Chemical Periodicity

**Goals: To gain an understanding of:**

1. Electron configurations
2. Periodicity.

The periodic law states that when the elements are arranged according to increasing atomic number there is a periodic pattern in their physical and chemical properties.

Although the periodic table is arranged by increasing atomic number (which indicates the number of protons), the electrons configuration is what really determines the physical and chemical properties of the elements. The periodic table can be divided into four groups based on electron configuration:

- The Noble gases (Group 0) - have their outermost s and p orbitals filled which creates a stable and non-reactive (inert) element.
- The representative elements - Group A elements - have their s and p orbitals being filled. These include:
  - Group 1A - Li, Na, K etc. - all very reactive with one electron in the outer s orbital
  - Group 2A - Be, Mg, Ca etc. - all quite reactive with 2 electrons filling their outer s orbital
  - Group 3A - Aluminum group - 3 electrons in outer energy level (2s and 1p) - properties vary from metallic to metalloid
  - Group 4A - Carbon group - 4 electrons in outer energy level (2s and 2p) - properties vary from nonmetallic to metalloid to metallic down the group
  - Group 5A - Nitrogen group - 5 electrons in outer energy level (2s and 3p) - properties vary from nonmetallic to metalloid to metallic
  - Group 6A - Oxygen group - 6 electrons in outer energy level (2s and 4p) - properties vary from nonmetallic to metalloid
  - Group 7A - Halogens - all have 7 electrons in the outer energy level (2s and 5p) - properties vary from nonmetallic to metalloid. Very reactive due to the outer energy level being almost filled.
- The transition metals - elements whose d orbitals are being filled - found in the "d-block." These are also called the Group B elements
- The Inner transition metals - These are the Lanthanide and Actinide series, element whose f orbitals are being filled.

The s, p, d, and f groups can be identified on the diagram below. The f block (inner transition metals) is usually shown separated and below the rest of the table.

Periodicity is the property of having periodic properties. The periodic table shows periodicity in the following properties:

- Atomic size - atoms of elements tend to increase as you go down a group (due to a greater number of energy levels) and atoms of elements tend to decrease in size across a period (greater positive nuclear charge which draws in electrons -energy levels are constant across a period).
- Ionization energy - the energy required to remove an electron from the gaseous state of the atom. Ionization energy decreases as you go down a group due to the outer electrons being further from the positive charge of the nucleus and being shielded from the nucleus' positive charge by the inner energy levels. Ionization energy increases as you move across a period. This is due to the increase in nuclear charge without the increase in number of energy levels.
- Electron affinity - this is the energy change associated with the addition of an electron to a gaseous atom. This trend is not as consistent as the others, but in general electron affinity decreases down a group and increases across a period.
- Ionic size - The size of ions increases as you go down a group, and decreases as you move across a period for the metals and for the nonmetals, for the same reasons as atomic size. However, the metallic ions are positive (have lost electrons, which make up the space or size of the atom) and are much smaller than the negative nonmetallic ions (have gained electrons which create the volume of atoms or ions).

- Electronegativity - The tendency of an atom to gain an electron(s) when combining with another element. Electronegativity decreases down a group (due to shielding) and increases as you move across a period (due to the increase in nuclear charge).

The three major forces affecting periodicity are:

- Nuclear charge - the greater the number of protons in the nucleus, the greater the positive charge and the stronger the electrons are held.
- Shielding - the effect of inner energy levels reducing the strength of the nuclear charge on the electrons in the outer energy levels.
- Electron configuration - atoms are most stable when their outer orbitals are filled (especially the s and p orbitals). This causes the Noble gases to be inert.

The representative groups of elements:

- Noble gases - Group 0, helium, neon, argon, krypton, xenon and radon
  - inert (unreactive) because of stable electron configuration (filled s and p orbitals)
  - helium is used in weather balloons
  - helium and neon are used to create artificial, unreactive environments (less soluble than nitrogen and therefore less likely to cause the bends
  - other noble gases are used to create unreactive environments in flashbulbs or aluminum welding
- Alkali metals - Group 1A, lithium, sodium, potassium, rubidium, cesium and francium
  - very reactive (one electron away from a filled s and p orbital)
  - low density
  - low melting point
  - good electrical conductivity
  - react with water to form strong bases (sodium hydroxide, lithium hydroxide etc.)
- Alkaline earth elements - Group 2A, beryllium, magnesium, calcium, strontium, barium and radium
  - very reactive (2 electrons away from a filled s and p orbital)
  - react with water to form hydroxides
  - used to form metal alloys
- Aluminum group - Group 3A - 3 electrons in outer energy level (2s and 1p) properties vary from metallic to metalloid
  - aluminum is the most useful metal of this group being lightweight and strong to make boats, aircraft etc.
- Group 4A - Carbon group - 4 electrons in outer energy level (2s and 2p) - properties vary from nonmetallic to metalloid down the group
  - diamond and graphite are forms of pure carbon
  - silicon and germanium are semiconductors used in electronics
  - tin and lead are useful metals
- Group 5A - Nitrogen group - 5 electrons in outer energy level (2s and 3p) - properties vary from nonmetallic to metalloid to metallic
  - nitrogen and phosphorus are elements necessary to form proteins and nucleic acids in living things
- Group 6A - Oxygen group - 6 electrons in outer energy level (2s and 4p) - properties vary from nonmetallic to metalloid
  - oxygen is the most abundant element on the earth
  - sulfur has many industrial uses (sulfuric acid is the most widely used industrial chemical)
- Group 7A - Halogens - all have 7 electrons in the outer energy level (2s and 5p) - properties vary from nonmetallic to metalloid. Very reactive due to the outer energy level being almost filled.
  - iodine is used as an antiseptic
  - chlorine is a bleaching agent and disinfecting agent
  - fluorine, as the fluoride ion, is used to maintain the health of our teeth
  - fluorine is used to make Teflon
The Periodic Law and Periodic Trends (Periodicity)

Electron Configuration and the Periodic Table

I. Periods and the Blocks of the Periodic Table
   A. Periods
      1. Horizontal rows on the periodic table
      2. Period number corresponds to the highest principal quantum number of the elements in the period
   B. Sublevel Blocks
      1. Periodic table can be broken into blocks corresponding to \( s, p, d, f \) sublevels

II. Blocks and Groups
   A. \( s \)-Block, Groups 1 and 2
      1. Group 1 - The alkali metals
         a. One \( s \) electron in outer shell
         b. Soft, silvery metals of low density and low melting points
         c. Highly reactive, never found pure in nature
      2. Group 2 - The alkaline earth metals
         a. Two \( s \) electrons in outer shell
         b. Denser, harder, stronger, less reactive than Group 1
         c. Too reactive to be found pure in nature
   B. \( d \)-Block, Groups 3 - 12
      1. Metals with typical metallic properties
      2. Referred to as "transition" metals
      3. Group number = sum of outermost \( s \) and \( d \) electrons
   C. \( p \)-Block elements, Groups 13 - 18
      1. Properties vary greatly because it contains both metals and nonmetals
         a. Metals
            (1) softer and less dense than \( d \)-block metals
            (2) harder and more dense than \( s \)-block metals
         b. Metalloids
            (1) Brittle solids with some metallic and some nonmetallic properties
            (2) Semiconductors
         c. Nonmetals
            (1) Halogens (Group 17) are the most reactive of the nonmetals
   D. \( f \)-Block, Lanthanides and Actinides
      1. Lanthanides are shiny metals similar in reactivity to the Group 2 metals
      2. Actinides
         a. All are radioactive
         b. Plutonium (94) through Lawrencium (103) are man-made

Electron Configuration and Periodic Properties

I. Atomic Radii
   A. Atomic Radius
      1. One half the distance between nuclei of identical atoms that are bonded together
   B. Trends
      1. Atomic radius tends to decrease across a period due to increasing positive nuclear charge
      2. Atomic radii tend to increase down a group due to increasing number energy levels (outer electrons are farther from the nucleus) and the shielding of the outer electrons from the full pull of the nucleus by the inner electrons

II. Ionization Energy
   A. Ion
      1. An atom or a group of atoms that has a positive or negative charge
   B. Ionization
      1. Any process that results in the formation of an ion
C. Ionization Energy
1. The energy required to remove one electron from a neutral atom of an element, forming a positive ion (cation) measured in kilojoules/mole (kJ/mol) \[ A + \text{energy} \rightarrow A^+ + e^- \] (endothermic)

D. Trends
1. Ionization energy of main-group elements tends to increase across each period
   a. Atoms are getting smaller, electrons are held closer to the nucleus
2. Ionization energy of main-group elements tends to decrease as atomic number increases in a group
   a. Atoms are getting larger, electrons are farther from the nucleus
   b. Outer electrons become increasingly more shielded from the nucleus by inner electrons
3. Metals have characteristically low ionization energies
4. Nonmetals have high ionization energies
5. Noble gases have very high ionization energies

E. Removing Additional Electrons

1. Ionization energy increases for each successive electron: \[ \text{IE}_1 < \text{IE}_2 < \text{IE}_3 \ldots \]
2. Each electron removed experiences a stronger effective nuclear charge
3. The greatest increase in ionization energy comes when trying to remove an electron from a stable, noble gas configuration

III. Electron Affinity
A. Electron Affinity
1. The energy change that occurs when an electron is acquired by a neutral atom, forming a negative ion (anion) measured in kJ/mol \[ A + e^- \rightarrow A^- + \text{energy} \] (exothermic)
   a. Most atoms release energy when they acquire an electron
   b. Some atoms must be forced to gain an electron

B. Trends
1. Halogens have the highest electron affinities
2. Metals have characteristically low electron affinities
3. Electron affinity tends to increase across a period
   a. Irregularities are due to the extra stability of half-filled and filled sublevels
4. Electron affinity tends to decrease down a group

IV. Ionic Radii
A. Cations
1. Positive ions
2. Smaller than the corresponding atom (they have lost electrons)
   a. Protons outnumber electrons (greater effective nuclear charge)
   b. Less shielding of electrons

B. Anions
1. Negative ions
2. Larger than the corresponding atoms (they have gained electrons)
   a. Electrons outnumber protons (weaker effective nuclear charge)
   b. Greater electron-electron repulsion

C. Trends
1. Ion size tends to increase downward within a group

V. Valence Electrons
A. Valence Electrons
1. The electrons available to be lost, gained, or shared in the formation of chemical compounds
2. Main group element valence electrons are in the outermost \(s\) and \(p\) sublevels

<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>
</tr>
</thead>
<tbody>
<tr>
<td>Number of valence Electrons</td>
<td>1</td>
<td>2</td>
<td>3</td>
<td>4</td>
<td>5</td>
<td>6</td>
<td>7</td>
<td>8</td>
</tr>
</tbody>
</table>
VI. Electronegativity

A. Electronegativity
   1. A measure of the ability of an atom in a chemical compound to attract electrons
   2. Elements that do not form compounds are not assigned electronegativities

B. Trends
   1. Nonmetals have characteristically high electronegativities
      a. Highest in the upper right corner (F has the highest electronegativity)
   2. Metals have characteristically low electronegativities
      a. Lowest in the lower left corner of the table
   3. Electronegativity tends to increase across a period
   4. Electronegativity tends to decrease down a group of main-group elements

VII. Periodic Properties of the $d$- and $f$-Block Elements

A. Atomic Radii
   1. Smaller decrease in radius across a period within the $d$- Block than within the main-group elements
      a. Added electrons are partially shielded from the increasing positive nuclear charge
   2. Little change occurs in radius across an $f$-block of elements

B. Ionization Energy
   1. Tends to increase across $d$- and $f$-Blocks

C. Ion Formation and Ionic Radii
   1. Electrons are removed from the outermost energy level $s$-sublevel first
      a. Most $d$-block elements form $+2$ ions (losing 2 $s$ electrons)
   2. Ions of $d$- and $f$-Blocks are cations, smaller than the corresponding atoms

D. Electronegativity
   1. Characteristically low electronegativity of metals
   2. Electronegativity increases as atomic radius decreases

Fill in the periodic trends on this table