CHEMISTRY 1E03 **13 NOVEMBER 2009**

Instructor: J. Barbier

MCMASTER UNIVERSITY - TERM TEST # 2 - DURATION: 120 minutes

This test contains 17 pages and **30 multiple-choice questions**. The last pages include extra blank pages for rough work, a page with some useful data and equations, and a Periodic Table. You may tear off the last five pages. **You are responsible for ensuring that your copy of the question paper is complete**. Bring any discrepancy to the attention of your invigilator.

Questions 1-25 are each worth 2 marks, questions 26–30 are each worth 3 marks. There is no penalty for incorrect answers. **Answer all questions on the answer sheet, in pencil**. Instructions for entering multiple-choice answers are given on page 2. **Select one answer only for each question.**

These question sheets must be returned with your answer sheet. However, no work written on the question sheets will be marked. You must enter your full name and student number on this question sheet, as well as on the answer sheet. **Your invigilator will be checking your student card for identification.**

Make sure to enter the correct version number of your test (shown at the bottom of each page) in the correct column on the answer sheet (see instructions on page 2).

Only **Casio FX 991 electronic calculators** may be used; but they must NOT be transferred between students. Use of periodic tables or any aids other than those provided, is not allowed.

Note: Academic dishonesty may include, among other actions, communication of any kind (verbal, visual, etc.) between students, sharing of materials between students, copying or looking at other students' work. If a problem arises, please ask an invigilator to deal with it for you.

QUESTIONS 1–25 ARE WORTH 2 MARKS EACH.

1. Under constant pressure conditions, for which one of the following reactions will ΔH be significantly **greater** than ΔU (i.e. $\Delta H - \Delta U > 0$)?

A) $2\ NO_2(g) \rightarrow N_2O_4(g)$
B) $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
C) $HCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + HNO_3(aq)$
D) $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$
E) $H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$

2. The manufacture of pure silicon involves the <u>unbalanced</u> reaction below. If 0.175 mol of pure $SiCl_4(l)$ is reacted with 7.00 g of Mg(s), what is the **mass (in grams)** of silicon produced?

$SiCl_4(l) + Mg(s) \rightarrow Si(s) + MgCl_2(s)$

A) 4.92
B) 8.09
C) 16.18
D) 9.83
E) 4.05

3. Identify the **correct** equation that applies to a system on which work is done by its surroundings during an adiabatic transformation (i.e. a transformation without heat transfer).

A) $\Delta U > w > 0$
B) $\Delta U = 0$
C) $\Delta U = \Delta H$
D) $\Delta U < 0$
E) $w = \Delta U$

4. Use the data below to calculate the enthalpy change **(in kJ)** for the reduction of molybdenum trioxide by hydrogen according to the reaction: $MoO_3(s) + 3\ H_2(g) \rightarrow Mo(s) + 3\ H_2O(g)$

$2\ Mo(s) + 3\ O_2(g) \rightarrow 2\ MoO_3(s) \quad \Delta H_1 = -1600$ kJ
$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g) \quad \Delta H_2 = -480$ kJ

A) +160
B) +80
C) +1120
D) -1120
E) +560

5. A chemical reaction with an enthalpy change $\Delta H° = -200$ kJ is carried out in a calorimeter containing 1500 cm^3 of pure water initially at 26.0 °C. If the heat capacity of the calorimeter itself is neglected, what will be the **final temperature of the water (in °C)**?

A) 57.9
B) 79.5
C) 75.9
D) 31.9
E) 59.7

6. When the following redox reaction in <u>acidic</u> solution is balanced with the smallest integer coefficients, what is the coefficient for the **H^+ ion**?

$MnO_4^-(aq) + HSO_3^-(aq) \rightarrow Mn^{2+}(aq) + SO_4^{2-}(aq)$

A) 2
B) 5
C) 1
D) 3
E) 4

7. Which of the following contains the **smallest** number of atoms?

A) 15 g Ne
B) 73 g Pb
C) 2 mol N_2
D) 1 L H_2 gas at 25 °C and 8 atm
E) 0.5 mol H_2O

8. Determine the number of resonance Lewis structures (with minimized formal charges) and the average bond order in the **chlorate (ClO_3^-) anion**.

A) 1, 5/3
B) 1, 3/5
C) 3, 3/5
D) 1, 5/2
E) 3, 5/3

9. Which one of the following atoms has **at least two unpaired electrons** in its ground-state electron configuration?

A) Sr
B) As
C) Ga
D) Xe
E) Cl

10. Identify the **true** statement regarding the phosgene molecule, $COCl_2$ (C is the central atom, all other atoms are bonded to C).

A) Non-zero formal charges are present in the molecule
B) The molecule contains six non-bonding electron pairs
C) The molecule has two resonance structures
D) The molecule is planar
E) The molecule is non-polar

11. Select the set of quantum numbers (n, l, m_l, m_s) that can be associated with a valence electron of a **tin (Sn) atom in its ground state**.

A) 5, 1, 1/2, 0
B) 5, 1, 2, -1/2
C) 5, 1, -1, 1/2
D) 5, -1, 0, -1/2
E) 5, 2, 0, 1/2

12. According to the VSEPR model, what is the shape of the AlH_4^- anion (all H's bonded to Al)? Which atom carries a formal charge?

A) square pyramidal; charge carried by Al
B) tetrahedral; charge carried by H
C) tetrahedral; charges carried by both Al and H
D) square pyramidal; charge carried by H
E) tetrahedral; charge carried by Al

13. Choose the **false** statement about the Lewis structure of the **peroxide anion, O_2^{2-}**.

A) The oxygen-oxygen bond is a single bond.
B) There are two resonance structures.
C) Each oxygen atom carries a formal charge of -1.
D) Each oxygen atom has 3 non-bonding electron pairs.
E) Each oxygen atom obeys the octet rule.

14. The standard enthalpy of formation for $NH_3(g)$ is –46 kJ/mol. If the energy of the N≡N triple bond is 946 kJ/mol and that of the H-H bond is 436 kJ/mol, calculate the energy of the **N-H bond (in kJ/mol)**.

A) 422
B) 360
C) 1127
D) 1173
E) 391

15. In a multi-electron atom, the energy of an electron is associated with

A) both, the principal and angular momentum quantum numbers.
B) the magnetic quantum number (m_l).
C) both, the angular momentum and magnetic quantum numbers.
D) the angular momentum quantum number (l).
E) the principal quantum number (n).

16. Identify the **correct bond polarities** (δ+ δ–) as given below where the arrow should point toward the negative charge (δ–).

(i) N → O
(ii) F → Si
(iii) N → B
(iv) O → P
(v) S → O

A) ii, iv
B) i
C) i, v
D) ii, iii
E) ii, iv, v

17. Identify the **non-polar** molecule in the list.

A) CH_3Cl
B) SF_4
C) NF_3
D) XeF_4
E) OF_2

18. According to the VSEPR model, the **smaller F-Xe-F angles** in O<u>Xe</u>F_4 (Xe is the central atom) should be close to

A) 109º
B) 120º
C) 180º
D) 60º
E) 90º

19. Select the **correct** statement regarding the lattice enthalpy (ΔH_{latt}) of sodium chloride.

A) $\Delta H_{latt}(NaCl) - \Delta H_{latt}(LiF) > 0$
B) ΔH_{latt} is the enthalpy change for the reaction $NaCl(s) + H_2O \rightarrow Na^+(aq) + Cl^-(aq)$
C) $\Delta H_{latt} < 0$ and its magnitude represents the bond energy in NaCl(s)
D) ΔH_{latt} is the enthalpy change for the reaction $NaCl(s) \rightarrow Na^+(g) + Cl^-(g)$
E) ΔH_{latt} is the enthalpy change for the reaction $Na(s) + 1/2\ Cl_2(g) \rightarrow Na^+(g) + Cl^-(g)$

20. Choose the **true** statement regarding the <u>emission</u> spectrum of excited H atoms in which the electrons occupy the energy level n = 3.

A) One of the lines in the spectrum corresponds to the transition $n = 2 \rightarrow n = 3$.
B) The spectrum contains three lines including one in the visible range.
C) The transition to n = 1 produces photons with the longest wavelength.
D) An H atom loses $\Delta E = 193.6 \times 10^{-20}$ J in the transition to n = 1.
E) The frequency of each line is given by $\Delta E \times h$, where ΔE is the energy lost by an H atom.

21. An Fe^{3+} cation in its ground-state has ___ unpaired electrons and is _____.

A) 3, diamagnetic
B) 5, paramagnetic
C) 5, diamagnetic
D) 3, paramagnetic
E) 6, diamagnetic

22. The equilibrium constant K_p for the reaction $C(s) + CO_2(g) \leftrightarrow 2\ CO(g)$ is equal to 1.52 at 700 °C. If the partial pressure of CO in an equilibrium mixture at 700 °C is 1.30 atm, what is **the partial pressure of CO_2 (in atm)** and **the total pressure (in atm)**?

A) 0.650, 3.25
B) 0.650, 1.95
C) 1.11, 2.41
D) 1.11, 3.71
E) 1.71, 3.01

23. Consider the equilibrium: $H_2O(g) + CO(g) \leftrightarrow CO_2(g) + H_2(g)$. Use the data below to calculate ***K*** at 25 °C.

$2\ CO(g) + O_2(g) \leftrightarrow 2\ CO_2(g)$ $\quad K_1 = 3.3 \times 10^{91}$ at 25 °C
$2\ H_2(g) + O_2(g) \leftrightarrow 2\ H_2O(g)$ $\quad K_2 = 9.1 \times 10^{80}$ at 25 °C

A) 1.8×10^{10}
B) 5.3×10^{-6}
C) 3.6×10^{10}
D) 3.2×10^{5}
E) 1.9×10^{5}

24. Ethane gas (C_2H_6) can be synthesized by reacting acetylene gas (C_2H_2) with hydrogen according to the reaction
$C_2H_2(g) + 2\ H_2(g) \leftrightarrow C_2H_6(g)$ $\quad \Delta H°_{rxn} = -311$ kJ
What conditions of temperature and pressure will favour **a large yield of ethane**?

A) low temperature, low pressure
B) high temperature, low pressure
C) high temperature, high pressure
D) high temperature, pressure has no effect
E) low temperature, high pressure

25. The dissociation of solid silver sulfide, $Ag_2S(s)$, in water to produce $Ag^+(aq)$ and $S^{2-}(aq)$ ions has an equilibrium constant of 1.4×10^{-16} at room temperature. Calculate the **equilibrium concentration of sulfide ions (in mol/L)** in solution.

A) 3.3×10^{-6}
B) 5.2×10^{-6}
C) 7.0×10^{-17}
D) 2.8×10^{-16}
E) 8.2×10^{-6}

QUESTIONS 26-39 ARE WORTH 3 MARKS EACH.

26. Calculate the **lattice enthalpy (in kJ mol^{-1})** of potassium iodide from the data. (*Note that $I_2(s)$ is the the standard state of iodine.*)

Enthalpy of sublimation of K(s) = + 89 kJ mol^{-1}
First ionization energy of K(g) = + 418 kJ mol^{-1}
Enthalpy of sublimation of $I_2(s)$ = + 62 kJ mol^{-1}
Bond energy in $I_2(g)$ = + 108 kJ mol^{-1}
Electron affinity of I(g) = –295 kJ mol^{-1}
Enthalpy of formation of KI(s) = –328 kJ mol^{-1}

A) –1420
B) –1250
C) –625
D) –710
E) –679

27. The equilibrium constant, K_p, for the reaction $CH_4(g) + H_2O(g) \leftrightarrow CO(g) + 3\ H_2(g)$ at 900°C is equal to 0.63. A rigid cylinder at 900°C contains 1.20 atm of methane, 0.50 atm of water vapor, 0.80 atm of carbon monoxide, and 0.90 atm of hydrogen. Select the **false** statement(s) below:

(i) The system is at equilibrium because reactants and products have equal total pressures.
(ii) A reverse shift will occur to establish equilibrium.
(iii) $K_p < 1$ implies that the system is not at equilibrium.
(iv) The partial pressure of hydrogen will be smaller than 0.9 atm at equilibrium.
(v) The partial pressures of methane and water vapor will be equal at equilibrium.

A) i, ii, iv
B) i, iii
C) iii
D) i, iii, v
E) ii, v

28. Which of the following statements are **false**?

i) $1s^2 2s^2 2p^6$ is the ground-state electron configuration of the oxide anion.
ii) $1s^2 2s^2 2p^6 3p^1$ can be an excited-state electron configuration of a magnesium atom.
iii) The 2s and 2p orbitals of a hydrogen atom correspond to identical energy levels.
iv) The 2s and 2p orbitals of a helium cation (He^+) correspond to different energy levels.
v) Carbon atoms in their ground state are diamagnetic.

A) ii, v
B) i, iii, iv
C) ii, iv, v
D) i, iii
E) v

29. Which of the following statements are **false** according to the VSEPR model? (Central atoms are underlined.)

(i) $\underline{P}H_3$ and $\underline{Br}O_3^-$ have the same shape.
(ii) $\underline{I}Cl_4^+$ is tetrahedral.
(iii) $\underline{I}Cl_4^-$ is square-planar.
(iv) $\underline{O}F_2$ has a smaller molecular dipole moment than $\underline{Xe}F_2$.
(v) $\underline{S}Cl_2$ is a non-polar molecule.

A) ii, v
B) ii, iii
C) i, ii
D) ii, iv, v
E) i, iii, iv

30. Which of the following statements are **true**?

i) Copper metal does not react with hydrochloric acid.
ii) Zinc metal reacts with hydrochloric acid and chlorine gas is evolved.
iii) A redox reaction takes place between zinc metal and copper(II) ions in aqueous solution.
iv) Copper metal reacts with nitric acid producing copper nitrate and hydrogen gas.
v) When aqueous sodium hydroxide is added to aqueous copper(II) sulfate and the solution is heated, a black solid is formed.

A) iii, v
B) ii, iii, iv
C) i, iv
D) i
E) i, iii, v

- **Some general data are provided on this page.**
- **A Periodic Table with atomic weights is provided on the next page.**

STP = 273.15 K, 1 atm
$R = 8.3145\ J\ K^{-1}\ mol^{-1} = 0.08206\ L\ atm\ K^{-1}\ mol^{-1}$
$c = 2.9979 \times 10^{8}\ m/s$
$m_e = 9.10 \times 10^{-31}\ kg$
Specific heat of water = $4.184\ J\ K^{-1}\ g^{-1}$

$F = 96485\ C/mol$
$N_A = 6.022 \times 10^{23}\ mol^{-1}$
$h = 6.6256 \times 10^{-34}\ Js$
density(H_2O, l) = 1.00g/mL

1 atm = 101.325 kPa
$1\ J = 1\ kg\ m^2\ s^{-2} = 1\ kPa\ L = 1\ Pa\ m^3$
$1\ cm^3 = 1\ mL$
1 Hz = 1 cycle/s

0°C = 273.15 K
$1\ m = 10^9\ nm = 10^{10}$ Å
$1\ g = 10^3\ mg$

$\lambda = h / mv = h / p$
$E_n = (-2.178 \times 10^{-18} / n^2)$
$\Delta G = \Delta G^o + RT\ lnQ = RT\ ln(Q/K)$
$\Delta G = -n\ F\ E_{cell}$
$E_{cell} = E^o_{cell} - (RT / nF)\ lnQ = E^o_{cell} - (0.0257 / n)\ lnQ = E^o_{cell} - (0.0592 / n)\ logQ$

Soluble Ionic Compounds:

1. All common compounds of Group 1A(1) and ammonium (NH_4^+) ions.
2. All common nitrates (NO_3^-), acetates (CH_3COO^-), and most perchlorates (ClO_4^-).
3. All common chlorides (Cl^-), bromides (Br^-), and iodides (I^-), *except* those of Ag^+, Pb^{2+}, Cu^+, and Hg_2^{2+}.
4. All common sulfates (SO_4^{2-}), *except* those of Ca^{2+}, Sr^{2+}, Ba^{2+}, and Pb^{2+}.

Insoluble Ionic Compounds:

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

Periodic Table of the Elements

MAIN–GROUP ELEMENTS (groups 1A–2A; 3A–8A)

- Metals (main-group)
- Metals (transition)
- Metals (inner transition)
- Metalloids
- Nonmetals

TRANSITION ELEMENTS (groups 3B–2B)

Period	1A (1)	2A (2)	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	8B (9)	8B (10)	1B (11)	2B (12)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
1	1 H 1.008																	2 He 4.003
2	3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.3	57 La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 (269)	111 (272)	112 (277)		114 (285)				

As of early 2002, elements 110–112 and 114 have not been named. Recently, element 116 was shown conclusively to exist.

INNER TRANSITION ELEMENTS

Period															
6	Lanthanides	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	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
7	Actinides	90 Th 232.0	91 Pa (231)	92 U 238.0	93 Np (237)	94 Pu (242)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

END OF TEST