Name: ____________________ **Student number:** ______________

Chemistry 1E03 **Final Exam** **Dec. 19, 2015**

McMaster University ***VERSION 1***

Instructors: Drs. R.S. Dumont, 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 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 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.

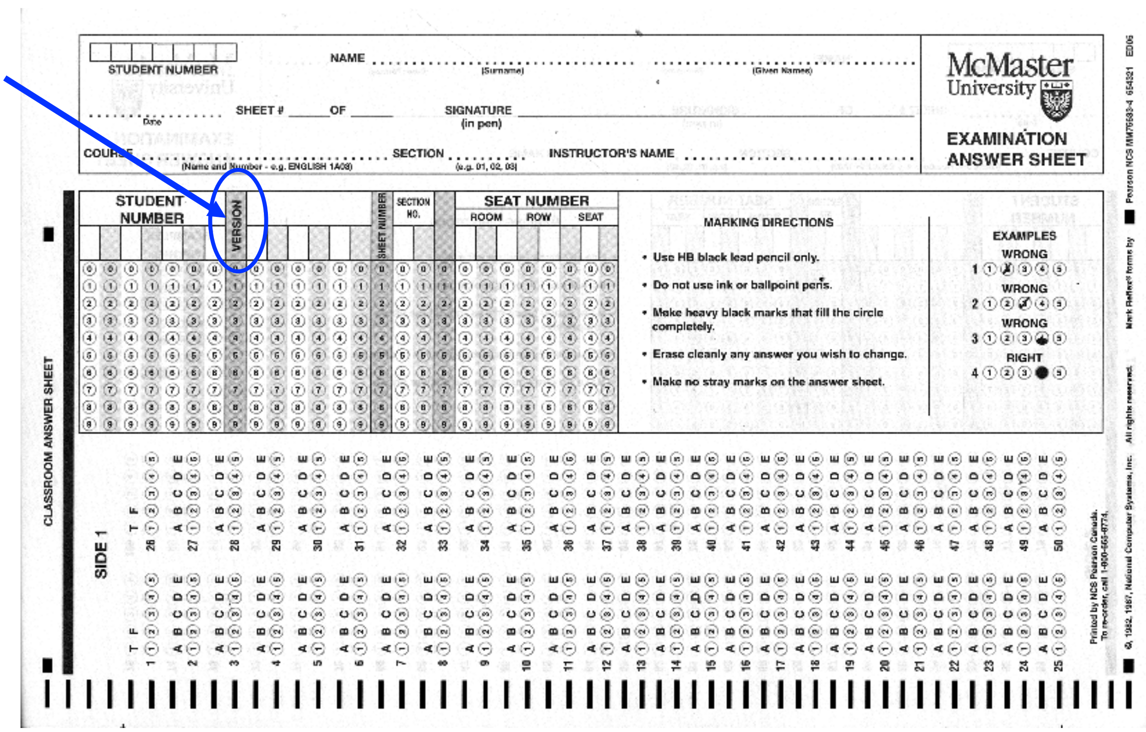
Name: ____________________ Student number: ______________

OMR EXAMINATION – STUDENT INSTRUCTIONS

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



 Continued on next page

1. Identify the **nonpolar** molecule from among the following.

A) BiF_3
B) CH_2Cl_2
C) XeF_4
D) OF_2
E) SF_4

2. Which of the following species has the **most negative average formal charge** on the oxygen atoms?

A) ClO_3^{-}
B) SeO_3^{2-}
C) PO_4^{3-}
D) SeO_4^{2-}
E) SO_2

 Continued on next page

3. What is the **maximum wavelength (in nm)** of **visible** light observed in an emission spectrum from excited hydrogen atoms?

A) 364
B) 121
C) 91
D) 1210
E) 657

4. Which of the following statements about periodic trends are **TRUE**?
(i) The correct sequence for decreasing ionic radius for Rb^+, Sr^{2+} and Br^- is
$Br^- > Rb^+ > Sr^{2+}$.
(ii) The ground-state electron configuration of Si has no unpaired electrons.
(iii) The oxide of calcium is a basic oxide.
(iv) Rb loses electrons more easily than Na.
(v) The electronegativity of chlorine is smaller than that of phosphorus.

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

 Continued on next page

5. How many **charge-minimized** resonance structures are required to portray the bonding in ClO_4^-?

A) 5
B) 3
C) 4
D) 1
E) 2

6. Balance the following (unbalanced) redox reaction in **basic** solution. Use lowest possible whole number coefficients.

$$Cr(OH)_3(s) + ClO^-(aq) + OH^-(aq) \rightarrow CrO_4^{2-}(aq) + Cl^-(aq) + H_2O(l)$$

When this has been done correctly, the stoichiometric **coefficients** for the **reactant** species, in order from **left** to **right** are as follows:

A) 2, 3, 4
B) 2, 4, 6
C) 2, 3, 7
D) 1, 2, 3
E) 3, 4, 5

 Continued on next page

7. Identify the following **types of reactions**:
   (i) $Ba(s) + 2HNO_3(aq) \rightarrow Ba(NO_3)_2(aq) + H_2(g)$
   (ii) $Ba(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow BaCO_3(s) + 2NaNO_3(aq)$
   (iii) $BaCO_3(s) + 2HCl(aq) \rightarrow BaCl_2(aq) + H_2CO_3(aq)$

A) (i) acid-base (ii) redox (iii) precipitation
B) (i) acid-base (ii) precipitation (iii) redox
C) (i) redox (ii) acid-base (iii) precipitation
D) (i) precipitation (ii) acid-base (iii) redox
E) (i) redox (ii) precipitation (iii) acid-base

8. Which of the following statements are **true**?
   (i) Alkali metal hydrides produce $H_2(g)$ upon reacting with water.
   (ii) The formation of $H_2(g)$ is a characteristic of the reaction of *all* metals with dilute acids.
   (iii) Redox reactions between alkali metals (Group 1A) and halogens (Group 7A) produce ionic compounds which are soluble in water.
   (iv) HF(aq) is a stronger acid than HCl(aq).
   (v) Mixing saturated aqueous solutions of phosphoric acid and calcium hydroxide produces no visible reaction.

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

9. A vessel is filled with N_2O_4(g) to an initial pressure of 5.00 bar. Some of the N_2O_4(g) decomposes. The equilibrium constant for the decomposition of N_2O_4(g) into NO_2(g) is 0.133 at the temperature of the vessel. What is the **equilibrium partial pressure** (in bar) of $\mathbf{NO_2}$(g)?

A) 0.277
B) 1.07
C) 0.492
D) 3.41
E) 0.783

10. Given the following reduction potentials, what **cell potential** can be generated from the associated half cells with the least deviation from standard conditions?

$Au^{3+} + 3e^- \longrightarrow Au\ (s)$ $\quad E°_{red} = 1.52$ V
$Ag^{+} + e^- \longrightarrow Ag\ (s)$ $\quad E°_{red} = 0.800$ V

A) 0.40 V
B) 2.30 V
C) 2.00 V
D) 0.95 V
E) 0.65 V

 Continued on next page

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

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

12. Why is **zinc** is used in the **cathodic protection** of iron?
.
A) Zn is a better reducing agent than Fe.
B) The reaction $Fe + Zn^{2+} \longrightarrow Fe^{2+} + Zn$ is spontaneous under standard conditions.
C) Fe is a better reducing agent that Zn.
D) Zn is difficult to oxidize.
E) Zn will serve as the cathode in the electrochemical corrosion process.

13. The cell potential (in V) of the following *non-standard* electrochemical cell is +0.559 V.
$Cu(s)$ | $Cu^{2+}(aq, 1.00\ M)$ || $PbI_2(aq, sat.)$ | $I_2(s)$ | $Pt(s)$
The lead iodide solution at the cathode is saturated - it is in contact with $PbI_2(s)$ (not shown above because it does not participate in the cell reaction).

What is the **solubility** of PbI_2 in mol L^{-1}?

Standard reduction potentials:

$I_2(s) + 2\ e^- \rightarrow 2\ I^-(aq)$ $\qquad E^o_{red} = 0.535$ V

$Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s)$ $\qquad E^o_{red} = 0.153$ V

.

A) 3.81×10^{-4}
B) 2.17×10^{-3}
C) 1.57×10^{-3}
D) 7.27×10^{-3}
E) 5.10×10^{-4}

14. The equilibrium constant for the reaction $N_2(g) + O_2(g) \longrightarrow 2\ NO(g)$ is $1.7x10^{-1}$ at an elevated temperature. A reaction vessel at this temperature contains $N_2(g)$ at a partial pressure of 0.25 bar, $O_2(g)$ at a partial pressure of 0.25 bar, and NO(g) at a partial pressure of $4.2x10^{-1}$ bar. Select the one **true** statement for this system.

A) The reaction quotient *Q* is larger than *K*, and the reaction proceeds to the products side.
B) The reaction quotient *Q* is smaller than *K*, and the reaction proceeds to the reactants side.
C) The reaction quotient *Q* is smaller than *K*, and the reaction proceeds to the products side.
D) The reaction quotient *Q* is larger than *K*, and the reaction proceeds to the reactants side.
E) The system is at equilibrium.

 Continued on next page

15. 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) 0.198
B) 0.0267
C) 5930
D) 0.0615
E) 86.2

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

	E^{o}_{red}
$2\ Hg^{2+}(aq) + 2\ e^{-} \longrightarrow Hg_2^{2+}(s)$	+0.92 V
$N_2(g) + 5\ H^{+}(aq) + 4\ e^{-} \longrightarrow N_2H_5^{+}(aq)$	–0.23 V
$Sn^{4+}(aq) + 2\ e^{-} \longrightarrow Sn^{2+}(aq)$	+0.13 V

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

 Continued on next page

17. The reaction of hydrogen gas and chlorine gas to form hydrogen chloride gas is a chain reaction. The relevant bond enthalpies are: Cl-Cl: 243 kJ/mol, H-H: 436 kJ/mol, H-Cl: 431 kJ/mol. Balance the reaction and calculate the enthalpy of reaction.
If 0.729 g of HCl gas are formed in the reaction, assuming all the heat (enthalpy) released is used to heat 100.0 g of water, what is the **temperature change** of the water, in °C?

A) 2.29
B) 4.38
C) 8.75
D) 11.9
E) 5.93

18. Use the following data to calculate the **electron affinity of O^{-}(g),** i.e., the **enthalpy change**, for the reaction,

$$O^{-}\,(g) + e^{-} \rightarrow O^{2-}\,(g).$$

Lattice energy of Na_2O =	-2481 kJ mol^{-1}
Formation enthalpy of Na_2O(s) =	-279.3 kJ mol^{-1}
Sublimation enthalpy of Na =	$+107.76$ kJ mol^{-1}
First ionization energy of Na =	$+500.$ kJ mol^{-1}
Formation enthalpy of O(g) =	$+249.2$ kJ mol^{-1}
Electron affinity of O(g) =	-141 kJ mol^{-1}

A) $+878$ kJ mol^{-1}
B) $+1486$ kJ mol^{-1}
C) $+1110.$ kJ mol^{-1}
D) -1486 kJ mol^{-1}
E) -878 kJ mol^{-1}

 Continued on next page

19. Calculate the **enthalpy of formation**, $\Delta H_f°$ in kJ mol^{-1}, of one mole of H_2O_2(g) from the relevant bond enthalpy data. (Hint: draw the Lewis structure for H_2O_2)

Bond	Bond enthalpy	Bond	Bond enthalpy
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) -260
B) -130
C) +260
D) -490
E) +130

20. The Bombardier beetle uses a hot chemical discharge as a defensive measure. The chemical reaction that releases the heat is the oxidation of hydroquinone by hydrogen peroxide to produce quinone and water:

$$C_6H_4(OH)_2(aq) + H_2O_2(aq) \rightarrow C_6H_4O_2(aq) + 2H_2O(l).$$

Calculate the **enthalpy change** (in kJ mol^{-1}) for this reaction, given the data below.

	$\Delta H°$ / kJ mol^{-1}
$C_6H_4O_2(aq) + H_2(g) \rightarrow C_6H_4(OH)_2(aq)$	–177.4
$H_2(g) + O_2(g) \rightarrow H_2O_2(aq)$	–191.2
$H_2(g) + ½O_2(g) \rightarrow H_2O(g)$	–241.8
$H_2O(g) \rightarrow H_2O(l)$	–43.8

A) –939.8
B) –654.2
C) –299.4
D) –202.6
E) –585.0

21. The reaction $2\ H_2(g) + O_2(g) \longrightarrow 2\ H_2O(g)$ is spontaneous **except** at very high temperature. Which of the following statements is/are **FALSE** for this reaction?
    (i) $\Delta H < 0$
    (ii) $\Delta S < 0$
    (iii) $H_2(g)$, $O_2(g)$ and $H_2O(g)$, all at 1 bar partial pressure, are in equilibrium at some very high temperature.
    (iv) $\Delta G < 0$ except at very high temperature.

A) iii
B) None of the statements are false.
C) i
D) i and ii
E) ii and iv

22. Consider the proposed reaction $H_2(g) + S(s) \longrightarrow H_2S(g)$ for which $\Delta H° = -20\ \text{kJ mol}^{-1}$ and $\Delta S° = +43\ \text{J K}^{-1}\ \text{mol}^{-1}$ at 298 K. Choose the **ONE TRUE** statement.

A) The reaction is driven by enthalpy only.
B) The reaction is not spontaneous at 298 K.
C) The reaction in not spontaneous at any temperature.
D) The reaction is spontaneous at all temperatures.
E) The reaction is spontaneous only above a temperature of 465 K.

 Continued on next page

23. Consider the reaction $H_2O_2(l) \longrightarrow H_2O(l) + 1/2\ O_2(g)$ for which $\Delta G° = -119$ kJ at 25°C. Select the **true** statement(s) below.
    (i) The partial pressure of oxygen gas is much less than 1 bar at equilibrium.
    (ii) This reaction is accompanied by an increase in entropy of the system.
    (iii) Hydrogen peroxide will decompose at 25°C and 1 bar pressure.

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

24. The melting point of H_2O is 0.00°C. The enthalpy of fusion (melting) for H_2O is 6.01 kJ mol^{-1}. What is the **entropy of fusion (J K^{-1} mol^{-1}) for H_2O**?
.

A) -45.5 J K^{-1} mol^{-1}
B) $+22.0$ J K^{-1} mol^{-1}
C) $+45.5$ J K^{-1} mol^{-1}
D) -22.0 J K^{-1} mol^{-1}
E) $+1.64$ J K^{-1} mol^{-1}

 Continued on next page

25. When NH_3(g) reacts spontaneously with HCl(g), NH_4Cl(s) is formed. This is accompanied by the following **changes**:

A) $\Delta G < 0$, $\Delta H < 0$, $\Delta S < 0$
B) $\Delta G > 0$, $\Delta H > 0$, $\Delta S > 0$
C) $\Delta G > 0$, $\Delta H < 0$, $\Delta S < 0$
D) $\Delta G < 0$, $\Delta H > 0$, $\Delta S > 0$
E) $\Delta G < 0$, $\Delta H > 0$, $\Delta S < 0$

26. Use the following data to determine **$\Delta G_f°$ of HBr(g)** (in kJ mol^{-1}) at 25.00°C.
Data:

	$\Delta H_f°$ / (kJ mol^{-1})	$S°$ / (J K^{-1} mol^{-1})
H_2(g)	0	130.684
Br_2(l)	0	152.2
HBr(g)	−36.40	198.695

.

A) −70.5
B) −53.5
C) −17.0
D) −106.9
E) −35.3

 Continued on next page

27. A 1.00 M solution of weak acid has a pH of 2.70. What is the associated value of $\boldsymbol{K_a}$?

A) 6.7×10^{-4}
B) 4.0×10^{-6}
C) 8.0×10^{-4}
D) 3.1×10^{-6}
E) none of these

28. What is the **pH** of an aqueous 0.0800 M NH_3 solution? [$K_b(NH_3) = 1.80\times10^{-5}$]

A) 12.15
B) 5.11
C) 11.08
D) 8.56
E) 7.93

 Continued on next page

29. You have 0.10 M aqueous solutions of $HClO_3$, $NaClO_3$, $HClO_2$ and $NaClO_2$. Which answer below correctly lists these solutions in **order of increasing pH**?

A) $HClO_3$ < $HClO_2$ < $NaClO_2$ < $NaClO_3$
B) $HClO_3$ < $NaClO_3$ < $HClO_2$ < $NaClO_2$
C) $HClO_3$ < $HClO_2$ < $NaClO_3$ < $NaClO_2$
D) $HClO_2$ < $HClO_3$ < $NaClO_3$ < $NaClO_2$
E) $HClO_2$ < $HClO_3$ < $NaClO_2$ < $NaClO_3$

30. A student is titrating NaOH against HCl to determine the unknown concentration of NaOH. The student accidentally uses a 5.00 ml volumetric pipette to pipette the 0.2451 M HCl thinking they used a 10.00 ml volumetric pipette. The student determines the concentration of NaOH to be 0.1871 M. What is the **actual concentration** of NaOH?

A) 0.2451 M
B) 0.09355 M
C) 0.3742 M
D) 0.1871 M
E) This cannot be determined without the volume of NaOH used

 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:
$E_n = -R_H / n^2 = -2.178 \times 10^{-18}$ J / n^2

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

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

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

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					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]