Unit 5.2 Periodic Trends

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Objectives

- Trends in
  - Electron configuration
  - Atomic Radii
  - Electronegativity
  - Ionization Energy
  - Electron Affinity

Periodic Table

e- configuration from the periodic table
Periodic Table and Subshells

Periodic Table as a Map

Metals, Nonmetals and Metalloids

- Metals are on the bottom left side of the table and represent the vast majority of the total number of elements.
- Nonmetals are on the top right side of the table.
- Metalloids are along the stair case, except aluminum, which is a metal.

- The periodic table can be classified by the behavior of their electrons.
Atomic Radius

- The atomic radius of an element is half of the distance between the centers of two atoms of that element that are just touching each other. Generally, the atomic radius decreases across a period from left to right and increases down a given group. The atoms with the largest atomic radii are located in Group I and at the bottom of groups.

Electronegativity

- Electronegativity is a measure of the attraction of an atom for the electrons in a chemical bond. The higher the electronegativity of an atom, the greater its attraction for bonding electrons. Electronegativity is related to ionization energy (IE) and electron affinity.

\[ \text{electronegativity} = \frac{1}{2} (\text{IE} + \text{E.A}) \]

- Electrons with low ionization energies have low electronegativities because their nuclei do not exert a strong attractive force on electrons. Elements with high ionization energies have high electronegativities due to the strong pull exerted on electrons by the nuclei. In a group, the electronegativity decreases as atomic number increases, as a result of increased distance between the valence electron and nucleus (greater atomic radius). An example of an electropositive (i.e., low electronegativity) element is cesium; an example of a highly electronegative element is fluorine.

Ionization Energy

- The ionization energy, or ionization potential, is the energy required to completely remove an electron from a gaseous atom or ion. The closer and more tightly bound an electron is to the nucleus, the more difficult it will be to remove, and the higher its ionization energy will be. The first ionization energy is the energy required to remove one electron from the parent atom. The second ionization energy is the energy required to remove a second valence electron from the univalent ion to form the divalent ion, and so on. Successive ionization energies increase. The second ionization energy is always greater than the first ionization energy. Ionization energies increase moving from left to right across a period (decreasing atomic radius). Ionization energy decreases moving down a group (increasing atomic radius). Group I elements have low ionization energies because the loss of an electron forms a stable octet.
Electron Affinity

- Electron affinity reflects the ability of an atom to accept an electron. It is the energy change that occurs when an electron is added to a gaseous atom. Atoms with stronger effective nuclear charge have greater electron affinity. Some generalizations can be made about the electron affinities of certain groups in the periodic table. The Group IIA elements, the alkaline earths, have low electron affinity values. These elements are relatively stable because they have filled s subshells. Group VIIA elements, the halogens, have high electron affinities because the addition of an electron to an atom results in a completely filled shell. Group VIII elements, noble gases, have electron affinities near zero, since each atom possesses a stable octet and will not accept an electron readily. Elements of other groups have low electron affinities.

1. Electronegativity Trend

Electronegativity Values of Selected Elements

<table>
<thead>
<tr>
<th>Metallic Elements</th>
<th>Nonmetallic Elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li (1.0) Be (1.5)</td>
<td>B (2.0) C (2.5) O (3.5) F (4.0)</td>
</tr>
<tr>
<td>Na (1.0) Mg (1.3) Al (1.5)</td>
<td>P (2.1) S (2.5) Cl (3.0)</td>
</tr>
<tr>
<td>K (1.9) Ca (1.0) Sr (1.5)</td>
<td>Ba (2.4) Br (3.0)</td>
</tr>
</tbody>
</table>

- Metals < 1.6
- Nonmetals > 2.0

http://library.kcc.hawaii.edu/external/chemistry/electronegativity.html

2. Trend in Atomic Radius

Relative Atomic Sizes of the Representative Elements

- The atomic radius decreases from left to right across a period due to the increase in nuclear charge.
- The atomic radius increases from top to bottom in a group due to the increase in atomic number.

(Images and diagrams of atomic size trends)
3. Trend in Ionization Potential

Ionization potential:
The energy required to remove the valence (outermost) electron from an atomic species. Largest toward NE corner of PT since these atoms hold on to their valence e- the tightest.

4. Trend in Electron Affinity

Electron Affinity:
The energy release when an electron is added to an atom. Most favorable toward NE corner of PT since these atoms have a great affinity for e-.

Trends in Groups: Shells

- The trends down a group can best be explained by the increasing number of shells of electrons.
- Each element in a group has the same number of valence electrons, but they are always in shells that have increasing n values and are further from the nucleus.
Core Charge

- The trends in periods from left to right can best be explained by
  - the increasing core charge

Core charge = # protons - inner shell electrons

<table>
<thead>
<tr>
<th>Group</th>
<th>1</th>
<th>2</th>
<th>13</th>
<th>14</th>
<th>15</th>
<th>16</th>
<th>17</th>
<th>18</th>
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<td>Pb</td>
<td>Bi</td>
<td>Po</td>
<td>At</td>
<td>Rn</td>
</tr>
</tbody>
</table>

Core charge = atomic number - # core electrons

Summary of Trend

1. Electronegativity: Largest toward NE corner of PT
2. Atomic Radius: Largest toward SW corner of PT
3. Ionization energy: Largest toward NE corner of PT