Periodic Trends

Although the periodic table was constructed with very limited understanding of the subatomic structure of the atom, the trends in properties that were used for grouping elements in columns are determined by the structure of atoms.

Key Terminology

- Outermost electrons: the electrons in the outer layer of an atom which are responsible for atomic reactivity (often called valence electrons)
- Cation: has less electrons than protons and thus is a positively charged ion
- Anion: has more electrons than protons and thus is a negatively charged ion
- Atomic Radius: an estimation of the distance from the nuclei to the "edge" of the outermost orbital
- Covalent Radius: half the distance between like-atoms that are bonded together in a molecule
- Ionic Radius: The effective radius of ions in solids such as NaCl
- Ionization Energy: The energy required to remove an electron from an atom by overcoming the electrostatic forces between the nucleus and the electron.
- Electron Affinity: the energy released (or absorbed) when an electron is incorporated by an atom, forming an anion.
- Effective Nuclear Charge (Zeff): the sum of all attractions and repulsions of the nucleus upon a specific electron
- Shielding: the effect of the inner orbitals have on decreasing the effect of Zeff on the outermost electrons
- Core Charge: (# protons) – (# all inner, or non-valence, electrons)

Trends in Atomic Radii (size)

- Atomic size is difficult to measure and results vary, but the overall trends in size remain the same regardless of whether size is determined by the minimum pore size an atom can fit through or a measurement determined from the volume of a known number of atoms of a substance
- Scientists use covalent radius, or half the distance between the nuclei of 2 bonded atoms.

Factors affecting the INCREASE in size of atoms down a group

- Atoms get larger primarily due to the addition of another energy level at each period
- Each full inner level of electrons (core electrons) also **shields** the outermost electrons from the attractive force of the protons. This shielding increases the electron repulsion of the valence electrons, effectively increasing the size of the atom.
- Note that core charge remains the same down a group (eg. +1 core charge for all group 1 atoms)

Factors affecting the DECREASE in size going across (in same period)

- Electrons are added to the same energy level, increasing electron-electron repulsion, but only to a small degree.
- Nuclear charge (core charge) increases going to the right. This greatly increases the force of attraction on the outer electrons, pulling this level closer to the nucleus.

Trends in Ionic Radius

Formation of Cations: Cations are **smaller** than their neutral atom counterparts

- Many atoms lose an energy level by forming cations, making them smaller.
- Formation of cations results in more protons than electrons
- Decrease in repulsive forces between electrons

- Increase in effective nuclear force felt by the valence electrons.

Formation of Anions: Anions are **larger** than their neutral atom counterparts

- Formation of anions results in more valence electrons than the corresponding atom.
- Increase in repulsion between electrons moving the outermost energy level further away
- Decrease in effective nuclear force felt by valence electrons

Trends in Ionization Energy (I.E.)

I.E. is the amount of energy (in kJ) required to remove an electron from the ground state of a gaseous atom or ion. The atom has been "ionized" or charged.

- First ionization energy is that energy required to remove first electron. [$Na_{(g)}$ + energy → $Na^{1+}_{(g)}$ + e^-]
- Second ionization energy is that energy required to remove second electron, etc.
- It requires more energy to remove each successive electron.
- When all *valence* electrons have been removed, the ionization energy takes a large leap.

Factors affecting the DECREASE in I.E. down a group

- Core charge (or Z_{eff}) is the same down the same group
- Shielding increases (more inner electron levels)
- Outermost electrons move further from the nucleus

This results in valence electrons that are held less tightly and require less energy to remove. The larger the atom is, the easier its electrons are to remove.

Factors affecting INCREASE in I.E. going across (in same period)

- Nuclear charge (core charge) increases.
- Outer electrons are held more tightly and pulled closer to the nucleus (smaller atomic radius)
- Electron repulsion increases but this has a negligible effect relative to the increase in Z_{eff}

In general, this results in valence electrons that are held more tightly and require more energy to remove.

I.E. Calculations

An atom has the following ionization energies. Determine which group in the periodic table it belongs to.

$I.E._1$ = 590 kJ/mol $I.E._2$ = 1145 kJ/mol $I.E._3$ = 4936 kJ/mol $I.E._4$ = 6752 kJ/mol

To determine this, you must identify the largest jump between successive ionization energies. You must divide a higher I.E. by the one that precedes it. Obviously this is not done for the first ionization energy since there is no preceding I.E.

$I.E._2/I.E._1 = 1145/590 = 1.94$ $I.E._3/I.E._2 = 4936/1145 = 4.31$ $I.E._4/I.E._3 = 6752/4936 = 1.37$

The biggest jump occurs between $I.E._2$ and $I.E._3$ and therefore $I.E._3$ is energy required to remove the electron from an inner energy level. Thus, $I.E._1$ and $I.E._2$ represent the energy required to remove electrons from the outer level; therefore, this atom has two outermost electrons and is in group 2 of the periodic table.

Trends in Electron Affinity (E.A.)

Electron affinity is amount of energy released when a neutral atom gains an electron. A release of energy would be a positive gain in energy to the surroundings.

Generally, the higher the ionization energy, the higher the electron affinity (Or atoms that accept an electron easily won't give one up without a fight)

Decrease in E.A. down a group

- The added electron is further from the nucleus moving down a group
- The shielding is increased
- Thus electron-electron repulsion is greater for the extra electron further down a group

Overall: Atoms further down a group have increasing resistance to receiving an extra electron and thus less positive electron affinities.

Increase in E.A across a period

- Core charge increases to the right of a period
- Therefore, the extra electron is more easily added
- Moving to the left, the electron-electron repulsion is more substantial due to the reduced nuclear force

Atoms further to the left in a period also have less positive electron affinities because of increasing resistance to receiving an extra electron. Since there will be less nuclear force acting on these outermost electrons (moving to the left), electron-electron repulsion becomes more substantial when an electron is added to each atom. Generally, the higher the ionization energy, the higher the electron affinity.

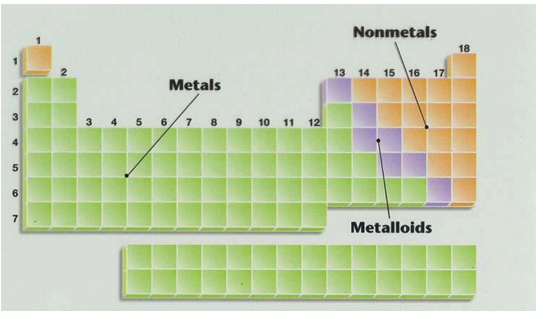
Trends in Reactivity

Reactivity ties all the previous trends together. It is easiest to look separately at metals and non-metals for overall reactivity.

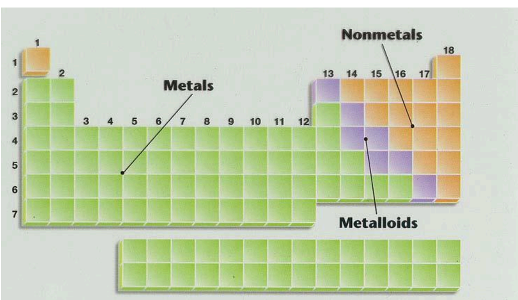
- The most reactive metals are the largest since they are the best electron givers.
  - Lowest I.E., lowest E.A.
- The most reactive nonmetals are the smallest ones, the best electron takers.
  - Highest E.A, highest I.E.

Summary Page

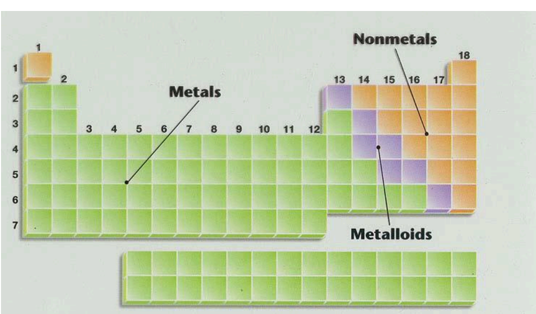
Atomic Radii



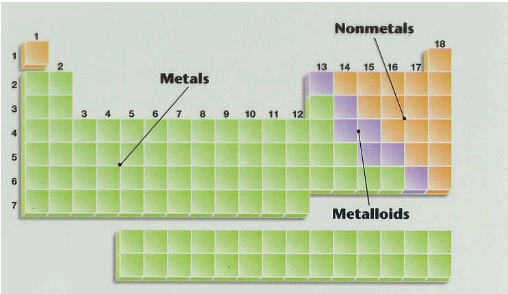
Ionization Energy



Electron Affinity



Reactivity


Nonmetals
Metals
Metalloids
1
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