in this reaction that occurs in an acidic aqueous solution. Be sure to complete the problem.

MnO4-1 + CH3OH 🡪 Mn+2 + H(CO2H)

**5.2: Voltaic cells**

49. In a voltaic cell, electrons flow from the \_\_\_\_\_\_\_\_\_ to the \_\_\_\_\_\_\_\_\_\_\_\_.

**5.3: Batteries**

50. Write an equation for the half reaction that occurs at the anode of the balanced

chemical equation shown below:

3MnO4-1 + 24H+1 + 5Fe 🡪 3Mn+2 + 5Fe+3 + 12H2O

51. Calculate the E°Cell value of the voltaic cell that uses the reaction shown here:

2Cr + 3Fe+2 🡪 3Fe + 2Cr+3 reaction V

Cr+3 + 3e- 🡪 Cr -0.74

Fe+2 + 2e- 🡪 Fe -0.440

Fe+3 + 3e- 🡪 Fe+2 +0.771

Sn+4 + 2e- 🡪 Sn+2 +0.154

**5.5: Corrosion**

52. Corrosion of iron is retarded by \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

a. salt

b. high pH

c. low pH

d. salt & high pH

e. salt & low pH

53. Describe how the process of cathodic protection works.

**5.6: Electrolysis**

54. What input is required for electrolysis?

55. In electrolysis, metals are deposited at the \_\_\_\_\_\_\_\_\_\_\_\_.

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| ***EXAM 3: Lectures 6 & 7*** |

**Lecture 6: Sub-atomic & quantum structure**

**6.1: Atomic properties from electron configuration**

56. Other than having the same number of valence electrons, elements in the same column of the periodic table share \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ and \_\_\_\_\_\_\_\_\_\_\_\_\_.

(I’m looking for two answers here)

**6.2: The true nature of the atom?**

**6.3: Developing a new physics for the atom**

**6.4: Bohr’s quantum planetary model**

57. In a hydrogen atom, an electron in a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ orbital can absorb a photon, but cannot emit a photon.

a. s

b. p

c. d

d. f

58. In the Bohr Model of the atom:

a. electrons travel in circular orbits around the nucleus

b. electrons can have any energy level

c. electron energies are quantitized

d. electron paths are controlled by probability

e. both A and C

**6.5: Applying quantum mechanics to the atom**

59. Define the term “orbital”.

60. What is required to move electrons from one orbital to another?

**6.6: Orbital filling and electron configuration**

61. What is the electron configuration of chlorine?

62. What is common to the electron configuration of elements found in the same row of the periodic table?

**Lecture 7: Chemical bonding**

**7.1: Basic bonding background information**

63. What group (or column) would these fake atoms be found in?



**7.2: Ionic bonding**

64. How does a metal's valence electron number determine how many ionic bonds that atom will form in an ionic bond with chloride?

65. Lattice energy is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

a. the energy required to convert a mole of ionic solid into its constituent ions in the gas phase

b. the energy given off when gaseous ions combine to form one mole of ionic solid

c. the energy required to produce one mole of ionic compound from its constituent elements in their standard states

d. the sum of ionization energies of the components in an ionic solid

e. the sum of electron affinities of the components in an ionic solid

66. Draw a series of Lewis structure diagrams to illustrate the formation of aluminum

chloride.

**7.3: Covalent bonding**

67. Which statement about bonds between carbon atoms (single, double and triple) is

TRUE?a. A triple bond is longer than a single bond.b. A double bond is stronger than a triple bond.c. A single bond is stronger than a triple bond.d. A double bond is longer than a triple bond. e. A single bond is stronger than a double bond.

68. Of these bonds, C-N, C=N & C=N, the first is the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

a. strongest/shortestb. strongest/longestc. weakest/shortestd. weakest/longest

e. intermediate in both strength & length

69. Draw the Lewis structure of water.

70. What is a dipolar charge?

71. Three possible Lewis structures are shown here for the same compound, NCS-1:

N – C = S N = C = S N = C - S

1. Complete the structures by adding unbonded electrons.
2. Calculate formal charges for each atom for each structure shown.
3. Circle the ‘best’ structure for the compound

**7.4: Resonance**

72. Which is TRUE of resonance structures?

a. All of the resonance structures exist in various proportions.b. One resonance structure corresponds to the observed structure.c. The observed structure is an average of the various resonance structures.

d. The same atoms need not be bonded to each other in all resonance forms.e. There cannot be more than two resonance structures for a given species.

**7.5: Exceptions to the octet rule**

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| ***EXAM 4: Lectures 8, 10 & 11 (abridged)*** |

**Lecture 8: Thermochemistry**

**8.1: Kinetic & potential energy**

**8.2: Transferring energy as heat & work**

73. Name two forms of energy that kinetic energy can be transformed into.

**8.3: System vs. surroundings**

74. You are doing an aqueous neutralization reaction in a calorimeter. List components that would be considered part of the system, and those that would constitute the surroundings.

**8.4: First law of thermodynamics**

75. The 1st Law of Thermodynamics states that: (circle one)

a. All spontaneous processes are accompanied by an increase in disorder.b. Energy is conserved during all processes.

c. The entropy of a pure, crystalline substance at absolute zero is zero.d. The amount of work done during a change in independent of the path of that change.e. None of the above.

76. What is the ∆E (in J) of a system that releases 12.4 J of heat and does 4.2 J of work on the surroundings?

**8.5: Enthalpy**

77. The value of ∆H° for the reaction below is –72 kJ. How many kJ of heat will be

evolved (given off) when 1.0 mole of HBr is formed by this reaction?

H2 (gas) + Br2 (gas) --> 2HBr (gas)

78. How much heat (kJ) is transferred when 5.10 g of H2(g) is formed as shown in this thermochemical equation:

CH3OH(l) 🡪 CO(g) + 2H2(g) ΔH = +128.1 kJ

79. Calculate the enthalpy of reaction for reaction 3 using Hess’s Law and thermochemical equations 1 & 2:

(1) P4 + 3O2 -> P4O6 ∆H = - 1640.1 kJ

(2) P4 + 5O2 -> P4O10 ∆H = - 2940.1 kJ

(3) P4O6 + 2O2 -> P4O10 ∆H = ?

80. Given the information below, determine ΔHrxn for this rxn:

2H2O(l) 🡪 2H2(g) + O2(g)

H2O(l) 🡪 H2O(g) ΔH = +44.01 kJ

2H2(g) + O2(g) 🡪 2H2O(g) ΔH = -483.64 kJ

**8.6: Calorimetry**

81. When a sample of aluminum absorbed 9.86 J of heat its temperature increased

from 23° to 30.5°. The specific heat of aluminum is 0.90 J/g-K. What is the mass of

the aluminum sample?

82. 50.0 mL of 1.0 M HCl were mixed with 50.0 mL of 1.0 M NaOH in a ‘coffee cup’ calorimeter. The solution warmed from 23.0° to 29.8°. Assume the solution had the density and specific heat of pure water.

Which of the following is ***false***?

1. The reaction was exothermic.
2. 2.8 kJ of heat was transferred.
3. ΔHrxn was extensive.
4. Heat flow was measured at constant pressure.
5. 1.4 kJ of heat was transferred.

**~~8.7: Enthalpy of formation~~**

~~83. Calculate the enthalpy of combustion of benzene (C6H6) using the enthalpies of formation given here: ∆Hf (kJ/mol)~~

~~C6H6 + 49.0~~

~~CO2 - 393.5~~

~~H2O - 285.8~~

~~84. Given the table of enthalpies of formation shown here, calculate the enthalpy of reaction for:~~

~~Ca(OH)2 + 2H3AsO4 🡪 Ca(H2AsO4)2 + 2H2O~~

~~ΔHf° (kJ/mol)~~

~~Ca(OH)2 -986.6~~

~~H3AsO4 -900.4~~

~~Ca(H2AsO4)2~~~~-2346.0~~

~~H2O~~~~-258.9~~

**Lecture 10: Kinetics *(abridged)***

**10.1: Reaction rates**

85. What 4 parameters (or factors) can affect the kinetics of chemical reactions?

86. Which does **not** play a role in determining the rate of a chemical reaction?

1. temperature
2. concentration of reactants
3. presence of a catalyst
4. surface area of solid or liquid reactants
5. the equilibrium constant

**~~10.4: Temperature & reaction rates~~**

~~90. Which of the following will lower the activation energy of a reaction?~~

1. ~~increasing concentration of reactants~~
2. ~~raising the temperature of the reaction~~
3. ~~adding a suitable catalyst~~
4. ~~all of the above~~
5. ~~none of the above~~

**~~1o.5: Rate-limiting step~~**

**10.6: Catalysis**

**Lecture 11: Equilibrium *(abridged)***

**11.1: Concept of equilibrium**

91. At equilibrium, which of the following statements is true?

1. All chemical processes have ceased.
2. The rate of the forward reaction equals the rate of the reverse reaction.
3. The rate constants of the forward and reverse reaction are equal.
4. Both rates of reaction and rate constants of the forward and reverse reactions are equal.
5. None of the above.

**11.2: Equilibrium constant**

92. Write the equilibrium expression for this reaction in which all species are gaseous:

CO + 3H2  CH4 + H2O

93. What is the balanced chemical equation represented by the following equilibrium expression. Kc = [H2]2[O2]  
 [H2O]2

**11.3: Working with equilibrium expressions**

94. If the volume of the system is decreased, which way will this reaction go (right or

left)?

2SO3 (g) 2SO2 (g) + O2 (g)

95. For a given chemical reaction, Kc for the forward reaction equal 4.2x10-4. Calculate Kc for the reverse reaction.

96. Consider this equilibrium reaction: 2SO2(g) + O2(g) 2SO3(g)

From which starting condition is it impossible to reach equilibrium?

1. 0.25 mole SO2 (g) and 0.25 mole O2 (g) in a 1.0-L container
2. 0.75 mole SO2 (g) in a 1.0-L container
3. It is possible to reach equilibrium from all starting conditions described.
4. 0.50 mole O2 (g) and 0.50 mole SO3 (g) in a 1.0-L container
5. 1.0 mole SO3 (g) in a 1.0-L container

97. At a given temperature, a flask at equilibrium contains 0.0114 M HCl, 0.0931 M Cl2, and 0.0154 M H2. What is the value of Kc at this temperature for the following reaction?

2HCl(g) Cl2(g) + H2(g)

**11.4: Le Châtelier’s Principle**

98. What action will cause the endothermic reaction below to form more reactant?

CH3CH=CH2(g) C3H6(g)

1. increasing temperature
2. decreasing temperature
3. increasing pressure
4. decreasing pressure
5. decreasing both temperature and pressure

99. How would you alter system temperature and pressure to maximize the yield of CO produced using this reaction?

2CO2(g) 2CO(g) + O2(g) ΔH = -514 kJ

**11.5: Catalysts & equilibrium**

100. What effect do catalysts have on equilibrium reactions?