Periodic Trends

Elemental Properties and Patterns

The Periodic Law

- Dimitri Mendeleev was the first scientist to publish an organized periodic table of the known elements.

Henry Moseley

- Discovered the proton (atomic number) of the elements and arranged them according to atomic number.
- Modern periodic table is arranged by atomic number
- Built on Mendeleev’s table which was arranged by atomic mass

The Periodic Law

- Mendeleev understood the ‘Periodic Law’ which states:
- When arranged by increasing atomic number, the chemical elements display a regular and repeating pattern of chemical and physical properties.

The Periodic Law

- Atoms with similar properties appear in groups or families (vertical columns) on the periodic table.
- They are similar because they all have the same number of valence (outer shell) electrons, which governs their chemical behavior.

Periodic Trends

- There are several important atomic characteristics that show predictable trends that you should know.
- The first and most important is atomic radius.
- Radius is the distance from the center of the nucleus to the “edge” of the electron cloud.

Atomic Radius

- Since a cloud’s edge is difficult to define, scientists use define covalent radius, or half the distance between the nuclei of 2 bonded atoms.
- Atomic radii are usually measured in picometers (pm) or angstroms (Å). An angstrom is 1×10^{-10} m.

Atomic Radius

- The trend for atomic radius in a vertical column is to go from smaller at the top to larger at the bottom of the family.
- Why?
- With each step down the family, we add an entirely new PEL (Principal energy level) to the electron cloud, making the atoms larger with each step.

Atomic Radius

- The trend across a horizontal period is less obvious.
- What happens to atomic structure as we step from left to right?
- Each step adds a proton and an electron (and 1 or 2 neutrons).
- Electrons are added to existing PELs or sublevels.

Atomic Radius

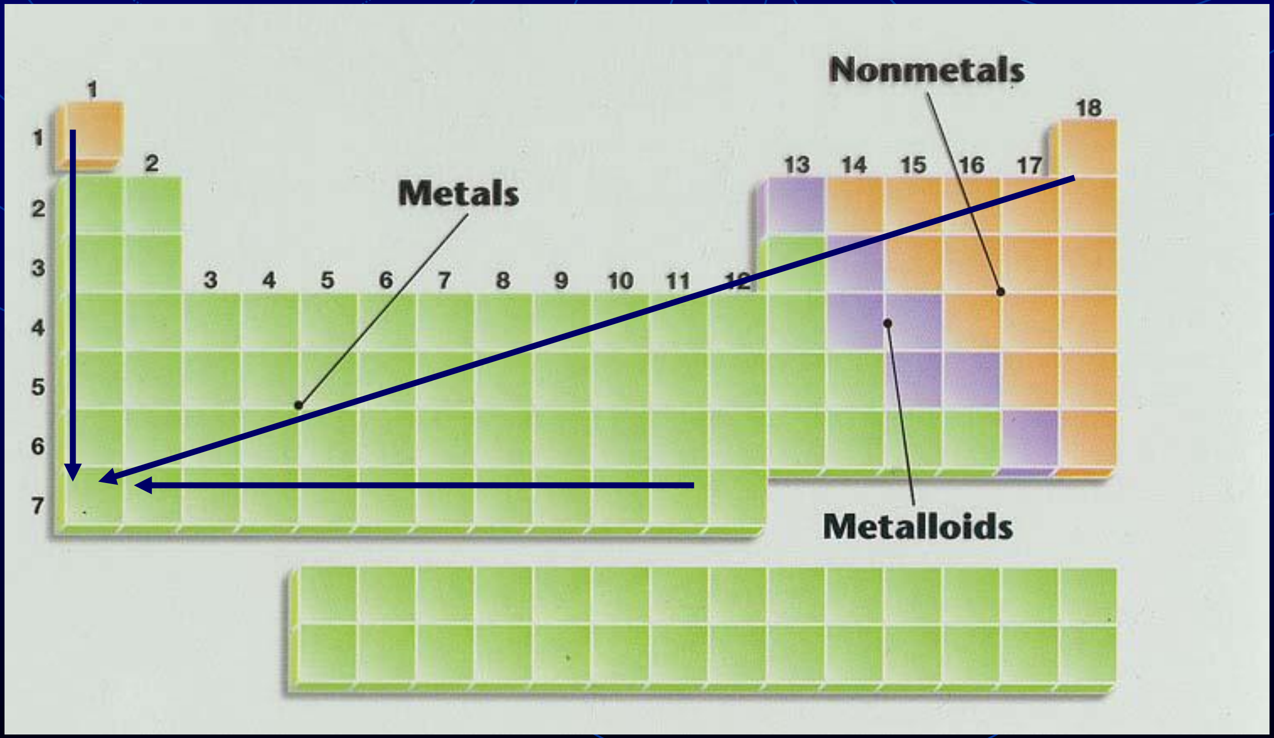
- The effect is that the more positive nucleus has a greater pull on the electron cloud.
- The nucleus is more positive and the electron cloud is more negative.
- The **<u>increased attraction pulls the cloud in</u>**, making atoms smaller as we move from left to right across a period.

Effective Nuclear Charge

- What keeps electrons from simply flying off into space?
- **Effective nuclear charge** is the pull that an electron "feels" from the nucleus.
- The closer an electron is to the nucleus, the more pull it feels.
- As effective nuclear charge increases, the electron cloud is pulled in tighter.

Atomic Radius

- On your periodic table sheet, draw arrows like this:



Shielding

- As more PELs are added to atoms, the inner layers of electrons **shield** the outer electrons from the nucleus.
- The effective nuclear charge (enc) on those outer electrons is less, and so the outer electrons are less tightly held.

Ionization Energy

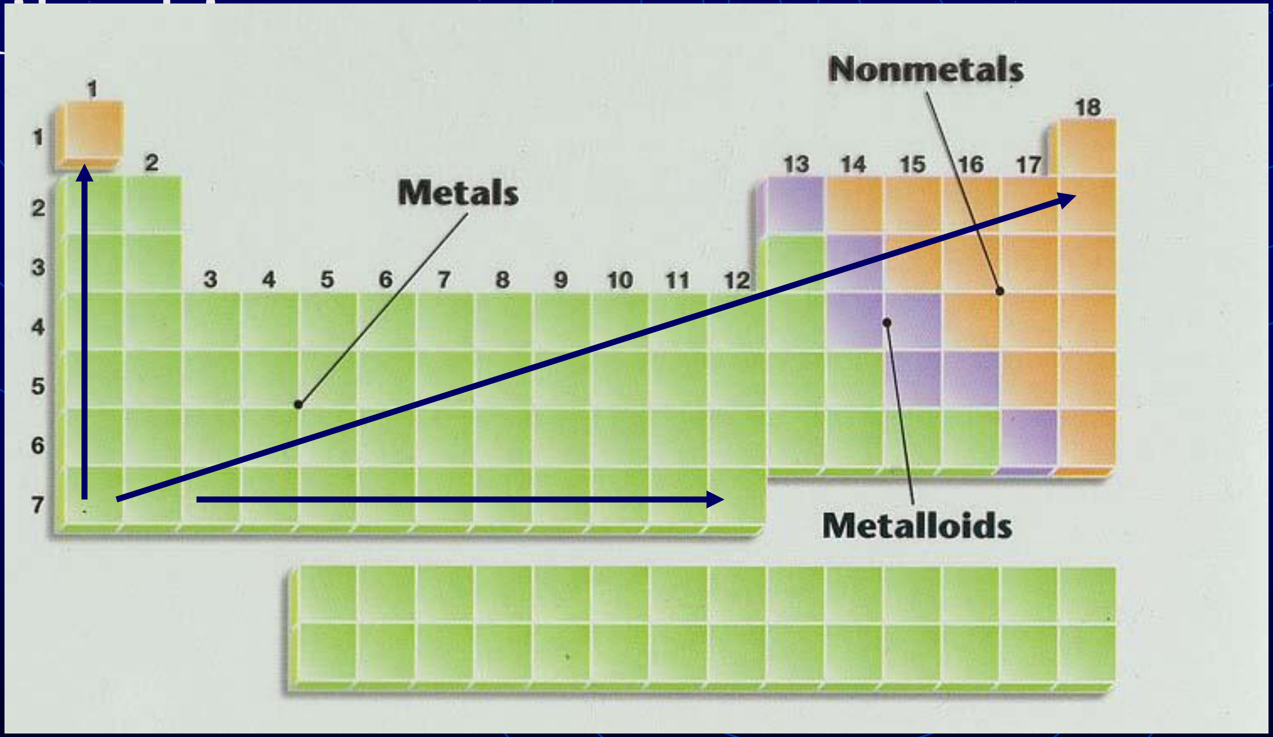
- This is the second important periodic trend.
- If an electron is given enough energy (in the form of a photon) to overcome the effective nuclear charge holding the electron in the cloud, it can leave the atom completely.
- The atom has been “ionized” or charged.
- The number of protons and electrons is no longer equal.

Ionization Energy

- The energy required to remove an electron from an atom is ionization energy. (measured in kilojoules, kJ)
- The larger the atom is, the easier its electrons are to remove.
- **Ionization energy and atomic radius are inversely proportional.**

Ionization Energy (Potential)

- Draw arrows on your periodic table sheet like this:

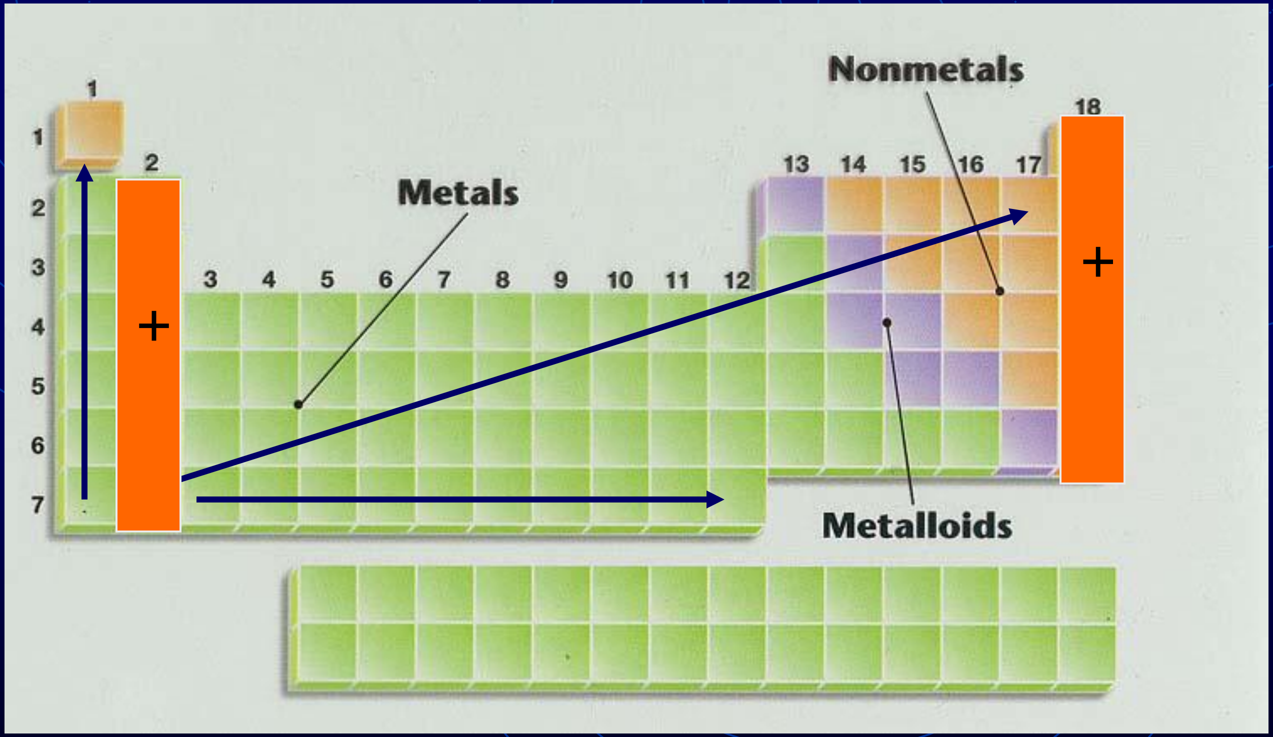


Electron Affinity

- What does the word ‘affinity’ mean?
- Electron affinity is the energy **<u>change</u>** that occurs when an atom **<u>gains an electron</u>** (also measured in kJ).

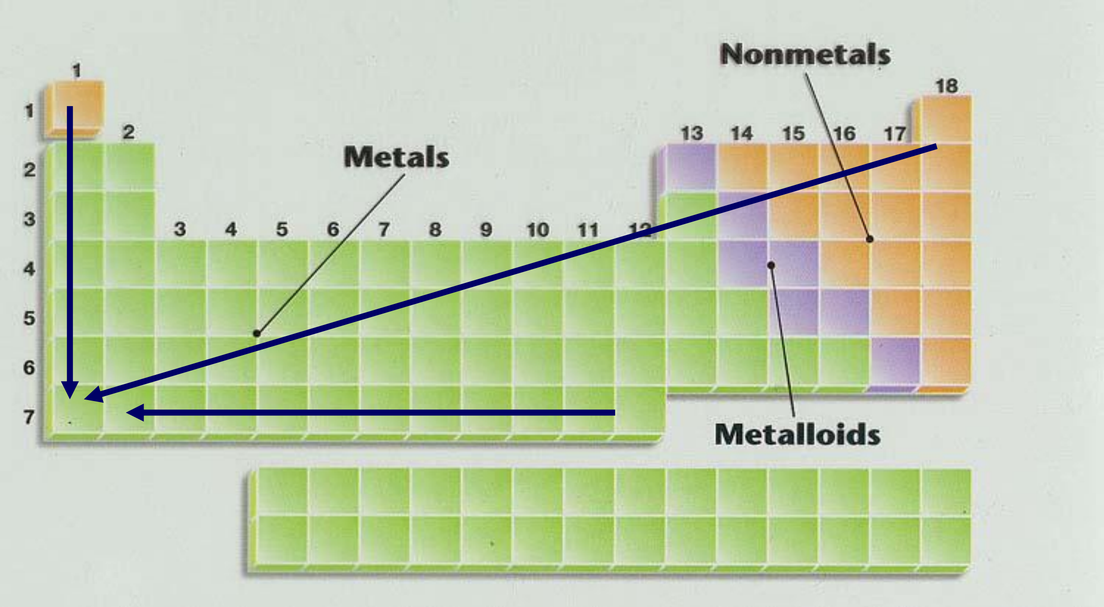
Electron Affinity

- Your periodic table should look like this:



Metallic Character

- This is simply a relative measure of how easily atoms lose or give up electrons.
- Your help sheet should look like this:

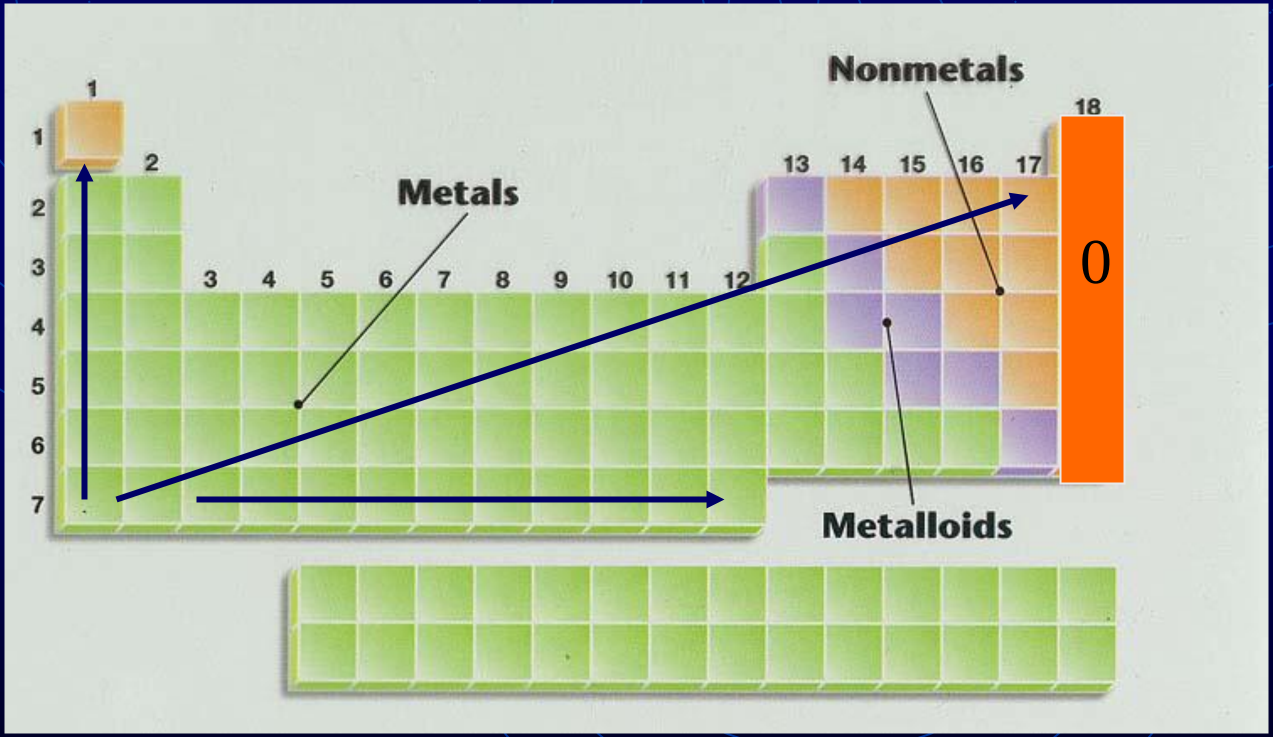


Electronegativity

- Electronegativity is a measure of an atom's attraction for another atom's electrons.
- It is an arbitrary scale that ranges from 0 to 4.
- Generally, metals are electron givers and have low electronegativities.
- Nonmetals are electron takers and have high electronegativities.
- What about the noble gases?

Electronegativity

- Your periodic table should look like this:

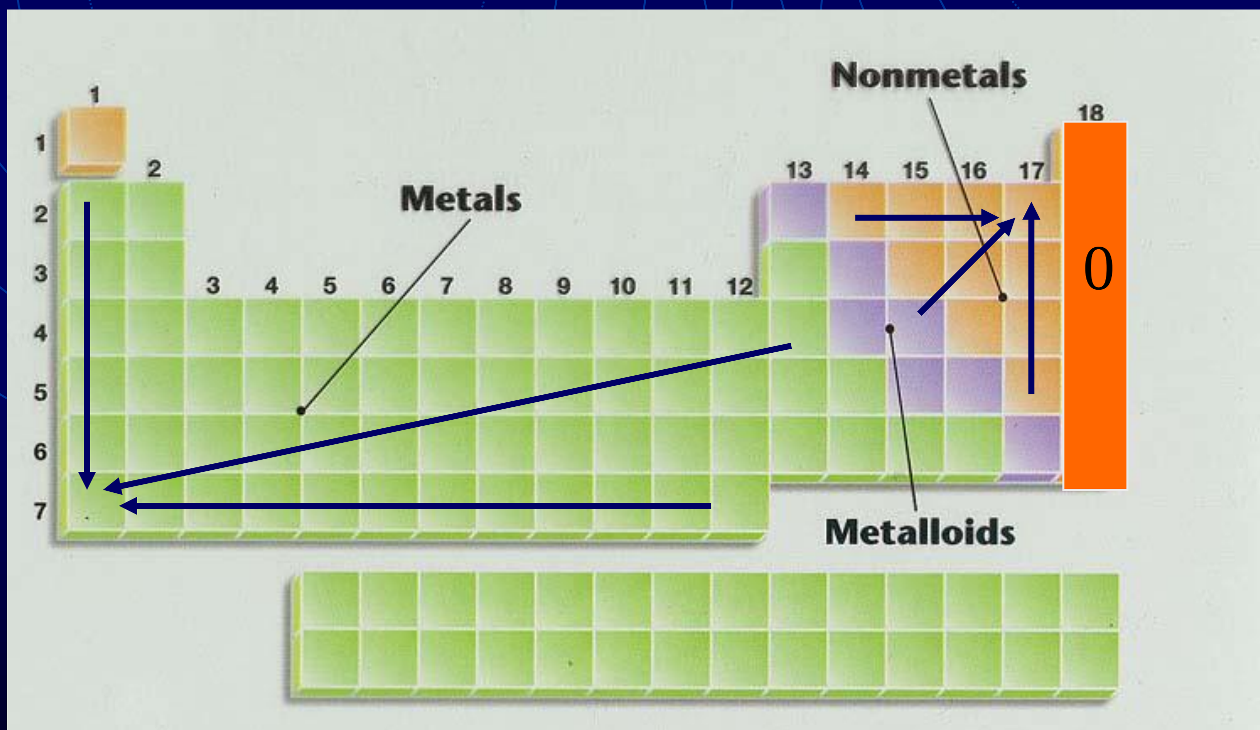


Overall Reactivity

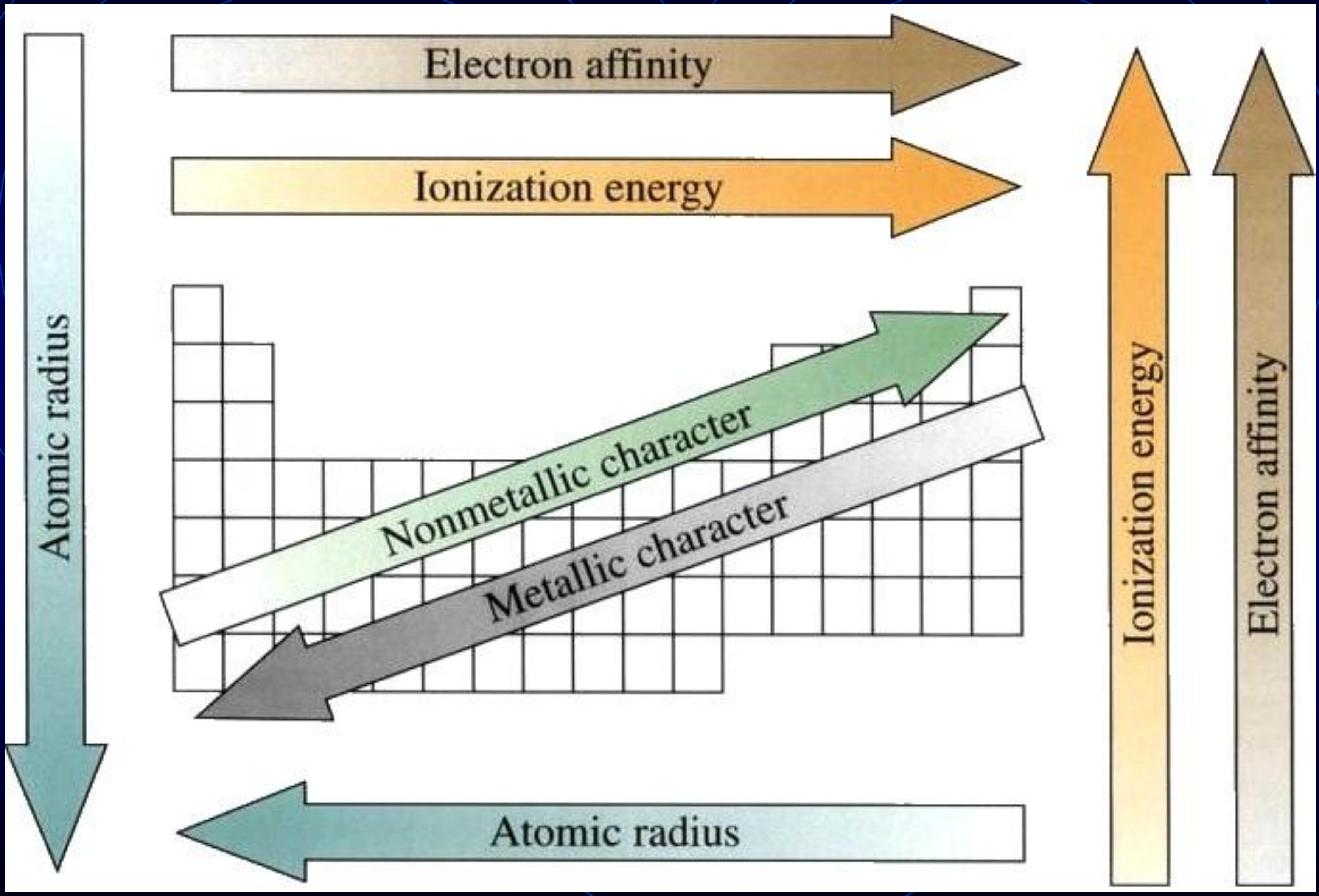
- This ties all the previous trends together in one package.
- However, we must treat metals and nonmetals separately.
- The most reactive metals are the largest since they are the best electron givers.
- The most reactive nonmetals are the smallest ones, the best electron takers.

Overall Reactivity

- Your periodic table will look like this:



Recap!



Electron affinity
Ionization energy
Atomic radius
Nonmetallic character
Metallic character
Ionization energy
Electron affinity
Atomic radius

Must Memorize!

- Up and to the right toward little Fluorine we have increasing: ionization energy, electron affinity, and electronegativity
- Down and to the left toward big Francium we have increase atomic radius and metallic character.

The Octet Rule

- The “goal” of most atoms (except H, Li and Be) is to have an octet or group of 8 electrons in their valence energy level.
- They may accomplish this by either giving electrons away or taking them.
- Metals generally give electrons, nonmetals take them from other atoms.
- Atoms that have gained or lost electrons are called ions.

Ions

- When an atom gains an electron, it becomes negatively charged (more electrons than protons) and is called an **anion.**
- In the same way that nonmetal atoms can gain electrons, metal atoms can lose electrons.
- They become positively charged **cations**.

Ions

- Here is a simple way to remember which is the cation and which the anion:

This is Ann Ion.

She's unhappy and negative.

This is a cat-ion.

He's a "plussy" cat!

Ionic Radius

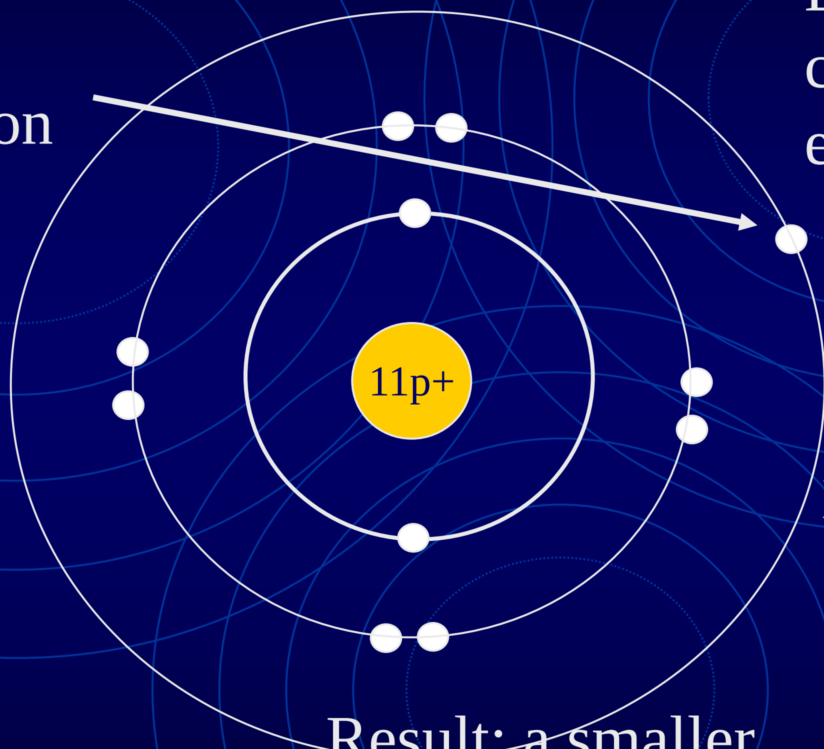
- Cations are always smaller than the original atom.
- The entire outer PEL is removed during ionization.
- Conversely, anions are always larger than the original atom.
- Electrons are added to the outer PEL.

Cation Formation

Na atom

1 valence electron

Effective nuclear charge on remaining electrons increases.



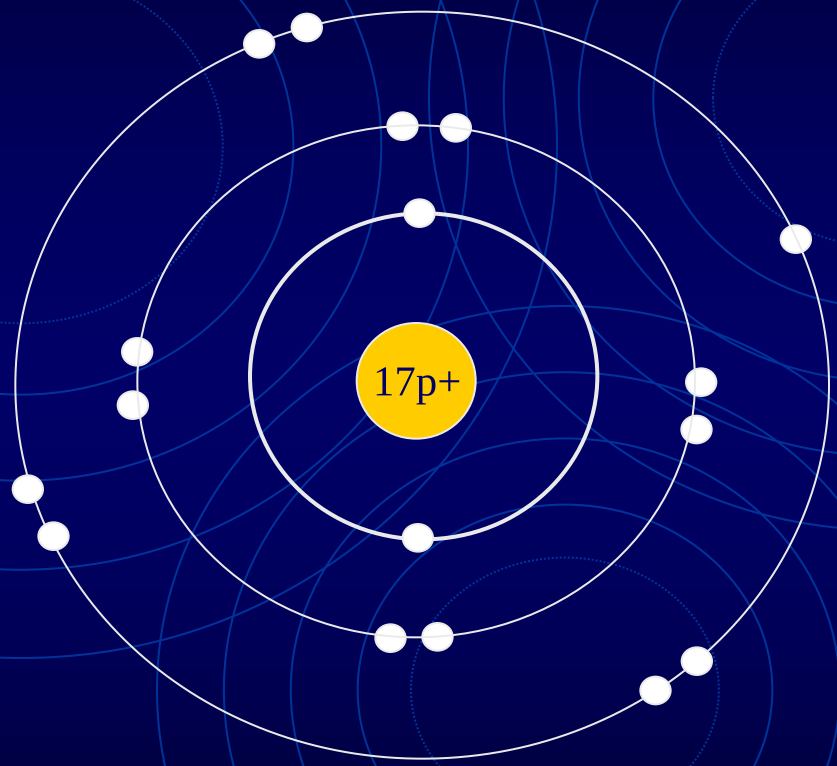
Remaining e- are pulled in closer to the nucleus. Ionic size decreases.

Valence e- lost in ion formation

Result: a smaller sodium cation, Na^+

Anion Formation

Chlorine atom with 7 valence e-



A chloride ion is produced. It is larger than the original atom.

One e- is added to the outer shell.

Effective nuclear charge is reduced and the e- cloud expands.