The elements in the periodic table are arranged in order of increasing atomic number. All of these elements display several other trends and we can use the periodic law and table formation to predict their chemical, physical, and atomic properties. Understanding these trends is done by analyzing the elements electron configuration; all elements prefer an octet formation and will gain or lose electrons to form that stable configuration.

**Atomic Radius**

We can never determine the atomic radius of an atom because there is never a zero probability of finding an electron, and thus never a distinct boundary to the atom. All that we can measure is the distance between two nuclei (internuclear distance). A covalent radius is one-half the distance between the nuclei of two identical atoms. An ionic radius is one-half the distance between the nuclei of two ions in an ionic bond. The distance must be apportioned for the smaller cation and larger anion. A metallic radius is one-half the distance between the nuclei of two adjacent atoms in a crystalline structure. The noble gases are left out of the trends in atomic radii because there is great debate over the experimental values of their atomic radii. The SI units for measuring atomic radii are the nanometer (nm) and the picometer (pm). $1 \text{ nm} = 1 \times 10^{-9} \text{ m}$; $1 \text{ pm} = 1 \times 10^{-12} \text{ m}$.

Figure 1: Covalent Radii  Figure 2: Ionic Radii  Figure 3: Metallic Radii
This is the covalent radius. It is ½ the distance between two nuclei of similar atoms that are bonded together by a single covalent bond.

The distance between two nuclei.

This is the ionic radius. It is ⅓ of the distance between the nuclei of two atoms that are bonded by an ionic bond.

This is the distance between the two nucleuses of the atoms. The positive sign represents the cations and the negative sign represents the anion.
To explain this trend, the concept of screening and penetration must be understood. Penetration is commonly known as the distance that an electron is from the nucleus. Screening is defined as the concept of the inner electrons blocking the outer electrons from the nuclear charge. Within this concept we assume that there is no screening between the outer electrons and that the inner electrons shield the outer electrons from the total positive charge of the nucleus. In order to comprehend the extent of screening and penetration within an atom, scientists came up with the effective nuclear charge, \(Z_{\text{eff}}\). The equation for calculating the effective nuclear charge is shown below.

\[
Z_{\text{eff}} = Z - S
\]

In the equation, \(S\) represents the number of inner electrons that screen the outer electrons. Students can easily find \(S\) by using the atomic number of the noble gas that is one period above the element. For example, the \(S\) we would use for Chlorine would be 10 (the atomic number of Neon). \(Z\) is the total number of electrons in the atom. Since we know that a neutral atom has an identical number of protons and electrons, we can use the atomic number to define \(Z\). For example, Chlorine would have a \(Z\) value of 17 (the atomic number of Chlorine). Continuing to use Chlorine as an example, the 10 inner electrons (\(S\)) would screen out the positive charge of ten protons. Therefore there would be an effective nuclear charge of 17-10 or +7. The effective nuclear charge shows that the nucleus is pulling the outer electrons with a +7 charge and therefore the outer electrons are pulled closer to the nucleus and the atomic radii is smaller. In summary, the greater the nuclear charge, the greater pull the nucleus has on the outer electrons and the smaller the atomic radii. In contrast, the smaller nuclear charge, the lesser pull the nucleus has on the outer electrons, and the larger atomic radii. Additionally, as the atomic number increases, the effective nuclear charge also increases. Figure 3 depicts the effect that the effective nuclear charge has on atomic radii.
Now we are ready to describe the atomic radius trend in the periodic table. The atomic number increases moving left to right across a period and subsequently so does the effective nuclear charge. Therefore, moving left to right across a period the nucleus has a greater pull on the outer electrons and the atomic radii decreases. Moving down a group in the periodic table, the number of filled electron shells increases. In a group, the valence electrons keep the same effective nuclear charge, but now the orbitals are farther from the nucleus. Therefore, the nucleus has less of a pull on the outer electrons and the atomic radii are larger.

We can now use these concept to explain the atomic radius differences of cations and anions. A cation is an atom that has lost one of its outer electrons. Cations have a smaller radius than the atom that they were formed from. With the loss of an electron, the positive nuclear charge out powers the negative charge that the electrons exert. Therefore, the positive nucleus pulls the electrons tighter and the radius is smaller. An anion is an atom that has gained an outer electron. Anions have a greater radius than the atom that they were formed from. The gain of an electron does not alter the nuclear charge, but the addition of an electron causes a decrease in the effective nuclear charge. Therefore, the electrons are held more loosely and the atomic radius is increased.
Ionization Energy (ionization potential)

Expelling an electron from an atom requires enough energy to overcome the magnetic pull of the positive charge of the nucleus. Therefore, ionization energy (I.E. or I) is the energy required to completely remove an electron from a gaseous atom or ion. The Ionization Energy is always positive.

$Na_{(g)} \rightarrow Na^+_{(g)} + e^-_{(g)}$

$Na^+_{(g)} \rightarrow Na^{2+}_{(g)} + e^-$

Ionization energies increase relative to high effective charge. The highest ionization energies are the noble gases because they all have high effective charge due to their octet formation and require a high amount of energy to destroy that stable configuration. The highest amount of energy required occurs with the elements in the upper right hand corner. Additionally, elements in the left corner have a low ionization energy because losing an electron allows them to have the noble gas configuration. Therefore, it requires less energy to remove one of their valence electrons.

**Table 1:** Ionization Energies of certain elements (1st IE, 2nd IE, etc)

<table>
<thead>
<tr>
<th></th>
<th>1st</th>
<th>2nd</th>
<th>3rd</th>
<th>4th</th>
<th>5th</th>
<th>6th</th>
<th>7th</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Na 496 4562
Mg 738 1451 7733
Al 577 1817 2745 11580
Si 786 1577 3232 4356 16090
P 1060 1903 2912 4957 6274 21270
S 999.6 2251 3361 4564 7013 8496 27110
Cl 1256 2297 3822 5158 6542 9362 11020
Ar 1520 2666 3931 5771 7238 8781 12000

These are the ionization energies for the period three elements. Notice how Na after in the second I.E, Mg in the third I.E., Al in the fourth I.E., and so on, all have a huge increase in energy compared to the proceeding one. This occurs because the proceeding configuration was in a stable octet formation; therefore it requires a much larger amount of energy to ionize.

Ionization Energies increase going left to right across a period and increase going up a group. As you go up a group, the ionization energy increases, because there are less electron shielding the outer electrons from the pull of the nucleus. Therefore, it requires more energy to out power the nucleus and remove an electron. As we move across the periodic table from left to right, the ionization energy increases, due to the effective nuclear charge increasing. This is because the larger the effective nuclear charge, the stronger the nucleus is holding onto the electron and the more energy it takes to release an electron.

Figure 7: Courtesy of Jessica Thornton (UCD)

The ionization energy is only a general rule. There are some instances when this trend does not prove to be correct. These can typically be explained by their electron configuration. For example, Magnesium has a higher ionization energy than Aluminum. Magnesium has an electron configuration of [Ne]3s2. Magnesium has a high ionization energy because it has a filled 3s orbital and it requires a higher amount of energy to take an electron from the filled orbital.
Electron Affinity

Electron affinity (E.A.) is the energy change that occurs when an electron is added to a gaseous atom. Electron affinity can further be defined as the enthalpy change that results from the addition of an electron to a gaseous atom. It can be either positive or negative value. The greater the negative value, the more stable the anion is.

- (Exothermic) The electron affinity is positive

\[ X_{(g)} + e^- \rightarrow X^- + \text{Energy} \]

- (Endothermic) The electron affinity is negative

\[ X_{(g)} + e^- + \text{Energy} \rightarrow X^- \]

It is more difficult to come up with trends that describe the electron affinity. Generally, the elements on the right side of the periodic table will have large negative electron affinity. The electron affinities will become less negative as you go from the top to the bottom of the periodic table. However, Nitrogen, Oxygen, and Fluorine do not follow this trend. The noble gas electron configuration will be close to zero because they will not easily gain electrons.

Electronegativity

Electronegativity is the measurement of an atom to compete for electrons in a bond. The higher the electronegativity, the greater its ability to gain electrons in a bond. Electronegativity will be important when we later determine polar and nonpolar molecules. Electronegativity is related with ionization energy and electron affinity. 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 by the positive nucleus on the negative electrons. Therefore the electronegativity increases from bottom to top and from left to right.
Metallic Character

The metallic character is used to define the chemical properties that metallic elements present. Generally, metals tend to lose electrons to form cations. Nonmetals tend to gain electrons to form anions. They also have a high oxidation potential therefore they are easily oxidized and are strong reducing agents. Metals also form basic oxides; the more basic the oxide, the higher the metallic character.

As you move across the table from left to right, the metallic character decreases, because the elements easily accept electrons to fill their valance shells. Therefore, these elements take on the nonmetallic character of forming anions. As you move up the table, the metallic character decreases, due to the greater pull that the nucleus has on the outer electrons. This greater pull makes it harder for the atoms to lose electrons and form cations.

Other Trends

Melting Points: Trends in melting points and molecular mass of binary carbon-halogen compounds and hydrogen halides are due to intermolecular forces. Melting destroys the arrangement of atoms in a solid, therefore the amount of heat necessary for melting to occur depends on the strength of attraction between the atoms. This strength of attraction increases as the number of electrons increase. Increase in electrons increases bonding.
Example: Melting point of HF should be approximately -145 °C based off melting points of HCl, HBr, and HI, but the observed value is -83.6°C.

Heat and electricity conductivity vary regularly across a period. Melting points may increase gradually or reach a peak within a group then reverse direction.

Example: Third period elements Na, Mg, and Al are good conductors of heat and electricity while Si is only a fair conductor and the nonmetals P, S, Cl and Ar are poor conductors.

---

**Redox Potentials**

**Oxidation Potential**

Oxidation is a reaction that results in the loss of an electron. Oxidation potential follows the same trends as the ionization energy. That is because the smaller the ionization energy, the easier it is to remove an electron. (e.g)

\[K_{(s)} \rightarrow K^+ + e^-\]

**Reduction Potential**

Reduction is a reaction that results in the gaining of an electron. Reduction potentials follow the same trend as the electron affinity. That is because the larger, negative electron affinity, the easier it is to give an electron. Example of Reduction:

\[F_{(s)} + e^- \rightarrow F^-\]

**Uses in knowing the Periodic Properties of Elements**

1. Predicting greater or smaller atomic size and radial distribution in neutral atoms and ions
2. Measuring and comparing ionization energies
3. Comparing electron affinities and electronegativities
   - Predicting redox potential
   - Comparing metallic character with other elements; its ability to form cations
   - Predicting what reaction may or may not occur due to the trends
   - Determining greater cell potential (sum of oxidation and reduction potential) between reactions
   - Completing chemical reactions according to trends