Homework for 6.1 Electromagnetic Energy

1. The light produced by a red neon sign is due to the emission of light from excited neon atoms. Qualitatively describe the spectrum produced by passing light from a neon lamp through a prism.

   The spectrum consists of colored lines, at least one of which (probably the brightest) is red.

2. An FM radio station found at 103.1 on the FM dial broadcasts at a frequency of 1.031×10^{8} s^{-1} (103.1 MHz). What is the wavelength of these radio waves in meters?

   2.91 m
   (use $c = \lambda v$)

3. FM-95, an FM radio station, broadcasts at a frequency of 9.51×10^{7} s^{-1} (95.1 MHz). What is the wavelength of these radio waves in meters?

   3.15 m
   (use $c = \lambda v$)

4. A bright violet-blue line occurs at 435.8 nm in the emission spectrum of mercury vapor. What amount of energy, in joules, must be released by an electron in a mercury atom to produce a photon of this light?

   4.56×10^{-19} J
   (use $E = hv$ and $c = \lambda v$)

5. Light with a wavelength of 614.5 nm looks orange. What is the energy, in joules, per photon of this orange light? What is the energy in eV (1 eV = 1.602×10^{-19}J)?

   3.24×10^{-19} J; 2.02 eV
   (use $E = hv$ and $c = \lambda v$)

6. Heated lithium atoms emit photons of light with an energy of 2.961×10^{-19} J. Calculate the frequency and wavelength of one of these photons. What is the total energy in 1 mole of these photons? What is the color of the emitted light?

   $v = 4.47 \times 10^{14}$ s^{-1}; λ = 671 nm (6.71×10^{-7} m); 178 kJ mol^{-1}; red
   (use $E = hv$ and $c = \lambda v$; multiply the energy of one photon by Avogadro's number; recall that violet light is approximately 400-450 nm and red approximately 630-750 nm)

7. A photon of light produced by a surgical laser has an energy of 3.027×10^{-19} J. Calculate the frequency and wavelength of the photon. What is the total energy in 1 mole of photons? What is the color of the emitted light?

   $v = 4.568 \times 10^{14}\ s^{-1}$; $\lambda = 656.3$ nm; 182 kJ mol^{-1}; red
   (use $E = hv$ and $c = \lambda v$; multiply the energy of one photon by Avogadro's number; recall that violet light is approximately 400-450 nm and red approximately 630-750 nm)

8. When rubidium ions are heated to a high temperature, two lines are observed in its line spectrum at wavelengths (a) 7.9×10^{-7} m and (b) 4.2×10^{-7} m. What are the frequencies of the two lines? What color do we see when we heat a rubidium compound?

   $v = 3.8 \times 10^{14}\ s^{-1}$ and $v = 7.1 \times 10^{14}\ s^{-1}$; red and violet; purply-red

9. The emission spectrum of cesium contains two lines whose frequencies are (a) 3.45×10^{14} Hz and (b) 6.53×10^{14} Hz. What are the wavelengths and energies per photon of the two lines? What color are the lines?

   (a) $\lambda = 8.69 \times 10^{-7}$ m; $E = 2.29 \times 10^{-19}$ J; (b) $\lambda = 4.59 \times 10^{-7}$ m; $E = 4.33 \times 10^{-19}$ J; The color of (a) is infrared (colorless to the human eye); (b) is blue.

10. One of the radiographic devices used in a dentist's office emits an X-ray of wavelength 2.090×10^{-11} m. What is the energy, in joules, and frequency of this X-ray?

    $E = 9.502 \times 10^{-15}$ J; $v = 1.434 \times 10^{19}\ s^{-1}$

11. The eyes of certain reptiles pass a single visual signal to the brain when the visual receptors are struck by photons of a wavelength of 850 nm. If a total energy of 3.15×10^{-14} J is required to trip the signal, what is the minimum number of photons that must strike the receptor?

    E of one photon = 2.34×10^{-19} J; number of photons = 1.35×10^{5}

12. RGB color television and computer displays use cathode ray tubes that produce colors by mixing red(~700 nm), green (~550 nm), and blue (~450 nm) light. If we look at the screen with a magnifying glass, we can see individual dots turn on and off as the colors change. What is the frequency and energy of a photon of each of these colors?

    Red: 700 nm; 4.29×10^{14} Hz; 2.84×10^{-19} J
    Green: 550 nm; 5.45×10^{14} Hz; 3.62×10^{-19} J
    Blue: 450 nm; 6.67×10^{14} Hz; 4.42×10^{-19} J

    (somewhat different numbers are also possible if slightly different wavelengths are chosen)

13. Answer the following questions about a Blu-ray laser:

(a) The laser on a Blu-ray player has a wavelength of 405 nm. In what region of the electromagnetic spectrum is this radiation? What is its frequency?

Visible (violet); $\nu = 7.41 \times 10^{14}\ s^{-1}$

(b) A Blu-ray laser has a power of 5 milliwatts (1 watt = 1 J s^{-1}). How many photons of light are produced by the laser in 1 hour?

E of one photon = $h\nu = 4.91 \times 10^{-19}$ J; 5 mW = 0.005 J s^{-1} × 60 × 60 = 18 J per hour; $18/4.91 \times 10^{-19} = 3.67 \times 10^{19}$ photons

(c) The ideal resolution of a player using a laser (such as a Blu-ray player) dictates how close together data can be stored on a compact disk. The ideal resolution is determined using the following formula: Resolution = 0.60 (λ/NA), where λ is the wavelength of the laser and NA is the numerical aperture. Numerical aperture is a measure of the size of the spot of light on the disk; the larger the NA, the smaller the spot. In a typical Blu-ray system, NA = 0.95. If the 405 nm laser is used in a Blu-ray player, what is the closest that information can be stored on a Blu-ray disk?

Resolution = 0.60 (405 nm/0.95) = 256 nm apart

14. What is the threshold frequency for sodium metal if a photon with frequency $6.66 \times 10^{14}\ s^{-1}$ ejects an electron with 7.74×10^{-20} J kinetic energy? Will the photoelectric effect be observed if sodium is exposed to orange light?

$5.49 \times 10^{14}\ s^{-1}$; no

Homework for 6.2 The Bohr Model

1. Why is the electron in a Bohr hydrogen atom bound less tightly when it has a quantum number of 3 than when it has a quantum number of 1?

   It is on average further from the nucleus.

2. What does it mean to say that the energy of the electrons in an atom is quantized?

   Quantized energy means that the electrons can possess only certain discrete energy values; values between those quantized values are not permitted.

3. Using the Bohr model, determine the energy, in joules, necessary to ionize a ground-state hydrogen atom. Show your calculations.

   $E = E_\infty - E_1 = -2.18 \times 10^{-18} (1/n_\infty^2 - 1/n_1^2)\ J = -2.18 \times 10^{-18} (0 - 1/1) = 2.18 \times 10^{-18}\ J$

4. The electron volt (eV) is a convenient unit of energy for expressing atomic-scale energies. It is the amount of energy that an electron gains when subjected to a potential of 1 volt; 1 eV = 1.602×10^{-19} J. Using the Bohr model, determine the energy, in electron volts, of the photon produced when an electron in a hydrogen atom moves from the orbit with $n = 5$ to the orbit with $n = 2$. Show your calculations.

$E = E_2 - E_5 = -2.18 \times 10^{-18} (1/(n_2^2 - 1/n_5^2)$ J $= -2.18 \times 10^{-18} (1/4 - 1/25) = 4.576 \times 10^{-19}$ J

$4.576 \times 10^{-19}/1.602 \times 10^{-19} = 2.86$ eV

5. Using the Bohr model, determine the lowest possible energy for the electron in the He^+ ion. *(Covered in textbook, but not in lecture. Not examinable)*

$E = -2.18 \times 10^{-18} Z(1/n^2) = -2.18 \times 10^{-18} (2)(1/1^2) = -8.72 \times 10^{-18}$ J

6. Using the Bohr model, determine the energy of an electron with $n = 8$ in a hydrogen atom.

-3.41×10^{-20} J

7. Using the Bohr model, determine the energy in joules of the photon produced when an electron in a Li^{2+} ion moves from the orbit with $n = 2$ to the orbit with $n= 1$. *(Covered in textbook, but not in lecture. Not examinable)*

1.47×10^{-17} J

8. Consider a large number of hydrogen atoms with electrons randomly distributed in the $n = 1, 2, 3$, and 4 orbits.

(a) How many different wavelengths of light are emitted by these atoms as the electrons fall into lower-energy orbitals?

6: 2→1, 3→1, 3→2, 4→1, 4→2, 4→3

(b) Calculate the lowest and highest energies of light produced by the transitions described in part (a).

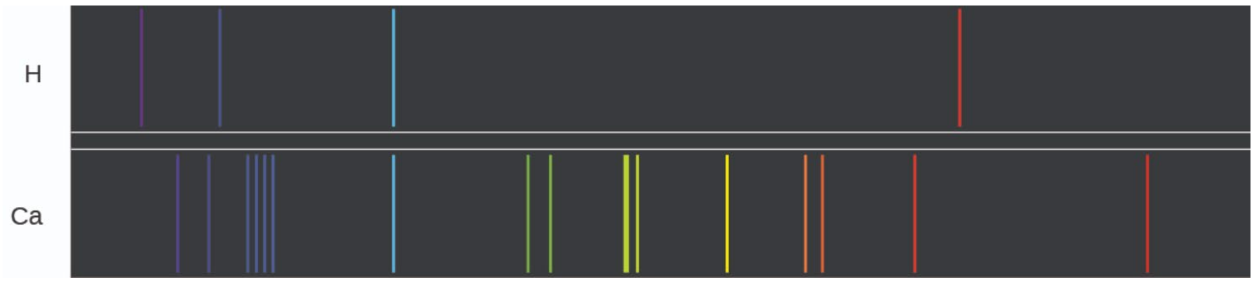
Highest = largest drop = 4→1
$E = E_1 - E_4 = -2.18 \times 10^{-18} (1/(n_1^2 - 1/n_4^2)$ J $= -2.18 \times 10^{-18} (1/1 - 1/16) = -2.04 \times 10^{-18}$ J

Lowest = smallest drop = 4→3
$E = E_3 - E_4 = -2.18 \times 10^{-18} (1/(n_3^2 - 1/n_4^2)$ J $= -2.18 \times 10^{-18} (1/9 - 1/16) = -1.06 \times 10^{-19}$ J

(c) Calculate the frequencies and wavelengths of the light produced by the transitions described in part (b).

v = 3.08 × 10^{15} s-1, λ = 97.0 nm (ultraviolet)
v = 1.60 × 10^{14} s-1, λ = 1.88 × 10^{-6} 1880 nm (infrared)

9. The spectra of hydrogen and of calcium are shown below. What causes the lines in these spectra? Why are the colors of the lines different? Suggest a reason for the observation that the spectrum of calcium is more complicated than the spectrum of hydrogen.



These are emission lines caused from excited electrons dropping into lower energy orbitals. The colours depend on the energy change during the emission. Calcium is more complicated than hydrogen because it has more than one electron and so there is less degeneracy (i.e. in H, the 4*s* and 4*p* orbitals have the same energy, but this is not true in Ca; the 4*s* are lower in energy thanks to being shielded less from the nucleus by the other electrons).

Homework for 6.3 Development of Quantum Theory

1. What are the allowed values for each of the four quantum numbers: *n*, *l*, m_l, and m_s?

n must be a positive integer; *l* = 0, 1,2,...(n-1); m_l = integer values from –l to +l ; m_s = ±½

2. An electron has four quantum numbers: *n*, *l*, m_l, and m_s. Describe the electron properties associated with each quantum number.

   n determines the general range for the value of energy and the probable distances that the electron can be from the nucleus. *l* determines the shape of the orbital. m_1 determines the orientation of the orbitals of the same *l* value with respect to one another. m_s determines the spin of an electron.

3. Identify the subshell in which electrons with the following quantum numbers are found:

   (a) *n* = 2, *l* = 1

   (b) *n* = 4, *l* = 2

   (c) *n* = 6, *l* = 0

(a) 2*p*; (b) 4*d*; (c) 6*s*

4. Which of the subshells described in the previous question contain degenerate orbitals? How many degenerate orbitals are in each?

   2*p*; 3. 4*d*; 5.

5. Identify the subshell in which electrons with the following quantum numbers are found:

   (a) $n = 3, l = 2$

   (b) $n = 1, l = 0$

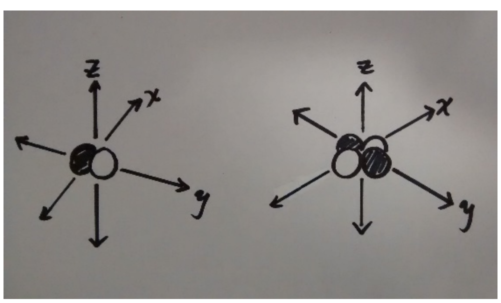
   (c) $n = 4, l = 3$

   (a) 3*d*; (b) 1*s*; (c) 4*f*

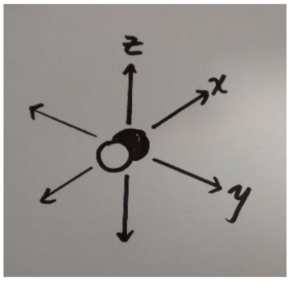
6. Which of the subshells described in the previous question contain degenerate orbitals? How many degenerate orbitals are in each?

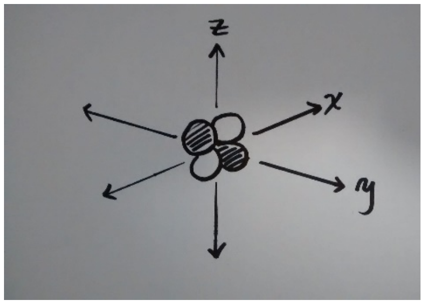
   3*d*; 5. 4*f*; 7.

7. Sketch the boundary surface of a p_y orbital and a d_{x2-y2} orbital. Be sure to show and label the axes.

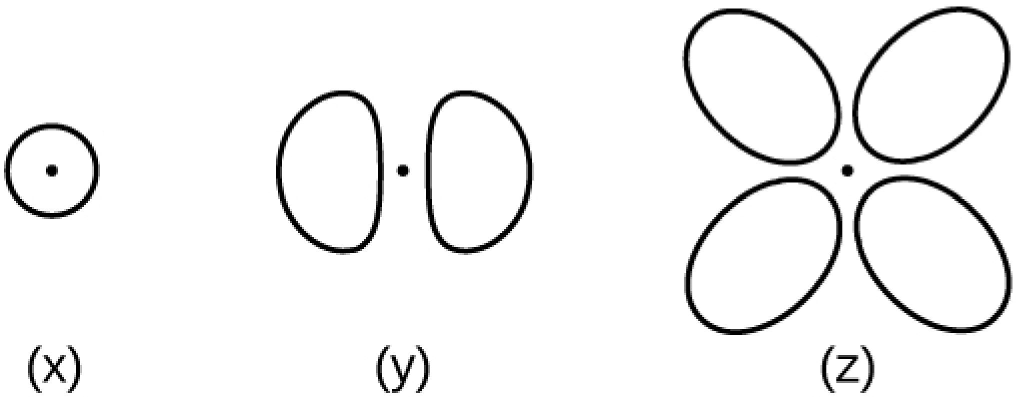


8. Sketch the p_x and d_{xz} orbitals. Be sure to show and label the coordinates.





9. Consider the orbitals shown here in outline.



(a) What is the maximum number of electrons contained in an orbital of type (x)? Of type (y)? Of type (z)?

(b) How many orbitals of type (x) are found in a shell with n = 2? How many of type (y)? How many of type (z)?

(c) Write a set of quantum numbers for an electron in an orbital of type (x) in a shell with n = 4. Of an orbital of type (y) in a shell with n = 2. Of an orbital of type (z) in a shell with n = 3.

(d) What is the smallest possible n value for an orbital of type (x)? Of type (y)? Of type (z)?

(e) What are the possible l and m_l values for an orbital of type (x)? Of type (y)? Of type (z)?

(a) x. 2, y. 2, z. 2
(b) x. 1, y. 3, z. 0
(c) x. 4 0 0 ½, y. 2 1 0 ½, z. 3 2 0 ½ (other combinations are possible for y and z)
(d) x. 1, y. 2, z. 3
(e) x. l = 0, m_l = 0, y. l = 1, m_l = –1, 0, or +1, z. l = 2, m_l = –2, –1, 0, +1, +2

10. How many electrons could be held in the second shell of an atom if the spin quantum number m_s could have three values instead of just two? (Hint: Consider the Pauli exclusion principle.)

    12

11. Write a set of quantum numbers for each of the electrons with an n of 4 in a Se atom.

n	l	m_l	m_s
4	0	0	+½
4	0	0	–½
4	1	−1	+½
4	1	0	+½
4	1	+1	+½
4	1	−1	–½

Homework for 6.4 Electronic Structure of Atoms

1. Using complete subshell notation ($1s^22s^22p^6$, and so forth), predict the electron configuration of each of the following atoms:

   (a) C, (b) P, (c) V, (d) Sb, (e) Sm

   (a) C: $1s^22s^22p^2$; (b) P: $1s^22s^22p^63s^23p^3$; (c) V: $1s^22s^22p^63s^23p^64s^23d^3$;
   (d) Sb: $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^3$;
   (e) Sm: $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^6$

2. Using complete subshell notation ($1s^22s^22p^6$, and so forth), predict the electron configuration of each of the following atoms:

(a) N, (b) Si, (c) Fe, (d) Te, (e) Tb

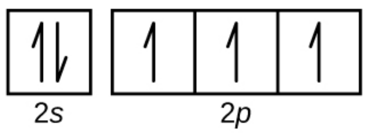
(a) $1s^22s^22p^3$; (b) $1s^22s^22p^63s^23p^2$; (c) $1s^22s^22p^63s^23p^64s^23d^6$;
(d) $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^4$; (e) $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^9$

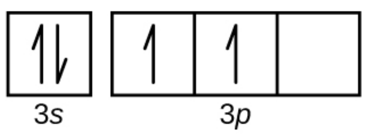
3. What additional information do we need to answer the question “Which ion has the electron configuration $1s^22s^22p^63s^23p^6$”?

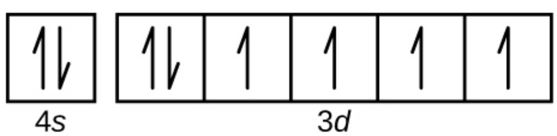
The charge on the ion.

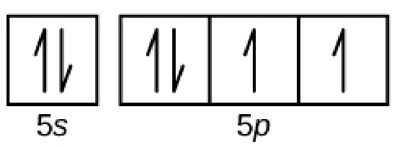
4. Use an orbital diagram to describe the electron configuration of the valence shell of each of the following atoms:

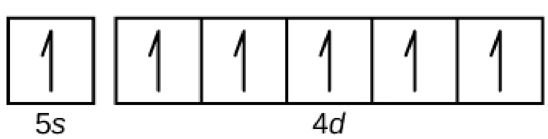
(a) N; (b) Si; (c) Fe; (d) Te; (e) Mo

(a) 

(b) 

(c) 

(d) 

(e) 

5. Using complete subshell notation ($1s^22s^22p^6$, and so forth), predict the electron configurations of the following ions.

(a) N^{3-}; (b) Ca^{2+}; (c) S^-; (d) Cs^+; (e) Cr^{2+}; (f) Gd^{3+}

(a) $1s^22s^22p^6$; (b) $1s^22s^22p^63s^23p^6$; (c) $1s^22s^22p^63s^23p^5$;
(d) $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6$; (e) $1s^22s^22p^63s^23p^63d^4$; (f) $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^64f^7$

6. The label below, from a commercial multivitamin product, contains numerous elements. Write the condensed electron configurations of (a) the cations Mg^{2+}, Fe^{3+}, Mn^{2+}, Cr^{3+}, and K^+; (b) the elements P, Si, B, and V; and (c) the anions I^-, Se^{2-}, and Cl^-.

Supplement Facts

Serving Size 1 Tablet

Each Tablet Contains	% Daily Value
Vitamin A 897 mcg (17% as Beta-Carotene)	99%
Vitamin C 90 mg	100%
Vitamin D3 10 mcg (400 IU)	50%
Vitamin E 13 mg	87%
Vitamin K 25 mcg	21%
Thiamin (Vitamin B-1) 1.5 mg	125%
Riboflavin (Vitamin B-2) 1.7 mg	131%
Niacin 20 mg	125%
Vitamin B6 2 mg	118%
Folate 850 mcg DFE (500 mcg folic acid)	213%
Vitamin B12 6 mcg	250%
Biotin 30 mcg	100%
Pantothenic Acid 10 mg	200%
Calcium 200 mg	15%
Iron 18 mg	100%
Phosphorus 109 mg	9%

Each Tablet Contains	% Daily Value
Iodine 150 mcg	100%
Magnesium 100 mg	24%
Zinc 11 mg	100%
Selenium 55 mcg	100%
Copper 0.9 mg	100%
Manganese 2.3 mg	100%
Chromium 35 mcg	100%
Molybdenum 45 mcg	100%
Chloride 72 mg	3%
Potassium 80 mg	2%
Silicon 2 mg	**
Lycopene 300 mcg	**
Lutein 250 mcg	**
Boron 150 mcg	**
Vanadium 10 mcg	**
Nickel 5 mcg	**

**Daily Value not established.

(a) Mg^{2+}: [Ne]; Fe^{3+}: [Ar]$3d^5$; Mn^{2+}: [Ar]$3d^5$; Cr^{3+}: [Ar]$3d^3$; K^+: [Ar]
(b) P: [Ne]$3s^23p^3$; Si: [Ne]$3s^23p^2$; B: [He]$2s^22p^1$; V: [Ar]$4s^23d^3$
(c) I^-: [Xe]; Se^{2-}: [Kr]; Cl^-: [Ar]

7. Which atom has the electron configuration $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^2$?

Zr

8. Which atom has the electron configuration $1s^22s^22p^63s^23p^64s^23d^7$?

Co

9. Which ion with a +1 charge has the electron configuration $1s^22s^22p^63s^23p^63d^{10}4s^24p^6$? Which ion with a –2 charge has this configuration?

Rb^+, Se^{2-}

10. Which of the following atoms contains only three valence electrons: Li, B, N, F, Ne?

    B

11. Which of the following has two unpaired electrons?

    (a) Mg; (b) Si; (c) S; (d) Both Mg and S; (e) Both Si and S.

    Although both (b) and (c) are correct, (e) encompasses both and is the best answer.

12. Which atom would be expected to have a half-filled 6*p* subshell?
    Bi

13. Which atom would be expected to have a half-filled 4*s* subshell?
    K

14. In one area of Australia, the cattle did not thrive despite the presence of suitable forage. An investigation showed the cause to be the absence of sufficient cobalt in the soil. Cobalt forms cations in two oxidation states, Co^{2+} and Co^{3+}. Write the condensed electron structure of the two cations.

    Co^{2+}: $[Ar]3d^7$; Co^{3+}: $[Ar]3d^6$

15. Thallium was used as a poison in the Agatha Christie mystery story "The Pale Horse." Thallium has two possible cationic forms, +1 and +3. The +1 compounds are the more stable. Write the electron structure of the +1 cation of thallium.

    $1s^22s^22p^63s^23p^63d^{10}4s^24p^65s^24d^{10}5p^66s^24f^{14}5d^{10}$

16. Cobalt–60 and iodine–131 are radioactive isotopes commonly used in nuclear medicine. How many protons, neutrons, and electrons are in atoms of these isotopes? Write the complete electron configuration for each isotope.

    ^{60}Co has 27 protons, 27 electrons, and 33 neutrons: $1s^22s^22p^63s^23p^64s^23d^7$
    ^{131}I has 53 protons, 53 electrons, and 78 neutrons: $1s^22s^22p^63s^23p^63d^{10}4s^24p^65s^24d^{10}5p^5$

17. Write a set of quantum numbers for each of the electrons with an *n* of 3 in a Sc atom.

    n, l, m_l, m_s

    3,0,0,+½ 3,0,0,-½

    3,1,-1,+½ 3,1,-1,-½ 3,1,0,+½ 3,1,0,-½ 3,1,1,+½ 3,1,1,-½

    3,2,0,+½ (other correct sets are possible)

Homework for 6.5 Periodic Variations in Element Properties

1. Based on their positions in the periodic table, predict which has the smallest atomic radius: Mg, Sr, Si, Cl, I.
   Cl
2. Based on their positions in the periodic table, predict which has the largest atomic radius: Li, Rb, N, F, I.

   Rb

3. Based on their positions in the periodic table, predict which has the largest first ionization energy: Mg, Ba, B, O, Te.

   O

4. Based on their positions in the periodic table, predict which has the smallest first ionization energy: Li, Cs, N, F, I.

   Cs

5. Based on their positions in the periodic table, rank the following atoms in order of increasing first ionization energy: F, Li, N, Rb

   Rb < Li < N < F

6. Based on their positions in the periodic table, rank the following atoms in order of increasing first ionization energy: Mg, O, S, Si

   Mg < Si < S < O

7. Atoms of which group in the periodic table have a valence shell electron configuration of ns^2np^3?

   15

   [strictly true for n = 2 and 3 only; n = 4 and 5 have $ns^2(n\text{-}1)d^{10}np^3$, n = 6 and 7 have $ns^2(n\text{-}2)f^{14}(n\text{-}1)d^{10}np^3$. But in these cases, the d and f orbitals are strongly contracted and act like core electrons]

8. Atoms of which group in the periodic table have a valence shell electron configuration of ns^2?

   2

9. Based on their positions in the periodic table, list the following atoms in order of increasing radius: Mg, Ca, Rb, Cs.

   Mg < Ca < Rb < Cs

10. Based on their positions in the periodic table, list the following atoms in order of increasing radius: Sr, Ca, Si, Cl.

   Cl < Si < Ca < Sr

11. Based on their positions in the periodic table, list the following ions in order of increasing radius: K^+, Ca^{2+}, Al^{3+}, Si^{4+}.

   $Si^{4+} < Al^{3+} < Ca^{2+} < K^+$

12. List the following ions in order of increasing radius: Na^+, Mg^{2+}, Br^-, Te^{2-}.

   $Mg^{2+} < Na^+ < Br^- < Te^{2-}$

13. Which atom and/or ion is (are) isoelectronic with Br^+: Se^{2+}, Se, As^-, Kr, Ga^{3+}, Cl^-?

   Se, As^-

14. Which of the following atoms and ions is (are) isoelectronic with S^{2+}: Si^{4+}, Cl^{3+}, Ar, As^{3+}, Si, Al^{3+}?

   Si, Cl^{3+}

15. Compare both the numbers of protons and electrons present in each to rank the following ions in order of increasing radius: As^{3-}, Br^-, K^+, Mg^{2+}.

   $Mg^{2+} < K^+ < Br^- < As^{3-}$

16. Of the five elements Al, Cl, I, Na, Rb, which has the most exothermic reaction involving gain or loss of an electron? Write the equation for the general reaction (E represents an atom.) What name is given to the energy for the reaction? Hint: note the process depicted does *not* correspond to electron affinity

   $E^+_{(g)} + e^- \longrightarrow E_{(g)}$

   Cl *(process described is the inverse of the ionization energy)*

17. Of the five elements Sn, Si, Sb, O, Te, which has the most endothermic reaction involving gain or loss of an electron? Write the equation for the general reaction (E represents an

atom.) What name is given to the energy for the reaction?

$E_{(g)} \longrightarrow E^+_{(g)} + e^-$

O *(first ionization energy)*

18. The ionic radii of the ions S^{2-}, Cl^-, and K^+ are 184, 181, 138 pm respectively. Explain why these ions have different sizes even though they contain the same number of electrons.

    They have different numbers of protons in the nucleus.

19. Which main group (*s* or *p* block) atom would be expected to have the lowest second ionization energy?

    Ra

20. Explain why Al is a member of group 13 rather than group 3?

    Group 13 have in common one *p* electron and either 0 or 10 *d* electrons. Group 3 has in common one *d* electron. Al has zero *d* electrons.