Chemistry 103, Dr. Hamers
Help Sheet #9, Chapter 7 Concepts
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ELECTRON CONFIGURATION

Electron configurations: $1s^2 2s^2 2p^6$ … etc.

- ***Valence Electrons***: Valence electrons are the s and p electrons just past the last noble gas.
- ***Core Electrons***: Electrons that are not the valence electrons; the noble gas electrons when using the shorthand notation
- ***Shorthand Notation*** with noble gases
- ***Hund's Rule***: The lowest energy configuration (ground state configuration) is the one with the greatest spin
- ***Charged Species*** (e.g., Cl^-): add or remove electrons from last subshell filled
- ***Transition Metals***: remove s electrons first when ionizing
- ***Half-filled and Filled Subshells***: More stable than partially-filled subshells
- ***Exceptions***: Cr: $[Ar]4s^1 3d^5$ and Cu: $[Ar]4s^1 3d^{10}$ families
- ***Diamagnetic:*** All electrons paired; not magnetic – not attracted to a magnet
- ***Paramagnetic***: At least 1 unpaired electron; magnetic – attracted to a magnet
- ***Isoelectronic***: Having the same electron configuration; having the same number of electrons
- ***Number of unpaired electrons***

Aufbau Diagram (Figure 7.2, page 307): electrons fill up the orbitals in a particular order:

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14}$

the superscripts represent the number of electrons within that subshell

Better to use Periodic Table to determine electron configuration!

BE ABLE TO WRITE ELECTRON CONFIGURATIONS FROM THE PERIODIC TABLE.

Shielding: Shielding is when inner electrons "block" or "shield" the nuclear charge from outer electrons

Penetration: How close an electron is to the nucleus; as penetration increases, the nuclear charge the electron experiences increases; Penetration: s > p > d > f in the same energy level (i.e., 4s penetrates more than 4p)

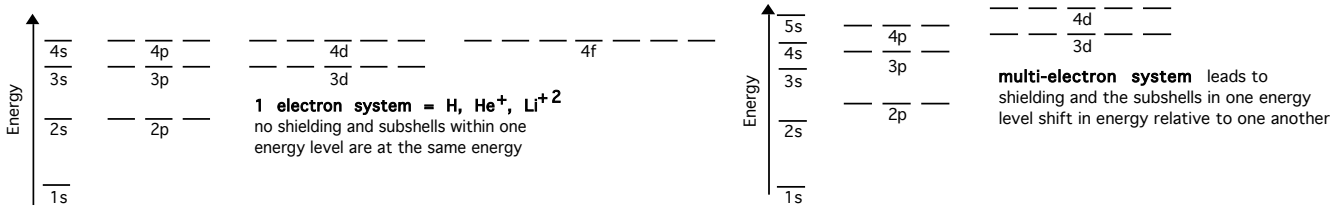
Effective Nuclear Charge: Z_{eff} or Z^*: The nuclear charge the electron experiences taking into account the shielding from other electrons (which lowers the Z^*) and the penetration of the electron (which raises the Z^*); the Z^* would therefore be an average nuclear charge that the electron experiences;

penetration↑ ⇒ Z^*↑; shielding↑ ⇒ Z^*↓

whichever subshell fills before another subshell, the one filling first typically will penetrate further and have a greater effective nuclear charge (Z_{eff});

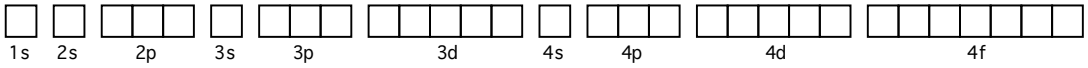
$Z_{eff}(2s) > Z_{eff}(2p)$ because the 2s electrons fill first and penetrate more

ORBITAL ENERGY DIAGRAM:



Energy difference between subshells gets smaller as n increases

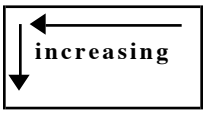
ORBITAL BOX DIAGRM:



(Note: The boxes have been placed in order of n value, **not** the order of filling!)

PERIODIC TABLE TRENDS: Organized by chemical reactivity of elements which is based on the valence e^- .

- Atomic Radii: Size of the atoms; down: size of quantum number n; left: effective nuclear charge



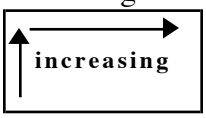
Increases down and to the left on Periodic Table.

- First Ionization Energy: The energy required to remove the <u>first</u> electron from a gaseous atom.

  Removing the 1st e^- (1^{st} ionization energy) $X(g) \rightarrow X^+(g) + e^-$

  Removing the 2nd e^- (2^{nd} ionization energy) $X^+(g) \rightarrow X^{+2}(g) + e^-$

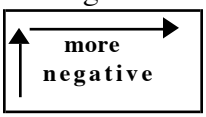
  Removing the 3rd e^- (3^{rd} ionization energy) $X^{+2}(g) \rightarrow X^{+3}(g) + e^-$



Generally increases up and to the right on Periodic Table.
Exceptions: occur at filled or half-filled subshells; Be and N columns

- Electron Affinity: The energy change associated with the addition of an electron to a neutral gaseous atom to form a negatively charged gaseous atom.

  Adding an e^- to a neutral atom $X(g) + e^- \rightarrow X^-(g)$ ΔE



Generally more negative up and to the right on the Periodic Table (more negative is more favorable). Negative value → energy given off. Trend has lots of exceptions. Cl has highest EA.

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1. Draw the d_{yz} orbital. Draw in where the nodal surfaces lie.

2. Write the ground state electron configuration (no abbreviations) for
a. Si b. Cl c. Fe d. Kr e. Ge f. Sb g. K^+ h. Si^{+3} i. S^{-2} j. Cu k. Cr

3. a. What are the valence electrons for P? b. What are the core electrons for Cl?

4. Which atom or ion has the following electron configuration, $1s^22s^22p^63s^23p^63d^5$?
a. Mn b. V c. Ti^{+2} d. Mn^{+2} e. Co^{-2}

5. Which of the following atoms has the largest number of valence electrons?
a. P b. Al c. Sr d. Sb e. Te

6. Which of the following is the correct electron configuration for strontium?
a. $[Kr]3s^2$ b. $[Kr]4s^2$ c. $[Kr]5s^2$ d. $[Kr]4d^2$ e. $[Kr]4f^2$

7. For any atom other than the hydrogen atom, which orbital is lowest in energy?
a. 2s b. 4f c. 4d d. 2p e. 3d

8. In the ground state of K, which subshell would have the smallest effective nuclear charge?
a. 1s b. 2s c. 2p d. 3s e. 3p

9. The second ionization energy for any element is
a. always about the same as the first ionization energy.
b. larger or smaller than the first ionization energy depending on the element.
c. always smaller than the first ionization energy.
d. predicted using Hund's Rule.
e. always larger than the first ionization energy.

10. In each set, select those elements/ions/compounds which are isoelectronic?
a. Cl^-, Ar, F^-, Br^-, Ca^{+1} b. CN^-, Si, S^{-2}, Mg^{-2}, Ca^{+2}

11. Which atom is the **most** paramagnetic?
a. C b. Ti c. Cu d. N e. Ar

12. Which atom has the **most** unpaired electrons in the ground state? Cr, Ni, Al, P, O

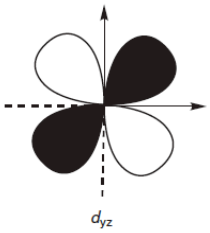
13. a. Write the electron configuration for Mn. Do not use abbreviations.
b. How many unpaired electrons does Mn have?
c. How many unpaired electrons does Mn^{+3} have?

14. I. Which of the following elements will have the highest second ionization energy?
a. Be b. Li c. K d. Mg e. Na
II. Which of the following atoms has the lowest second ionization energy?
a. Mg b. Na c. Sr d. Li e. Rb

15. On the basis of periodic trends rank the elements Al, Rb, and Ca in order of their increasing atomic radii.
a. Rb < Ca < Al b. Ca < Rb < Al c. Al < Ca < Rb d. Rb < Al < Ca e. Al < Rb < Ca

16. Which of the following will release the most energy in the process: $E + e^- \rightarrow E^-$?
a. I b. B c. Cl d. Cs e. Zn

ANSWERS

1. 

2. a. $1s^22s^22p^63s^23p^2$ b. $1s^22s^22p^63s^23p^5$ c. $1s^22s^22p^63s^23p^64s^23d^{16}$
d. $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$ e. $1s^22s^22p^63s^23p^64s^23d^{10}4p^2$ f. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^3$
g. $1s^22s^22p^63s^23p^6$ h. $1s^22s^22p^63s^1$ i. $1s^22s^22p^63s^23p^6$ j. $1s^22s^22p^63s^23p^64s^13d^{10}$
k. $1s^22s^22p^63s^23p^64s^13d^5$

3. a. $3s^23p^3$ b. $1s^22s^22p^6$

4. d {With charged transition metals, remove the last s electrons first.}

5. e 6. c 7. a 8. e 9. e 10. a. Cl^-, Ar, Br^- b. CN^-, Si,, Mg^{-2}

11. d {C: 2; Ti: 2; Cu: 1; N: 3; Ar: 0}

12. Cr {Cr has 4 unpaired 3d e^-; Ni, 2 unpaired 3d e^-; Al, 1 unpaired 3p e^-; P, 3 unpaired 3p e^-; O, 2 unpaired 2p e^-}

13. a. $1s^22s^22p^63s^23p^64s^23d^5$ b. 5 c. 4 14. I. b II. c 15. c 16. c