Electron Configuration and Periodic Trends HW

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2. spin +1/2 When filling p orbitals (or any equal energy orbitals) electrons with the same spin occupy each of the orbitals before electrons of opposite spin fill any orbital.

3. a. Oxygen: If the atom has eight electrons and it has no charge, it must have eight protons, which makes it oxygen.

   b. Each of the 2p orbitals has one electron with spin +1/2. The fourth p electron has opposite spin, filling one of the p orbitals.

   c. The Pauli exclusion principle states that only two electrons of opposite spin can occupy any given orbital. In this orbital diagram, no orbital has more than two electrons and the electrons that do occupy the same orbital are of opposite spin.

4. a. In neon, the first and second energy levels are filled: $1s^2 2s^2 2p^6$
   The next energy level to be filled would be n = 3. Thus, the electrons following the [Ne] would have to be 3s electrons, not 4s.

   b. Krypton ends the fourth period so electron configurations that begin with [Kr] are in the fifth period. In the fifth period, the 4d orbitals begin to fill before the 5p orbitals. Therefore $5p^6$ cannot follow $5s^2$.

   c. Bromine does not have all filled orbitals in the outer energy level so it cannot be used as a base for any electron configuration.

7. a. The s^1 puts the elements in Group 1, so the elements would be H, Li, Na, K, Rb, Cs, and Fr.

   b. The p^6 puts the elements in Group 18 and the absence of any d electrons in groups above Group 2 eliminates elements in Period 4 and above, so the elements would be Ne and Ar.

   c. The p^4 puts the elements in Group 16 and the presence of 10 d electrons in the $n - 1$ period places the elements in Period 4 and above, so the elements would be Se, Te, and Po.

10. a. Be, Ca, Ba: Atomic radii increase going down a group (more energy levels).

b. Cl, Al, Na: Atomic radii increase going right to left across a period (less protons in the nucleus therefore less nuclear charge).

c. Sn, Li, Rb: Rb is larger than Sn because it has less nuclear charge than Sn, Rb is larger than Li because it has more energy levels. Need to calculate Zeff to predict the order of Sn and Li

11. Z represents the actual charge of the nucleus. In atoms of elements with two or more filled inner shells, the negative charges of the electrons slightly cancel the effect on the positive charges of the nucleus on the outer electrons. Zeff is the effective charge of the nucleus on the outer electrons or the charge that the nucleus appears to have with respect to the outer electrons.

12. a. The first ionization energy is the amount of energy needed to remove an electron from a neutral atom. If the first ionization energy is low, that means that it is easy to remove the electron and thus the atom's chemical reactivity is high.

b. The first ionization energy tends to increase across a period from left to right. As you go across a period, the charge on the nucleus becomes greater and the outer electrons are closer to the nucleus, thus it takes more energy to remove an electron. The first ionization energy tends to decrease down a group. The nuclear charge is increasing down a group and the outer energy level is getting farther away from the nucleus. The force of attraction varies inversely as the square of the distance between the charges. The distance factor has a greater effect and as the radius gets larger down the group ionization energy decreases.

c. As the radius of an atom becomes larger, the ionization energy decreases.

13. Atoms are most stable when orbitals are half-filled or completely filled. For B ([He]$2s^2 2p^1$) and Al ([Ne]$3s^2 3p^1$), removal of the only p orbital electron in each will produce more stable filled-orbital electron configurations for these atoms. For O ([He]$2s^2 2p^4$) and S ([Ne]$3s^2 3p^4$), removal of one p orbital electron in each will produce more stable half-filled orbital electron configurations for these atoms. Therefore, less energy is needed to remove these electrons.