Name: _____ Student number: _____

Chemistry 1A03 **Exam** **December 8, 2014**

McMaster University **VERSION 1**

Instructors: D. Brock, G. Goward, J. Landry

Duration: 180 minutes

This test contains 24 numbered pages printed on both sides. There are **35** multiple-choice questions appearing on pages numbered 3 to 20. Pages 21 and 22 are extra space for rough work. Page 23 includes some useful data and equations, and there is a periodic table on page 24. 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.

All questions are worth 2 marks the total marks available are 76. 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.

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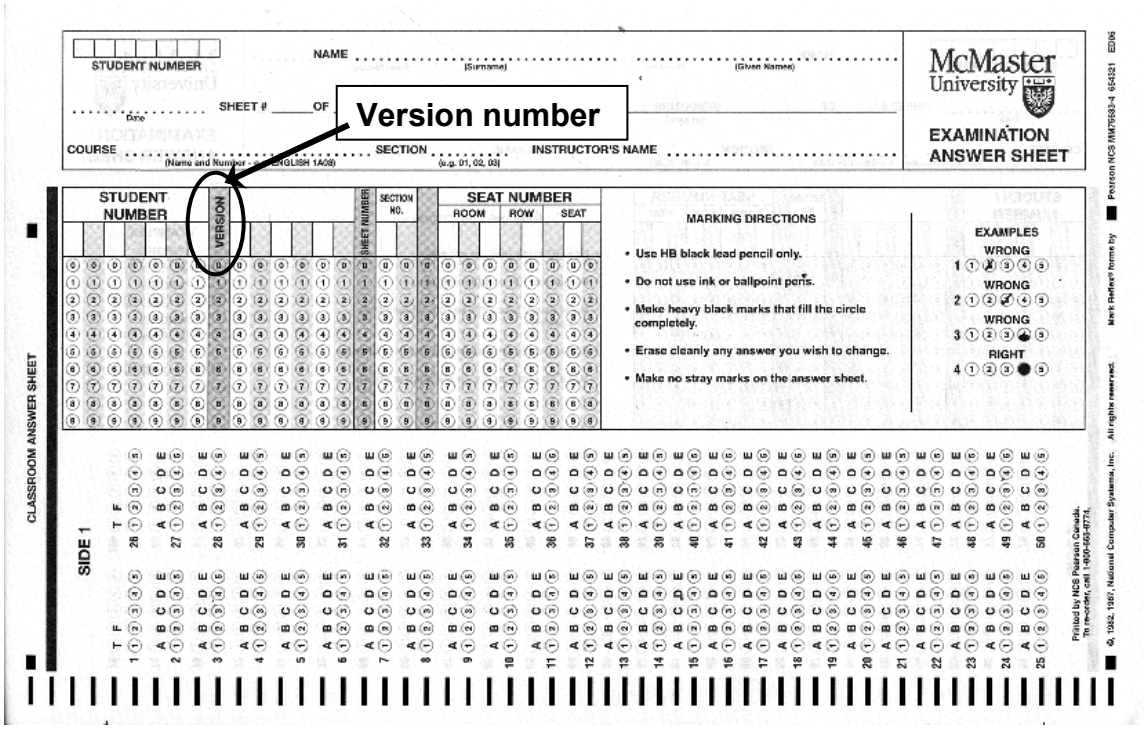
OMR EXAMINATION – STUDENT INSTRUCTIONS

NOTE: IT IS YOUR RESPONSIBILITY TO ENSURE THAT THE ANSWER SHEET IS PROPERLY COMPLETED: YOUT 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. Marks made with a pen will **NOT** be sensed. 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, 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** in the space provided and <u>fill in the corresponding bubble numbers underneath</u>.
3. Answers: mark only **ONE** choice from the alternatives (A,B,C,D,E) provided for each question. 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".



1. Which element has the **largest atomic radius**?

   A) Li
   B) Ga
   C) Cs
   D) F
   E) I

2. Under which of the following situations is it **not appropriate** to hit the panic alarm?

   A) Another student is having a seizure.
   B) A large fire has started and is spreading.
   C) A large spill (20 L) of concentrated nitric acid has occurred.
   D) Your TA has injured themselves and is bleeding profusely.
   E) More than one of the above.

3. Which atom has the **greatest first ionization energy (IE_1)**?

   A) Li
   B) Ga
   C) Ba
   D) Be
   E) B

4. Graphite and carbon dioxide are kept at constant volume (1.00 L vessel) and 1000. K until the following reaction has come to equilibrium. The initial pressure of $CO_2(g)$ is 1.00 bar, the mass of graphite is 12.0 g, and *K* is 1.746. What is the **equilibrium pressure of CO(g)**?

$$C(\text{graphite}) + CO_2(g) \rightleftharpoons 2CO(g)$$

   A) 0.726 bar
   B) 0.955 bar
   C) 0.889 bar
   D) 0.231 bar
   E) 0.524 bar

5. Determine the **FALSE** statement:

   A) A group and period intersect at a right angle on the periodic table.
   B) The atomic number of an element is always smaller than its average atomic mass.
   C) Considering two metals of equal volume, the metal with the highest density will have the lowest mass.
   D) Concentration is an intensive property.
   E) The oxidation number of P in PO_4^{3-} is +5.

6. In order to form ozone (O_3), the bonds of oxygen molecules must first be broken by sunlight. The minimum energy required to break the oxygen-oxygen bond is 495 kJ mol^{-1}. What is the **maximum wavelength** (in nm) of sun light that can break the oxygen-oxygen bond?

   A) 837
   B) 242
   C) 19.4
   D) 619
   E) 356

7. Which of the following statements is **FALSE** regarding HPO_4^{2-} and $H_2PO_4^-$? (H atoms are attached to O atoms)

   A) Both anions have a tetrahedral electron pair geometry around P.
   B) The two anions form a Bronsted-Lowry conjugate acid base pair.
   C) The average formal charges on the terminal oxygen atoms are −2/3 and −1/2, for HPO_4^{2-} and $H_2PO_4^-$, respectively.
   D) Both anions are represented by the same number of resonance structures.
   E) The average P-O bond order for terminal oxygen atoms is 4/3 for HPO_4^{2-} and 3/2 for $H_2PO_4^-$.

8. Consider the following equilibria

$$2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g) \qquad K = 5.7 \times 10^{-6}$$
$$2NO_3^-(aq) \rightleftharpoons 2NO_2^-(aq) + O_2(g) \qquad K = 3.5 \times 10^{-2}$$

   What is the **equilibrium constant, *K*,** for the reaction:

$$SO_2(g) + NO_3^-(aq) \rightleftharpoons NO_2^-(aq) + SO_3(g) \quad K = ?$$

   A) 4.7×10^{3}
   B) 9.2×10^{-9}
   C) 3.1×10^{-5}
   D) 6.5×10^{-7}
   E) 7.8×10^{1}

9. Determine the **FALSE** statement:

   A) NH_4^+ is the conjugate acid of NH_3.
   B) A Brønsted base is a proton acceptor.
   C) H_2F^+ would be a strong acid.
   D) H^+ is a Lewis base.
   E) $Mg(OH)_2$ would react to completion with HCl to produce a neutral solution.

10. Vitamin B9, also known as folic acid (FAH), helps the body convert carbohydrates into glucose for energy. It has a pK_a of 4.65. The conjugate base, Folate, FA^-, is found in many vitamin supplements.

    A 1.60 mg sample of folate (molecular mass = 441.40 g mol^{-1}) was dissolved in water and made up to a final volume of 10.0 mL. **What is the % dissociation** of this solution?

    A) 18
    B) 3.4
    C) 6.9
    D) 0.11
    E) 1.5

11. What is the **pH** of a 0.80 M solution of monochloroacetic acid ($CH_2ClCOOH$)?
$K_a = 1.35 \times 10^{-3}$

A) 1.59
B) 1.49
C) 2.87
D) 4.17
E) 1.30

12. Which of the following is the **strongest acid**?

A) OH^-
B) F^-
C) CH_4
D) HSO_4^-
E) NH_2^-

13. Which of the following salts would produce a solution with the **highest pH** (assume all are 1.0 M)?

A) NH_4I
B) $NaNH_2$
C) NH_4F
D) NH_4OH
E) NH_3

14. Consider a weak acid, HA with $K_a = 1.0 \times 10^{-4.}$ Which of the following is **FALSE** with respect to a 1.0 M solution of HA.

A) The small x approximation is valid when solving for the pH.
B) $K_a < [H_3O^+]$ in solution.
C) $[OH^-]$ in solution is 10^{-10} M.
D) pK_b of the conjugate base = $-\log 10^{-10}$.
E) The conjugate base is a weak base.

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15. The painkiller, Advil® contains the active ingredient ibuprofen (IB), which has a pK_b of 5.20. If the pH of a solution containing a dissolved tablet of Advil® was found to be 8.20, what was the **equilibrium concentration (mol L^{-1}) of ibuprofen**?

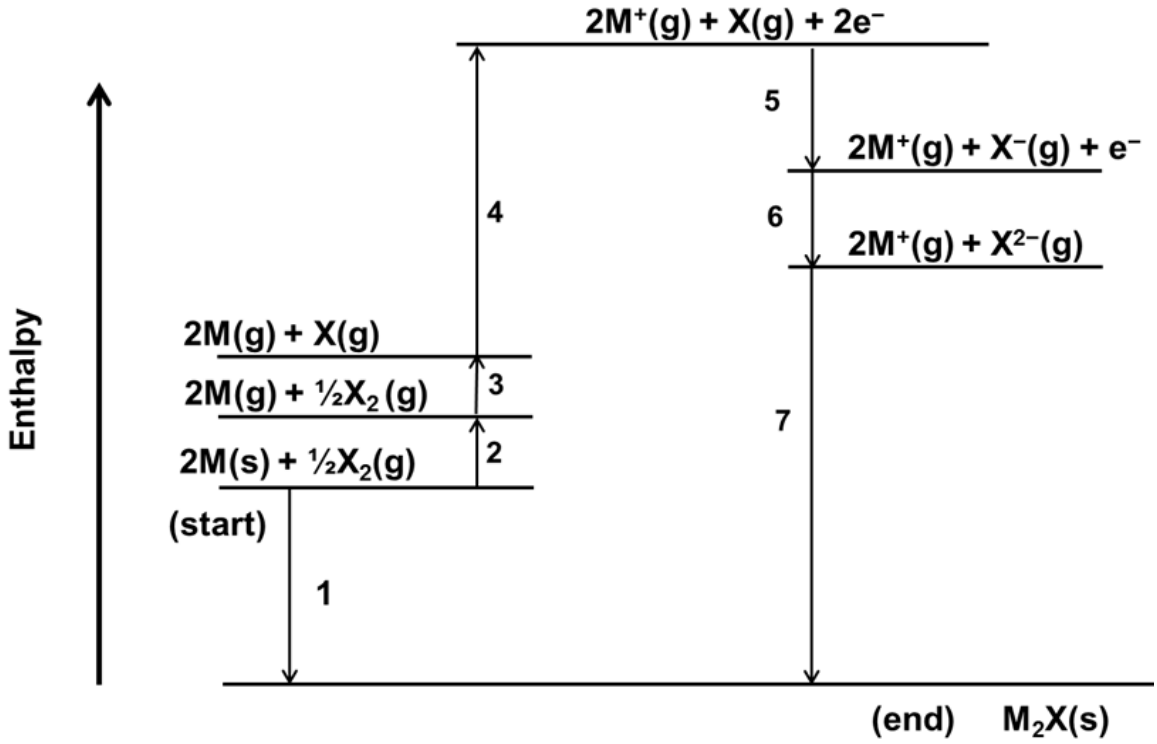
$$IB + H_2O \rightleftharpoons IBH^+ + OH^-$$

A) 3.7×10^{-4}
B) 5.5×10^{-5}
C) 2.8×10^{-7}
D) 4.0×10^{-7}
E) 9.2×10^{-7}

16. Using an ice calorimeter such as that in experiment # 4, it was determined that the reaction of 0.18 g of magnesium with excess HCl (aq), caused 0.68 g of ice to melt. What was the **heat of reaction per mole of Mg** (kJ mol^{-1})? [ΔH_{fus}(ice) = 333 J g^{-1}]

A) −180
B) 590
C) −95
D) 240
E) −31

17. Consider an ionic compound M_2X where M is a metal that forms a cation of +1 charge, and X is a nonmetal that forms an anion of –2 charge. A Born-Fajans-Haber cycle for M_2X is given below. Each step in this cycle has been assigned a number (1-7).



Identify the **FALSE** statement:

A) Step 5 is the electron affinity of X and is exothermic.
B) Step 3 is the bond enthalpy of X_2 and is endothermic.
C) Step 1 is the enthalpy of formation of M_2X and is exothermic.
D) Step 7 is the lattice enthalpy of M_2X and is exothermic.
E) Step 2 is twice the enthalpy of sublimation of M and is endothermic.

18. A reaction of 6.085 g of Na with excess water was carried out in constant pressure and constant volume calorimeters at 25.00 ^{o}C. What is the **difference in energy** between q_P and q_V.

    A) q_P and q_V are the same
    B) q_V is greater than q_P by 328 J
    C) q_V is greater than q_P by 947 J
    D) q_P is greater than q_V by 947 J
    E) q_P is greater than q_V by 328 J

19. Identify the **FALSE** statement among the following:

    A) If a system is heated during a process for which $\Delta U = 0$, then the system did work on its surroundings during this process.
    B) The thermite reaction, $Fe_2O_3(s) + 2\ Al(s) \rightarrow\ 2\ Fe(l) + Al_2O_3(s)$, is exothermic.
    C) The breaking of a chemical bond is always endothermic.
    D) For any compound, $\Delta H_f°$(liquid) is less than $\Delta H_f°$ (gas).
    E) If 100 J of work is done on a system during a process for which $\Delta U = 50$ J, then 50 J of heat flowed into the system from its surroundings during this process.

20. Find the **standard enthalpy of formation of ethylene, $C_2H_4(g)$ (in kJ),** given the following data:

$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$ $\Delta H_f^o = -1411$ kJ

$C(\text{graphite}) + O_2(g) \rightarrow CO_2(g)$ $\Delta H_f^o = -393.5$ kJ

$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$ $\Delta H_f^o = -285.8$ kJ

A) −2090
B) +486
C) −1090
D) +52
E) +732

21. Calculate the **enthalpy of formation, $\Delta H_f°$ in kJ,** of hydrogen peroxide, $H_2O_2(g)$, from the relevant bond energy data.

Bond	Bond energy	Bond	Bond energy
O-H	463 kJ mol^{-1}	H-H	436 kJ mol^{-1}
O=O	498 kJ mol^{-1}	O-O	138 kJ mol^{-1}

A) +130
B) –130
C) +285
D) –380
E) –285

22. A sample of aluminum, initially at 5.0 °C, was placed in 18.0 g of water, initially at 25.0 °C. The final temperature of the water and the metal was 21.0 °C. The specific heat of water is 4.184 J g^{-1} $°C^{-1}$ and aluminum is 0.940 J g^{-1} $°C^{-1}$. Ignore the heat capacity of the container. What is the **mass (in g) of the aluminum**?

A) 1.7
B) 20.
C) 68
D) 34
E) 59

23. A gas is allowed to expand at constant temperature from a volume of 1.0 L to 10.1 L against an external pressure of 0.50 bar. If the gas absorbs 250 J of heat from the surroundings, what are the values of ***q*, *w*, and Δ*U* (in J)**?

	q	***w***	**Δ*U***
A)	−250	−4.55	+245
B)	−250	+460	−210
C)	+250	−460	−210
D)	+250	+460	+710
E)	−250	−460	−710

24. Consider the following two reactions, with thermodynamic data at 298.15 K:

(1) $Pb(s) + CO(g) \rightarrow PbO(s) + C(graphite)$ $\Delta H° = -107$ kJ mol^{-1} $\Delta G° = -51$ kJ mol^{-1}
(2) $2C(graphite) + O_2(g) \rightarrow 2CO(g)$ $\Delta H° = -221$ kJ mol^{-1} $\Delta G° = -274$ kJ mol^{-1}

Which of the following statements are **FALSE**? (Assume $\Delta H°$ and $\Delta S°$ are independent of temperature.)

(i) $\Delta G_f°$ [PbO(s)] = +188 kJ mol^{-1}.
(ii) Both reactions are spontaneous under standard conditions at room temperature.
(iii) $\Delta S°$ for reaction (2) is +178 J mol^{-1} K^{-1} at 298.15 K.
(iv) Reaction (2) is spontaneous at 500 °C when the partial pressures of O_2 and CO are 1 bar each.
(v) Reaction (1) will become non-spontaneous as temperature is decreased.

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

25. Choose the **FALSE** statement regarding entropy.

A) $\Delta S° < 0$ for the reaction $H_2(g) + O_2(g) \rightarrow HOOH(g)$.
B) Entropy is a property of state; its changes do not depend on the path taken by the system.
C) Entropy is an intensive property, in other words, it does not depend on the amount of substance present.
D) For a spontaneous chemical reaction at 298.15 K and 1 bar, $\Delta G° < 0$.
E) For a system at equilibrium, $\Delta H_{sys} = T\Delta S_{sys}$.

26. The standard elemental form of mercury at 300. K is Hg(l). The standard enthalpy of formation for Hg(g) is +60.78 kJ mol^{-1}. The standard entropy of vapourization of mercury is +97.3 J mol^{-1} K^{-1}.

    Calculate the equilibrium **vapour pressure** (in bar) of mercury at 300. K.

    A) 21.7
    B) 3.16×10^{-6}
    C) 0.822
    D) 6.13×10^{5}
    E) 0.492

27. Identify the reaction with the **largest positive $\Delta S°$**.

    A) $PCl_5(g) \rightarrow PCl_3(g) + Cl_2(g)$
    B) $H_2O(s) \rightarrow H_2O(g)$
    C) $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$
    D) $KClO_4(s) + 4\ C(s) \rightarrow KCl(s) + 4\ CO(g)$
    E) $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$

28. Choose the expected order of **increasing molar entropy**.

A) $NH_4NO_3(aq) < C(diamond)(s) < C(graphite) < H_2O(l) < C_4H_{10}(g)$
B) $NH_4NO_3(aq) < H_2O(l) < C(diamond)(s) < C(graphite) < C_4H_{10}(g)$
C) $C(diamond)(s) < C(graphite) < NH_4NO_3(aq) < C_4H_{10}(g) < H_2O(l)$
D) $C(diamond)(s) < C(graphite) < H_2O(l) < NH_4NO_3(aq) < C_4H_{10}(g)$
E) $C(graphite)(s) < C(diamond) < H_2O(l) < NH_4NO_3(aq) < C_4H_{10}(g)$

29. The reaction of nitrogen with oxygen under high pressure and temperature is a source of nitric oxide, NO, precursor to smog. Calculate $\Delta G°$ (in kJ) for the following reaction at 25°C:

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g) \qquad \Delta H_f°[NO, g] = +90.4 \text{ kJ mol}^{-1}$$

$S°[N_2, g] = 191.5 \text{ J mol}^{-1}\text{ K}^{-1}$
$S°[O_2, g] = 205.0 \text{ J mol}^{-1}\text{ K}^{-1}$
$S°[NO, g] = 210.6 \text{ J mol}^{-1}\text{ K}^{-1}$

A) 528.5
B) 263.8
C) 173.4
D) −893.2
E) 101.3

30. In the lab you create a concentration cell. What **ratio of concentrations** is needed for this concentration cell to produce a voltage of 0.100 V (z = 2, T = 298 K)? Consider the solutions to be both $M^{2+}(aq)/M(s)$ where M = a metal.

A) 6.95×10^{6}
B) 7.59×10^{5}
C) 1.56×10^{4}
D) 4.15×10^{-4}
E) 2.26×10^{-5}

31. Given the following reduction potentials, what species, either reactant or product has the **greatest tendency to be oxidized**?

$S(s) + 2H^{+} + 2e^{-} \rightarrow H_2S(g)$	$E^{o} = 0.144$ V
$O_2(g) + 4H^{+} + 4e^{-} \rightarrow 2\ H_2O(l)$	$E^{o} = 1.229$ V
$Mg^{2+}(aq) + 2e^{-} \rightarrow Mg(s)$	$E° = -2.356$ V

A) $Mg^{2+}(aq)$
B) O_2 (g)
C) $H_2O(l)$
D) Mg(s)
E) S(s)

32. Determine the **FALSE** statement regarding electrochemistry.

A) Electrons travel towards the cathode in a galvanic cell.
B) E_{anode} must always be lower than $E_{cathode}$ in a spontaneous electrochemical cell.
C) In a concentration cell, electrons flow from the higher to lower concentration half-cell.
D) During electrolysis, electricity is needed to drive a non-spontaneous redox reaction.
E) The purpose of an inert electrode is to facilitate the transfer of electrons.

33. When the following redox reaction is balanced in basic solution, what is the **coefficient in front of hydroxide**?

$$NH_3\,(aq) + OCl^-\,(aq) \rightarrow Cl_2\,(g) + N_2H_4\,(aq)$$

A) 3
B) 1
C) 6
D) 2
E) 4

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34. How many **electrons are gained** (+) or lost (–) when oxide (O^{2-}) turns into peroxide (O_2^{2-})? Use the lowest whole number coefficients to balance the half-reaction.

A) –2
B) 0
C) +1
D) –1
E) +2

35. Three electrochemical cells are built based on the following reductions which are listed in order of decreasing reduction potential (1 is highest, 3 lowest)

1) Hg^{2+} (aq) + 2 e^- → Hg (s)
2) Cu^+ (aq) + 1 e^- → Cu (s)
3) In^{3+} (aq) + 3 e^- → In (s)

If all concentrations used are 0.20 M, given the following generated potentials, determine **$E°_{red}$ for each metal** (V).

Reduction	Oxidation	Potential (V)
Hg^{2+}/Hg	Cu/Cu^+	0.35
Hg^{2+}/Hg	In/In^{3+}	1.18
Cu^+/Cu	In/In^{3+}	0.82

A) Hg = 0.99; Cu = 0.48; In = 0.22
B) Hg = 0.67; Cu = –0.14; In = –0.55
C) Hg = 0.75; Cu = 0.33; In = –0.16
D) Hg = 0.88; Cu = 0.25; In = –0.30
E) Hg = 0.85; Cu = 0.52; In = –0.34

Extra space for rough work:

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- **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
$N_A = 6.022 \times 10^{23}$ mol^{-1}
$h = 6.6256 \times 10^{-34}$ Js
density(H_2O, l) = 1.00g/mL
F = 96485 C/mol
$c = 2.9979 \times 10^{8}$ m/s
$m_e = 9.109 \times 10^{-31}$ kg
$\Delta H^o_{vap}[H_2O] = 44.0$ kJ mol^{-1}
Specific heat of water = 4.184 J / g·°C
$R = 8.3145$ J K^{-1} mol^{-1} = 0.08206 L atm K^{-1} mol^{-1} = 0.083145 L bar K^{-1} mol^{-1}

1 bar = 100.00 kPa = 750.06 mm Hg = 0.98692 atm
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 = 10^6 µm = 10^9 nm = 10^{10} Å
1 g = 10^3 mg

De Broglie wavelength:
$\lambda = h / mu = h / p$

Hydrogen atom energy levels:
$E_n = -R_H / n^2 = -2.179 \times 10^{-18}$ J / n^2
$KE = ½mu^2$

Nernst Equation:

$$E = E° - \frac{RT}{zF}\ln Q \quad = E° - \frac{0.0257\ \text{V}}{z}\ln Q \quad = E° - \frac{0.0592\ \text{V}}{z}\log_{10} Q$$

Entropy change: $\Delta S = \frac{q_{rev}}{T}$

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 and the NH_4^+ cation are soluble . Except LiF and Li_2CO_3 which are insoluble.
2. Nitrates, acetates, bicarbonates, and perchlorates are soluble.
3. Salts of silver, lead and mercury (I) are insoluble. Except AgF which is soluble.
4. Fluorides, chlorides, bromides, and iodides are soluble. Except Group 2 fluorides which are insoluble
5. Carbonates, phosphates, chromates, sulfides, oxides, and hydroxides are insoluble. Except Group 2 sulfides and hydroxides of Ca^{2+}, Sr^{2+}, and Ba^{2+} which are soluble.).
6. Sulfates are soluble except for those of calcium, strontium, and barium.

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PERIODIC TABLE OF THE ELEMENTS

ALDRICH®

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					Transition Metals						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]												

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]

Continued on next page

END OF EXAM