

## 5-3 Periodic Properties

### Section 5-3 Periodic Properties

Atomic Radius – half of the distance between the nuclei of identical atoms that are bonded together.

#### Trends

Radii get smaller across a period from the increasing positive charge of the nucleus. More protons pull the electrons in closer to make atom smaller.

Radii get larger going down a group from each higher energy level that is added.

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### Section 5-3 (Cont'd 2)

First ionization energy – amount of energy required to remove one electron from a neutral atom.

ion – atom or group of bonded atoms having a positive or negative charge. A charge is formed from unequal amounts of protons and electrons in an atom.

ionization – Any process that results in the formation of an ion

The energies for removal of additional electrons would be 2<sup>nd</sup> ionization energy, 3<sup>rd</sup> ionization energy, etc.

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### Section 5-3 (Cont'd 3)

#### Trends

ionization energy increases across each period. Having a greater positive nuclear charge makes it harder to take electrons away.

ionization energy decreases going down each group. It is easier to pull off electrons that are farther away from the nucleus.

Each additional ionization energy will increase as more electrons are removed. The positive nuclear charge will have a tighter grip on the remaining electrons.

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### Section 5-3 (Cont'd 4)

Electron Affinity - Energy change when a neutral atom acquires an electron. Most atoms release energy when acquiring an electron.

#### Trends

Halogens gain electrons most readily. They have high negative numbers for electron affinity. The ease of gaining electrons is a major reason for the high reactivity.

Electron affinity does not follow a regular trend following periods or groups.

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### Section 5-3 (Cont'd 5)

Ionic Radius – the radius of an ion

#### Cation

- positive ion (ex.  $\text{Li}^+$ ,  $\text{Ca}^{2+}$ )
- has more protons than electrons
- smaller than a neutral atom
- often seen with metals

#### Anion

- negative ion (ex.  $\text{Cl}^-$ ,  $\text{O}^{2-}$ )
- has more electrons than protons
- larger than a neutral atom
- often seen with nonmetals

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### Section 5-3 (Cont'd 6)

#### Trends

Metals towards left form cations

Nonmetals in upper right form anions

Cationic radii decrease across a period

Anionic radii decrease across a period

Gradual increase of atomic radii down a group

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# 5-3 Periodic Properties

## Section 5-3 (Cont'd 7)

Valence electrons – the electrons available to be lost, gained, or shared in the formation of a chemical compound.

The valence electrons are the electrons in the s and p orbitals on the highest occupied energy level.

Group	Valence Electrons
1	1
2	2
13	3
14	4
15	5
16	6
17	7
18	8

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## Section 5-3 (Cont'd 8)

Electronegativity – measure of the ability of an atom in a chemical compound to attract electrons

Electronegativity was created by Linus Pauling

The most electronegative element is Fluorine. It has a value of 4.00

Trends

Electronegativity tends to increase across the periods

Electronegativity tends to decrease going down a group

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## Section 5-3 (Cont'd 9)

Properties of the d and f blocks tend to vary less than the main group elements

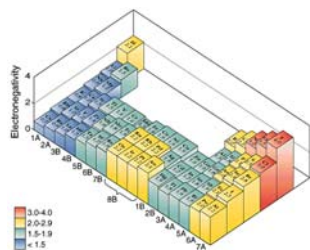
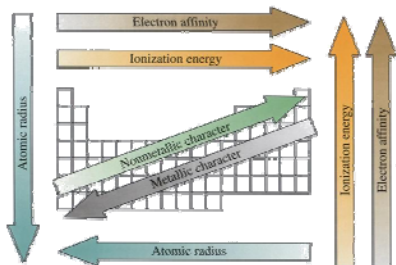
Atomic radii – the d-block slightly decreases and then slightly increases due to electron repulsion.

Ionization energy – the d-block and f-block increase down a group because outer electrons are less shielded from increasing nuclear charge

Electrons are removed from the highest occupied energy level first

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## Section 5-3 (Cont'd 10)



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