MASTERING Periodic Trends

Perfect your performance with periodicity!

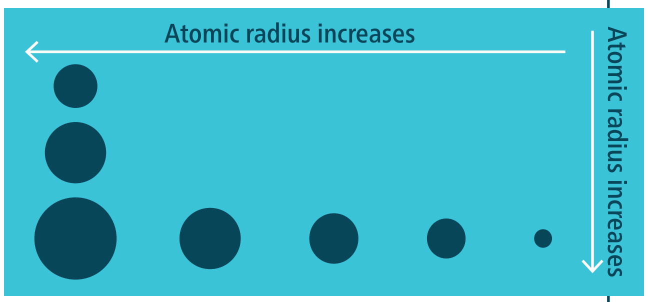
Important Trend Terms

Effective nuclear charge: the net positive charge from the nucleus that an electron can "feel" attractions from. The core electrons are said to shield the valence electrons from the full attractive forces of the protons in the nucleus.

Shielding: core (nonvalence) electrons shield the valence electrons from the full attractive forces of the protons in the nucleus.

Electron-electron repulsions: due to their like charges, electron pairs orient themselves as far away as possible from each other, causing the electron cloud to expand (justifies trends across a period).

1. Atomic Radius



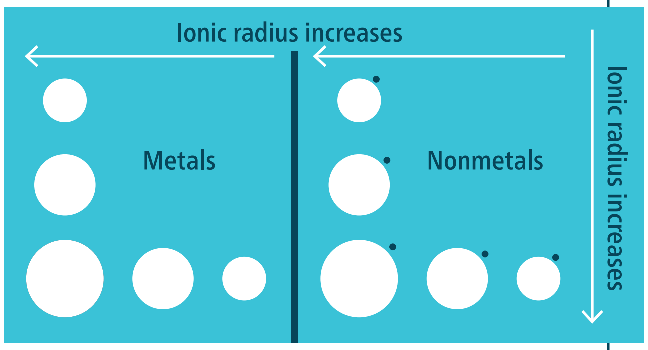
Atomic radius is the distance from the atom's nucleus to the outer edge of the electron cloud.

In general, atomic radius decreases across a period and increases down a group.

Across a period, effective nuclear charge increases as electron shielding remains constant. A higher effective nuclear charge causes greater attractions to the electrons, pulling the electron cloud closer to the nucleus which results in a smaller atomic radius.

Down a group, the number of energy levels (n) increases, so there is a greater distance between the nucleus and the outermost orbital. This results in a larger atomic radius.

2. Ionic Radius



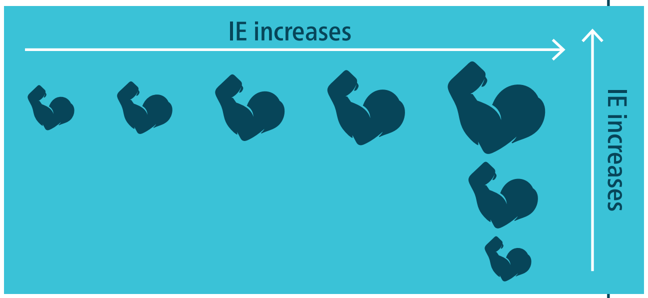
Ionic radius is the distance from the nucleus to the outer edge of the electron cloud of an ion.

The same trend of atomic radius applies once you divide the table into metal and nonmetal sections.

A cation has a smaller radius than its neutral atom because it loses valence electrons. The "new" valence shell is held closer to the nucleus, resulting in a smaller radius for the cation.

An anion has a larger radius than the neutral atom because it gains valence electrons. There are added electron/electron repulsions in the valence shell that expand the size of the electron cloud, which results in a larger radius for the anion.

3. Ionization Energy



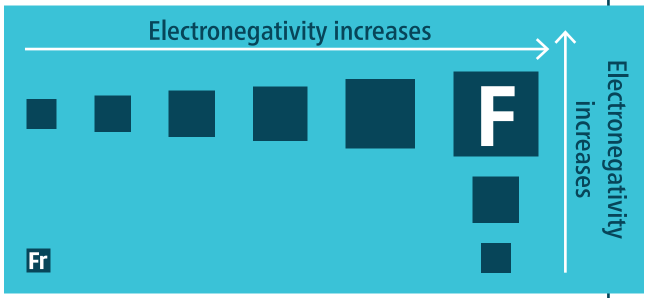
Ionization energy (IE) is the energy required to remove the highest-energy electron from a neutral atom.

In general, ionization energy increases across a period and decreases down a group.

Across a period, effective nuclear charge increases as electron shielding remains constant. This pulls the electron cloud closer to the nucleus, strengthening the nuclear attraction to the outer-most electron, and is more difficult to remove (requires more energy).

Down a group, the number of energy levels (n) increase and the distance is greater between the nucleus and highest-energy electron. The increased distance weakens the nuclear attraction to the outer-most electron, and is easier to remove (requires less energy).

4. Electronegativity



Electronegativity is the measure of the ability of an atom in a bond to attract electrons to itself.

Electronegativity increases across a period and decreases down a group.

Towards the left of the table, valence shells are less than half full, so these atoms (metals) tend to lose electrons and have low electronegativity. Towards the right of the table, valence shells are more than half full, so these atoms (nonmetals) tend to gain electrons and have high electronegativity.

Down a group, the number of energy levels (n) increases, and so does the distance between the nucleus and the outermost orbital. The increased distance and the increased shielding weaken the nuclear attraction, and so an atom can't attract electrons as strongly.

Fluorine is the most electronegative element, whereas francium is the least electronegative element.

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