Name: ____________________ **Student number:** ______________

Chemistry 1E03 *Final Exam* *Dec. 14, 2016*

McMaster University ***VERSION 1***

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

Duration: 150 minutes

This test contains 21 numbered pages printed on both sides. There are **30** multiple-choice questions appearing on pages numbered 3 to 17. Pages 18 and 19 provide extra space for rough work. Page 20 includes some useful data and equations, and there is a periodic table on page 21. 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 25. 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. Keep your eyes on your own paper – looking around the room may be interpreted as an attempt to copy answers.

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.

continued on next page

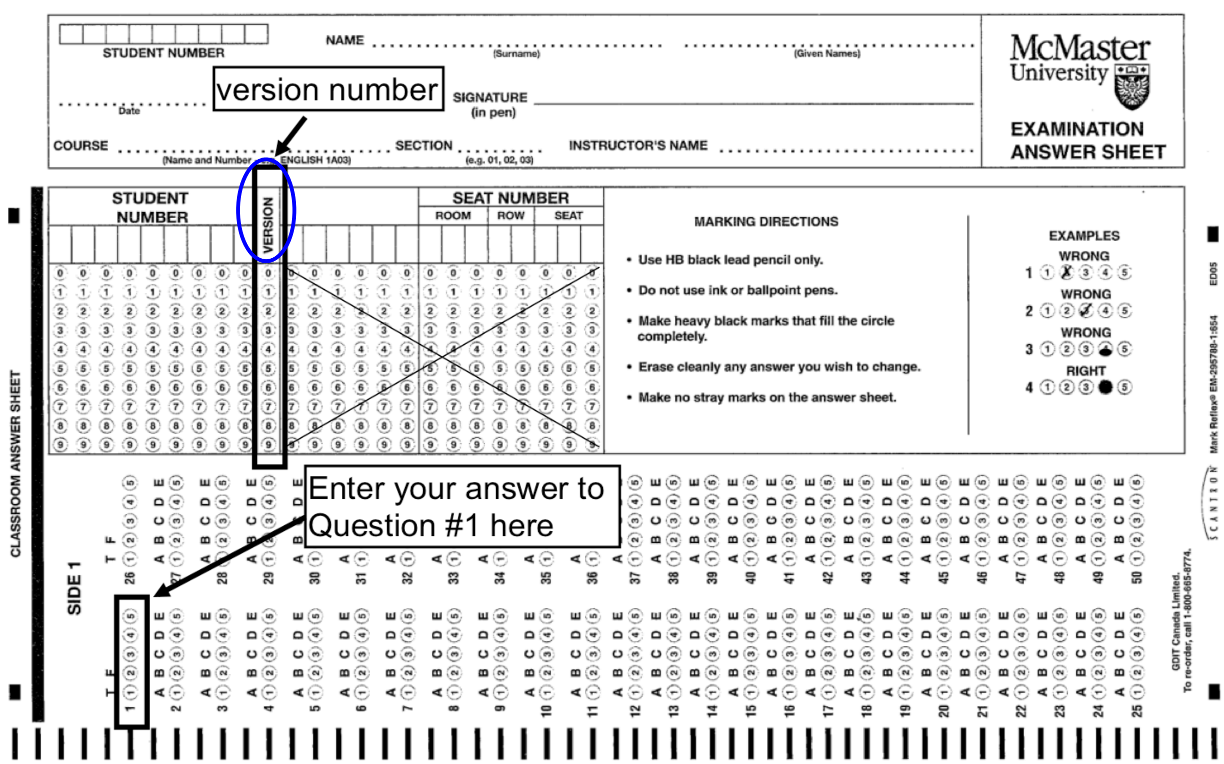
Name: ____________________ Student number: ______________

OMR EXAMINATION – STUDENT INSTRUCTIONS

NOTE: IT IS YOUR RESPONSIBILITY TO ENSURE THAT THE ANSWER SHEET IS PROPERLY COMPLETED: YOUR EXAMINATION 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".



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1. Using an ice calorimeter, it was determined that the reaction of 0.14 g of zinc with excess HCl (aq), caused 0.68 g of ice to melt. What was the **enthalpy of reaction** per mole of Zn (kJ mol^{-1})?
[ΔH_{fus}(ice) = 333 J g^{-1} K^{-1}]

A) −110
B) −95
C) 240
D) −180
E) 590

2. One of the radioisotopes for medical purposes produced at the McMaster Nuclear reactor is ^{125}I. How many **more neutrons** than electrons does an **iodide ion** of this isotope have?

A) 18
B) 19
C) 20
D) 21
E) 22

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3. Calculate the **wavelength of light**, in nanometers, emitted when an electron in a hydrogen atom makes a transition from the $n = 4$ to the $n = 2$ state.

A) 486
B) 4860
C) 810
D) 292
E) 2920

4. Which is the correct ordering according to **increasing electronegativity**, for the atoms Mg, Ca, O, P?

A) P < O < Mg < Ca
B) Ca < Mg < P < O
C) Mg < Ca < P < O
D) O< Mg < Ca < P
E) Ca < P < O < Mg

continued on next page

5. Identify the correct order of **decreasing molecular dipole moment** among the following:

A) $H_2O > H_2S > CO_2$
B) $CO_2 > H_2O > H_2S$
C) $H_2O > CO_2 > H_2S$
D) $H_2S > H_2O > CO_2$
E) $H_2S > CO_2 > H_2O$

6. For the species NO^+, NO_2^-, NO_3^-, what is the correct order of **decreasing average N-O bond length**?

A) $NO^+ > NO_2^- > NO_3^-$
B) $NO_3^- > NO_2^- > NO^+$
C) $NO_2^- > NO^+ > NO_3^-$
D) $NO_3^- > NO^+ > NO_2^-$
E) $NO_2^- > NO_3^- > NO^+$

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7. Which of the following pairs of substances would ***not*** **produce a gas** when reacting together?

A) $Ca(s)$ and $H_2O(l)$
B) $Cu(s)$ and $HNO_3(aq)$
C) $CuO(s)$ and $H_2SO_4(aq)$
D) $Zn(s)$ an $HCl(aq)$
E) $S(s)$ and $O_2(g)$

8. Complete and balance the following redox equation in acidic aqueous solution, using smallest integer coefficients for all species. Then select the correct **coefficient** for the **hydrogen sulfite ion (HSO_3^-)**.

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

A) 1
B) 2
C) 3
D) 5
E) 10

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9. Select the one **false statement** for the following equilibrium in a constant volume container:

$$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g) \qquad \Delta H = -198.2\ \text{kJ mol}^{-1}$$

The partial pressure of $SO_2(g)$ will **increase** if:

A) The temperature is increased.
B) An inert gas is added to increase the total pressure.
C) The volume is increased.
D) O_2 is removed.
E) SO_3 is added.

10. NOBr(g) is introduced in an evacuated container. It dissociates according to the following equilibrium:

$$2\ NOBr(g) \rightleftharpoons 2\ NO(g) + Br_2(g)$$

When equilibrium is established at 25°C, 66% of the initial NOBr remains and the total equilibrium pressure is 0.25 bar. What is K_p for this equilibrium?

A) 0.088
B) 0.21
C) 9.8×10^{-6}
D) 0.019
E) 0.0096

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11. A chemical reaction with an enthalpy change $\Delta H^\circ = -400$ kJ is carried out in a calorimeter containing 1500 mL of pure water initially at 25.0 °C. What is the **final temperature** (in °C) of the water? (Assume the calorimeter heat capacity is just the heat capacity of the water.)

A) 63.7
B) 88.8
C) −38.7
D) 336.7
E) −67.5

12. What is the **enthalpy change** (in kJ mol^{-1}) of the following gas-phase reaction? (Hint: write the Lewis structures of the molecules before using the given bond enthalpies.)

$2\ C_2H_4(g) \rightarrow C_2H_2(g) + C_2H_6(g)$

Bond enthalpies (kJ mol^{-1}): C-C 348; C=C 619; C≡C 812; C-H 413

A) +78
B) −78
C) −39
D) −156
E) +39

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13. Which ONE of the following ionic solids has the **smallest absolute value (magnitude) of lattice energy**?

A) LiCl
B) NaCl
C) $CaBr_2$
D) LiBr
E) KBr

14. Identify the **false statement**(s) from among the following:

(i) For any chemical reaction, $\Delta G^\circ < \Delta H^\circ$.
(ii) The endothermic dissolution of NH_4NO_3(s) in water is driven by entropy.
(iii) For a chemical reaction at equilibrium at temperature T, $\Delta H = T\Delta S$.

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

continued on next page

15. Consider the reaction Si(s) + 2 H_2(g) $\rightarrow$ SiH_4(g). Use the data below to identify the **true statement**(s).
   (i) $\Delta S° > 0$ for the forward reaction.
   (ii) The reverse reaction is spontaneous at all temperatures.
   (iii) If $P(H_2)$ = 100 bar at equilibrium at 25 °C, then $P(SiH_4) = 1.10 \times 10^{-6}$ bar.

Data:
$\Delta H_f°[SiH_4(g)] = 34.3$ kJ mol^{-1} $S°[SiH_4(g)] = 204.62$ J K^{-1} mol^{-1}
$S°[H_2(g)] = 130.68$ J K^{-1} mol^{-1} $S°[Si(s)] = 18.83$ J K^{-1} mol^{-1}

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

16. In which of the following processes does the **entropy of the system decrease**?

   (i) CH_4 (g) + 2 O_2 (g) $\rightarrow$ CO_2 (g) + 2 H_2O (l)
   (ii) B_2O_3 (s) + 3H_2O (g) $\rightarrow$ B_2H_6 (g) + 3 O_2 (g)
   (iii) H_2 (g) + ½ O_2 (g) $\rightarrow$ H_2O (l)
   (iv) I_2 (s) $\rightarrow$ I_2 (g)
   (v) Li^+ (aq) + F^- (aq) $\rightarrow$ LiF(s)

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

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17. Put these substances - $N_2(g)$, $CH_3OH(l)$, $SO_3(g)$, $Pb(s)$ and $PbO_2(s)$ - in order of **increasing molar entropy**, *S*º:

A) $N_2(g) < CH_3OH(l) < SO_3(g) < Pb(s) < PbO_2(s)$
B) $N_2(g) < SO_3(g) < CH_3OH(l) < Pb(s) < PbO_2(s)$
C) $Pb(s) < PbO_2(s) < CH_3OH(l) < N_2(g) < SO_3(g)$
D) $PbO_2(s) < Pb(s) < CH_3OH(l) < SO_3(g) < N_2(g)$
E) $PbO_2(s) < Pb(s) < SO_3(g) < CH_3OH(l) < N_2(g)$

18. Calculate $\Delta G°$ in kJ mol^{-1}, for the following reaction: $3\ NO_2(g) + H_2O(l) \rightarrow 2\ HNO_3(l) + NO(g)$.

Data:
$\Delta G_f°[H_2O(1)] = -237.2$ kJ mol^{-1} $\Delta G_f°[HNO_3(l)] = -79.9$ kJ mol^{-1}
$\Delta G_f°[NO(g)] = +86.7$ kJ mol^{-1} $\Delta G_f°[NO_2(g)] = +51.8$ kJ mol^{-1}

A) 8.7
B) 19.2
C) –41.4
D) –19.2
E) –15.5

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19. Determine the **boiling point** of CS_2, in °C, from the following data:

	$\Delta H_f°$ (kJ mol^{-1})	$S°$ (J mol^{-1} K^{-1})
$CS_2(g)$	115.3	237.8
$CS_2(l)$	87.9	151.0

A) 316
B) 811
C) 87
D) −11
E) 43

20. The standard elemental form of mercury at 300 K is Hg(l). The standard enthalpy of formation for Hg(g) is $\Delta H_f°$[Hg(g)] = +60.78 kJ mol^{-1}. The standard entropy of vaporization of mercury is $\Delta S_{vap}°$[Hg] = +97.3 J mol^{-1}K^{-1}. Calculate the **pressure** (in bar) of mercury **vapor** (gas) in equilibrium with the mercury liquid, at **300 K**. Hint: write an expression for the equilibrium constant.

A) 2.11×10^{-6}
B) 0.904
C) 3.16×10^{-6}
D) 11.3
E) 2.97×10^{-6}

continued on next page

21. Determine the **pH** of a 1.0×10^{-7} solution of HCl. (Hint: consider the autoionization of water.)

A) 7.00
B) 6.79
C) 6.00
D) 7.34
E) 6.21

22. Which of the following salts generates the **lowest pH (most acidic)** when dissolved in water, as a 1 mol L^{-1} solution?

A) Nicotine-HCl ($pK_b = 6.00$)
B) $NaNO_2$ ($pK_b = 10.70$)
C) $NaOOCH_3$ ($pK_b = 9.24$)
D) NaF ($pK_b = 10.80$)
E) Cocaine-HCl ($pK_b = 5.40$)

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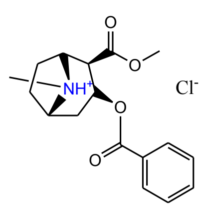
23. Calculate the **pH** of a **0.050 M Na_2CO_3** aqueous solution.
(for **H_2CO_3**, pK_{a1} = 6.4 and pK_{a2} = 10.3):

A) 10.2
B) 8.7
C) 7.4
D) 3.7
E) 11.5

24. Determine the **pH of 0.100 M cocaine as its HCl salt (BH^+)** when dissolved in aqueous solution if the free base (*B*) has a pK_b = 5.40:



A) 4.80
B) 10.54
C) 5.74
D) 8.65
E) 6.32

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25. Given the following data, identify the **best reducing agent**.

half-reaction	E^o_{red}
$2\ Hg^{2+}(aq) + 2\ e^- \rightarrow Hg_2^{2+}(s)$	+0.92 V
$N_2(g) + 5\ H^+(aq) + 4\ e^- \rightarrow N_2H_5^+(aq)$	+0.23 V
$Sn^{4+}(aq) + 2\ e^- \rightarrow Sn^{2+}(aq)$	+0.13 V

A) $Sn^{2+}(aq)$
B) $Hg^{2+}(aq)$
C) $N_2(g)$
D) $Hg_2^{2+}(aq)$
E) $N_2H_5^+(aq)$

26. **Zinc** is used in the **cathodic protection of iron** because:

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

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27. What is the value of the **equilibrium constant, *K*, at 25.00°C** for the electrochemical cell described by reaction $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$ for which $\boldsymbol{E^{o}_{cell} = +0.78\ V}$?

A) 30
B) 1.52×10^{13}
C) 2.35×10^{26}
D) 7.61
E) 9.88×10^{-16}

28. Determine the **voltage (V) of a galvanic cell** described below, where the concentration of KCl is 0.100 M, E^{o}_{red} for Cd = –0.380 V, and AgCl = +0.222V:

$$Cd\ (s)\ |\ Cd^{2+}\ (aq)\ ||\ Cl^{-}\ (aq)\ |\ AgCl\ (s)\ |\ Ag\ (s)$$

A) +0.413
B) +0.543
C) +0.238
D) +0.374
E) +0.192

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29. The standard reduction potential of the zinc cation is –0.76 V. The measured cell potential for the cell Zn(s) | Zn^{2+} (aq) || H^{+} (aq, 1 M) | H_2 (g, 1 bar) | Pt(s) is +0.81V at 25ºC. What is the **concentration of Zn^{2+}(aq) (in mol L^{-1}) ?**

A) 0.020
B) 0.20
C) 2.0
D) 3.8
E) 0.012

30. Given the following data:

$Cu^{2+}(aq) + 2\,e^{-} \rightarrow Cu(s)$	E^{o}_{red} = +0.34 V
$NO_3^{-}(aq) + 4\,H^{+}(aq) + 3\,e^{-} \rightarrow NO(g) + 2\,H_2O(l)$	E^{o}_{red} = +0.96 V

Which of the **following statements is/are true** for the balanced spontaneous redox reaction involving the above species? The balanced equation must contain the smallest integers as stoichiometric coefficients.

(i) E_{cell} = +1.30 V
(ii) Δ G = –359 kJ
(iii) Δ S > 0

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

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Extra space for rough work:

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Extra space for rough work:

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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} =0.08314 L bar K^{-1} mol^{-1}

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

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

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

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

1 bar = 100.0 kPa

0°C = 273.15 K

1 J = 1 kg m^2 s^{-2} = 0.01 L bar = 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:
$E_n = -R_H / n^2 = -2.178 \times 10^{-18}$ J / n^2

Density of water:
1.00 g mL^{-1}

Specific heat capacity of water:
4.18 J K^{-1} g^{-1}

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

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{zF}\ln Q \quad = E^{\circ}_{\text{cell}} - \frac{0.025691\text{V}}{z}\ln Q \quad = E^{\circ}_{\text{cell}} - \frac{0.059156\text{ V}}{z}\log_{10} Q$$

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

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

Faraday constant: F = 96485 C mol^{-1}

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

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 are *insoluble*, *except* for the alkali metal and ammonium salts.

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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											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 TEST