Doc M's Key

Chapter 5 Day 1 (Sections 5.1 – 5.4) (Unit 3) 4 October 2017

$h = 6.626 \times 10^{-34}$ J s

$c = \lambda\upsilon = 2.998 \times 10^{8}$ m/s

$\Delta E_{\text{per mol photon}} = \Delta E_{\text{per photon}} \times N_A$

$N_A = 6.023 \times 10^{23}$ mol^{-1}

$E = -2.178 \times 10^{-18}\text{J}(1/n^2)$

$\Delta E = E_f - E_i = -2.178 \times 10^{-18}\text{J}(1/n_f^2 - 1/n_i^2)$

$1/\lambda = 1.097 \times 10^{-2}\ \text{nm}^{-1}(1/n_f^2 - 1/n_i^2)$

1. What is the frequency of a helium-neon laser light with a wavelength of 632.8 nm?

$c = \lambda\upsilon$

$\upsilon = c/\lambda$

$= \frac{3.0\times10^{8}\ \text{m/s}}{632.8\times10^{-9}\ \text{m}}$

$= 4.74\times10^{+14}\ \text{s}^{-1}$

2a. What is the energy of a wavelength of light of 550 nm?

$E = \frac{hc}{\lambda}$

$= \frac{6.626\times10^{-34}\ \text{J·s} \mid 3\times10^{8}\ \text{m/s}}{550\times10^{-9}\ \text{m}}$

$= 3.6\times10^{-19}\ \text{J}$

2b. Continuing on, what is this energy in kJ/mol?

217.6 kJ/mol

3. The MRI body scanners used in hospitals operate with 400 MHz frequency energy. What is the wavelength of this frequency? Note: 1 Hz = 1 s^{-1} (or 1/s)

$400\times10^{6}\ \text{s}^{-1}$

$\lambda = \frac{c}{\upsilon} = \frac{3\times10^{8}\ \text{m/s}}{400\times10^{6}\ \text{s}^{-1}}$

0.75 m

4. Arrange the following spectral regions in order of increasing wavelength:

infrared, microwave, ultraviolet, visible

UV VIS IR MICRO

5. In the Balmer series, a red line with wavelength of 486.1 results when an electron drops from the n = 4 level. What is the energy of this electron transition in kJ/mol? Make sure you indicate whether it is + or -. Hint: Look up what the Balmer series is exactly.

Exothermic ΔE is (−)

$E = \frac{hc}{\lambda} = 4.09\times10^{-19}$

$\Delta E = -246$ kJ/mol

Questions in final exam format (multiple choice):

6. Wave (a) is on the left and (b) on the right

(a) (b)

Wave (a) has the

A. longer wavelength and higher energy than wave (b).

B. longer wavelength and lower energy than wave (b).

C. shorter wavelength and higher energy than wave (b).

D. shorter wavelength and lower energy than wave (b).

7. According to the Balmer-Rydberg equation, electromagnetic radiation with the shortest wavelength will be emitted when an electron undergoes which of the following transitions?

A. $m = 1 \rightarrow n = 2$ B. $m = 2 \rightarrow n = 3$

C. $n = 2 \rightarrow m = 1$ D. $n = 3 \rightarrow m = 2$

8. The four lines observed in the visible emission spectrum of hydrogen tell us that

A. The hydrogen molecules they came from have the formula H4.

B. We could observe more lines if we had a stronger prism.

C. There are four electrons in an excited hydrogen atom.

D. Only certain energies are allowed for the electron in a hydrogen atom.

Now try these problems from the book:

Section 5.1. (Electromagnetic spectrum) Problems 2, 24, 32

Section 5.2. (Electromagnetic spectrum) Problems 1, 36 38, 40, 42 and 44

Section 5.3. no assigned problems

Section 5.4. (Bohr model) Problems 25, 4, 46, 48, 50, 114, and 116.

Doc M's Key

Chapter 5 Day 2 (Sections 5.5 – 5.8) (Unit 3) 6 October 2017

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110	111	112		114		116		118

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

1. Which of the following sets of quantum numbers is not allowed? Explain what is wrong.

a. n = 4; l = 2; m_l = 2 OK	b. n = 1; l = 1; m_l = 0 $l < n$	c. n = 3; l = 3; m_l = -2 $l < n$
d. n = 4; l = 0; m_l = 1 $\lvert m_l \rvert \leq l$	e. n = 5; l = 2; m_l = -2 OK	f. n = 4; l = 3; m_l = -3 OK

2. Write the following sets of quantum numbers using the s, p, d, and f designations. The first one is done for you.

a. n = 4; l = 2; m_l = 2 4d	b. n = 5; l = 3; m_l = -3 5f	c. n = 3; l = 0; m_l = 0 3s
d. n = 2; l = 1; m_l = -1 2p	e. n = 1; l = 0; m_l = 0 1s	f. n = 3; l = 2; m_l = -2 3d

3. What values for l are allowed for each of these values of n? The first one is done for you.

a. n = 4; l = ? l = 0...3	b. n = 5; l = ? l = 0,,,4	c. n = 3; l = ? l = l = 0,,,2
d. n = 2; l = ? l = 0, 1	e. n = 1; l = l = 0	f. n = 74; l = l = 0,,,73

4. What values for n are allowed for each of these values of l? The first one is done for you.

(a) n = ?; l = 3 n > 3	(b) n = ?; l = 1 n > 1	(c) n = ?; l = 0 n ≥ 1
(d) n = ?; l = 4 n > 4	(e) n = ?; l = 2 n > 2	(f) n = ?; l = 37 n > 37

5. What values for n and l are allowed for each of these values of m_l =? The first one is done for you.

a. n = ?; l = ?; m_l = -2 n > l and $l \geq 2$	b. n = ?; l = ?; m_l = 3 $l \geq 3$ n > 3	c. n = ?; l = ?; m_l = 0 $l \geq 0$ n > 0
d. n = ?; l = ?; m_l = -1 $l \geq 1$ n > 1	e. n = ?; l = ?; m_l = 4 $l \geq 4$ n > 4	f. n = ?; l = ?; m_l = 15 $l \geq 15$ n > 15

6. List the orbitals in order of increasing energy (aufbau principle) up through barium, atomic number 56. The series has been started for you:

1s, 2s, 2p, 3s 3p 4s 3d 4p 5s 4d 5p 6s

Questions in final exam format (multiple choice):

7. What are the possible values of l if n = 5?

A. 5 (B) 0, 1, 2, 3, or 4

C. -4, -3, -2, -1, 0, +1, +2, +3, or +4

D. -5, -4, -3, -2, -1, 0, +1, +2, +3, +4, or +5

8. The subshell designations follow the alphabet after *f*. What is the first shell in which an *h* orbital would be allowed?

A. fifth (B. sixth) C. seventh D. eighth

Now try these problems from the book:

Section 5.5. (de Broglie) and Section 5.6. (Heisenberg) Read sections, but no assigned problems

Section 5.7. (Quantum mechanical model) Problems 11, 70, 74, 76, 78, 80, and 82

Section 5.8. (Orbital shapes) Problems 13, 14, and 26

Chapter 5 Day 3 (Sections 5.9 – 5.14) (Unit 3) 9 October 2017

This has been changed somewhat from the worksheet handed out in class.

(You will need a periodic table for this worksheet.)

1. 1. Write the electron configuration for each of the following. You may use core notation for Z > 18

(a) Na $1s^2 2s^2 2p^6 3s^1$	(b) P $1s^2 2s^2 2p^6 3s^2 3p^3$
(c) Co [Ar] $4s^2 3d^7$	(d) Sn [Kr] $5s^2 4d^{10} 5p^2$
(e) Zr [Kr] $5s^2 4d^2$	(f) Bi [Xe] $6s^2 4f^{14} 5d^{10} 6p^3$

2. How many unpaired electrons does each of the elements in Question 1 possess?

(a) Na 1	(b) P 3
(c) Co 3	(d) Sn 2
(e) Zr 2	(f) Bi 3

3. Circle the largest member of each set. In some cases the correct answer is "Can't predict"

(a) Na or P	(b) P or Bi	(c) Co or Zr
(d) Sn or P	(e) Sn or Bi cannot predict	(f) Li or Na

4. Circle the smallest member of each set. In one case the correct answer is "Can't predict"

(a) Na K Rb	(b) P S Cl	(c) Cu Ag Au
(d) Sn As S	(e) Pt Ag Zn	(f) C P Se cannot predict

5. True/False

T F Z_{eff} is always < Z. except for H

T F The atomic radius is proportional to Z_{eff} within a shell. ↑ directly

T F Atomic radius decreases within a subshell.

T F The maximum number of possible unpaired electrons increases as *l* increases.

T F The atomic orbitals within a subshell have degenerate energies.

↑ the same

Questions in final exam format (multiple choice):

6. How many orbitals are there in the seventh shell?

A. 6 B. 7 C. 21 D. 49

7. An element in a ground state electron configuration has 4 electrons in the 4p orbitals. Which of the following statements can **not** describe the electron configurations in this atom?

A. At least one electron has an orbital angular momentum (*l*) of 2.
B. Six electrons are in the n=4 shell.
C. The valence electron configuration is identical to carbon.
D. No electrons have an orbital angular momentum (l) of 3.3.

8. Molybdenum has an anomalous electron configuration. Write the electron configuration of Mo using shorthand notation.

A. [Kr] $5s^0\ 4d^6$
B. [Kr] $5s^0\ 4d^0\ 5p^6$
C. [Kr] $5s^1\ 4d^5$
D. [Kr] $5s^2\ 4d^4$

9. 58) Rank these elements in order of increasing effective nuclear charge: Na, Mg, Cl, and S.

A. S < Cl < Mg < Na
B. Cl < Mg < Na < S
C. Mg < Na < Cl < S
D. Na < Mg < S < Cl

Now try these problems from the book:

Section 5.9. (Quantum number *m*) Problems 80 and 82

Section 5.10. (Orbital energy levels) Problems 84 and 90

Section 5.11 & 5.12. (electron configurations) Problems 8, 92, 94, 96, 98, 110 and 111

Section 5.13. (Periodic table) Problems 8, 15, 16, 96, 97, 98, 100 and 102

Section 5.14. (atomic radii) Problems 86, 104, 106, 108 and 110

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Chapter 6 Day 1 (Sections 6.1 – 6.4) (Unit 3) 11 October 2017

1. Write the electron configuration for each of the following. You may use core notation for all ions with Z > 18.

(a) Na^+	$1s^2 2s^2 2p^6$
(b) O^{2-}	$1s^2 2s^2 2p^6$
(c) Sr^{2+}	[Kr] or [Ar] $4s^2 3d^{10} 4p^6$
(d) Sn^{2+}	[Kr] $5s^2 4d^{10} 5p^0$
(e) Cr^{3+}	[Ar] $4s^0 3d^3$
(f) P^{3-}	[Ar] or [Ne] $3s^2 3p^6$

2. How many unpaired electrons does each of the elements in Question 1 possess?

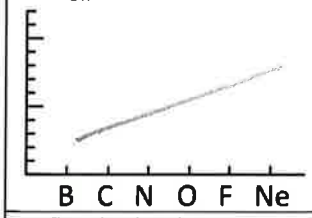
(a) Na^+ 0	(b) O^{2-} 0	(c) Sr^{2+} 0
(d) Sn^{2+} 0	(e) Cr^{3+} 3	(f) P^{3-} 0

3. List each set of atoms/ions from smallest to largest.

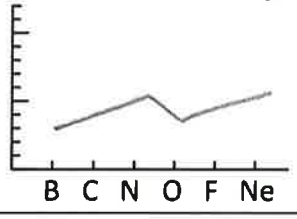
(a) Na^+ Mg^{2+} Al^{3+} — Al^{3+} Mg^{2+} Na^+	(b) Na^+ K^+ Rb^+ — as written
(c) S^{2-} Se^{2-} Te^{2-} — as written	(d) Sn Sn^{2+} — Sn^{2+} Sn
(e) F F^- — as written	(f) P^{3-} S^{2-} Cl^- — Cl^- S^{2-} P^{3-}

4. Graph the following periodic trends. The y-axis increases in value upward.

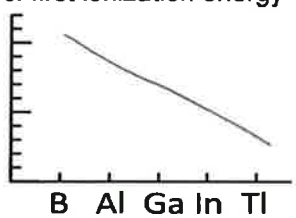
a. Z_{eff}



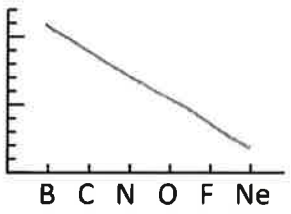
b. first ionization energy



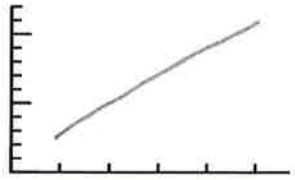
c. first ionization energy



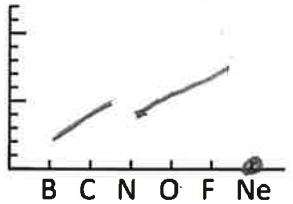
d. atomic radius



e. atomic radius



(f) electron affinity



Questions in final exam format (multiple choice):

5. Arrange the ions N^{3-}, O^{2-}, Mg^{2+}, Na^+, and F^- in order of increasing ionic radius, starting with the smallest first.
   A. Mg^{2+}, Na^+, F^-, O^{2-}, N^{3-}
   B. N^{3-}, Mg^{2+}, O^{2-}, Na^+, F^-
   C. N^{3-}, O^{2-}, Mg^{2+}, F^-, Na^+
   D. N^{3-}, O^{2-}, F^-, Na^+, Mg^{2+}

6. Of the following, which element has the highest first ionization energy?
   A. Li B. F C. Cs D. At

7. Which period 3 element has successive first through seventh ionization energies (kJ/mol) of E_{i1} = 578; E_{i2} = 1,817; E_{i3} = 2,745; E_{i4} = 11,575; E_{i5} = 14,830; E_{i6} = 18,376; and E_{i7} = 23,293?
   A. Mg B. Al C. S D. Cl

8. Which of the following atoms with the specified electronic configurations would have the lowest first ionization energy?
   A. [He] $2s^2 2p^3$
   B. [Ne] $3s^2 3p^4$
   C. [Xe] $6s^1$ Cs
   D. [Xe] $6s^2 4f^{14} 5d^{10} 6p^1$

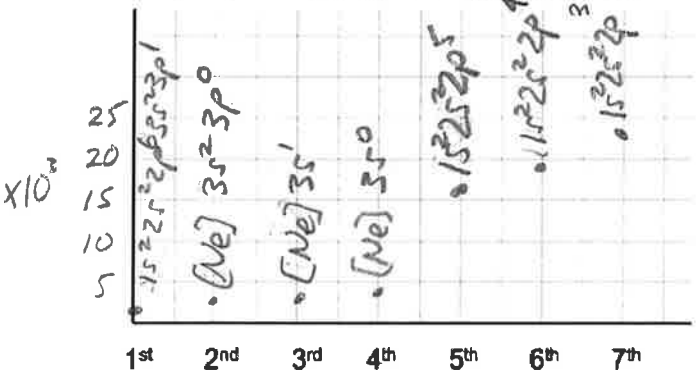
9. 20) List the elements Cs, Ca, Ne, Na, Ar in order of decreasing first ionization energy.
   A. Ar > Ca > Cs > Na > Ne
   B. Ne > Ar > Ca > Na > Cs
   C. Ne > Ar > Na > Cs > Ca
   D. Ne > Na > Cs > Ca > Ar

Now try these problems from the book:
Section 6.1. (Electron configurations of ions) Problems 1, 2, 20, 36, 38, 40, 42, and 44
Section 6.2. (Ionic radii) Problems 3, 4, 21, 46, and 48
Section 6.3. (Ionization energy) Problems 5, 6, 50, and 52
Section 6.4. (Higher ionization energies) Problems 7, 8, 23, 24, 54, 56, 58

Key

Chapter 6 Day 2 (Sections 6.5 – 6.8) (Unit 3) 13 October 2017

1. 1. The sequential ionization energies for silicon are (all in kJ/mol): E_{i1} = 787; E_{i2} = 1577; E_{i3} = 3231; E_{i4} = 4356; E_{i5} = 16091; E_{i6} = 19784; and E_{i7} = 23783. Plot these data on the graph paper below. (Make each horizontal line an increment of 5000 kJ) Write the electron configuration of the resulting ion near each data point – so for 1st, you would write $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$, and so on.



2a. What is the solubility rule involving Group 1 cations?

All soluble

2b. What is the solubility rule involving halides?

All soluble except for Ag^+ Pb^{2+} Hg_2^{2+}

2c. From the table in the book (and on the handout) what is the range of lattice energy values for Group 1 halides? Is this the energy associated with forming the lattice from gas phase ions or breaking the lattice into gas phase ions?

649 – 1036 kJ/mol

2d. What is energy, E, for this process? (Think carefully about the sign) $Na^+(g) + Cl^-(g) \rightarrow NaCl(s)$ E =

E = -787 kJ/mol

3a. The lattice energies of $CaSO_4$, $SrSO_4$, and $BaSO_4$ are 2653, 2603 and 2423 kJ/mol, respectively. Why do you think these lattice energies are larger than for Group 1 halides?

$E \propto \frac{Z^+ \times Z^-}{d}$

3b. How does lattice energy seem to relate to solubility?

Large lattice energy → low solubility

4a. The charge on the cation is called Z^+ and the charge on the anion is Z^-. Lattice energies are $\propto$ $Z^+ \times Z^-$. Complete this table of $Z^+ \times Z^-$ and give an example of each combination. Circle those that you know or suspect are insoluble from a solubility rule.

$\lvert Z^+ \times Z^- \rvert$	$Z^+ = 1$	$Z^+ = 2$	$Z^+ = 3$
$Z^- = -1$	-1, NaCl	-2	-3
$Z^- = -2$	-2	-4, $CaCO_3$	-6
$Z^- = -3$	-3	-6	-9, $AlPO_4$

4b. Can you make a generalization about the product of charges and lattice energies?

If product ≥ 4 insoluble

Questions in final exam format (multiple choice):

5. Which liberates the most energy?
   A. $Br(g) + e^- \rightarrow Br^-(g)$
   B. $Cl(g) + e^- \rightarrow Cl^-(g)$
   (C) $F(g) + e^- \rightarrow F^-(g)$ ← We would predict this
   D. $I(g) + e^- \rightarrow I^-(g)$

What is the ground-state electron configuration of the ion Cu^{2+}?

(A) [Ar] $3d^9$ B. [Ar] $4s^1\ 3d^8$
C. [Ar] $4s^2\ 3d^7$ D. [Ar] $4s^2\ 3d^{10}\ 4p^1$

6. Which ionic compound would be expected to have the highest lattice energy?
   A. Li_2O B. Na_2O
   C. K_2O_2 (D) Ca_3PO_4

Now try these problems from the book:

Section 6.5. (Electron affinity) Problems 9, 10, 62, 64, 66, and 96

Section 6.6. (Octet Rule) Problems 11, 12, 30, 68 and 70

Section 6.7. (Ionic Bonds/Ionic Solids) Problems 13, 14, 26, 80, 82, 94 and 98

Section 6.8. (Lattice energy) Problems 15, 16, and 28