UNIVERSITY OF SWAZILAND

SUPPLEMENTARY FINAL EXAMINATION

ACADEMIC YEAR 2010/2011

TITLE OF PAPER: INTRODUCTORY INORGANIC CHEMISTRY

COURSE NUMBER: C201

TIME ALLOWED: THREE (3) HOURS

INSTRUCTIONS: THERE ARE SIX (6) QUESTIONS. ANSWER ANY FOUR (4) QUESTIONS. EACH QUESTION IS WORTH 25 MARKS.

A PERIODIC TABLE AND A TABLE OF CONSTANTS HAVE BEEN PROVIDED WITH THIS EXAMINATION PAPER.

ELECTRONIC CALCULATORS MAY BE USED

PLEASE DO NOT OPEN THIS PAPER UNTIL AUTHORISED TO DO SO BY THE CHIEF INVIGILATOR.

Question One

a) Use the Rydberg equation to predict which of the following electronic transitions, in a hydrogen atom, produces the spectral line having the longest wavelength:

   i)   n=2 to n=1

   ii)   n=3 to n=2

   iii)   n=4 to n=3

   [NB: Full calculations are not needed to arrive at the answer]

                                                                           [6]

b) Explain how the existence of line spectra is consistent with Bohr's theory for energy of an electron in a hydrogen atom.

                                                                           [4]

c) What is the wavelength of an electron moving with a speed of 5.97×10^6 ms^{-1}?

                                                                           [4]

d) Use Heisenberg Uncertainty Principle to calculate the uncertainty in the position of

   i)   A 1.50-mg mosquito moving at a speed of 1.40 m/s if the speed has an uncertainty of ± 0.01 m/s

   ii)   An electron moving at a speed of $(3.00 \pm 0.01) \times 10^5$ m/s.

   What are the implications of these calculations to our model of the atom?

                                                                           [7]

e) Give the quantum number that governs

   i)   The shape of an orbital

   ii)   The energy of an orbital

   iii)   Magnetic properties of an atom or molecule

   iv)   Spatial orientation of an orbital

                                                                           [4]

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Question Two

a) Use Slater's rules to calculate Z_{eff} for 2p electrons for isoelectronic ions O^{2-} and F^-.

[8]

b) Based on Z_{eff} values calculated in a) above, arrange the ions in order of

   i)   Increasing size

   ii)   Increasing first ionization potential

   Explain each of your two arrangements above

[4]

c) For each of the orbitals given below, sketch the radial distribution function, indicating the location of the most probable radius.

   i)   4p   ii)   5g

[5]

d) Sketch angular functions of the orbitals given below. Nodal planes or surfaces, if present, will need to be labeled.

   i) $7p_y$ ii)   5s   iii) $6d_{x^2-y^2}$ iv) d_{z^2}

[8]

Question Three

a) Construct a Born-Haber cycle for the formation of $Al_2O_3(s)$ from its constituent elements in their standard states.

[9]

b) Explain the following trends in lattice energy:

   i) $CaF_2 > BaF_2$

   ii) $NaCl > RbBr > CsBr$

   iii) $BaO > KF$

[4]

c) Consider the molecules CO and CO_2.

  i)   Draw the Lewis structure of each of the species. For each of the molecules, give the average C-O bond order.

  ii) Based on the Lewis structures you have drawn in i) above, predict the ordering of C-O bond lengths in CO and CO_2. That is, indicate which of the molecules is expected to have a longer or shorter bond.

  [5]

d) An element crystallizes in the bcc lattice. The edge of each unit cell is 2.86×10^{-10} m long and the density of the crystal is 7.92 g.cm^{-3}. Calculate the atomic mass of the element.

  [7]

Question Four

a) Draw a molecular orbital energy level diagram for a diatomic molecule, C_2. With the help of the diagram, answer the questions that follow below.

  i)   Give electron configurations for C_2, C_2^{2-} and C_2^{4-}.

  ii) Calculate the bond order for each of the species in i) above

  iii) State whether the species in i) above are paramagnetic or diamagnetic. Explain each of your answers.

  iv) List the species in order of increasing bond strength

  [14]

b) Consider the molecule SO_2Cl_2 where the sulphur atom, S, is the central atom.

  i)   Draw at least three non-equivalent Lewis structures for the molecule

  ii) Use formal charges to suggest which of your Lewis structures is expected to be the most stable

  [8]

c) Use the information from b) above to determine the hybridization around the sulphur atom in SO_2Cl_2.

  [3]

4

Question Five

a) The Lewis structure of carbon disulphide is S=C=S. Use a suitable orbital diagram to illustrate how pi-type orbitals overlap to form pi bonds.

[4]

b) With the help of a suitable reaction equation, explain each of the following:

   i)   The oxide Cl_2O_7 dissolves in water to form an acidic solution, whereas the oxide Na_2O also dissolves in water but leads to the formation of a basic solution.

   ii)   Hydrofluoric acid cannot be stored in glass bottles

   iii)   Solutions of $Na_3PO_4^{3-}$ are quite basic.

[12]

c) Sodium trimetaphosphate, $Na_3P_3O_9$, and sodium tetrametaphosphate, $Na_3P_4O_{12}$, are used as water softening agents. Sketch the structures for these ions.

[4]

d) Predict whether each of the following oxides is ionic or molecular:

   SO_2, MgO, P_4O_{10}, Y_2O_3, N_2O, XeO_3

   Give reasons for your choices.

[3]

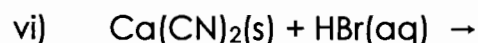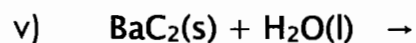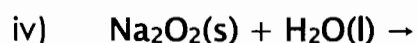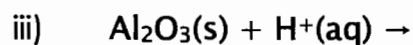e) Arrange the following oxides in order of increasing acidity, starting with the least acidic:
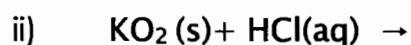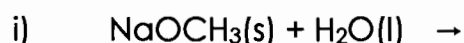
   SO_3, CaO, Al_2O_3

   Give an explanation for your answer.

[2]

Question Six

a) Complete and balance the following reactions:

      i) $NaOCH_3(s) + H_2O(l)\;\rightarrow$
      ii) $KO_2(s) + HCl(aq)\;\rightarrow$
      iii) $Al_2O_3(s) + H^+(aq)\;\rightarrow$
      iv) $Na_2O_2(s) + H_2O(l)\;\rightarrow$
      v) $BaC_2(s) + H_2O(l)\;\rightarrow$
      vi) $Ca(CN)_2(s) + HBr(aq)\;\rightarrow$

[12]

b) Identify the following hydrides as ionic, interstitial (i.e., metallic), or molecular:

      i) H_2Te ii) $Ti_{1.7}$ iii) BaH_2 iv) B_2H_6 v) RbH   vi) H_2O

[3]

c) Explain the following:

      i)   Magnesium forms stable compounds halogens whereas argon does not
      ii)   At room temperature H_2O is a liquid, whereas H_2S is a gas
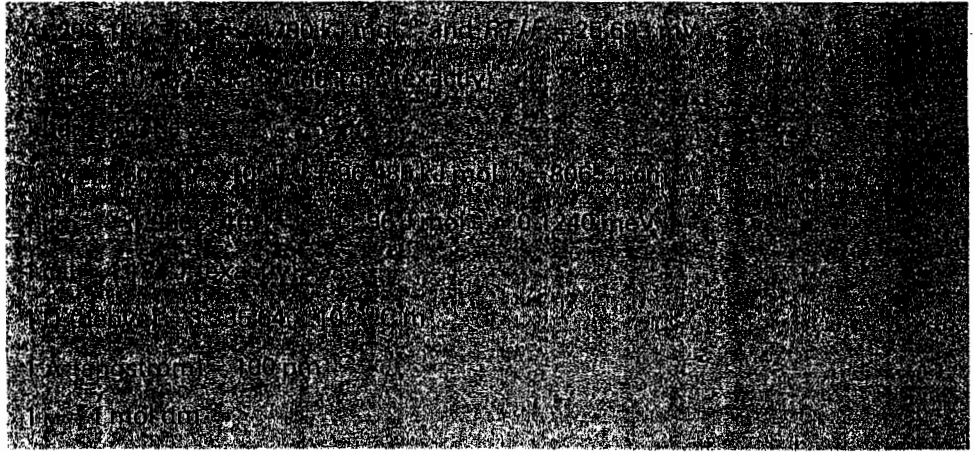   iii)   The halogens decrease in oxidizing power in the order $F_2 > Cl_2 > Br_2 > I_2$

[6]

d) Draw the structure of diborane, B_2H_6. Briefly describe the nature of bonding in this molecule.

[4]

Useful relations

(illegible)

$$pi\,(\pi) = 3.142$$

General data and fundamental constants

Quantity	Symbol	Value
Speed of light	c	
Elementary charge	e	
Faraday constant	$F = eN_A$	
Boltzmann constant	k	
Gas constant	$R = kN_A$	
Planck constant	h	
	$\hbar = h/2\pi$	
Avogadro constant	N_A	
Atomic mass unit	u	
Mass of electron	m_e	
Vacuum permittivity	ε_0	
	$4\pi\varepsilon_0$	
Bohr magneton	$\mu_B = e\hbar/2m_e$	
Bohr radius	$a_0 = 4\pi\varepsilon_0\hbar^2/m_e e^2$	
Rydberg constant	$R_\infty = m_e e^4/8h^3 c\varepsilon_0^2$	

Prefixes

(illegible)

The Periodic Table

Period

s block

d block

f block

Lanthanides

Actinides

Group	1	2	3	4	5	6	7	8	9	10	11	12	13/III	14/IV	15/V	16/VI	17/VII	18/VIII

1	3 Li 6.941	4 Be 9.012
2		

H 1.0079

11 Na 22.99	12 Mg 24.30
19 K 39.10	20 Ca 40.08
37 Rb 85.47	38 Sr 87.62
55 Cs 132.9	56 Ba 137.3
87 Fr 223.0	88 Ra 226.0

Lanthanides:

57 La 138.9	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm 146.9	62 Sm 150.4	63 Eu 152.0	64 Gd 157.2	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0

Actinides:

89 Ac 227.0	90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np 237.0	94 Pu 239.1	95 Am 241.1	96 Cm 244.1	97 Bk 249.1	98 Cf 252.1	99 Es 252.1	100 Fm 257.1	101 Md 256.1	102 No 259.1	103 Lr 260.1