Name: ____________________ **Student number:** ______________

Chemistry 1E03 **Final Exam** **Dec. 15, 2017**

McMaster University ***VERSION 1***

Instructors: Drs. R.S. Dumont, P. Kruse & L. Davis

Duration: 150 minutes

This test contains 20 numbered pages printed on both sides. There are **30** multiple-choice questions appearing on pages numbered 3 to 16. Pages 17 and 18 provide extra space for rough work. Page 19 includes some useful data and equations, and there is a periodic table on page 20. You may tear off the last pages 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 1 mark; the total marks available are 30. There is **no** additional penalty for incorrect answers.

BE SURE TO ENTER THE CORRECT VERSION NUMBER OF YOUR TEST (shown near the top of page 1), IN THE SPACE PROVIDED ON THE ANSWER SHEET.

<u>ANSWER ALL QUESTIONS</u> ON THE ANSWER SHEET, IN PENCIL.
Instructions for entering multiple-choice answers are given on page 2.
SELECT <u>ONE AND ONLY ONE</u> 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 MS or MS+ 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.

 Continued on next page

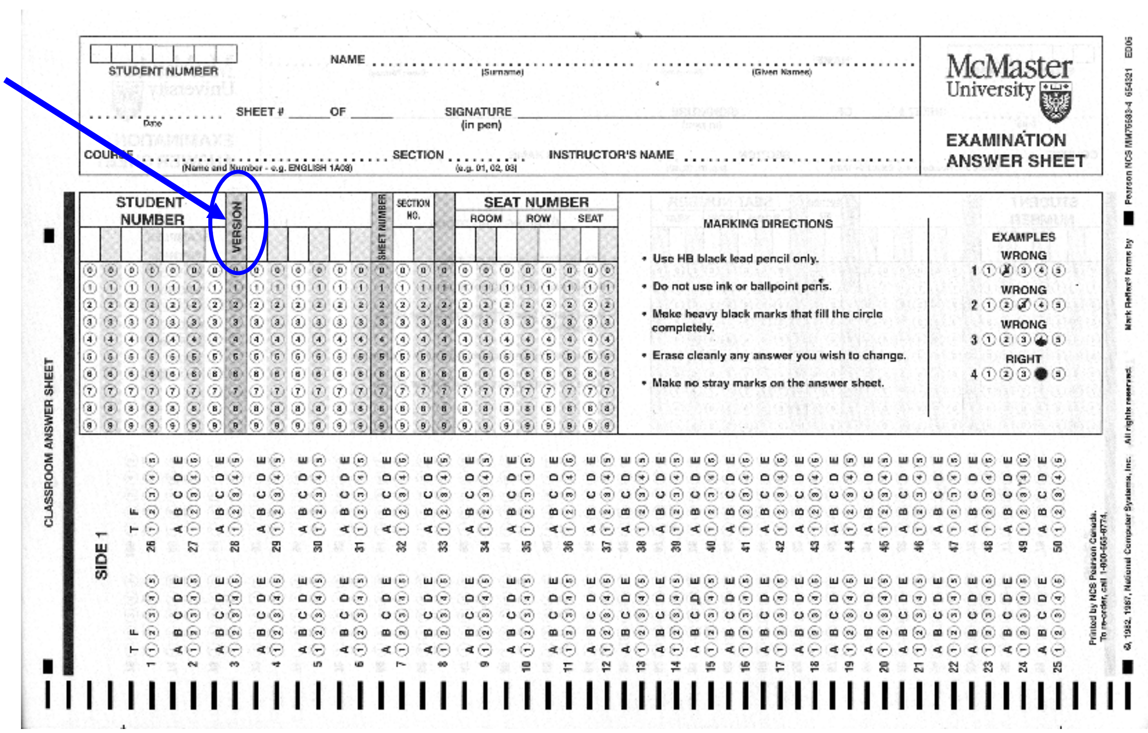
Name: ____________________ Student number: ______________

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 fill in the corresponding bubble numbers underneath.
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. Select the **correct** sequence of the elements Al, K, Mg and N in order of **increasing first ionization energy**:

   .

A) K < Al < N < Mg
B) K < Mg < Al < N
C) Al < Mg < K < N
D) K < Al < Mg < N
E) K < N < Mg < Al

2. Which one of the following atoms will **not** accommodate **more** than 8 electrons in its valence shell?

   .

A) P
B) Kr
C) S
D) C
E) Br

3. A hydrocarbon gas (C_xH_y) sample weighing 0.5653 g is completely burned in excess oxygen. The resulting gas mixture is cooled until all the water produced, in the reaction, is condensed and weighed. The result is 1.016 g of water. What is the molecular formula of the **hydrocarbon gas**?

   .

A) C_2H_4
B) CH
C) C_2H_6
D) C_3H_8
E) CH_4

 Continued on next page

4. Choose the **correct** statements from the following:

   (i) Sodium has a larger first ionization energy than potassium.
   (ii) Sodium has a larger atomic size than chlorine.
   (iii) Calcium has a larger first ionization energy than fluorine.
   (iv) Sulfur has a larger electronegativity than chlorine.
   .

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

5. The *emission* spectrum of atomic hydrogen can be divided into several well-separated series of lines, associated with particular transitions. Calculate the **longest wavelength** observed in series in the near infrared that contains all transitions ending at n = 3.
   .

A) 656 nm
B) 1.10 μm
C) 1.88 μm
D) 91.6 nm
E) 1.29 μm

6. Which **one** of the following will oxidize $H_2S(g)$?
   .

A) $NH_3(g)$
B) $H_2(g)$
C) Na(s)
D) $KMnO_4(aq)$
E) HBr(aq)

 Continued on next page

7. When the following pairs of solutions are mixed, which produces the **strongest electrolyte** solution?

.

A) $Pb(CH_3COO)_2$ (1 M) + NaI (2 M)
B) $Ba(OH)_2$ (3 M) + $CuSO_4$ (3 M)
C) HCl (2 M) + NaOH (2 M)
D) KCl (1 M) + NH_4NO_3 (1 M)
E) $Zn(NO_3)_2$ (1 M) + $AgClO_4$ (2 M)

8. Which of the following statements are **FALSE** according to the shapes predicted by the VSEPR model? (Central atoms are underlined.)

(i) $\underline{As}H_3$ 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 permanent dipole moment than $\underline{Xe}F_2$.
(v) $\underline{N}O_2$ is a non-polar molecule.

.

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

 Continued on next page

9. Which one of the following molecules has **no** net dipole moment? (Central atoms are underlined)

.

A) $\underline{C}H_3Cl$
B) $H_2\underline{Se}$
C) $\underline{N}O_2$
D) $\underline{N}F_3$
E) $\underline{S}O_3$

10. The standard enthalpy of formation of HCN(g) = +135.1 kJ/mol. The entropy change for the formation reaction of HCN(g) is +34.5 J/mol K. Calculate $\Delta G_f°$ for HCN(g) at 298 K (in kJ mol^{-1}).

.

A) +3.97
B) –10.1
C) +125
D) +145
E) +10.4

 Continued on next page

11. Calculate the standard **entropy change of the universe, $\Delta S_{univ}°$** (in J K^{-1}), at 25.00°C for

$$CO(g) + 2\ H_2(g) \rightarrow CH_3OH(l)$$

from the following data:

	$\Delta H_f°$ (kJ mol^{-1})	$S°$ (J K^{-1} mol^{-1})
$H_2(g)$	0	130.684
$CO(g)$	–110.525	197.674
$CH_3OH(l)$	–238.66	126.8

.

A) +97.52
B) +429.8
C) +1171
D) –332.2
E) +838.8

12. Identify the processes accompanied by a large **increase of entropy** of the system:

(i) $Ni(s) + 4\ CO(g) \rightarrow Ni(CO)_4(g)$
(ii) $Cu(s) + 1/2\ O_2(g) \rightarrow CuO(s)$
(iii) $H_2O(g) \rightarrow H_2O(l)$
(iv) $CO(g) + H_2O(g) \rightarrow CO_2(g) + H_2(g)$

.

A) none
B) i
C) ii, iii
D) all
E) iv

 Continued on next page

13. 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) $2\ C(graphite) + O_2(g) \rightarrow 2\ CO(g)$ $\Delta H° = -221$ kJ mol^{-1} $\Delta G° = -274$ kJ mol^{-1}

Which of the following statements is(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 K^{-1} mol^{-1} at 298.15 K.
(iv) Reaction 2 is spontaneous at 500°C when the partial pressures of both O_2 and CO are 1 bar.

.

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

14. Calculate the standard Gibbs free energy change, **$\Delta G°$ (in kJ mol^{-1})**, at a temperature of **425 K**, for the reaction,

$$2\ CO(g) + O_2(g) \rightarrow 2\ CO_2(g)$$

at 298 K:

	CO(g)	O_2(g)	CO_2(g)
$\Delta H_f°$ (kJ mol^{-1})	–110.5	0	–393.5
$S°$ (J mol^{-1}K^{-1})	198	205	214
$\Delta G_f°$ (kJ mol^{-1})	–137	0	–394

.

A) –514
B) –566.0
C) +72.9
D) –639.5
E) –492

 Continued on next page

15. The enthalpy of vaporization of HCN is 20.0 kJ/mol. 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 (in °C).

.

A) +225
B) +4.4
C) –48
D) +0.22
E) +89

16. Calculate the lattice energy (in kJ mol^{-1}) of silver chloride from the following data:

ΔH^o_f[Ag(g)] = 284 kJ mol^{-1}
ΔH^o_f[AgCl(s)] = –127 kJ mol^{-1}
ΔH^o_f[Cl(g)] = 122 kJ mol^{-1}
First ionization energy Ag(g) = 731 kJ mol^{-1}
Electron affinity Cl(g) = –349 kJ mol^{-1}

.

A) –915
B) –1037
C) –661
D) –1359
E) –1613

 Continued on next page

17. Put the following salts in the expected order of increasing magnitude of lattice energy: LiF, KBr, BaS, CaO, $MgCl_2$

.

A) LiF < $MgCl_2$ < CaO < KBr < BaS
B) $MgCl_2$ < BaS < CaO < LiF < KBr
C) KBr < LiF < $MgCl_2$ < BaS < CaO
D) KBr < LiF < CaO < BaS < $MgCl_2$
E) LiF < KBr < $MgCl_2$ < CaO < BaS

18. Pure $N_2O_4(g)$ is introduced in an evacuated container. It dissociates according to the following equilibrium:

$$N_2O_4(g) \rightleftharpoons 2\ NO_2(g)$$

When equilibrium is established at 25 °C, the partial pressure of N_2O_4 is 59.0% of its initial value, and the total equilibrium pressure is 14.1 bar. What is Kp for this equilibrium?

.

A) 6.85×10^{-6}
B) 88.3
C) 0.0197
D) 11.4
E) 0.211

 Continued on next page

19. The equilibrium constant K_p for

$$C(s) + CO_2(g) \rightleftharpoons 2\ CO(g)$$

is 1.52 at 700 °C. If the partial pressure of CO in an equilibrium mixture at 700 °C is 1.30 bar, what is the **partial pressure** of CO_2 (in bar)?

.

A) 1.11
B) 0.900
C) 1.17
D) 1.30
E) 0.860

20. The interior and exterior of a nerve cell behave as a concentration cell. If the concentration of K^+ outside the cell is 0.030 M and the concentration inside is 0.30 M, what is the potential difference across the cell membrane, in volts V? Assume normal body temperature, T = 37 °C and a one-electron transfer (n = 1).

.

A) 0.062
B) 5900
C) 86
D) 0.026
E) 0.19

 Continued on next page

21. Zinc is used in the cathodic protection of iron because:
.
A) Zn is a better reducing agent than Fe.
B) The reaction $Fe + Zn^{2+} \rightarrow Fe^{2+} + Zn$ is spontaneous under standard conditions.
C) Zn is difficult to oxidize.
D) Zn will serve as the cathode in the electrochemical process.
E) Fe is a better reducing agent that Zn.

22. 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) 1.5×10^{13}
B) 2.3×10^{26}
C) 7.6
D) 9.8×10^{-16}
E) 30

 Continued on next page

23. Cerium is a lanthanide and like most rare earth metals its most common oxidation state is +3. In an electrochemical cell, the +4 oxidation state oxidizes chloride ions to chlorine gas. The cell potential is measured to be 0.30 V at 25°C. If the concentrations of all aqueous ions are 2.0 M, what is the **pressure of chlorine gas (in bar)**? Hint: write an expression for the reaction quotient. The standard reduction potentials are given below, with the *unbalanced* half-reactions:

$$Cl_2(g) \rightarrow Cl^-(aq) \qquad E_{red}° = +1.36\ V$$
$$Ce^{4+}(aq) \rightarrow Ce^{3+}(aq) \qquad E_{red}° = +1.61\ V$$

A) 0.08
B) 3.9
C) 0.6
D) 0.02
E) 0.005

24. What potential, in volts, is developed by the following electrochemical cell at 298.15 K?

$$Al(s)\ |\ Al^{3+}(0.00100\ M)\ ||\ Ag^{+}(0.00100\ M)\ |\ Ag(s)$$

Data:

$$Al^{3+}(aq) + 3\ e^- \rightarrow Al(s) \qquad E_{red}° = -1.66\ V$$
$$Ag^{+}(aq) + e^- \rightarrow Ag(s) \qquad E_{red}° = 0.80\ V$$

A) 2.52
B) 2.58
C) 2.40
D) 2.46
E) 2.34

 Continued on next page

25. Select the **TRUE** statement(s) for the following electrochemical cell:

$$Cu(s)|Cu(NO_3)_2(aq)||AgNO_3(aq)\ |Ag(s)$$

(i) The silver I concentration, $[Ag^+]$, increases at the Ag electrode.
(ii) Electrons flow in the external circuit from the Cu to the Ag electrode.
(iii) Ag metal is a stronger reducing agent than Cu metal.
(iv) The Cu electrode is the cathode of the electrochemical cell.
(v) Nitrate anions migrate to the Cu electrode through the salt bridge.

.

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

26. **Order** the following acids to increasing strength - i.e., according to decreasing pK_a: H_2SO_4, H_2SO_3, H_2SeO_3 and $HClO_4$

.

A) $H_2SO_3 < H_2SeO_3 < H_2SO_4 < HClO_4$
B) $H_2SeO_3 < H_2SO_3 < H_2SO_4 < HClO_4$
C) $H_2SO_4 < H_2SO_3 < H_2SeO_3 < HClO_4$
D) $H_2SeO_3 < H_2SO_3 < HClO_4 < H_2SO_4$
E) $H_2SO_3 < H_2SO_4 < H_2SeO_3 < H_2SO_4$

 Continued on next page

27. Heroin, a derivative of morphine, is a powerful analgesic and a powerful narcotic agent. It is also a base with $K_b = 9.5 \times 10^{-7}$. What is the **pH** of a 1.00×10^{-3} M solution of heroin?
.
A) 9.48
B) 10.72
C) 9.77
D) 6.35
E) 8.11

28. What is the **pOH** of a 100 mL solution of 0.00059 M HNO_3?
.
A) 4.23
B) 9.77
C) 10.77
D) 7.00
E) 3.23

 Continued on next page

29. Your stomach (volume = 2.5 L) has a pH of 1.00 because of the presence of HCl. **How many grams of $Mg(OH)_2$** (58.3 g/mol) do you need to add to completely neutralize the acid in your stomach?

.

A) 21 g
B) 2.9 g
C) 15 g
D) 5.8 g
E) 7.3 g

30. What is the **pH** of a monochloroacetic acid ($CH_2ClCOOH$) solution that is **5.00% dissociated**? $K_a = 1.35 \times 10^{-3}$.

.

A) 4.17
B) 2.87
C) 0.29
D) 1.59
E) 1.30

 Continued on next page

Extra Space:

 Continued on next page

Extra Space:

 Continued on next page

Name: ____________________ Student number: ______________

Some general data are provided on this page.

A Periodic Table with atomic weights is provided on the next page.

R = 8.3145 J K^{-1} mol^{-1} = 0.083145 L bar K^{-1} mol^{-1}

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

$c = 2.9979 \times 10^{8}$ m s^{-1}

$h = 6.6256 \times 10^{-34}$ J s

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

density(H_2O, l) = 1.00 g mL^{-1}

Specific heat capacity of H_2O(l) = 4.18 J g^{-1} K^{-1}

F = 96485 C mol^{-1}

1 bar = 100.0 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

1 Hz = 1 cycle/s

De Broglie wavelength:

$\lambda = h / mv = h / p$

Hydrogen atom energy levels:

$E_n = -2.180 \times 10^{-18}$ J / n^2

Nernst Equation (the last two equations are for T = 298.15 K):

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{nF}\ln Q \quad = E^{\circ}_{\text{cell}} - \frac{0.0257\text{ V}}{n}\ln Q \quad = E^{\circ}_{\text{cell}} - \frac{0.0592\text{ V}}{n}\log_{10} Q$$

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

Gibbs free energy of reaction: $\Delta G = \Delta G^{\circ} + RT\ln Q$

The roots of quadratic equation, $a x^2 + b x + c = 0$, are given by $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

Solubility Guidelines for Common Ionic Solids

1. Alkali metal and ammonium salts are *soluble*.
2. Nitrate, chlorate, perchlorate, hydrogen carbonate and ethanoate salts are *soluble*.
3. Sulfate salts are *soluble*, *except* for the calcium, strontium, barium and lead salts which are *insoluble*.
4. Chloride, bromide and iodide salts are *soluble*, *except* for the silver, lead and mercury I salts which are *insoluble*.
5. Silver, lead and mercury I salts are *insoluble*, unless deemed soluble by rule 2 or 3.
6. Sulfide salts are *insoluble*, *except* for the alkali metal, ammonium, and alkaline earth salts which are *soluble*.
7. Oxide and hydroxide salts are *insoluble*, *except* for the alkali metal, ammonium, calcium, strontium and barium salts which are soluble.
8. Carbonate and phosphate salts are *insoluble*, *except* for the alkali metal and ammonium salts

 Continued on next page

Name: ____________________ Student number: ______________

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											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]

 END OF EXAM