Dougherty Valley HS AP Chemistry
Atomic Structure – Electron Configuration
Periodicity Study Questions

WORKSHEET #4

Name: **Date:** **Period:** **Seat #:**

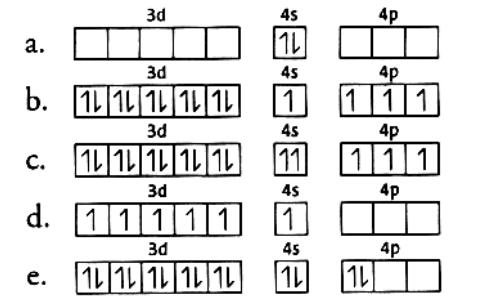
"No human investigation can be called real science if it cannot be demonstrated mathematically."
Leonardo da Vinci (1452 – 1519)

1. Write the electron configurations of the following elements using the shorthand notation for the noble gas cores.	
a.	phosphorus
b.	nickel
c.	osmium
d.	californium
e.	titanium

2. Which orbital is filled following these orbitals?	
a.	3d
b.	4s
c.	5p
d.	5f

3. How many electrons can be accommodated in	
a.	a d subshell
b.	a set of f orbitals
c.	the n = 4 shell
d.	the 7s orbital
e.	a p_x orbital?

4. What is wrong with the following ground state electron configurations?



a.	
b.	
c.	
d.	
e.	

5. How many unpaired electrons are there in	
a.	a nitrogen atom
b.	an iodine atom
c.	a nickel (II) cation
d.	an oxide ion?

6. Which of the following sets of quantum numbers describe an impossible situation? Explain why. **SKIP**	
a.	$n = 2, l = 1, m_l = 2, m_s = +½$
b.	$n = 5, l = 2, m_l = 1, m_s = –½$
c.	$n = 6, l = 5, m_l = 0, m_s = 0$
d.	$n = 3, l = 3, m_l = 1, m_s = –½$
e.	$n = 4, l = 2, m_l = 1, m_s = +½$

7. Arrange the elements S, Ge, P, and Si in order of increasing atomic size.		<		<		<	

8.	Arrange the ions Na^+, K^+, Cl^-, and Br^- in order of increasing size.		<		<		<	

9.	Arrange the elements Be, Ca, N, and P in order of increasing ionization energy.		<		<		<	

10. Which one of each of the following pairs would you expect to have the higher electron affinity? Explain why.		
a.	Cl or Cl^-	
b.	Na or K	
c.	Br or I	

12. Which elements fit the following descriptions?		
a.	the smallest alkaline earth metal	
b.	has a valence shell configuration $4f^{14}\ 5d^{10}\ 6s^1$	
c.	the halogen with the lowest ionization energy	
d.	has 13 more electrons than argon	
e.	the smallest non metal	
f.	the Group 4A element with the largest ionization energy	
g.	its 3+ ion has the electron configuration [Kr] $4d^{10}$	

13. Given this series of ionic radii:
C^{4-} 260 pm; N^{3-} 171 pm; O^{2-} 126 pm; F^- 119 pm;
Na^+ 116 pm; Mg^{2+} 86 pm; Al^{3+} 68 pm,
Estimate the atomic radius of neon. Do you think this is a fair estimate?

Dougherty Valley HS AP Chemistry
Atomic Structure – Electron Configuration
Periodicity Study Questions - **ANSWERS**

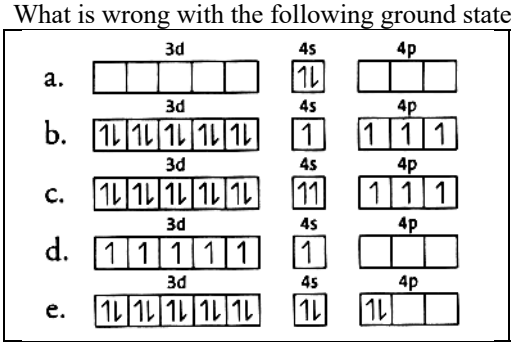
WORKSHEET #4

1. Write the electron configurations of the following elements using the shorthand notation for the noble gas cores.
   a. phosphorus [Ne] $3s^2\ 3p^3$
   b. nickel [Ar] $3d^8\ 4s^2$
   c. osmium [Xe] $4f^{14}\ 5d^6\ 6s^2$
   d. californium [Rn] $5f^{10}\ 7s^2$
   e. titanium [Ar] $3d^2\ 4s^2$

2. Which orbital is filled following these orbitals?
   a. 3d is followed by 4p
   b. 4s is followed by 3d
   c. 5p is followed by 6s
   d. 5f is followed by 6d

3. How many electrons can be accommodated in
   a. a d subshell 10; two in each of 5 orbitals
   b. a set of f orbitals 14; two in each of 7 orbitals
   c. the $n = 4$ shell 32; $2n^2$
   d. the 7s orbital 2
   e. a p_x orbital? 2

4. What is wrong with the following ground state electron configurations?



   a. OK, 4s is filled before 3d.
   b. 4s should be filled before 4p subshell.
   c. 4s cannot have 2 electrons with same spin.
   d. OK, stability of half-filled 3d subshell.
   e. 4p should first fill one orbital at a time.

5. How many unpaired electrons are there in
   a. a nitrogen atom 3 unpaired electrons ($2s^2\ 2p^3$)
   b. an iodine atom 1 unpaired electrons ($5s^2\ 5p^5$)
   c. a nickel (II) cation 2 unpaired electrons ($3d^8$)
   d. an oxide ion no unpaired electrons (closed shell [Ne])

6. Which of the following sets of quantum numbers describe an impossible situation? Explain why.
   a. $n = 2, \ell = 1, m_\ell = 2, m_s = +½$ maximum value for m_ℓ is +l
   b. $n = 5, \ell = 2, m_\ell = 1, m_s = –½$ OK
   c. $n = 6, \ell = 5, m_\ell = 0, m_s = 0$ m_s must be +½ or -½
   d. $n = 3, \ell = 3, m_\ell = 1, m_s = –½$ maximum value for $\ell = n$-1
   e. $n = 4, \ell = 2, m_\ell = 1, m_s = +½$ OK

7. Arrange the elements S, Ge, P, and Si in order of increasing atomic size. S < P < Si < Ge
   Increase from right to left and down from top to bottom.

8. Arrange the ions Na^+, K^+, $Cl^–$, and $Br^–$ in order of increasing size. $Na^+ < K^+ < Cl^– < Br^–$
   K^+ and $Cl^–$ are isoelectronic; the negative ion is the larger of the two (the nuclear charge is greater for the positive ion). $Br^–$ is larger than K^+ and $Cl^–$. Na+ is smaller than K^+ and $Cl^–$.

9. Arrange the elements Be, Ca, N, and P in order of increasing ionization. Ca < Be < P < N
   Elements lying close to a diagonal like Be and P are difficult to predict. Notice that Si, to the left of P, has an ionization energy less than Be.

10. Which one of each of the following pairs would you expect to have the highest electron affinity?
    a. Cl or Cl^-
       Cl; the Cl^- ion has a complete shell and zero electron affinity.
       Cl has room in the 3p shell with 17 protons attracting the electron.
    b. Na or K
       Na; you would expect the electron affinity to increase up a group.
    c. Br or I
       Br; you would expect the electron affinity to increase up a group; an exception is a decrease from Cl to F—attributed to electron-electron affinity in the very small fluoride ion.

11. Which ions would you expect to exist, and which wouldn't you expect to exist?

a.	K^{2+}	no, K+ ≈ [Ar]	e.	P^{3-}	yes
b.	Cl^-	yes	f.	Mn^{7+}	no, charge is far too high
c.	Al^{2+}	no, Al^{3+} ≈ [Ne]	g.	Fe^{2+}	yes, one of the stable ions of Fe
d.	Ar^+	no, [Ar] is stable	h.	Na^-	sodide ion; actually can be made; configuration $3s^2$

12. Which elements fit the following descriptions:

a.	the smallest alkaline earth metal	Be
b.	has a valence shell configuration $4f^{14}\ 5d^{10}\ 6s^1$	Au
c.	the halogen with the lowest ionization energy	At or I
d.	has 13 more electrons than argon	Ga [Ar] $3d^{10}\ 4s^1\ 4p^2$
e.	the smallest non metal (I think He – see the discussion for Ne's atomic radius on question 13!)	H
f.	the Group 4A element with the largest ionization energy	C
g.	its 3+ ion has the electron configuration [Kr] $4d^{10}$	In^{3+}

13. Given this series of ionic radii:
    C^{4-} 260 pm; N^{3-} 171 pm; O^{2-} 126 pm; F^- 119 pm;
    Na^+ 116 pm; Mg^{2+} 86 pm; Al^{3+} 68 pm,
    estimate the atomic radius of neon. Do you think this is a fair estimate?

    All these species are isoelectronic; all have the configuration [He] $2s^2\ 2p^6$. However, the nuclear charge increases along the series and the result is a decrease in the radius as the electrons are pulled closer to the nucleus.

Somewhere between 116 and 119 pm? This is where Ne fits in the isoelectronic series. Ne does not form compounds like the nonmetals in the preceding groups and measuring its atomic radius is therefore difficult. Some data could be obtained from the crystalline structure of solid neon but the bonding in such a solid is quite different from the bonding in elements like carbon or sulfur, or in compounds like sodium fluoride. There is wide variation in stated values for the atomic radius of neon—anywhere from 70 pm to 112 pm. Putting together a set of self-consistent data is not easy. For example, Na+ itself is listed as having an atomic radius anywhere from 95 and 116 pm. You would expect Ne to be larger than Na^+. F is listed usually at about 72 pm – you might have expected Ne to be smaller. However, in neon the interelectronic repulsion is high; the electrons are very crowded. As a result, neon might well be larger than fluoride or oxygen. A value between 100 and 110 pm seems reasonable