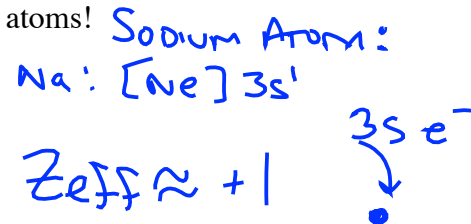


**Chem E-1a**  
**Friday Review Notes**  
**Chapter 8: The Periodic Table**

**Electron Configurations of Ions:**

- Start with the ground state electron configuration for the neutral atom, and then add on electron(s) for negative ions or remove the outermost electron(s) for positive ions.
  - For anions (negatively charged ions), add additional electrons in the normal filling order.
  - For cations (positively charged ions), remove electrons as follows:
    - Remove electrons from the subshell with the highest value of  $n$  first.
    - If there are two subshells both with the highest value of  $n$ , remove electrons from the subshell with the highest  $l$  value first.
    - These rules mean that, for instance, 4s electrons are removed before 3d electrons.
- Note that electrons are removed in forming cations in a different order than they are filled when you are writing electron configurations of neutral atoms!



**Periodic Properties:**

**General Considerations:**

- The general trends for periodic properties can typically be explained by considering  $n$ , the principal quantum number, and  $Z_{\text{eff}}$ , the effective nuclear charge.

**Principal Quantum Number,  $n$ :**

- Recall the  $n$  indicates the general size of an orbital. In other words, it roughly indicates how far away from the nucleus an electron will typically be found.
- $n$  increases as you move down a group on the periodic table.
- $n$  stays constant as you move across a period on the periodic table.

**Effective Nuclear Charge,  $Z_{\text{eff}}$ :**

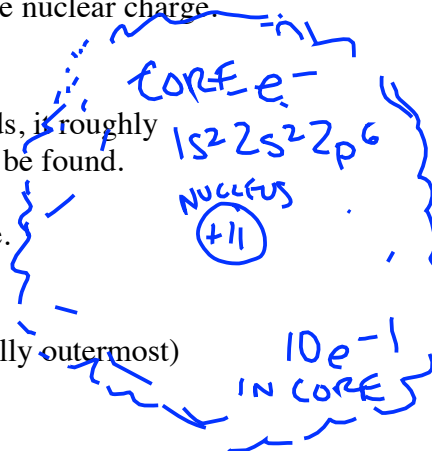
- $Z_{\text{eff}}$  represents the nuclear charge that is felt by a particular (usually outermost) electron in an atom.
- $Z_{\text{eff}}$  can be calculated as follows:

$$Z_{\text{eff}} = Z - \sigma$$

where:  $Z$  = nuclear charge = atomic number

$\sigma$  = shielding constant = approximately the number of core (non-valence) electrons.

- $Z_{\text{eff}}$  increases as you move left to right across a period because the number of core electrons, and thus the shielding constant, stays essentially the same, but the nuclear charge increases.
- You might expect that  $Z_{\text{eff}}$  remains constant as you move down a group on the periodic table. In actuality it increases, but the effect of this increase in  $Z_{\text{eff}}$  moving down a group is outweighed by the increase in  $n$  as you move down a group.



**Atomic Radius:**

- Atomic Radius decreases as you move left to right across a period because  $n$  stays constant but  $Z_{eff}$  increases so the electrons are held tighter to the nucleus, and thus radius decreases.
- Atomic Radius increases as you move down a group on the periodic table because  $n$  increases so the outermost electrons are in larger orbitals that are farther from the nucleus.

**Ionic Radius:**

- Anions are bigger than their corresponding neutral atoms because they have more electrons but the same nuclear charge.
- Cations are smaller than their corresponding neutral atoms because they have fewer electrons but the same nuclear charge.
- An isoelectronic series is several atoms or ions with the same number of electrons and the same electron configurations. For any atoms and ions in an isoelectronic series, the atom or ion with the greatest nuclear charge will be smallest, and the one with the smallest nuclear charge will be the largest. Since they have the same number of electrons, their size is just dependent on the number of protons in the nucleus.

**Ionization Energy:**

- Ionization Energy is defined as the energy required to remove the outermost electron from a gas-phase atom or ion in the ground state.
- The First Ionization Energy ( $IE_1$ ) is the energy required to remove an electron from a neutral atom to form a +1 ion:  
$$IE_1: A(g) \rightarrow A^+(g) + e^- \quad \Delta H = IE_1$$
- The Second Ionization Energy ( $IE_2$ ) is the energy required to remove an electron from a +1 ion to form a +2 ion:  
$$IE_2: A^+(g) \rightarrow A^{2+}(g) + e^- \quad \Delta H = IE_2$$
- The Third Ionization Energy ( $IE_3$ ) is the energy ... etc....
- Ionization energy increases as you move right across a period because  $Z_{eff}$  is increasing while  $n$  stays constant, and the outermost electrons are thus held tighter to the nucleus.
- Ionization energy decreases as you move down a group because  $n$  increases and the outermost electrons are farther from the nucleus and are not held as tightly.

### Electron Affinity:

- Electron Affinity is defined as the energy change when an electron is added to a gas phase atom or ion in the ground state.
- The First Electron Affinity ( $EA_1$ ) is the energy change when an electron is added to a neutral atom to form a  $-1$  ion:  
$$EA_1: A(g) + e^- \rightarrow A^-(g) \quad \Delta H = EA_1$$
- The Second Electron Affinity ( $EA_2$ ) is the energy change when an electron is added to a  $-1$  ion to form a  $-2$  ion:  
$$EA_2: A^-(g) + e^- \rightarrow A^{2-}(g) \quad \Delta H = EA_2$$
- The Third Electron Affinity ( $EA_3$ ) is the energy change ... etc ....
- Electron Affinity tends to increase as  $Z_{eff}$  increases (moving right across a period), but there are many exceptions to this pattern.
- There is not a very well-defined periodic trend in electron affinities when moving down a group.

### Exceptions to the Periodic Trends:

- Exceptions to the periodic trends typically occur due to electron configurations that are particularly stable. Forming one of these more stable configurations can be favorable, and disrupting one of these more stable configurations can be energetically costly.
- Noble gas electron configurations ( $s^2p^6$ ) are very stable.
- Filled subshells are very stable.
- Half-filled subshells are also somewhat stable.

### Explaining Periodic Trends:

(i.e. How to explain those pesky “observations” – some general guidelines.)

- If the observation follows the standard periodic trend, then the explanation for the observation is simply the explanation for the trend. You cannot simply cite the trend, you must explain why the trend occurs. This may involve discussions of things such as  $Z_{eff}$  and  $n$ . In general, you shouldn't explain something that follows the general periodic trend by citing a specific electron configuration. The electron configurations can contribute to the observation (and are not incorrect), but the explanation for the trend is usually necessary.
- If the observation goes against the standard periodic trend, then you should comment that it violates the trend and explain why. This is where electron configurations will typically come up—most of the standard deviations from the periodic trends can be explained through electron configurations and the stability of noble gas configurations, filled subshells, and half-filled subshells.
- These notes are just some rough guidelines! You sometimes may need to use your judgment to decide how best to explain a particular observation!