Name ______________________________

Chemistry 101 Friday Discussion Week 12

Electron Configuration & Periodic Trends

Part 1: Electron Configuration

1. Give the ground state long-form electron configuration in full form ($1s^2$, $2s^2$, etc) for each of the following atoms:
    a. Magnesium

    b. Lithium

    c. Oxygen

    d. Sulfur

2. Using the symbol of the previous noble gas to indicate the core electrons, write the electron configuration for each of the following elements.

    a. Arsenic

    b. Titanium

    c. Strontium

    d. Chlorine

3. Which element does each of the following electron configurations correspond to?
    a. $1s^22s^2$

    b. $1s^22s^22p^63s^23p^2$

    c. $[Ar]4s^23d^{10}4p^5$

    d. $[Kr]5s^24d^{10}5p^2$

4. Give the electron configuration of the following ions. You may use the noble gas core form, but do it in a way that shows all valence electrons (for example, K^+ would be $[Ne]3s^23p^6$ instead of [Ar]. Then, identify which noble gas has the same electron configuration.

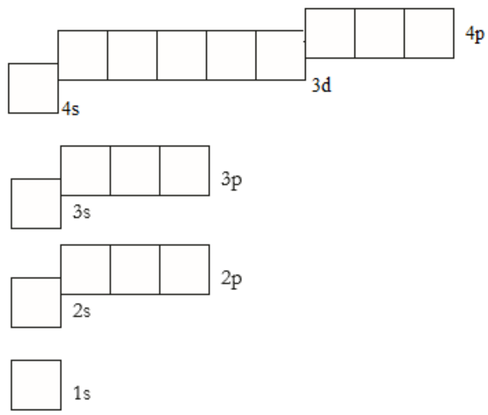
Letter	Ion	Electron Configuration	Noble Gas with Same Electron Configuration
A	Br^-		
B	Ca^{+2}		
C	Al^{+3}		
D	Te^{2-}		
E	The stable ion formed by an atom of strontium		
F	The stable ion formed by an atom of phosphorus		

5. Where are valence electrons found in an atom, and why are these particular electrons most important to the chemical properties of an atom?

6. How are valence electrons represented in an electron configuration? Give an example.

Part 3: Electron Filling Diagrams

7. Using the electron filling diagrams below, show the electron filling for iron and give the number of unpaired electrons:



8. An electron configuration for an excited state atom is $1s^2 2s^2 2p^3 3p^1$. Identify the atom and write the ground state configuration.

Part 4: Periodic Trends

9. Moving across the periodic table, ionization energy tends to increase. Moving down the periodic table, ionization energy tends to decrease. Explain this phenomena based on principles of energy levels (shielding) and the pull of the protons on valence electrons (effective nuclear charge).

10. Moving across the periodic table, atomic radius tends to decrease. Moving down the periodic table, atomic radius tends to increase. Explain this phenomena based on principles of energy levels and valence electrons.

11. Arrange the elements in each of the following sets in order from easiest to lose an electron to hardest to lose an electron.

    a. C, Li, O

    b. Fr, Rb, Na

    c. Ca, Mg, Ba

    d. Br, At, F

12. Arrange the following elements from smallest atomic radius to largest atomic radius.

    a. I, F, Cl

    b. Te, Xe, In

    c. S, Al, Ar, Na

13. Arrange the atoms or ions in each set below in order of increasing size:

    a. Ca^{+2}, Ca, Ca^{+}

    b. S^{-2}, S, S^{-}

    c. Justify your answer to part b below:

14. Each set of atoms and ions below forms an isoelectronic series. This means that although different elements are present, each element has the same number of electrons. Give the numbers of protons and electrons in each atom or ion in the following isoelectronic series.

    a. Ne, Na^{+}, F^{-}

    b. P^{-3}, S^{-2}, Cl^{-}

15. Arrange the atoms or ions in each set below in order of increasing size. Then, make a separate list ranking them in order of increasing first ionization energy.

    a. Ne, Na^{+}, F^{-}

    b. P^{-3}, S^{-2}, Cl^{-}

Justify your answer to part a) below, explaining both trends in radius as well as in ionization energy: