Name: ____________________ Student number: ______________

Chemistry 1A03 **FINAL EXAM** **December, 2008**

McMaster University **VERSION 1**

Instructors: Drs. P. Britz-McKibbin, G. Goward, P. Lock

Duration: 3 hours

This test contains **28** numbered pages printed on both sides. There are **35** multiple-choice questions appearing on pages numbered 3 to 25. Page 26 is extra space for rough work. Page 27 includes some useful data and equations, and there is a periodic table on page 28. You may tear off the last page to view the periodic table and the data provided.

You must enter your 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. **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 to 25 are each worth 2 marks, questions 26 – 35 are each worth 3 marks; the total marks available are 80. There is **no** additional penalty for incorrect answers.
BE SURE TO ENTER THE CORRECT VERSION OF YOUR TEST (shown near the top of page 1), IN THE SPACE PROVIDED ON THE ANSWER SHEET.

ANSWER ALL QUESTIONS ON THE ANSWER SHEET, IN PENCIL.
Instructions for entering multiple-choice answers are given on page 2.
SELECT ONE AND ONLY ONE ANSWER FOR EACH QUESTION from the answers **(A)** through **(E). No work written on the question sheets will be marked**. The question sheets may be collected and reviewed in cases of suspected academic dishonesty.

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 you have a problem please ask the invigilator to deal with it for you. Do not make contact with other students directly. Try to keep your eyes on your own paper – looking around the room may be interpreted as an attempt to copy.

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.

Name: ____________________ Student number: ______________

OMR EXAMINATION – STUDENT INSTRUCTIONS

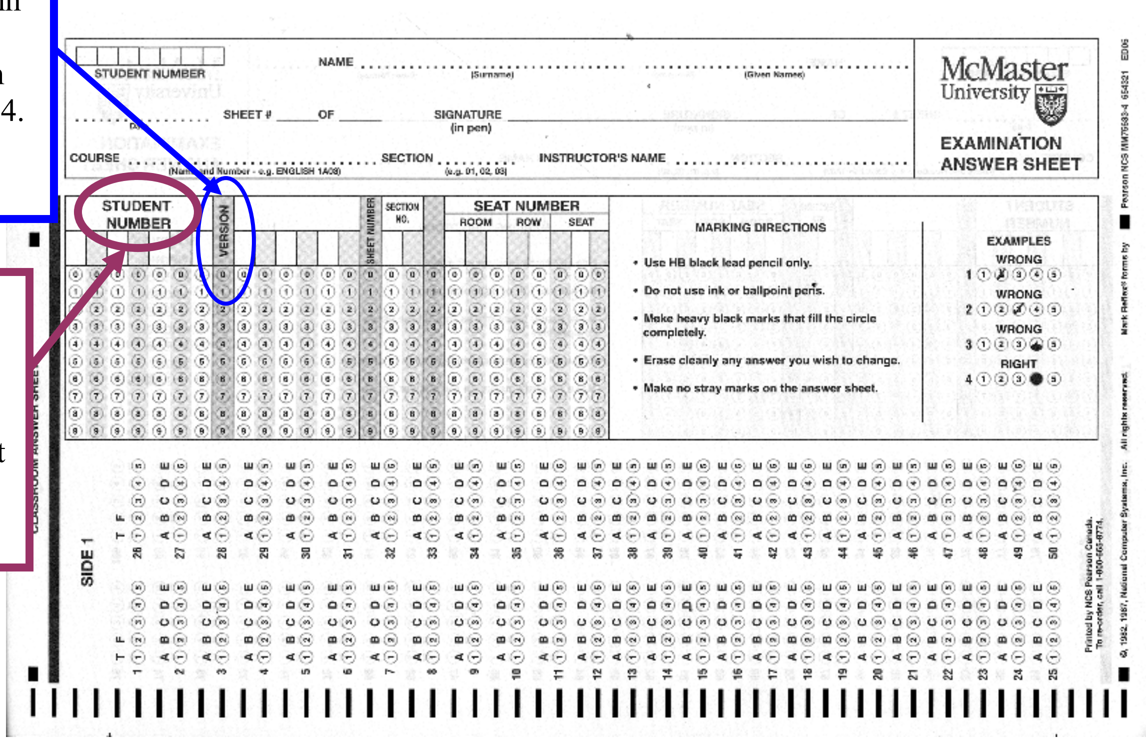
NOTE: IT IS YOUR RESPONSIBILITY TO ENSURE THAT THE ANSWER SHEET IS PROPERLY COMPLETED: YOUR EXAMINIATION RESULT DEPENDS UPON PROPER ATTENTION TO THESE INSTRUCTIONS.

The scanner, which reads the sheets, senses the bubble shaded areas by their non-reflection of light. A heavy mark must be made, completely filling the circular bubble, with an HB pencil. Do not use pen to bubble answers. Erasures must be thorough or the scanner will still sense a mark. Do **NOT** use correction fluid on the sheets. Do **NOT** put any unnecessary marks or writing on the sheet.

1. On SIDE 1 (**red side**) of the form, in the top box, *in pen*, print your student number, name, course name, ~~(section number, instructor name)~~ and the date in the spaces provided. Then you **MUST** write your signature, in the space marked SIGNATURE.
2. In the second box, *with a pencil*, mark your student number, **exam version number** (mark "1", "2", "3" or "4" and **do not** use the "0" bubble) (~~and course section number~~) in the space provided and fill in the corresponding bubble numbers underneath.
3. Answers: mark only **ONE** choice from the alternatives (~~1,2,3,4,5 or~~ A,B,C,D,E) provided for each question. ~~If there is a True/False question, enter response o 1 (or A) as True, and 2 (or B) as False.~~ The question number is to the left of the bubbles. Make sure that the number of the question on the scan sheet is the same as the number on the test paper.
4. Pay particular attention to the Marking+ Directions on the form.
5. Begin answering the question using the first set of bubbles, marked "1".

Use this column for version number. Fill in only 1, 2, 3 or 4. Do not use the 0.

Make sure your student number is filled in, and that the correct numbers are bubbled in.



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See pages 1 and 2 of the exam for instructions.

Questions 1-25 are worth two (2) marks each.

1. Which ONE of the following choices lists the halogen and halide species in order of **increasing size** (from smallest to largest)?

(A) $F^- < F < Cl$

(B) $F < F^- < Cl^-$

(C) $I^- < I < Br$

(D) $F^- < Cl^- < Cl$

(E) $Cl^+ < Cl^- < Cl$

2. In the upper atmosphere ozone can be destroyed with help from Cl radicals. These radicals can be generated from photolytic dissociation of Cl bonds in various molecules, including $Cl_2(g)$.
From the Cl-Cl bond energy (243 kJ mol^{-1}), calculate the **minimum frequency of light (in Hz, or s^{-1})** which will dissociate a Cl_2 molecule into Cl atoms.

(A) 3.67×10^{38}

(B) 7.63×10^{-38}

(C) 9.06×10^{15}

(D) 5.62×10^{-19}

(E) 6.09×10^{14}

 Continued on next page

3. You are put in charge of designing a system to filter Cd^{2+}(aq) cations from solution. Using the data below, which of the following salts would **be best suited to remove Cd^{2+}** from solution?

K_{sp} $Cd_3(PO_4)_2 = 1.2 \times 10^{-33}$; K_{sp} $CdF_2 = 6.44 \times 10^{-3}$; K_{sp} $Cd(OH)_2 = 17.2 \times 10^{-15}$

(A) $Cu_3(PO_4)_2$; $K_{sp} = 1.2 \times 10^{-37}$

(B) MgF_2; $K_{sp} = 5.16 \times 10^{-11}$

(C) $Zn(OH)_2$; $K_{sp} = 17.2 \times 10^{-17}$

(D) LiF; $K_{sp} = 1.84 \times 10^{-3}$

(E) Ag_3PO_4; $K_{sp} = 8.89 \times 10^{-17}$

4. The production of chlorine (Cl_2) by t**he chloralkali process is a** key industrial reaction that is widely used in water disinfection programs. However, chlorine must be isolated from hydroxide generated as a by-product in the reaction in order to avoid the following **disproportionation reaction** involving Cl_2:

$$Cl_2\ (aq) \rightarrow Cl^-\ (aq) + OCl^-\ (aq)$$

Determine the **smallest whole number coefficients** for the following species based on a **balanced** reaction under **alkaline (basic)** conditions: **Cl_2, OCl^- and OH^-.**

	Cl_2	OCl^-	OH^-
(A)	2	1	1
(B)	3	2	1
(C)	1	1	2
(D)	2	5	9
(E)	1	3	6

 Continued on next page

5. A major source of **fluoride** in natural ground water supplies in many countries is derived from the **dissolution** of calcium fluoride**, CaF_2 (s)** that is found in the bedrock. Calculate the **equilibrium concentration (mM) of fluoride** in water if K_{sp} (CaF_2) is 5.3×10^{-9} assuming that there are no other ions present in solution.

(A) 0.056
(B) 0.34
(C) 120
(D) 2.2
(E) 25

6. Which of the following substances will form **acidic solutions** when placed into water?

NaF CH_3COOLi $(CH_3)_2NH_2Br$ KNO_3 SO_3

(A) NaF, CH_3COOLi
(B) CH_3COOLi, SO_3
(C) $(CH_3)_2NH_2Br$, KNO_3
(D) KNO_3, NaF
(E) SO_3, $(CH_3)_2NH_2Br$

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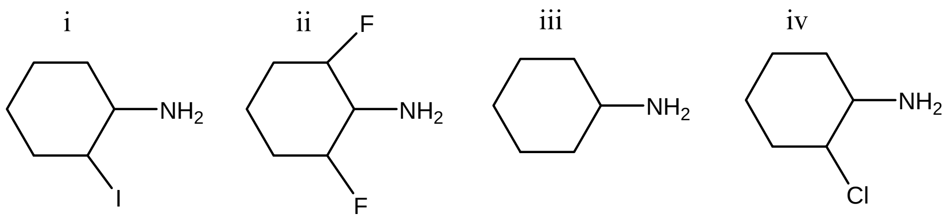
7. For the following **weak bases**, select the correct order of **INCREASING** base strength from **weakest base → strongest base** (or high pK_b → low pK_b):



(A) ii < iii < iv < i

(B) iii < ii < i < iv

(C) iv < i < iii < ii

(D) i < iv < ii < iii

(E) ii < iv < i < iii

8. Dissolving 3.90 g of CaF_2 in 50.0 mL of pure water (density = 1.00 g mL^{-1}) causes the temperature of the solution to decrease from 20.00 to 16.79 °C. What is the **<u>molar</u> enthalpy of dissolution of CaF_2 (in kJ mol^{-1}) ?** Assume that the specific heat of the solution equals 4.184 J g^{-1} K^{-1}, and that there is no solution volume change upon dissolution of the salt.

(A) –6.81

(B) –2.40

(C) +13.4

(D) +6.05

(E) +11.7

9. 0.40 mol of a gas is heated at constant pressure, so that its temperature increases from 21.00ºC to 79.83ºC. As a result, the gas expands from 10.0 L to 12.0 L against an external pressure of 0.965 atm. The molar heat capacity of the gas, at constant pressure, is 20.8 J mol^{-1} K^{-1}. What is the total energy change, **ΔU (in J) , for the gas?**

(A) -470
(B) –210
(C) +290
(D) +590
(E) +785

10. Which one of the following statements **is FALSE**?

(A) The reaction $CO(g) \rightarrow C(g) + O(g)$ causes work to be done on the surroundings.
(B) For the process $H_2O(g) \rightarrow H_2O(l)$, heat flows from system to surroundings.
(C) If excess gas is produced in a reaction occurring at constant temperature, then ΔH is greater than ΔU, (where greater means more positive, or less negative).
(D) In any process, the change of enthalpy depends on the path taken by the system to go from the initial to the final state.
(E) The enthalpy change for $O_2(g) \rightarrow 2\ O(g)$ is positive.

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11. Which one of the following processes is **endothermic**?

(A) Photosynthesis
(B) Cellular respiration
(C) Combustion of gasoline
(D) Detonation of an airbag in a car
(E) A body cooling down in frigid water

12. Calculate the standard enthalpy of formation of liquid benzene, C_6H_6, in kJ mol^{-1}, from its standard enthalpy of combustion which equals -3268 kJ mol^{-1}. Data:

$\Delta H_f°(H_2O,l) = -285.8$ kJ mol^{-1} $\Delta H_f°(CO_2,g) = -393.5$ kJ mol^{-1}
$\Delta H_f°(C, \text{diamond}) = 1.90$ kJ mol^{-1} $\Delta H_f°(CO,g) = -110.5$ kJ mol^{-1}

(A) -62 kJ mol^{-1}
(B) 53 kJ mol^{-1}
(C) -48 kJ mol^{-1}
(D) -42 kJ mol^{-1}
(E) 50 kJ mol^{-1}

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13. Choose the **FALSE** statement about phase changes. (Recall "fusion" = melting).

(A) The phase transition for the melting of paraffin wax requires energy, and therefore temperature decreases during the transition.
(B) In a phase change material energy is released by the material upon freezing and absorbed upon melting.
(C) The heat capacity of a material is determined by, for example, the vibrational and rotational degrees of freedom available to the molecules.
(D) The greater the amount of material, the greater the energy available as it freezes.
(E) At pressure = 1 atm, freezing occurs when $T < \Delta H^o(\text{fus}) / \Delta S^o(\text{fus})$.

14. Your body can extract all the energy from glucose ($C_6H_{12}O_6$) by aerobic metabolism (using oxygen; the net reaction is combustion). However, (without oxygen, when under stress) *anaerobic metabolism* converts one mole of glucose into two moles of lactic acid ($C_6H_{12}O_6 \rightarrow 2C_3H_6O_3$). How much **energy (in kJ) does the body extract** from the **anaerobic metabolism of one mole of glucose**? The energy of combustion of lactic acid is –321 kJ mol^{-1} and the energy of combustion for glucose is –2808 kJ mol^{-1}.

(A) 1183
(B) 2166
(C) 2287
(D) 3116
(E) 3050

15. A galvanic electrochemical cell (like those seen in experiment # 5) is constructed with both half-cell reactions involving the same metal (M) and different concentrations of $M^+(aq)$. Choose the **CORRECT** statement about the cell.

(A) The half-cell with the lower $[M^+]$ will contain the anode.

(B) A salt bridge is not needed in this cell.

(C) The weight of the cathode will decrease over time.

(D) A voltage of twice $E°_{red}$ for $M^+(aq)/M(s)$ will be observed.

(E) Electrons will flow from the half-cell of higher $[M^+]$ to the half-cell with lower $[M^+]$.

16. Given the following half-reactions, identify the **strongest reducing agent**.

	E^o_{red}
$Al^{3+} + 3e^- \rightarrow Al(s)$	-1.66
$AgBr(s) + e^- \rightarrow Ag(s) + Br^-$	$+ 0.07$
$Sn^{4+} + 2e^- \rightarrow Sn^{2+}$	$+ 0.14$
$Fe^{2+} + 2e^- \rightarrow Fe$	$- 0.44$

(A) $Sn^{4+}(aq)$

(B) $Al(s)$

(C) $Fe(s)$

(D) $Br^-(aq)$

(E) $Al^{3+}(aq)$

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17. Calculate E^0_{cell} (in volts) for the reaction as written:

$$4\ Al(s) + 3\ O_2(g) + 12\ H^+(aq) \rightarrow 4\ Al^{3+}(aq) + 6\ H_2O$$

Data: $O_2 + 4\ H^+ + 4\ e^- \rightarrow 2\ H_2O$ $E^0_{red} = +1.22\ V$

$Al^{3+} + 3\ e^- \rightarrow Al$ $E^0_{red} = -1.66\ V$

(A) +5.98

(B) +2.88

(C) +1.22

(D) -0.56

(E) -1.98

18. What is the value of the equilibrium constant, K, at 25 °C for the electrochemical cell described by the reaction $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$ for which $E°_{cell} = 0.78$ V?

(A) 3.0×10^{5}

(B) 1.9×10^{13}

(C) 2.4×10^{26}

(D) 7.2×10^{-3}

(E) 9.1×10^{-16}

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19. Knowing that $E°_{red}$ (Mg^{2+}/Mg) = -2.37 V, calculate the standard free energy change ($\Delta G°$ in kJ) for the following reaction:

$$MgCl_2(aq) + H_2(g) \rightarrow Mg(s) + 2\ HCl(aq)$$

(A) +457
(B) +209
(C) +174
(D) -269
(E) –356

20. Calculate the **cell potential (in V)** at 25ºC for a fuel cell using the combustion of hydrogen gas to form liquid water, when the pressure of H_2 gas is 0.50 atm and the pressure of O_2 gas is 0.25 atm. The standard cell potential for this fuel cell is 1.23 V. Hint: start by balancing the reaction.

(A) 1.16
(B) 1.21
(C) 1.24
(D) 1.27
(E) 1.31

 Continued on next page

21. The freezing point of ammonia ($NH_3(l) \rightarrow NH_3(s)$) is −78°C at 1 atm. Predict the **signs of ΔH, ΔS, ΔG** (in that order) for the freezing of ammonia at −80°C and 1 atm.

	ΔH	ΔS	ΔG
(A)	−	+	−
(B)	−	−	−
(C)	+	+	−
(D)	−	−	+
(E)	+	−	+

22. Which of the following processes are likely to have $\Delta S_{sys} > 0$?

(i) $NaCl(s) \rightarrow NaCl(aq)$

(ii) $HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$

(iii) $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

(iv) $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$

(v) $HCl(aq) + NaHCO_3(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$

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

23. The enthalpy of vaporization of HCN is 20.0 kJ J mol^{-1}. The molar entropy of the gas is 201.7 J mol^{-1} K^{-1}, and the molar entropy of the liquid is 112.8 J mol^{-1} K^{-1}. Calculate the **normal boiling point** of HCN.

(A) –48ºC
(B) –3.6ºC
(C) +5.9ºC
(D) –78ºC
(E) +22ºC

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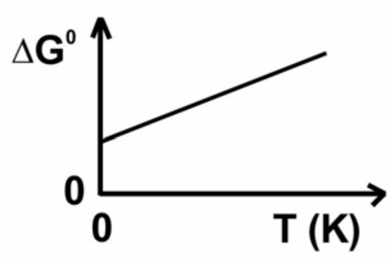
24. The figure below shows $\Delta G°$ for a chemical reaction plotted against the absolute temperature. Select the **TRUE** statement below.
(Hint: Consider $\Delta G° = \Delta H° - T\Delta S°$ as the equation of a line).



(A) The reaction is spontaneous at all temperatures.
(B) The reaction has a positive entropy.
(C) The reaction reaches equilibrium at very low temperature.
(D) The reaction is spontaneous at temperatures at T << 0.
(E) The reaction is endothermic.

25. Milk of magnesia, $Mg(OH)_2$, is a widely used antacid for treatment of acid reflux. Although it is a sparingly soluble salt in aqueous solution at neutral pH, its solubility is greatly enhanced under acidic conditions. With use of the data below, determine the **free energy change ($\Delta G°_{rxn}$)** at 25°C for the dissolution of Mg $(OH)_2$ and neutralization of excess stomach acid based on the following equation:

$$Mg(OH)_2\,(s) + 2H^+\,(aq) \overset{K}{\rightleftharpoons} Mg^{2+}\,(aq) + 2H_2O\,(l)$$

Species	$Mg(OH)_2(s)$	Mg^{2+} (aq)	$H_2O(l)$	$H^+(aq)$
$\Delta G_f°$ (kJ mol^{-1})	-833.5	-454.8	–237.1	0

(A) –95.5 kJ
(B) 126.1 kJ
(C) -148.3 kJ
(D) 17.6 kJ
(E) –278.4 kJ

Questions 26-35 are worth three (3) marks each.

26. Which of the following ions or molecules has the **fewest number of equivalent charge-minimized structures**?

(A) ClO_4^-

(B) SO_3^{2-}

(C) ClO_2^-

(D) NO_3^-

(E) AsO_3^{3-}

27. K_P = 2.00 for the reaction $2\ COF_2(g) \rightleftharpoons CO_2(g) + CF_4(g)$. All three gases are placed in a 5.00 L flask at 298K. The initial amounts are 0.145 moles of COF_2, 0.074 moles of CO_2 and 0.074 moles of CF_4. What is the **partial pressure** of CO_2 (g) at equilibrium?

(A) 0.529 atm
(B) 0.0217 atm
(C) 0.392 atm
(D) 0.0573 atm
(E) 0.118 atm

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28. "SO_x" gases contribute to the production of acid rain. Consider the following equilibrium for production of SO_3, and select the **TRUE statements** for the equilibrium in a constant volume container:

$$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g) \qquad \Delta H = -198.2\ \text{kJ};\ K_P = 1.02 \times 10^5 \text{ at } 425°\text{C}.$$

(i) At 425 °C, $\Delta G°_{rxn} = -66.9$ kJ.

(ii) The concentration of $SO_2(g)$ will decrease if temperature is increased.

(iii) Addition of water to form sulfuric acid ($SO_3(g) + H_2O(l) \rightarrow H_2SO_4(l)$) will lead to a decrease in the partial pressure of oxygen gas.

(iv) At 425 °C, if $P_{eq}(SO_2) = P_{eq}(SO_3) = 1.50$ atm, then $P_{eq}(O_2) = 3.1 \times 10^{-3}$ atm.

(v) Increasing the container volume will cause a decrease in the partial pressure of $SO_3(g)$.

(A) i, ii

(B) i, v

(C) ii, iv

(D) i, iii, v

(E) ii, iii, iv

 Continued on next page

29. Sodium monofluoroacetate ($NaOOCCH_2F$) is a highly toxic chemical produced by some classes of plants. Calculate the **pH** of a solution containing 600 mg of sodium monofluoroacetate dissolved in 20 mL water. (pK_a for $HOOCCH_2F$ = 2.6)

(A) 5.2
(B) 7.0
(C) 9.6
(D) 6.5
(E) 8.0

 Continued on next page

30. One rection in the creation of photochemical smog is **$O_3(g) + NO(g) \rightarrow NO_2(g) + O_2(g)$.** Use the following average bond enthalpies (BE) to estimate **the enthalpy change, ΔH (in kJ),** for this reaction.

Bond	BE (kJ mol^{-1})
O-O	142
O=O	498
N-O	222
N=O	590

(A) –468

(B) –59

(C) –80

(D) +46

(E) +82

 Continued on next page

31. A calorimeter is designed to monitor the energy absorbed during the sublimation of CO_2 (sublimation is $CO_2(s) \rightarrow CO_2(g)$, heat of sublimation = 570 J g^{-1}) at a constant temperature of 194.7 K. The oxidation of calcium metal (0.268 g) by excess oxygen released enough heat to cause CO_2 to sublime, producing a 2.34 atm pressure in a 1.00 L vessel. What is the **enthalpy of reaction in kJ mol^{-1}** for the oxidation of calcium metal? (Disregard any work done during the process).

(A) −550

(B) −215

(C) −75.6

(D) −891

(E) −348

32. Identify the **CORRECT** statements regarding the following electrochemical cell:

$Pt(s) \mid Fe(NO_3)_2(aq), Fe(NO_3)_3(aq) \parallel AgNO_3(aq) \mid Ag(s)$.

Data: $E°_{red}\ (Ag^+/Ag) = +0.799\ V$ $E°_{red}\ (Fe^{3+}/Fe^{2+}) = +0.771\ V$

(i) The spontaneous cell reaction is $Fe^{3+}(aq) + Ag(s) \rightarrow Fe^{2+}(aq) + Ag^+(aq)$.

(ii) $E°_{cell} = +0.028\ V$.

(iii) In the outer circuit, the electrons flow from the Pt to the Ag electrode.

(iv) Inside the cell, nitrate ions migrate from the Ag side to the Fe side.

(v) $\Delta G° = -5.40$ kJ for the reaction $Fe^{2+}(aq) + Ag^+(aq) \rightarrow Fe^{3+}(aq) + Ag(s)$.

(A) i, ii, v

(B) ii, iii, iv

(C) i, iv, v

(D) ii, iii

(E) ii, v

 Continued on next page

33. Identify the **TRUE** statement(s):

(i) A spontaneous endothermic reaction is driven by entropy.

(ii) $P_4(s) + 5\ O_2(g) \rightarrow P_4O_{10}(s)$ is an entropy-driven reaction.

(iii) For a (reactants + products) mixture at equilibrium, $\Delta G_{rxn} = 0$.

(iv) The reaction $C(graphite) + CO_2(g) \rightarrow 2\ CO(g)$ is spontaneous at temperatures below 600 °C. ($\Delta H_{rxn} = +172.5$ kJ, $\Delta S_{rxn} = +175.7$ J K^{-1}).

(v) At T = 0 K, a perfect crystal has zero entropy.

(vi) In a cold pack, the dissolution of ammonium nitrate is an exothermic process that leads to an increase in entropy, thus it is a spontaneous process.

(A) i, iii, iv, vi

(B) i, v

(C) ii, iii, vi

(D) i, iii, v

(E) ii, iv, vi

 Continued on next page

34. The **Haber Bosch** process is an important industrial chemical process for the synthesis of **ammonia** that is performed under high temperature and high pressure conditions for optimum performance. Calculate ***K*** K_P **at 400°C** for the reaction based on the standard state thermodynamic parameters below, assuming that the $\Delta H_f°$ and S° terms are unchanged from 25°C to 400°C

$$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$$

$\Delta H_f°\,[NH_3] = -46.11\ \text{kJ mol}^{-1}$; $S°\,[NH_3] = 192.5\ \text{J mol}^{-1}\,\text{K}^{-1}$

$S°\,[N_2] = 191.6\ \text{J mol}^{-1}\,\text{K}^{-1}$

$S°[H_2] = 130.7\ \text{J mol}^{-1}\,\text{K}^{-1}$

(A) 3.09×10^{5}

(B) 6.00×10^{-4}

(C) 1.27×10^{-10}

(D) 2.28×10^{12}

(E) 4.20×10^{-1}

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35. The formation of **ATP** is a major way that energy is stored chemically during metabolism based on the reaction below:

$$ADP(aq) + H_2PO_4^-(aq) + H^+(aq) \rightleftharpoons ATP(aq) + H_2O(l)$$

If $\Delta G° = -9.2$ kJ mol^{-1} for this reaction under standard conditions, what is the apparent **free energy (ΔG)** for the same reaction **inside a cell** under normal biological conditions, assuming [ADP] = 50 μM, [$H_2PO_4^-$] = 10 mM, [ATP] = 5.0 mM, pH = 7.4 at 37°C?

(A) +58 kJ mol^{-1}

(B) -25 kJ mol^{-1}

(C) -0.48 kJ mol^{-1}

(D) +130 kJ mol^{-1}

(E) +26 kJ mol^{-1}

 Continued on next page

This page is for rough work only.

Name: ____________________ Student number: ______________

Some general data are provided on this page and the next page. Other data appear with the questions.

A periodic table is provided on the next page.

STP = 273.15 K, 1 atm

F = 96485 C/mol

R = 8.3145 J/K·mol = 0.08206 L·atm/K·mol

$N_A = 6.022 \times 10^{23}$ mol^{-1}

1 atm = 760 mm Hg = 101.325 kPa

0°C = 273.15 K

1 J = 1 kg m^2 s^{-2} = 1 kPa·L = 1 Pa·m^3

1 m = 10^9 nm = 10^{10} Å

1 cm^3 = 1 mL

1 g = 10^3 mg = 10^{-3} kg

μM = 10^{-6} M

mM = 10^{-3} M

1 Hz = 1 cycle/s

c = 2.9979×10^8 m/s

h = 6.6256×10^{-34} J·s

$m_e = 9.10 \times 10^{-31}$ kg

λ = h / mu = h / p

$E_n = -R_H / n^2 = -2.179 \times 10^{-18} J / n^2$ (R_H is the energy form of the Rydberg constant for H)

$K_w = 1.0 \times 10^{-14}$ (25 °C)

$$\Delta G = \Delta G^o + RT \ln Q$$

$$E_{cell} = E^o{}_{cell} - \frac{RT}{nF} \ln Q = E^o{}_{cell} - \frac{0.0592}{n} \log_{10} Q$$

Solubility guidelines for Common Ionic Solids

Follow the lower-numbered guideline when two guidelines are in conflict. This leads to the correct prediction in most cases.

1. Salts of group 1 cations (with some exceptions for Li^+) and the NH_4^+ cation are soluble.
2. Nitrates, acetates, and perchlorates are soluble.
3. Salts of silver, lead, and mercury(I) are insoluble.
4. Chlorides, bromides, and iodides are soluble.
5. Carbonates, phosphates, sulfides, oxides, and hydroxides are insoluble (sulfides of group 2 cations and hydroxides of Ca^{2+}, Sr^{2+}, and Ba^{2+} are slightly soluble).
6. Sulfates are soluble except for those of calcium, strontium, and barium.

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Name: ____________________ Student number: ______________

PERIODIC TABLE OF THE ELEMENTS

I 1	II 2	3	4	5	6	7	8	9	10	11	12	III 13	IV 14	V 15	VI 16	VII 17	VIII 18
1 H 1.0079																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.086	15 P 30.974	16 S 32.066	17 Cl 35.453	18 Ar 39.948
19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.88	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.847	27 Co 58.933	28 Ni 58.69	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.922	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.468	38 Sr 87.62	39 Y 88.906	40 Zr 91.224	41 Nb 92.906	42 Mo 95.94	43 Tc [98]	44 Ru 101.07	45 Rh 102.91	46 Pd 105.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.75	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57 *La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po [209]	85 At [210]	86 Rn [222]
87 Fr [223]	88 Ra 226.03	89 **Ac 227.03	104 Unq [261]	105 Unp [262]	106 Unh [263]												

Transition Metals: groups 3–12

Atomic weights are based on $^{12}C = 12$ and conform to the 1987 IUPAC report values rounded to 5 significant digits.
Numbers in [ ] indicate the most stable isotope.

* Lanthanides	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm [145]	62 Sm 150.36	63 Eu 151.97	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
** Actinides	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np 237.05	94 Pu [244]	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 [262]

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