Name: _____ Student number: _____

Chemistry 1A03 **Exam** **Dec 12, 2015**

McMaster University **VERSION 1**

Instructors: D. Brock, G. Goward, A. Hitchcock, L. Davis

Duration: 150 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 70. There is **no** 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 any aids other than those provided, is not allowed.

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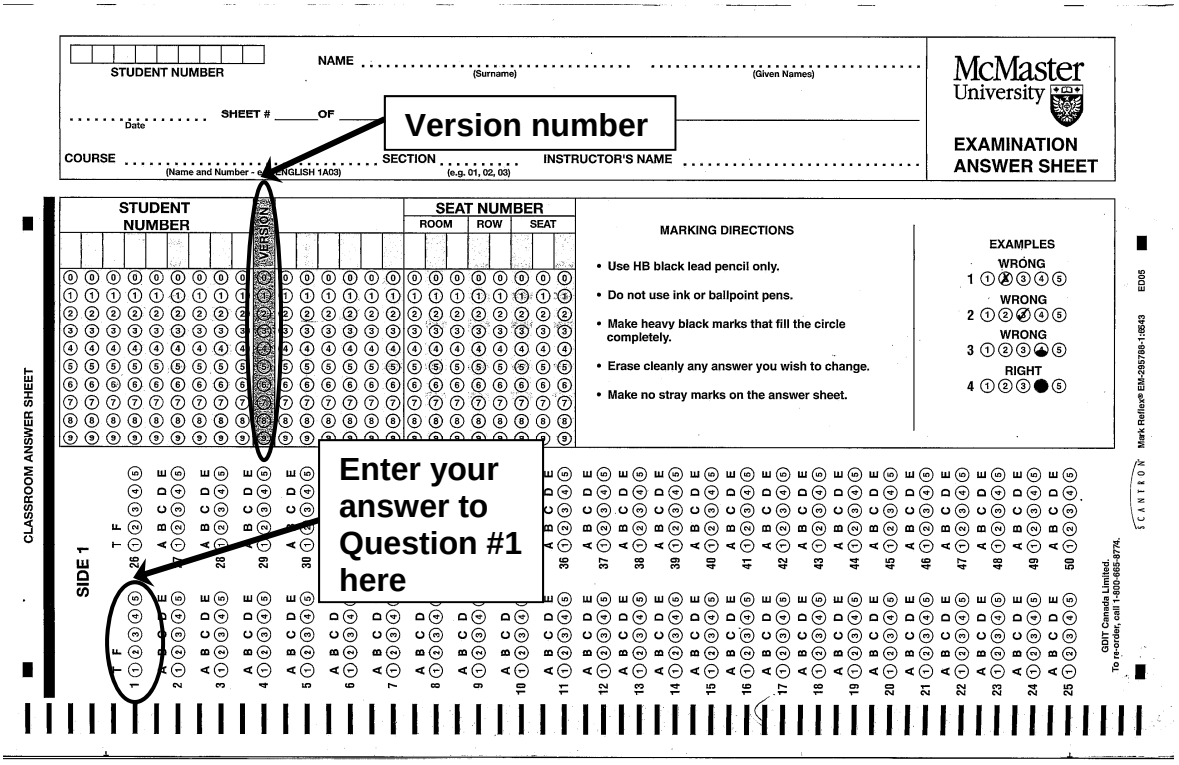
OMR EXAMINATION – STUDENT INSTRUCTIONS

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 and the **exam version number** in the space provided. If your student number does **NOT** begin with a 4, put "00" before your student number. Then fill in the corresponding bubble numbers underneath. Do not put in a leading zero when bubbling in your version number.
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".



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1. Which atomic property **decreases down** a group?

   A) the ionic radius
   B) the atomic radius
   C) the metallic character
   D) the core charge
   E) the first ionization energy

2. Which one of the following statements about $SeOF_3^-$ and $BrO_2F_2^-$ is **FALSE**?

   A) $SeOF_3^-$ and $BrO_2F_2^-$ are both polar molecules
   B) The average Br-O bond order is greater for $BrO_2F_2^-$ than the average Se-O bond order in $SeOF_3^-$.
   C) $SeOF_3^-$ has more resonance structures than $BrO_2F_2^-$.
   D) The average formal charge on O is more negative for $SeOF_3^-$ than $BrO_2F_2^-$.
   E) The average Se-O bond length is greater in $SeOF_3^-$ than the average Br-O bond length in $BrO_2F_2^-$.

3. Which one of the following statements about IOF_3 is **FALSE**?

   A) The electron pair geometry is trigonal bipyramidal.
   B) The molecule is non-polar.
   C) The formal charge on each atom is 0.
   D) The bond angles are non-ideal.
   E) The molecular shape is seesaw.

4. The four quantum numbers used to designate electrons in atoms are the set (n, ℓ, m_ℓ, m_s). Consider **the compound AX**. The quantum numbers of the 6 highest energy electrons in their ground states in A are:
   (5, 1, 1, ½), (5,1,1,−½), (5, 1,0, ½), (5,1,0, −½), (5, 1,−1, ½), (5,1,−1,−½).

   The quantum numbers of the 6 highest energy electrons in their ground states in X are:
   (2, 1, 1, ½), (2,1,1,−½), (2, 1, 0, ½), (2, 1,0,−½), (2, 1,−1, ½), (2,1,−1,−½).

   What are the elements A and X?

   A) Cs and F
   B) K and Cl
   C) Xe and Ne
   D) Cs and N
   E) Rb and O

5. In the reaction of hydrofluoric acid with water to produce hydronium ion and fluoride, **water** is acting as?

i) Arrhenius Base
ii) Brønsted-Lowry Acid
iii) Lewis Base
iv) Brønsted-Lowry Base
v) Lewis Acid
vi) Arrhenius Acid

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

6. Which one of the following would be the **strongest acid**?

A) CH_3OH
B) CF_3COH
C) CF_3COOH
D) CCl_3OH
E) CCl_3COOH

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7. A 2.5 M solution of formic acid (HCOOH) has a pH of 1.67. What is the **pK_b** of the formate ion ($HCOO^-$) ?

   A) 11.12
   B) 10.27
   C) 4.56
   D) 8.59
   E) 9.44

8. $\Delta G^\circ = 25.5$ kJ mol^{-1} for the dissociation of a weak monoprotic acid (HA) in water at 298. K. What would be the **pH** for this acid at standard conditions?

   A) 2.69
   B) 1.99
   C) 3.13
   D) 2.24
   E) 2.15

9. A student reacts 5.00 mL of 0.00255 M $Fe(NO_3)_3$ (aq) with 1.00 mL of 0.00375 M KSCN (aq) resulting in a total volume of 6.00 mL. The student measures an absorbance of 0.824 for the product. If the slope of the calibration curve relating the absorbance of $FeSCN^{2+}$ (aq) to the concentration of $FeSCN^{2+}$ (aq) is 4120 M^{-1}, what is the **equilibrium constant** for the reaction?

   A) 173
   B) 244
   C) 68.3
   D) 371
   E) 89.1

10. Consider the following reaction mixture, where the mass of SnO_2 is 10. g, and the partial pressures of CO and CO_2 at temperature, T, are P_{CO} = 1 bar, and P_{CO2} = 3 bar.

    $SnO_2(s) + 2CO(g) \rightleftharpoons Sn(s) + 2CO_2(g)$ K = 11 at temperature T

    Select the **TRUE** statement:

    A) $Q < K$ and the reaction proceeds in the forward direction
    B) $Q < K$ and the reaction proceeds in the reverse direction
    C) $Q = K$ and the reaction is at equilibrium
    D) $Q > K$ and the reaction proceeds in the forward direction
    E) $Q > K$ and the reaction proceeds in the reverse direction

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11. Consider the gas phase equilibrium between hydrogen and iodine.

$$H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \qquad K = 33 \text{ at temperature, T.}$$

Suppose a vessel is filled with 1.00 bar $H_2(g)$ and 0.100 bar $I_2(g)$. What is the **partial pressure of HI** when equilibrium is established at temperature T?

A) 0.078 bar
B) 0.136 bar
C) 0.178 bar
D) 0.050 bar
E) 0.197 bar

12. Which of the following statements is **FALSE**?

A) For any process, the change of total energy for the system is equal to the sum of work and heat.
B) In any process, the change of enthalpy depends on the path taken by the system to go from the initial to the final state.
C) When reacting HCl(aq) with Zn(s) the measured value of q would be different in a coffee cup calorimeter when compared to a bomb calorimeter.
D) When two objects at different temperatures are in contact, heat will flow between them.
E) Entropy is a property of state; its changes do not depend on the path taken by the system.

13. In which of the following reactions/processes that were presented in class would you expect **ΔH to be positive**?

    A) Thermite reaction, $Fe_2O_3(s) + Al(s) \rightarrow Al_2O_3(s) + Fe(l)$.
    B) Conversion of smog (NO_2) into N_2O_4.
    C) Rocket fuel, sugar and potassium chlorate.
    D) Crystallization of sodium acetate.
    E) Dissolution of solid ammonium nitrate in water.

14. Given the following catalyzed reaction,

    $$S_8(s) + 12O_2(g) \rightarrow 8SO_3(g)$$

    if 1.00 mol of S_8 is burned at 298 K and 1.00 bar pressure, what is the value for **work** in (kJ)?

    A) −1.86
    B) −3.27
    C) +9.91
    D) −9.91
    E) +3.27

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15. A 20.0 g sample of copper is heated from 25 °C to 1083 °C and **half** of it is melted at that temperature. The specific heat (sometimes referred to as the specific heat capacity) of copper is 0.385 J g^{-1} K^{-1} and the heat of fusion of copper is 205 J g^{-1}. How much **heat, in kJ**, is absorbed by the sample of copper?

A) +6.43
B) +10.2
C) –6.43
D) +9.66
E) –10.2

16. Using ice calorimetry, the heat of reaction was monitored when 2.00 mL of 5.241 M standardized HNO_3 (aq) reacted with excess NaOH (aq). This reaction caused 0.570 g of ice to melt. What was the **heat of reaction per mole of HNO_3** (in kJ mol^{-1})? (ΔH_{fus}(ice) = 333 J g^{-1})

A) +18.1
B) +5.21
C) −18.1
D) −31.0
E) +31.0

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17. Calculate the **lattice enthalpy** (in kJ mol^{-1}) of potassium iodide from the following data:

ΔH^{o}_{f} [K(g)] = 89 kJ mol^{-1}
First ionization energy of K(g) = 418 kJ mol^{-1}
ΔH^{o}_{f} [I(g)] = 107 kJ mol^{-1}
Electron affinity of I(g) = –295 kJ mol^{-1}
ΔH^{o}_{f} [KI(s)] = –328 kJ mol^{-1}

A) –1237
B) –581
C) –647
D) –1532
E) –754

18. Which one of the following processes is expected to have **ΔH more positive than ΔU**?

A) $CO_2(g) \rightarrow CO_2(s)$
B) $H_2O(g) \rightarrow H_2O(l)$
C) $Cl_2(g) \rightarrow 2Cl(g)$
D) $Li(g) + F(g) \rightarrow Li^+(g) + F^-(g)$
E) $Cl_2(aq) + 2KBr(aq) \rightarrow 2KCl(aq) + Br_2(aq)$

19. The reaction of hydrogen gas and chlorine gas to form hydrogen chloride gas is a chain reaction. The bond energies involved are: Cl-Cl: 243 kJ mol^{-1}, H-H: 436 kJ mol^{-1}, H-Cl: 431 kJ mol^{-1}. Balance the reaction with lowest whole number coefficients and calculate the **enthalpy of reaction**.

A) +183
B) –250.
C) 1110
D) –183
E) +250.

20. Calculate the standard **entropy of reaction, $\Delta S°$** (in $J\ K^{-1}\ mol^{-1}$), for:

$$Na_2CO_3(aq) + 2\ HCl(aq) \rightarrow 2\ NaCl(aq) + H_2O(l) + CO_2(g)$$

given the following data:

	$Na_2CO_3(aq)$	$HCl(aq)$	$NaCl(aq)$	$H_2O(l)$	$CO_2(g)$
$S°$ ($J\ K^{-1}\ mol^{-1}$)	134.98	56.5	115.5	69.91	213.74

A) +266.7
B) –207.7
C) –266.7
D) +207.7
E) +323.2

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21. Calcium sulfate can be used as a dessicant to take up water, due to its ability to absorb water to form calcium sulfate dihydrate:

$$CaSO_4\,(s) + 2H_2O(g) \rightleftharpoons CaSO_4 \cdot 2H_2O(s)$$

for this hydration process: $\Delta H^o = -104.9$ kJ mol^{-1} and $\Delta S^o = -290.2$ J K^{-1} mol^{-1}. What is the **value of *K* at 25°C**?

A) 7.53×10^5
B) 1.69×10^3
C) 4.69×10^3
D) 3.28×10^8
E) 2.72×10^9

22. Calculate the **temperature of the boiling point (in °C)** of liquid bromine (to the nearest 1 °C), at P = 1 bar given the following data:

$\Delta H_f°\ [Br_2(g)] = 30.907$ kJ mol^{-1}
$S°\ [Br_2(l)] = 152.2$ J K^{-1} mol^{-1}
$S°\ [Br_2(g)] = 245.463$ J K^{-1} mol^{-1}

A) 355
B) 4
C) 71
D) 58
E) 262

23. Order the following set of substances according to **increasing molar entropy**.

$CO_2(g)$, $CO_2(s)$, $CS_2(g)$, $CS_2(s)$

A) $CO_2(g) < CO_2(s) < CS_2(g) < CS_2(s)$
B) $CS_2(s) < CO_2(s) < CS_2(g) < CO_2(g)$
C) $CS_2(s) < CO_2(s) < CO_2(g) < CS_2(g)$
D) $CO_2(s) < CS_2(s) < CS_2(g) < CO_2(g)$
E) $CO_2(s) < CS_2(s) < CO_2(g) < CS_2(g)$

24. Select the **FALSE** statement from the following:

A) $\Delta S° < 0$ for the reaction $H_2(g) + O_2(g) \rightarrow HOOH(g)$.
B) For a system at equilibrium, $\Delta H_{sys} = -T\Delta S_{sys}$.
C) The free energy of a system is a measure of its disorder.
D) The standard entropy change for a chemical reaction can be calculated from the absolute entropies of reactants and products.
E) For a spontaneous chemical reaction at 298 K and 1 bar, $\Delta G° < 0$.

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25. For which of the following reactions would you predict $\Delta H_{rxn}^{\circ} < 0$ and $\Delta S_{rxn}^{\circ} > 0$?

i) $2NI_3(s) \rightarrow 3I_2(g) + 2N_2(g)$
ii) $Pb^{2+}(aq) + 2I^{-}(aq) \rightarrow PbI_2(s)$
iii) $N_2O_4(g) \rightarrow 2NO_2(g)$
iv) $5C_{12}H_{22}O_{11}(s) + 48KNO_3(s) \rightarrow 24K_2CO_3(s) + 36CO_2(g) + 55H_2O(l) + 24N_2(g)$
v) $NH_4NO_3(s) + H_2O(l) \rightarrow NH_4^{+}(aq) + NO_3^{-}(aq) + H_2O(l)$

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

26. Determine ΔS_{univ} **in J K mol^{-1}** for the formation reaction of one mole of liquid hydrazine (N_2H_4) at 25°C, given:

$\Delta H^{o}_{f}\,[N_2H_4(l)] = +50.6$ kJ mol^{-1}
$\Delta S^{o}_{f}\,[N_2H_4(l)] = +121.2$ J K^{-1} mol^{-1}

A) +48.5
B) −384
C) +384
D) −75.6
E) −48.5

27. Consider the reaction:

$$H_2(g) + S(s) \rightarrow H_2S(g)$$

for which $\Delta H° = -20$ kJ mol^{-1} and $\Delta S° = +43$ J K^{-1} mol^{-1} at 298.15 K.

Choose the **TRUE** statement.

A) The reaction is spontaneous at all temperatures.
B) The reaction is not spontaneous at 298.15 K.
C) The reaction is driven by enthalpy only.
D) The reaction is not spontaneous at any temperature.
E) The reaction is spontaneous only above a temperature of 465 K.

28. A concentration cell is created based on two half cells containing Co^{2+}/Co. Which one of the following statements is **FALSE** ?

A) Electrons are being transferred from the concentrated half-cell to the diluted half-cell.
B) $E°_{cell} = 0.0$ V
C) The flow of electrons will stop when the concentrations of Co^{2+} (aq) become equal in both half-cells.
D) Co^{2+} (aq) in the concentrated half-cell is being reduced
E) Co (s) in the diluted half-cell is being oxidized.

29. 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) 2
B) 3
C) 4
D) 6
E) 1

30. Identify the **one FALSE** statement regarding the electrochemical cell

$$Co(s)\,|\,CoSO_4\,(1.00\ M)\,||\,Fe(NO_3)_3\,(1.00\ M)\,,\,Fe(NO_3)_2\,(1.00\ M)\,|\,Pt(s)$$

for which $E_{cell} = +1.05$ V. The cell contains a KCl salt bridge.

A) Pt(s) is the cathode.
B) Decreasing the concentration of $CoSO_4$(aq) reduces the cell potential.
C) The Co^{2+} concentration increases during operation of the cell.
D) Cl^- ions from the salt bridge migrate to the anode.
E) Fe^{3+} ions migrate toward the Pt(s) electrode.

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31. A student creates an electrochemical cell based on 5.00 x10^{-3} mol of both O_2 (g) and H_2S (g) in separate 2.0 L vessels equipped with Pt (s) electrodes in acidic solution at 298 K. What would be the **voltage** produced ?

$S(s) + 2H^+ + 2e^- \rightarrow H_2S\ (g)$ $\quad E^o = +0.144$ V
$O_2(g) + 4H^+ + 4e^- \rightarrow 2\ H_2O\ (l)$ $\quad E^o = +1.229$ V

A) +0.932 V
B) +1.145 V
C) +1.002 V
D) +1.031 V
E) +1.088 V

32. Given the following reduction potentials, which **combination of metal ions** (with electrodes made of the same metal as the ions in their solutions) would give a potential closest to **+0.013 V** at standard conditions?

$Cu^{2+}(aq) + 2e^- \rightarrow Cu\ (s)$; $\quad E°_{red} = +0.340$ V
$Zn^{2+}(aq) + 2e^- \rightarrow Zn\ (s)$; $\quad E°_{red} = -0.763$ V
$Pb^{2+}(aq) + 2e^- \rightarrow Pb\ (s)$; $\quad E°_{red} = -0.125$ V
$Sn^{2+}(aq) + 2e^- \rightarrow Sn\ (s)$; $\quad E°_{red} = -0.137$ V

A) Sn^{2+}/Zn^{2+}
B) Cu^{2+}/Sn^{2+}
C) Pb^{2+}/Sn^{2+}
D) Cu^{2+}/Pb^{2+}
E) Pb^{2+}/Zn^{2+}

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33. What is the **equilibrium constant** for the reaction taking place in an electrochemical cell based on $I_2/2I^-$ and Pb^{2+}/Pb for which $E°_{cell} = +0.66$ V at 298 K?

A) 2.0×10^{22}
B) 6.3×10^{12}
C) 7.2×10^{4}
D) 56
E) 5.1×10^{-3}

34. Determine $\mathbf{E°_{cell}}$ for the reaction: $2\ Ag^+ + Mg \rightarrow 2\ Ag + Mg^{2+}$. The half reactions are:

$$Mg^{2+}(aq) + 2\ e^- \rightarrow Mg(s) \qquad E° = -2.356\ V$$
$$Ag^{+}(aq) + e^- \rightarrow Ag(s) \qquad E° = +0.800\ V$$

A) +1.556 V
B) −2.356 V
C) −1.556 V
D) +3.156 V
E) +3.956 V

35. 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) $Mg(s)$
C) $O_2\ (g)$
D) $H_2O(l)$
E) $S(s)$

Name: _____ Student number: _____

Extra space for rough work:

Name: _____ Student number: _____

Extra space for rough work:

Name: _____ Student number: _____

- **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
Specific heat of water = 4.184 J / g·°C

$F =$ 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}

$R = 8.3145$ 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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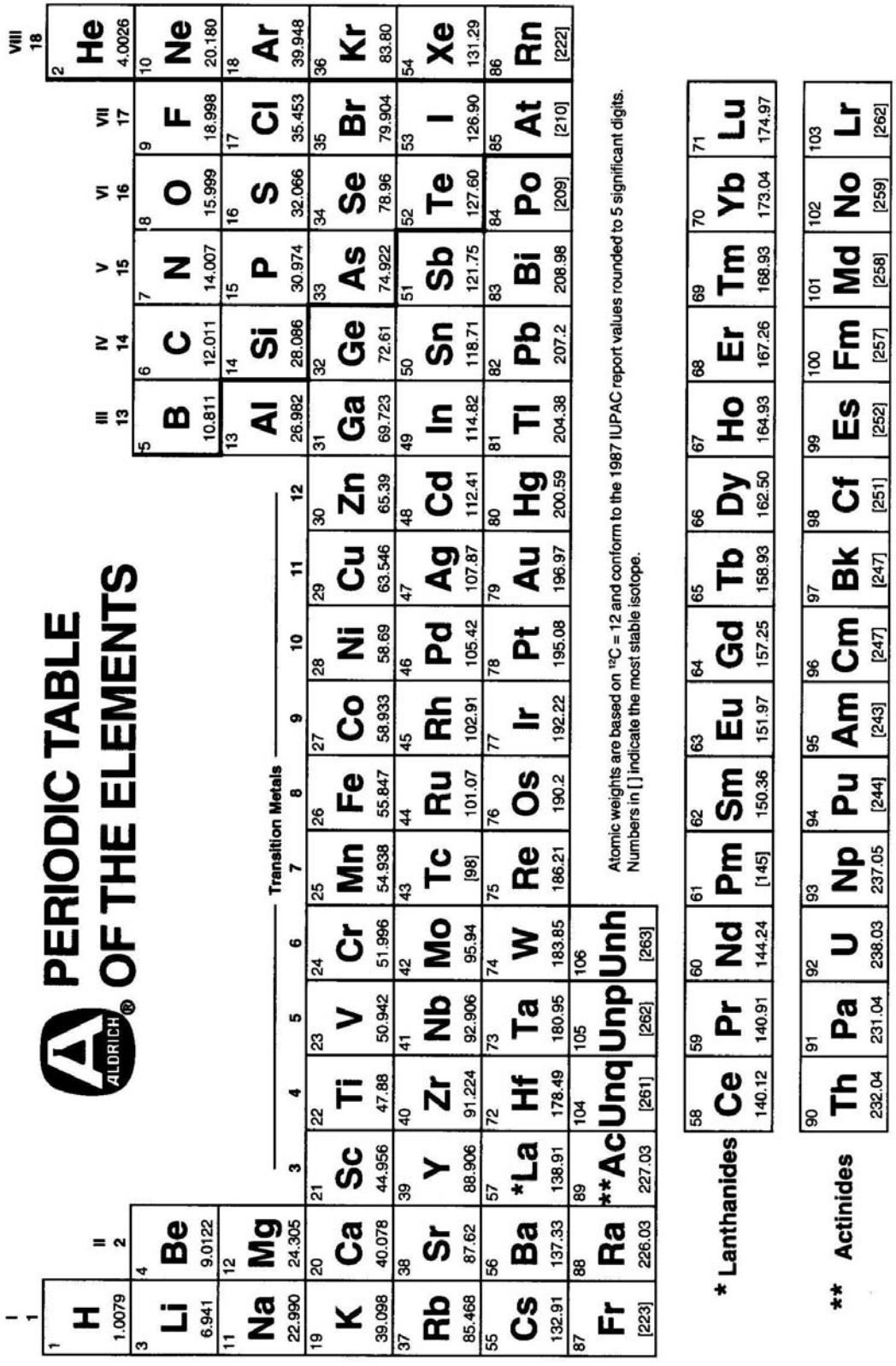

END OF EXAM