Waves and Calculations (5 points)

1. What type of relationship (direct or inverse) exists for the relationships between wavelength, frequency, and photon energy?

$\lambda \uparrow \nu \downarrow E \downarrow$

$\lambda \downarrow \nu \uparrow E \uparrow$

inverse → direct

2. The laser in an audio CD player uses light with a wavelength of 7.80 x 10² nm. Calculate the frequency of this light.

$c = \lambda \cdot \nu$ $\quad 7.80 \times 10^2 nm \times \frac{1 m}{10^9 nm} = 7.80 \times 10^{-7} m$

$\frac{c}{\lambda} = \nu$ $\quad \frac{3.0 \times 10^8 m/s}{7.80 \times 10^{-7} m} = 3.85 \times 10^{14} s^{-1}$

3. Calculate the following values associated with light and the electromagnetic spectrum:
   a. Calculate the energy of a photon of electromagnetic radiation whose frequency is 6.75 x 10¹² Hz.

   $\Delta E_{photon} = h \cdot \nu$ $\quad 6.75 \times 10^{12} s^{-1} \cdot 6.626 \times 10^{-34} J \cdot s = 4.47 \times 10^{-21} J$

   b. Calculate the energy of a photon of radiation whose wavelength is 322 nm.

   ① $322 nm \times \frac{1 m}{10^9 nm} = 3.22 \times 10^{-7} m$ ② $\nu = \frac{c}{\lambda} = \frac{3.0 \times 10^8 m/s}{3.22 \times 10^{-7} m} = 9.32 \times 10^{14} s^{-1}$ ③ $E = 6.626 \times 10^{-34}$ ... 6.17×10^{-19}

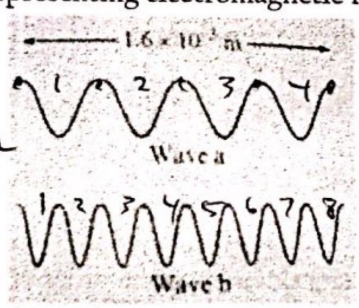
   c. What wavelength of radiation has photons of energy 2.87 x 10⁻¹⁸ J?

   $E = h \times \nu$, $\lambda = \frac{c}{\nu}$

   $\frac{2.87 \times 10^{-18} J}{6.626 \times 10^{-34} J \cdot s} = \frac{6.626 \times 10^{-34} J \cdot s \cdot \nu}{6.626 \times 10^{-34} J \cdot s}$ $\quad \nu = 4.33 \times 10^{15} s^{-1}$

   $\lambda = \frac{3.0 \times 10^8 m/s}{4.33 \times 10^{15}}$ ... 6.92×10...

4. Consider the following waves representing electromagnetic radiation:



... 2x wavelength as b, ½ energy & frequency

$\frac{1.6 \times 10^{-3} m}{4} = 4.0 \times 10^{-4} m$

$\frac{1.6 \times 10^{-3} m}{8} = 2.0 \times 10^{-4} m$

a) Which wave has the longer wavelength? Calculate this wavelength.

wave a

b) Which wave has the higher frequency and photon energy? Calculate these values.

wave b $\nu = \frac{c}{\lambda}$ $\quad \nu = \frac{3.0 \times 10^8 m/s}{4.0 \times 10^{-4} m} = 7.5 \times 10^{11} s^{-1} \cdot 6.6$... $= 5.0 \times 10$...

c) Which wave has the greater speed (velocity)?

both are moving at same speed $3.0 \times 10^8 m/s$

d) Which type of electromagnetic radiation does each wave represent?

infared

wave B $\quad \nu = \frac{3.0 \times 10^8 m/s}{2.0 \times 10^{-4} m} = 1.5 \times 10^{12} s^{-1} \cdot 6.626 \times 10^{-34} J \cdot s$

$9.9 \times 10^{-22} J$

Light and Atomic Emission

Figure 7.7 see page 310

5. Using vertical lines, indicate the following energy transitions for the hydrogen atom on an energy-level diagram:

   a. n = 3 → n = 2 — 5, 4, 3, 2, 1 — green light

   b. n = 4 → n = 2 — 5, 4, 3, 2 — blue light

   c. n = 2 → n = 1 — red? — 2, 1

6. Describe briefly why the study of electromagnetic radiation has been important to our understanding of the arrangement of electrons in atoms.

   different amounts of E R have been emitted from & added to atoms and give us evidence atom's electrons exist at different energy levels with different quanta of energy required

7. An electron is excited from the n = 1 ground state to the n = 3 state in a hydrogen atom. Which of the following statements are true? Correct the false statements to make them true.

   a. It takes ~~more~~ less energy to ionize (completely remove) the electron from n = 3 than from the ground state

      because the e⁻ in the n=3 is further from nucleus

   b. The electron is farther from the nucleus on average in the n = 3 state than in the n = 1 state.

      T

   c. The wavelength of light emitted when the electron returns to the ~~ground state~~ n=1 from n = 3 will be the same as the wavelength of light absorbed to go from n = 1 to n = 3.

   d. For n = ~~3~~ 2, the electron is in the first excited state.

Electromagnetic Radiation

8. Arrange the following kinds of electromagnetic radiation in order of increasing wavelength:
   infrared, green light, red light, radio waves, X-rays, ultraviolet light

   Short → long

   X-rays, UV, Green, Red, infrared radio

carbon single bond found in organic compounds is 347 kJ/mol. Would X rays and/or radio waves be able to disrupt organic compounds by breaking carbon-carbon single bonds?

X-Rays: $\frac{3.0\times10^{8}\ m/s}{1.0\times10^{-10}\ m}\cdot 6.626\times10^{-34}\ J\cdot s = \frac{1.99\times10^{-15}\ J}{1\ photon}\times\frac{6.022\times10^{23}\ photons}{1\ mol}\times\frac{1\ kJ}{1000\ J}$

X Rays | $\frac{1.2\times10^{9}\ kJ}{1\ mol}$

10. The rays of the Sun that cause tanning and burning are in the ultraviolet portion of the UV spectrum. These rays are categorized by wavelength. So-called UV-A radiation has wavelengths in the 320-380 nm range, whereas UV-B radiation has wavelengths in the range of 290-320 nm.

a. Calculate the frequency of light that has a wavelength of 320 nm.

$320\ nm\times\frac{1\ m}{10^{9}\ nm}=3.2\times10^{-7}\ m$ $\nu=\frac{3.0\times10^{8}\ m/s}{3.2\times10^{-7}\ m}=9.4\times10^{14}\ s^{-1}$

b. Calculate the energy of a mole of 320-nm photons.

$9.4\times10^{14}\ s^{-1}\times 6.626\times10^{-34}\ J\cdot s=\frac{6.2\times10^{-19}\ J}{1\ photon}\times\frac{6.022\times10^{23}\ photons}{1\ mol}=374000\ \frac{J}{mol}=374\ \frac{kJ}{mol}$

c. Which has more energy: photons of UV-A radiation or photons of UV-B radiation?

UV-B shorter wavelength, higher frequency, higher energy

d. The UV-B radiation from the Sun is considered a greater cause of sunburn in humans than UV-A radiation. Is this observation consistent with your answer to part (c)?

higher energy would cause more possible damage to molecules in cells

#9

Radio: $\frac{3.0\times10^{8}\ m/s}{1.0\times10^{4}\ m}\cdot 6.626\times10^{-34}\ J\cdot s=1.99\times10^{-29}\ J$

$E=h\cdot\frac{c}{\lambda}$

$\frac{1.99\times10^{-29}\ J}{1\ photon}\times\frac{6.022\times10^{23}\ photons}{1\ mol}\times\frac{1\ kJ}{1000\ J}=$ $1.2\times10^{-8}\ \frac{kJ}{mol}$

X-rays would be able to disrupt 347 kJ/mol bond because the energy is much higher, @ 1.2×10^{9} kJ/mol

radio would not because it is below 347 kJ/mol @ 1.2×10^{-8} kJ/mol

Atomic Theory
Periodic Trends

Name ______________________
Period ____ Date ______________

Periodicity

1. Group the following electron configurations in pairs that would represent similar chemical properties of their atoms:
   a. $1s^2 2s^2 2p^5$
   b. $1s^2 2s^1$
   c. $1s^2 2s^2 2p^6$
   d. $1s^2 2s^2 2p^6 3s^2 3p^5$
   e. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
   f. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

   a, d halogens
   b, e alkali metals
   c, f noble gases

2. Element 106 has been named seaborgium, Sg, in honor of Glenn Seaborg, discoverer of the first transuranium element.
   a. Write the expected electron configuration of element 106.

   [Rn] $5f^{14} 6d^4 7s^2$

   b. What other element would be most like element 106 in its properties?

   W, element above it in same group

3. Answer the following questions concerning nuclear charge:
   a. What is meant by the term *effective nuclear charge*?

   net positive charge experienced by an e^- ex. Mg $12p^+$) $2e^-$) $8e^-$) $2e^-$ $12p^+$ − $10e^-$ = $2p^+$ attract $2e^-$ valence

   b. How does the effective nuclear charge experienced by the valence electrons of an atom vary going from left to right across a period of the periodic table?

   Al $13p^+$) $2e^-$) $8e^-$) $3e^-$ ($10e^-$) $13p^+$ − $10e^-$ = $3+$ attract 3 valence e^-

   effective nuclear charge increases across a period because inner e^- stay same (shielding stays same)

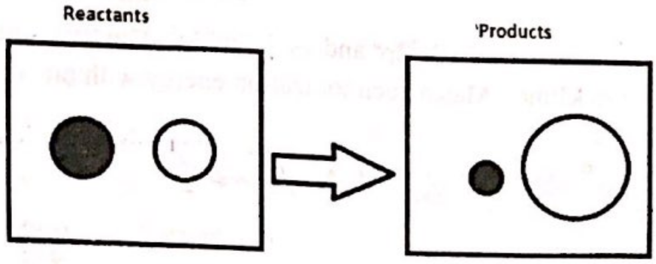
Atomic Radius

4. Using only the periodic table, arrange each set of atoms in order of increasing radius:
   a. Ba, Ca, Na   Na, Ca, Ba
   b. Sn, Sb, As   As, Sb, Sn
   c. Al, Be, Si   Al [illegible]

   (actual Be, Si, Al)

Scanned with CamScanner

5. In the reaction:



which sphere represents a metal and which represents a non-metal? Explain your answer.

black, non-metal, loses e^- to form + ion (smaller)
white, metal, gains e^- to form − ion (larger)

6. For each of the following groups, place the atoms and/or ions in order of decreasing size:

a. Cu, Cu^+, Cu^{2+} large Cu, Cu^+, Cu^{+2} small

b. Ni^{2+}, Pd^{2+}, Pt^{2+} large Pt^{2+}, Pd^{2+}, Ni^{2+} small

c. O, O^-, O^{2-} large O^{2-}, O^{-1}, O small

d. Te^{2-}, I^-, Cs^+, Ba^{2+}, La^{3+} La^{3+}, Ba^{2+}, Cs^{+1}, I^{-1}, Te^{2-} small ← flip!! large

Te^{2-}, I^{-1}, Cs^{+1}, Ba^{2+}, La^{3+}
smallest nucleus (p^+) most e^- ... largest nucleus (p^+) least e^-

7. Explain the following variations in atomic or ionic radii:

a. $I^- > I > I^+$

~~[crossed out]~~ all with same nucleus charge, I^{-1} gains e^- must be biggest

b. $Ca^{2+} > Mg^{2+} > Be^{2+}$

same family/group
Ca^{+2} has most energy levels

c. $Fe > Fe^{2+} > Fe^{3+}$

same nuclear charge Fe^{3+} has lost $3e^-$

Ionization Energies

8. Answer the following questions about ionization energies:

a. Why does Li have a larger first ionization energy than Na? Li $3p^+$)) $2e^-$ $1e^-$ closer to nucleus

~~[crossed out]~~
e^- removed from energy level → Na $11p^+$))) $2e^-$ $8e^-$ $1e^-$
further from nucleus

b. The difference between the third and fourth ionization energies of scandium is much larger than the difference between the third and fourth ionization energies of titanium. Why?

Titanium
[Ar] $4s^2 3d^2$
[Ar] $4s^2 3d^1$
Scandium

taking a 4th e^- away from Scandium removes that e^- from a level closer to nucleus which requires much more energy

9. Two atoms have the electron configurations $1s^2 2s^2 2p^6$ and $1s^2 2s^2 2p^6 3s^1$. The first ionization energy of one is 2080 kJ/mol, and that of the other is 496 kJ/mol. Match each ionization energy with one of the given electron configurations. Justify your choice.

$1s^2 2s^2 2p^6$ = 2080 kJ/mol e⁻ is taken from a lower energy level closer to nucleus (p+)

$1s^2 2s^2 2p^6 3s^1$ = 496 kJ/mol e⁻ is taken from a level further from nucleus (p+)

10. The first four ionization for the elements X and Y are shown below. The units are not kJ/mol.

	X	Y
First	170 ← smaller	200 ← larger
Second	350	400
Third	1800 } jump	3500 } jump
Fourth	2500	5000

Identify the elements X and Y. There may be more than one correct answer, so explain fully.

Y = Be ← larger 1st ionization energy, more energy req. to remove 1st e⁻, closer to nucleus (p+)

X = Mg

Both are group 2 because taking a 3e⁻ from the atom cause a huge jump in energy

11. Explain why a graph of ionization energy versus atomic number (across a row) is not linear. Where are the exceptions? Why are there exceptions?

Electronegativity

12. By referring only to the periodic table, select:

a. the most electronegative element in group 6A. Oxygen

b. the least electronegative element in the group (Al), Si, P.

c. the most electronegative element in the group Ga, P, (Cl), Na.

d. the element in the group K, C, Zn, (F) that is most likely to form an ionic compound with Ba. ← strongest puller of e⁻, highest electronegativity

13. Predict the order of increasing electronegativity in each of the following groups of elements:

a. C, N, O →

b. S, Se, Cl   Se, S, Cl

c. Si, Ge, Sn   Sn, Ge, Si

d. Tl, S, Ge   Tl, Ge, S

Electron Configurations – Applications

Name ______________________

Period ______ Date ______________

Orbital Theory and Periodic Table

1. Answer the following questions about orbitals:
   a. What are the similarities between the 1s and 2s orbitals of the hydrogen atom?

      Same shape

   b. What can you say about the average distance from the nucleus of an electron in the 2s orbital as compared with a 3s orbital?

      2s avg distance is shorter than 3s

   c. For the hydrogen atom, list the following orbitals in order of increasing energy (that is, most stable ones first): 4f, 6s, 3d, 1s, 2p

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

2. State where in the periodic table these elements appear:
   a. elements with the valence-shell electron configuration ns^2np^5 halogens / Gp 7A
   b. the elements that have three unpaired p electrons Group 5A atoms
   c. an element whose valence electrons are $4s^24p^1$ Ga
   d. the d-block elements transition metals

Electron Configurations and Orbital Diagrams

3. Write the condensed electron configurations for the following atoms, using appropriate noble-gas core abbreviations:
   a. Cs [Xe] $6s^1$
   b. Ni [Ar] $4s^2 3d^8$
   c. Se [Ar] $4s^2 3d^{10} 4p^4$
   d. Cd [Kr] $5s^2 4d^{10}$
   e. U [Rn] $6s^2 5f^3$

4. Write the expected last terms of the electron configurations for each of the following atoms:

Cl, Sb, Sr, W, Pb, Cf

Cl $3p^5$ Sr $5s^2$ Pb $6p^2$

Sb $5p^3$ W $5d^4$ Cf $5f^9$

5. Identify the specific element that corresponds to each of the following electron configurations and indicate the number of unpaired electrons for each:

a. $1s^22s^2$ Be no unpaired ↑↓ ↑↓

b. $1s^22s^22p^4$ Oxygen 2 unpaired ↑↓ ↑↓ ↑↓ ↑ ↑

c. $[Ar]4s^23d^4$ Cr 4 unpaired ↑↓ ↑ ↑ ↑ ↑ _

d. $[Kr]5s^24d^{10}5p^4$ Se 2 unpaired ↑↓ ↑ ↑

6. Identify the following elements:

a. An excited state of this element has the electron configuration $1s^22s^22p^53s^1$.

Ne

b. The ground-state electron configuration is $[Ne]3s^23p^4$.

Sulfur

c. An excited state of this element has the electron configuration $[Kr]5s^24d^65p^26s^1$.

47 ≠ 36 + 11e⁻ Ag

d. The ground-state electron configuration contains three unpaired 6p electrons.

$6p^3$ ↑ ↑ ↑ Bi

7. Using the periodic table, write the expected ground-state electron configurations for:

a. the third element in Group 5A.

$[Kr]5s^24d^{10}5p^3$

b. element number 116.

Lv $[Rn]6s^25f^{14}4d^{10}6p^4$.

c. an element with three unpaired 5d electrons.

↑ ↑ ↑ __ __ Tantalum

d. the halogen with electrons in the 6p atomic orbitals.

↑↓ ↑↓ ↑ At