THE COLLEGE OF THE BAHAMAS

FINAL EXAMINATION

SEMESTER 01-2011

FACULTY OF PURE AND APPLIED SCIENCES

SCHOOL OF CHEMISTRY, ENVIRONMENTAL & LIFE SCIENCES

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DATE AND TIME OF EXAMINATION: April 27, 2011, 7 pm, rooms MHEC 1A, 1B, 1C & 2A
DURATION: 3 HOURS

COURSE NUMBER: CHEM 135

COURSE TITLE: COLLEGE CHEMISTRY I

SECTION NUMBER: **NAS** 01 02 03 04 **FRP** 01 (Circle your coarse section number.)

STUDENT NAME:

STUDENT NUMBER:

LECTURER'S NAME: **Dr. F. Banks** **Dr. G. Holden** **Dr. A. Moxey** **Dr. D. Hepburn** (Circle your Lecturer.)

INSTRUCTIONS TO CANDIDATES: This paper has 10 pages and 2 sections. Please follow the instructions given with each section.

PERIODIC TABLE OF THE ELEMENTS

←Periods→	I	II	← Groups →										III	IV	V	VI	VII	0
1	1 **H** hydrogen 1.0																	2 **He** helium 4.0
2	3 **Li** lithium 6.9	4 **Be** beryllium 9.0											5 **B** boron 10.8	6 **C** carbon 12.0	7 **N** nitrogen 14.0	8 **O** oxygen 16.0	9 **F** fluorine 19.0	10 **Ne** neon 20.2
3	11 **Na** sodium 23.0	12 **Mg** magnesium 24.3											13 **Al** aluminium 27.0	14 **Si** silicon 28.1	15 **P** phosphorus 31.0	16 **S** sulfur 32.1	17 **Cl** chlorine 35.5	18 **Ar** argon 39.9
4	19 **K** potassium 39.1	20 **Ca** calcium 40.1	21 **Sc** scandium 45.0	22 **Ti** titanium 47.9	23 **V** vanadium 50.9	24 **Cr** chromium 52.0	25 **Mn** manganese 54.9	26 **Fe** iron 55.8	27 **Co** cobalt 58.9	28 **Ni** nickel 58.7	29 **Cu** copper 63.5	30 **Zn** zinc 65.4	31 **Ga** gallium 69.7	32 **Ge** germanium 72.6	33 **As** arsenic 74.9	34 **Se** selenium 79.0	35 **Br** bromine 79.9	36 **Kr** krypton 83.8
5	37 **Rb** rubidium 85.5	38 **Sr** strontium 87.6	39 **Y** yttrium 88.9	40 **Zr** zirconium 91.2	41 **Nb** niobium 92.9	42 **Mo** molybdenum 95.9	43 **Tc** technetium 98.9	44 **Ru** ruthenium 101.1	45 **Rh** rhodium 102.9	46 **Pd** paladium 106.4	47 **Ag** silver 107.9	48 **Cd** cadmium 112.4	49 **In** indium 114.8	50 **Sn** tin 118.7	51 **Sb** antimony 121.8	52 **Te** tellurium 127.6	53 **I** iodine 126.9	54 **Xe** xenon 131.3
6	55 **Cs** cesium 132.9	56 **Ba** barium 137.3	57 **La** lanthanum 138.9	72 **Hf** hafnium 178.5	73 **Ta** tantalum 180.9	74 **W** tungsten 183.85	75 **Re** rhenium 186.2	76 **Os** osmium 190.2	77 **Ir** iridium 192.2	78 **Pt** platinum 195.1	79 **Au** gold 197.0	80 **Hg** mercury 200.6	81 **Tl** thallium 204.4	82 **Pb** lead 207.2	83 **Bi** bismuth 209.0	84 **Po** polonium	85 **At** astatine	86 **Rn** radon
7	87 **Fr** francium	88 **Ra** radium	89 **Ac** actinium															

Key: 19 ← proton number; K ← symbol of element; potassium ← name of element; 39.1 ← relative atomic mass

Lanthanides:

58 **Ce** cerium 140.1	59 **Pr** praseodymium 140.9	60 **Nd** neodymium 144.2	61 **Pm** promethium	62 **Sm** samarium 150.4	63 **Eu** europium 152.0	64 **Gd** gadolinium 157.3	65 **Tb** terbium 158.9	66 **Dy** dysprosium 162.5	67 **Ho** holmium 164.9	68 **Er** erbium 167.3	69 **Tm** thulium 168.9	70 **Yb** ytterbium 173.0	71 **Lu** lutetium 175.0

Actinides:

90 **Th** thorium 232.0	91 **Pa** protoactinium 231.0	92 **U** uranium 238.0	93 **Np** neptunium 237.0	94 **Pu** plutonium	95 **Am** americium	06 **Cm** curium	97 **Bk** berkelium	98 **Cf** californium	99 **Es** einsteinium	100 **Fm** fermium	101 **Md** mendelevium	102 **No** nobelium	103 **Lr** lawrencium

Note: relative atomic masses are omitted for highly unstable elements.

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1atm = 101325 Pa = 760 mmHg = 760 torr, ,	1 dm^3 = 1 L	The molar volume of any gas at s.t.p. is 22.4 $dm^3 \cdot mol^1$	$pV = nRT$
density of H_2O = 1.0 $g \cdot cm^{-3}$ at 25°C	R= 8.31 J mol^{-1} K^{-1}= 0.0821 $L \cdot atm \cdot mol^{-1} \cdot K^{-1}$	Avogadro's constant = 6.02×10^{23} mol^{-1}.	$\Delta H = -mc\Delta T$

SECTION I: Multiple Choice Questions

Select the **SINGLE** best alternative in each of the following cases and mark the appropriate letter on the separate answer sheet by shading in the appropriate square with soft pencil. (1 mark each = **30** marks)

1) Which one of the following species in its ground state has three unpaired electrons in the 2p sub-shell?
A a hydrogen atom.
B an oxygen atom.
C a boron atom.
D a nitrogen atom.
E a phosphorus atom.

2) A real gas most closely approaches the behavior of an ideal gas at
A 0.50 atm and 500K
B 1 atm and 273K
C 1 atm and 298K
D 15 atm and 500K
E 15 atm and 200K

3) Which is the correct ground-state electron configuration for iron?
A [Ar] $3d^2 4s^2$
B [Ar] $3d^3 4s^2$
C [Ar] $3d^4 4s^2$
D [Ar] $3d^5 4s^2$
E [Ar] $3d^6 4s^2$

4) For the reaction below the $\Delta H° = +128.1$ kJ.
$CH_3OH\ (l) \rightarrow CO\ (g) + 2H_2\ (g)$
How many kJ of heat are consumed when 0.200 mol of H_2 (g) is formed as shown in the equation?
A 12.8
B 25.6
C 128
D 256
E 6.4

5) Which of the following is an isoelectronic series?
A B^{5-}, Si^{4-}
B F^-, Cl^-
C S, Cl
D Si^{2-}, P^{2-}
E O^{2-}, F^-

6) 2 mol of benzoic acid, C_6H_5COOH, contains
A 2 mol of oxygen atoms
B 6 mol of hydrogen atoms
C 10 mol of hydrogen atoms
D 12 mol of carbon atoms
E 14 mol of carbon atoms

7) How many grams of hydrogen are in 48 g of CH_4O?
A 2
B 4
C 6
D 8
E 10

8) Electronegativity increases in which of the following orders?
A Cl < S < P
B I < Br < Se
C C < N < O
D Cl < Br < I
E F < S <Se

9) Which gas would effuse more rapidly under the same conditions of temperature and pressure?
A CO_2 (RMM = 44)
B SO_2 (RMM = 64)
C C_2H_2 (RMM = 26)
D UF_6 (RMM = 352)
E Ne (RMM = 10)

10) Which one of the following molecules has a dipole moment?
A O_2
B N_2
C Cl_2
D F_2
E IF

11) 50 g of an ionic compound XY_2 (RFM=100) were dissolved in water and made up to 10 dm^3. What is the concentration of the solution with respect to Y^- ions?
A 0.05 M
B 0.10 M
C 0.15 M
D 0.20 M
E 0.40 M

12) For the following elements, the correct order of increasing first ionization energy is
A Be > Li
B Li > Be
C Na > Be
D O > F
E Na > Li

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13) A sample of CH_2F_2 with a mass of 5.2 g contains _____ atoms of F.
A 6.02×10^{23}
B 6.02×10^{22}
C 6.02×10^{21}
D 1.204×10^{23}
E 1.204×10^{22}

14) The molarity (M) of an aqueous solution containing 52.5 g of sucrose ($C_{12}H_{22}O_{11}$) RMM = 342 g/mol in 3550 mL of solution is nearest to
A 5.46×10^{-2}
B 1.48×10^{-2}
C 0.104×10^{-2}
D 4.32×10^{-2}
E 1.85×10^{-2}

15) The density (D = P*M*/RT) of a gas Z (RMM = 44) at s.t.p. in g L^{-1} is closest to
A 0.50
B 2.0
C 1.0×10^{2}
D 0.13
E 2.5×10^{-4}

16) The enthalpy change for which reaction represents the first ionization energy of magnesium?
A $Mg(s) \rightarrow Mg^{+}(g) + e^{-}$
B $Mg(g) \rightarrow Mg^{+}(g) + e^{-}$
C $Mg(g) \rightarrow Mg^{+}(s) + e^{-}$
D $Mg(s) \rightarrow Mg^{2+}(g) + 2e^{-}$
E $Mg(g) \rightarrow Mg^{2+}(g) + 2e^{-}$

17) The O-C-O bond angle in the CO_2 is approximately
A 90°
B 109.5°
C 120°
D 180°
E 60°

18) Successive ionization energies of an element increase because
A As each electron is removed, proton number increases.
B As each electron is removed, proton number decreases.
C As each electron is removed, mass number decreases.
D As each electron is removed the charge on the resulting ion increases.
E As each electron is removed there are a greater number of protons attracting fewer electrons.

19) Which element is expected to have the most negative electron affinity?
A S
B Cl
C Se
D Br
E I

20) Calcium carbide (CaC_2) reacts with water to produce acetylene (C_2H_2):
$CaC_2 (s) + 2H_2O (g) \rightarrow Ca(OH)_2 (s) + C_2H_2 (g)$
Production of 13 g of C_2H_2 requires consumption of _____ g of H_2O.
A 4.5
B 9.0
C 18
D 4.8×10^{2}
E 506

21) A sample of O_2 compressed to a smaller volume at constant temperature. Which statement(s) is correct regarding O_2 molecules?
A The average kinetic energy of O_2 molecules is increased.
B The average speed of O_2 molecules is decreased.
C The average distance between molecules increases.
D The number of collisions per unit time, of O_2 molecules with unit area of wall, increases.
E The average mass of the molecules increases.

22) The shell of lowest energy that contains f orbitals is the shell with n =
A 3
B 2
C 4
D 1
E 5

23) Electrons in the outer shell of which atom experiences the greatest effective nuclear charge?
A Li
B Be
C B
D C
E N

24) A chemical reaction occurs in a calorimeter containing 100 g of dilute solution producing a rise in temperature of 3.2°C. The specific heat of the contents is 4.2 $J \cdot g^{-1} \cdot K^{-1}$. ΔH in kJ for the amount of reaction that occurs is nearest to:
A +1.3
B -1.3
C 1300
D -1300
E 0.076

25) Which statement is <u>not</u> true for oxygen in its lowest energy state?
A It is in group VI.
B It is in period 2.
C It is a *d*-block element.
D The electrons of lowest energy are in a 1s-orbital.
E The oxygen atom has 2 unpaired electrons.

26) Which one of the following is a unit of pressure?

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A $J \cdot K^{-1}$
B $g \cdot cm^{-3}$
C $N \cdot m^{-2}$
D $g \cdot mol^{-1}$
E $mol \cdot dm^{-3}$

27) In the nitrite ion (NO_2^-),
A both bonds are single bonds
B both bonds are double bonds
C one bond is a double bond and the other is a single bond
D both bonds are the same because of resonance
E there is one fewer electron than the number of protons.

28) Ethane has a lower boiling point than water partly because
A it is a covalent compound.
B its molecules do not form hydrogen bonds.
C it has a larger molecular mass.
D it does not contain ionic bonds.
E it contains carbon atoms.

29) Which atom/ion has eight valence electrons?
A Ti^{4+}
B Kr
C Cl^-
D Na^+
E all of the above

30) The most polar bond is found in this compound
A HBr
B HI
C HCl
D HF
E H_2

SECTION B: STRUCTURED QUESTIONS

Answer each of the following questions in the spaces provided on the question paper. Full marks can only be given to calculations where working is included and the final answer is clearly indicated. Clear and concise expression is an essential part of a good answer. You may write your answers in pencil. There are a total of **65** marks for this section.

1) The following questions relate to qualitative analysis and net ionic equations.
   a) Three salt solutions were tested in an effort to identify the salts present. The results are presented in the table below. Use the results to identify the ions as completely as you can by filling in the spaces in the table with **ONE** possibility in each case. (6)

TEST	***Sample A***	***Sample B***	***Sample C***
Add sodium hydroxide solution dropwise	White *ppt.* forms which is soluble in excess sodium hydroxide solution	No change observed	Green *ppt.* insol. in excess. *ppt* turns brown on standing/ exposure to air
Add ammonia solution	White *ppt.* forms which is soluble in excess ammonia solution	No change observed	Green *ppt.* insol. in excess. *Ppt* turns brown on standing.
Add barium chloride solution followed by dilute/ strong nitric acid	No change observed.	A dense white *ppt.* forms which is soluble in dil. strong acid to release a gas which turns limewater cloudy	A dense white *ppt.* forms which is insoluble in dil. strong acid.
Add silver nitrate solution	Cream *ppt.* Insoluble in dilute ammonia, soluble in conc. ammonia	Test not necessary.	Test not necessary.
Add $MgCl_2$ solution	Test not necessary.	Immediate white *ppt.*	Test not necessary.
Flame test	No flame colour	Lilac flame	No flame colour
CATION in sample?			
ANION in sample?			

   b) Write a balanced net ionic equation, including states, to explain each of the following observations. If an equation is unbalanced it will attract zero marks: check carefully. Do not include spectator ions. Each answer line should be completed with an equation.
      i. Addition of dilute ammonia solution to a solution of magnesium nitrate results in a white precipitate.

      _____ (1)

      ii. The addition of dilute sodium hydroxide solution to a solution of aluminium sulfate results in a white precipitate which redissolves in excess to form a colourless solution.

      _____ (1)

      _____ (1)

      iii. The addition of dilute ammonia solution to a solution of zinc chloride results in a white precipitate which redissolves in excess.

      _____ (1)

      _____ (1)

      iv. The addition of dilute sodium hydroxide to a solution of silver nitrate results in a brown precipitate.

      _____ (1)

      _____ (1)

v. The addition of barium nitrate solution to a solution of sodium dichromate results in a yellow precipitate of barium chromate which redissolves in dilute hydrochloric acid.

_____ (1)

_____ (1)

vi. The addition of barium nitrate solution to a solution of sodium sulfite results in a white precipitate which redissolves on the addition of dilute nitric acid to produce a solution which smells of sulfur dioxide.

_____ (1)

_____ (1)

vii. The addition of silver nitrate solution to a solution of calcium iodide results in a yellow precipitate which is insoluble in concentrated ammonia solution.

_____ (1)

2) The following questions relate to thermochemistry.

a) Define the standard enthalpy of formation, ΔH_f° (2)

_____

_____

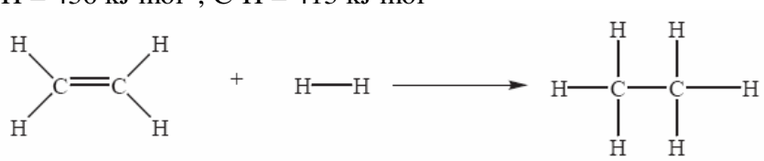
b) Write the equation for the complete combustion of **1 mol**, of ethane gas, C_2H_6, (to produce carbon dioxide and liquid water). (1)

_____

The standard enthalpy of combustion of ethane gas, C_2H_6 is $-1560\ kJmol^{-1}$.
Given that $\Delta H_f^\circ(CO_2, (g)) = -394\ kJ\cdot mol^{-1}$ and $\Delta H_f^\circ(H_2O, (l)) = -286\ kJ\cdot mol^{-1}$

c) Calculate ΔH_f° (C_2H_6 (g)) (4)

d) Use the average bond energies below to estimate a value for the standard enthalpy change, ΔH_f° for the hydrogenation of 1 mol of ethane. Mean bond energies are as follows: C-C = 348 $kJ\cdot mol^{-1}$, C=C = 615 $kJ\cdot mol^{-1}$, H-H = 436 $kJ\cdot mol^{-1}$, C-H = 413 $kJ\cdot mol^{-1}$ (4)



3) The following questions relate to stoichiometry:
An organic ester, X, contains 54.55% C, 9.09% H, and 36.36% O by mass.

a) Determine the empirical formula of X. (RAM of H = 1.00, C = 12.0, O = 16.0) (3)

b) 0.15 g of X, when vaporized in a suitable apparatus, occupied a volume of 57.0 cm^3 at 127°C and 1.00×10^5 Nm^{-2}. Using the expression PV = nRT (or otherwise) calculate the relative molecular mass of X.

(3)

c) What is the molecular formula of X? (2)

4) The following questions relate to bonding and electronic configuration.

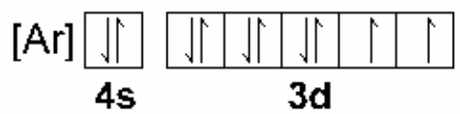
a) State:

i) the Aufbau (building-up) Principle, (2)

ii) the Pauli Exclusion Principle (2)

iii) Hund's Rule (2)

b) Draw orbital diagrams for the ground state electron configurations of chloride ($_{17}Cl^-$), chromium ($_{24}Cr$) and nitrogen, ($_7N$). For example, the orbital diagram of nickel, $_{28}Ni$ in the 4th period would be



(2)

$_{17}Cl^-$: **$_{24}Cr$:**

i) How many unpaired electrons are there in an atom of sodium, $_{11}Na$? (1)

ii) Show that the maximum number of electrons that can be accommodated in all the orbitals with principal quantum number n = 4 in a given atom is 32. (2)

5) Draw Lewis electron dot diagrams to represent the bonding in the molecules: PCl_5, BF_3, and NH_3.

a) Show all valence electrons. (3)

i. PCl_5 **ii. BF_3** **iii. NH_3**

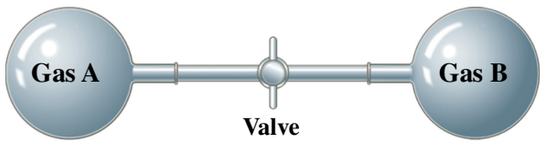
b) Draw diagrams to indicate the shapes of the above molecules according to VSEPR theory and write a descriptive term for each shape. (For example, for the methane molecule you might write "tetrahedral").(6)

6) The following question relates to gases.

a) Define *partial pressure* (2)

b) State Dalton's law of partial pressure (2)

c) A flask of volume 600 cm^3 contains a gas, A, at 120 mmHg that is connected via a valve to a second flask of volume 1200 cm^3 that contains a gas, B, at 255 mmHg. The stopcock is opened and the gases are allowed to mix completely. Find the final pressure of the mixture, assuming no change in temperature. (Neglect the volume of the tube connecting the bulbs.) (5)



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