Name: _____ Student number: _____

Chemistry 1A03 **EXAM** **9:00 – 11:30 Dec 16, 2017**

McMaster University **VERSION 1**

Instructors: L. Davis, D. Emslie, S. Greenberg, A.P. Hitchcock

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 1 mark - the total marks available are 35. There is **no** penalty for incorrect answers.
BE SURE TO ENTER THE CORRECT VERSION OF YOUR EXAM (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 MS** or **FX-991 MS+** electronic calculators may be used. They must NOT be transferred between students. Use of any aids other than those provided, is not allowed.

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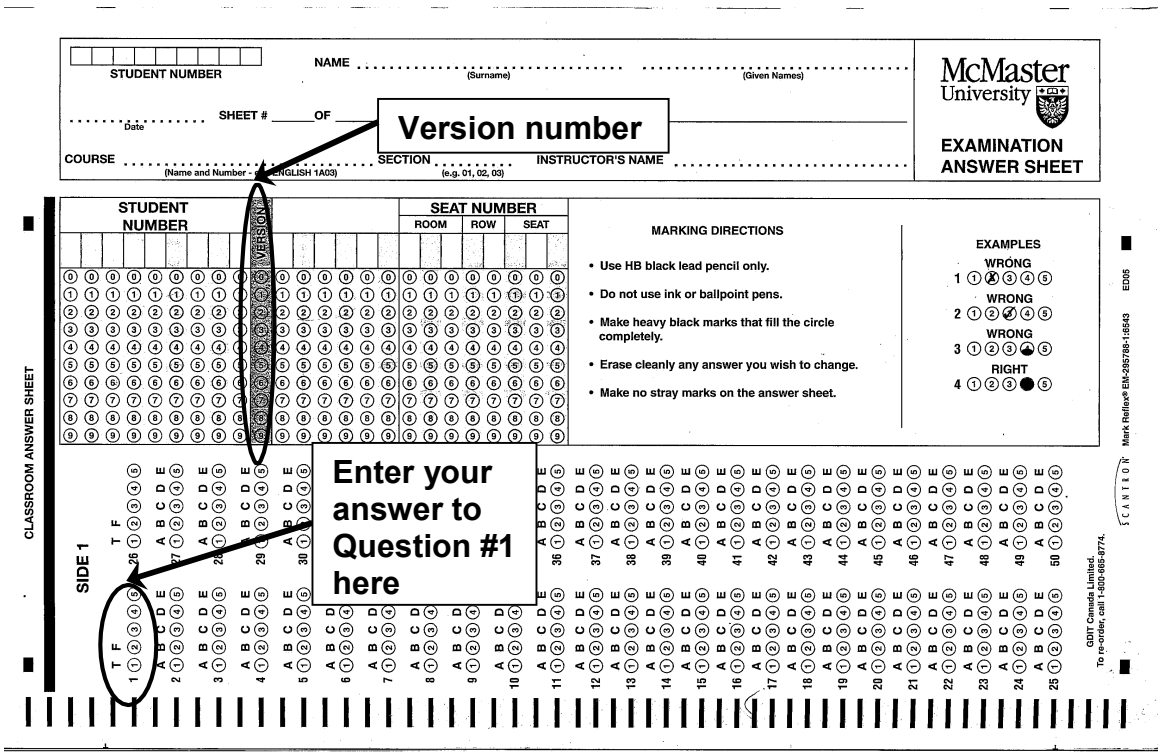
OMR EXAMINATION – STUDENT INSTRUCTIONS

NOTE: IT IS YOUR RESPONSIBILITY TO ENSURE THAT THE ANSWER SHEET IS PROPERLY COMPLETED. 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** 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.
3. Do NOT put in a leading zero when bubbling in your **exam version number**.
4. 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.
5. Pay particular attention to the marking directions on the form.
6. Begin answering the question using the first set of bubbles, marked "1".



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1. What is the **name** of the element with **atomic number** equal to 15 ?

   A) Phosphorus
   B) Nitrogen
   C) Beryllium
   D) Sulfur
   E) Potassium

2. An electron in a hydrogen atom relaxes to the n = 4 level, emitting light of 7.4×10^{13} Hz.What is the **value of n** for the level in which the electron originated?

   A) 7
   B) 5
   C) 3
   D) 2
   E) 8

3. Which one of the following statements about IF_3Cl_2 is **FALSE** ?

A) The I-F bonds are more polar than the I-Cl bonds.
B) The molecule can exist as three possible isomers.
C) One isomer of the molecule is non-polar.
D) The molecule contains 16 lone pairs of electrons.
E) In this compound, iodine does not obey the octet rule.

4. Which of the following statements regarding the reaction shown below are **FALSE** ?

$$SF_4\,(g) + F^-\,(g) \rightarrow SF_5^-\,(g)$$

i) SF_4 has a lone pair of electrons on sulfur.
ii) SF_4 is acting as a Lewis base in this reaction.
iii) The molecular geometry of the SF_5^- anion is trigonal bipyramidal.
iv) In SF_5^-, the oxidation state of sulfur is 4+, and the formal charge on sulfur is 1−.
v) The reaction has a negative ΔS.

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

5. Which one of the following molecules will have the **largest** net dipole moment ? Note: carbon is the central atom in CO_2, COS and COSe.

   A) I_2
   B) CO_2
   C) COS
   D) COSe
   E) XeF_4

6. Which of the following statement or statements are **FALSE** about ammonium nitrate ?

   i. The ammonia used to make ammonium nitrate is made from nitrogen using the Haber-Bosch process.
   ii. Ammonium nitrate is used as a fertilizer.
   iii. When ammonium nitrate is dissolved in water the solution gets warmer.
   iv. A solution of ammonium nitrate in water will turn phenolphthalein indicator pink.

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

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7. Consider the following equilibrium reaction:

$$2\ C\ (s) + O_2\ (g) \rightleftharpoons 2\ CO\ (g) \qquad \Delta H = 98.1\ kJ/mol$$

Which one of the following changes will cause the **position of equilibrium** to shift to the **right** ?

A) Decrease the pressure.
B) Cool the system down.
C) Remove oxygen gas.
D) Add solid carbon.
E) Add carbon monoxide gas.

8. In Experiment #3, the equilibrium constant for the following reaction was determined separately in three different trials. Which of the **factors** below would **NOT** impact the **reproducibility** of the experimentally determined value of *K* from one trial to next?

$$Fe^{3+}\ (aq) + SCN^{-}\ (aq) \rightleftharpoons Fe(SCN)^{2+}\ (aq)$$

i) Changing the initial concentrations of Fe^{3+} and SCN^{-}.
ii) Changing the temperature of the solution.
iii) Unknowingly using the wrong concentration of the Fe^{3+} stock solution.

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

9. When **equal volumes** of 2.0 M solutions of $Fe(NO_3)_3$ and KSCN are mixed, the following equilibrium is established.

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons Fe(SCN)^{2+}(aq)$$

At equilibrium, the concentration of $Fe(SCN)^{2+}$ is measured to be 0.80 M. What is the **equilibrium constant** for this reaction ?

A) 1.0
B) 40
C) 10
D) 20
E) 5.0

10. What is the **pH** of a 0.25 M solution of strontium hydroxide?

A) 7.00
B) 13.40
C) 0.30
D) 13.70
E) 0.60

11. Which statement(s) is/are **correct**?

(i) H_2O is a stronger acid than H_2S.
(ii) CH_3NH_2 is a stronger base than $ClNH_2$.
(iii) $HBrO_3$ is a stronger acid than $HBrO_2$.

A) i and ii
B) i, ii and iii
C) iii
D) ii and iii
E) none of these

12. At 25 °C, a 0.100 mol L^{-1} solution of lactic acid (a monoprotic acid) has a pH of 2.06. What is the **K_a of lactic acid**?

A) 8.7×10^{-3}
B) 8.3×10^{-4}
C) 7.6×10^{-4}
D) 1.1×10^{2}
E) 4.4×10^{-3}

13. Which one of the following statements is **FALSE** ?

A) An open system freely exchanges energy and matter with its surroundings.
B) *q* and *w* are both path functions.
C) Internal energy, U, is the total energy in a system.
D) The energy of an isolated system is not constant.
E) Internal energy, U, is a state function.

14. One mole of Ar gas was compressed by an external pressure of 2.00 atm, from an initial volume of 20.0 L to a final volume of 10.0 L. Calorimetry measurements determined that 1.00 kJ of energy was transferred as heat from the gas to the surroundings during the compression. For this process, what is the **change in internal energy** of the gas, $\Delta \boldsymbol{U}$ (in kJ) ?

A) +1.03
B) –1.03
C) –3.03
D) +3.03
E) –1.53

15. Calculate the **lattice enthalpy** (in kJ mol^{-1}) of solid **calcium bromide** from the following data:

State function	Value (kJ/mol)
Formation enthalpy of $CaBr_2$ (s)	–647.9
Sublimation enthalpy of Ca	192
Vaporization enthalpy of Br_2	30.9
Bond energy (enthalpy) of Br_2	192.9
First ionization energy of Ca	589.8
Second ionization energy of Ca	1145.4
Electron affinity of Br	–325

A) –2911
B) –2625
C) –3246
D) –2149
E) –1093

16. How much **energy (in kJ)** is required to heat a 20.00 g sample of ice (solid water) from -5.00 °C to 95.00 °C. Assume the specific heats of solid and liquid water are independent of temperature.

    specific heat H_2O (l) = 4.184 J g^{-1} $°C^{-1}$
    specific heat H_2O (s) = 2.060 J g^{-1} $°C^{-1}$
    ΔH_{fusion} = 333.55 J g^{-1}

    A) 19.35
    B) 25.68
    C) 34.21
    D) 18.92
    E) 14.83

17. Using bond energy data, estimate ΔH^o (in kJ) for the following reaction.

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

Bond	Bond Energy (kJ/mol)
H–H	432
C–H	413
C–C	347
C=C	614
C ≡ C	839

    A) +42
    B) –42
    C) –75
    D) +75
    E) –67

18. Which **ONE** of the following reactions corresponds to that of a **standard enthalpy of formation** ?

A) $N_2\,(g) + 3\,H_2\,(g) \rightarrow 2\,NH_3\,(g)$
B) $C\ (graphite) + O_2\,(g) \rightarrow CO_2\,(g)$
C) $CO\,(g) + 1/2\,O_2\,(g) \rightarrow CO_2\,(g)$
D) $O^{2-}\,(g) + Mg^{2+}\,(g) \rightarrow MgO\,(s)$
E) $H_2O_2\,(l) \rightarrow 2\,H_2O\,(l) + 1/2\,O_2\,(g)$

19. Using an ice calorimeter similar to the one used in Experiment # 4, it was determined that the reaction of 0.140 g of zinc with excess HCl (aq), caused 0.680 g of ice to melt. What was the **heat of reaction per mole of Zn** (in kJ mol^{-1}) ? Assume that no heat escaped the reaction chamber when $H_2(g)$ was evolved. [ΔH_{fus}(ice) = 333.55 J g^{-1}].

A) −106
B) −95.0
C) +240
D) −180
E) +590

20. When the following reaction proceeds in the forward direction, is it **endothermic** or **exothermic**, and is **work** done **on** or **by** the system ?

$$N_2O_4\,(g) \rightarrow 2\ NO_2\,(g)$$

A) endothermic, work done by the system
B) endothermic, work done on the system
C) exothermic, work done by the system
D) exothermic, work done on the system
E) not enough information

21. Identify the one **FALSE** statement regarding entropy and Gibbs energy.

A) $\Delta G^\circ_{\text{reaction}} = \Sigma n \Delta G^\circ_f\,(\text{products}) - \Sigma n \Delta G^\circ_f\,(\text{reactants})$.
B) $\Delta S > 0$ for melting of any substance.
C) At the temperature of a phase transition, ΔG for that phase transition is zero.
D) $\Delta G_{\text{reaction}} = T\Delta S_{\text{universe}}$.
E) A reaction is spontaneous if $\Delta G_{\text{reaction}} < 0$.

22. Which one of the following statements about **molar entropy** is **FALSE** ?

A) $S[CH_2I_2\ (l)]$ > $S[CH_2Cl_2\ (l)]$ at 25 °C.
B) $S[CH_3CCl_3\ (l)]$ at 60 °C > $S[CH_3CCl_3\ (l)]$ at 0 °C.
C) $S[CH_3(CH_2)_3CH_3\ (l)]$ > $S[CH_3(CH_2)_3CH_3\ (g)]$ at 25 °C.
D) $S[CH_3(CH_2)_3CH_3\ (l)]$ > $S[(CH_2)_5\ (l)]$ at 25 °C. $(CH_2)_5$ is cyclopentane.
E) $S[CO_2\ (g)]$ > $S[CO\ (g)]$ at 25 °C.

23. Which one of the following reactions will be **spontaneous** at <u>low temperature</u>, but **not spontaneous** at <u>high temperature</u> ?

A) $N_2\ (g) + 2\ O_2\ (g) \rightarrow N_2O_4\ (g)$
B) $Cl_2\ (g) \rightarrow 2\ Cl\ (g)$
C) $2\ NI_3\ (s) \rightarrow N_2\ (g) + 3\ I_2\ (g)$
D) graphite (s) → diamond (s)
E) $2\ Li\ (g) \rightarrow Li_2\ (g)$

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24. Bromine, Br_2, boils at 58.8 °C with an enthalpy of vaporization of 30.0 kJ mol^{-1}. At 58.8 °C, what is the **entropy of vaporization** (in J mol^{-1} K^{-1}) of Br_2 ?

    A) +66.2
    B) +154
    C) +90.4
    D) −66.2
    E) +22.7

25. A reaction has $\Delta H = -42.3$ kJ, and $\Delta S = +90$ J/K. **At what temperature(s)** would the reaction have an equilibrium constant **greater than 1.0** ?

    A) above 155°C
    B) all temperatures
    C) at 155°C only
    D) no temperature
    E) below 155°C

26. Choose the one **FALSE** statement regarding entropy, enthalpy, Gibbs energy, and equilibria.

A) The 3rd law of thermodynamics states that the entropy of a pure, perfect crystal of a substance is zero at a temperature of 0 K.
B) For a spontaneous process, ΔS could be positive or negative.
C) If Q for a reaction is greater than K, the reaction will proceed in the direction of reactants.
D) For a spontaneous process, ΔH could be positive or negative.
E) Above the fusion (i.e. melting) point of a substance, ΔG_{fusion} will have a positive value.

27. NI_3 decomposes as shown below:

$$2\ NI_3\ (s) \rightarrow N_2\ (g) + 3\ I_2\ (g)$$

Use the thermochemical information given below to calculate the **standard Gibbs energy change** (in kJ mol^{-1}) of this reaction at 298.15 K.

	NI_3 (s)	N_2 (g)	I_2 (g)
$\Delta H_f°$ (kJ mol^{-1})	287.0	0	62.40
$S°$ (J $mol^{-1}K^{-1}$)	230.0	191.6	260.7

A) –387.0
B) –540.0
C) +513.7
D) +119.0
E) –161.1

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28. Given the following data, calculate ΔG^{o} (in kJ) for the reaction:

$$6\ C\ (s) + 3\ H_2\ (g) \rightarrow C_6H_6\ (g)$$

$C_6H_6\ (l) + 7.5\ O_2\ (g) \rightarrow 6\ CO_2\ (g) + 3\ H_2O\ (l)$	$\Delta G^{o} = -3200$ kJ
$C\ (s) + O_2\ (g) \rightarrow CO_2\ (g)$	$\Delta G^{o} = -394$ kJ
$H_2\ (g) + 0.5\ O_2\ (g) \rightarrow H_2O\ (l)$	$\Delta G^{o} = -237$ kJ

A) +125
B) –2110
C) +411
D) –411
E) +256

29. Which one of the following reactions is **NOT** an oxidation-reduction reaction ?

A) $H_2\ (g) + Cl_2\ (g) \rightarrow 2\ HCl\ (g)$
B) $SO_3\ (g) + H_2O\ (l) \rightarrow H_2SO_4\ (aq)$
C) $Cl_2\ (aq) + H_2O\ (l) \rightarrow HCl\ (aq) + HOCl\ (aq)$
D) $2\ Al\ (s) + Fe_2O_3\ (s) \rightarrow A1_2O_3\ (s) + 2\ Fe\ (s)$
E) $3\ C\ (graphite) + 2\ KClO_3\ (s) \rightarrow 3\ CO_2\ (g) + 2\ KCl\ (s)$

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30. Sulfate can be converted to sulfite through the following reaction.

$$SO_4^{2-}(aq) \rightleftharpoons SO_3^{2-}(aq) + 0.5\ O_2(g)$$

How many **electrons** does **sulfur** gain or lose when the reaction proceeds in the forward direction ?

A) sulfur loses 4 electrons
B) sulfur loses 2 electrons
C) sulfur gains 4 electrons
D) sulfur gains 2 electrons
E) sulfur does not gain or lose electrons

31. Given the following E^o values, which species, **either reactant or product**, has the greatest tendency to be **oxidized** ?

$Cu^+(aq) + e^- \rightarrow Cu(s)$ $E^o = +0.520$ V
$O_2(g) + 4\ H^+ + 4\ e^- \rightarrow 2\ H_2O(l)$ $E^o = +1.229$ V
$Zn^{2+}(aq) + 2\ e^- \rightarrow Zn(s)$ $E^o = -0.763$ V

A) Cu (s)
B) H_2O (l)
C) Zn (s)
D) O_2 (g)
E) Zn^{2+} (aq)

32. Identify the one **FALSE** statement regarding the following electrochemical cell:

Cu (s) | Cu^{2+} (aq, 0.010 M) || Cu^{2+} (aq, 1.00 M) | Cu (s)

A) Electrons flow from the more dilute to the more concentrated half-cell.
B) The concentrated solution is at the cathode.
C) Q = [concentrated] / [dilute].
D) When [dilute] = [concentrated], the cell potential will be 0 V.
E) E_{cell} is always greater than or equal to 0.

33. An electrochemical cell is created based on the $Zn(NO_3)_2$ (aq) | Zn (s) and the $AgNO_3$ (aq) | Ag (s) half reactions. Note that $AgNO_3$ (s) is partially soluble in water. The $AgNO_3$ (aq) solution is saturated, and there is excess $AgNO_3$ (s) present. A cell created with $Zn(NO_3)_2$ (aq) at a concentration of 1.0×10^{-2} M gives a potential of 1.527 V at 298 K. Calculate the $\mathbf{K_{sp}}$ of **$AgNO_3$ (s)**.

Ag^+ (aq) + e^- → Ag (s) E° = +0.800 V
Zn^{2+} (aq) + 2 e^- → Zn (s) E° = −0.763 V

A) 2.1×10^{-3}
B) 7.8×10^{-12}
C) 1.3×10^{-14}
D) 6.9×10^{-6}
E) 6.1×10^{-4}

34. A Chemistry 1A03 student assembles an electrochemical cell based on the Cl_2 (g) | 2 Cl^- (aq) and Pb^{2+} (aq) | Pb (s) half reactions. When operated with all reagents under standard state conditions, she measures a cell voltage of 1.485 V at 298 K. What is **the equilibrium constant** for the reaction, when it is written to have the lowest whole number coefficients ?

A) 52
B) 1.5×10^{50}
C) 4.3×10^{12}
D) 3.1×10^{32}
E) 6.3×10^{4}

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

Fe^{3+} (aq) + OCl^- (aq) → FeO_4^{2-} (g) + Cl^- (aq)

A) 1
B) 3
C) 6
D) 10
E) 12

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Extra space for rough work

Name: _____ Student number: _____

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											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]												

(Columns 3–12: Transition Metals)

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