Chemical Bonding

TRENDS OF THE PERIODIC TABLE

CHEMISTRY 11

Trends of the Periodic Table

- 3 factors are usually discussed when explaining trends
  - nuclear charge
  - "n" value (outer most filled shell)
  - Inter-electron repulsion (usually not as important as the other 2)
- Important to remember that trends do not always work! – they are generalizations

Nuclear Charge

Along a Period (from left to right):

- atomic number increases (more +ve charges)
- # of shielding electrons stay the same (any electrons closer to the nucleus)

Shielding Electrons: - electrons in lower energy orbitals that block the pull of the protons

- campfire analogy – if someone/ something is between you and the fire, you feel less heat!

Overall: pull on electrons from positive charge increases

Nuclear Charge

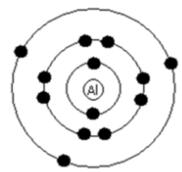
Down the Table (from Top to Bottom):

- atomic number increases (more +ve charges)
- # of shielding electrons increases (more people are blocking the fire)

Overall: pull on electrons from positive charge doesn't drastically increase

"n" Value

- As more electrons are added, more orbitals are occupied that are higher energy
- Higher energy orbitals are FARTHER away from the nucleus (remember Bohr's Theory!)



Inter-electron repulsion

- As you add more electrons to an atom or orbital, they repel each other (move farther apart)

Atomic Radius - Left to Right

- nuclear charge
  - More protons, pull electrons closer
    ∴ atom is smaller
- "n" value (outer most filled shell)
  - Stays the same in the same period
    ∴ no change
- Inter-electron repulsion
  - More electrons, electrons push farther away
    ∴ atom is larger

So what actually happens?

Overall: as you go from left to right along the periodic table...

→

The atomic radius DECREASES

Atomic Radius - Top to Bottom

- nuclear charge
  - More protons, but more shielding electrons
    ∴ atom ~ same size
- "n" value
  - Increases with each period you go down the table
    ∴ atom is larger
- Inter-electron repulsion
  - More electrons, electrons push farther away
    ∴ atom is larger

So what actually happens?

Overall: as you go from the top to the bottom of the periodic table…

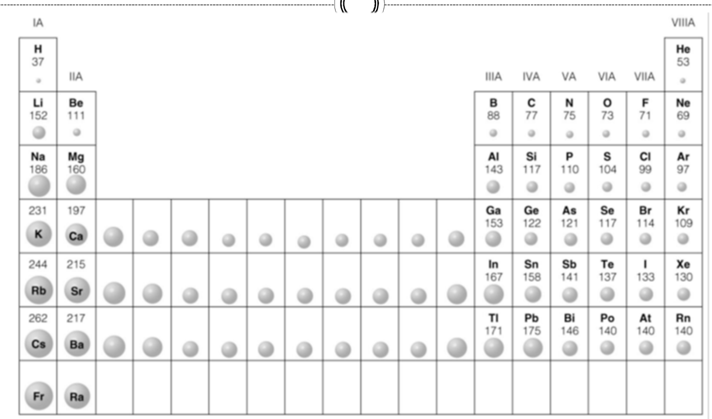
The atomic radius INCREASES ↓

Examples:

- What is the largest atom?

  Fr
- Which is larger:
  - F or B ? B
  - Ba or Be ? Ba
  - Sr or P ? Sr

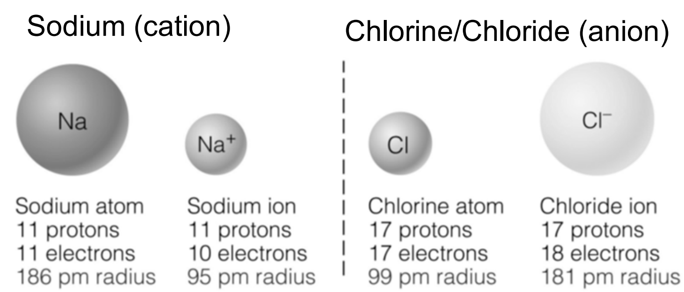
What happens in reality:



What About Ions?

- <u>Cations</u>:
  - Lose outer electrons
    - makes smaller, less electron repulsion, net positive charge pulls electrons closer
    - ∴ A CATION is SMALLER than the atom it forms from
- <u>Anions</u>:
  - Gain electrons
    - More electrons to repel each other
    - ∴ An ANION is LARGER than the atom it forms from

Examples:



2) Ionization Energy:

- Energy needed to remove an electron from an atom

  <u>First Ionization</u>:

  $$X_{(g)} \rightarrow X^{+}_{(g)} + e^{-}$$

- It ALWAYS takes energy to remove an electron (break its attraction to the nucleus)

Ionization Energy – Left to Right

- nuclear charge
  - More protons, attract electrons more
  - ∴ electron is harder to remove
- "n" value
  - Stays the same in the same period
  - ∴ no change
- Inter-electron repulsion
  - More electrons, electrons push farther away
  - ∴ electron is easier to remove

So what actually happens?

Overall: as you go from left to right along the periodic table…

The Ionization Energy INCREASES

Ionization Energy - Top to Bottom

- nuclear charge
  - More protons, but more shielding electrons
    ∴ same attraction to electrons
- "n" value
  - Increases with each period you go down the table
    ∴ electrons are farther away from protons and easier to remove
- Inter-electron repulsion
  - More electrons, electrons push farther away
    ∴ electrons easier to remove

So what actually happens?

Overall: as you go from the top to the bottom of the periodic table…

The Ionization energy DECREASES

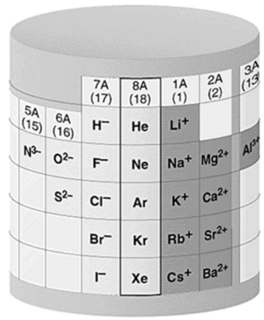
Ionization Energy Cont.

- MINIMUM for ALKALI METALS (easiest to remove electrons)
  - Have 1 electron in outer orbital, if removed are left with a full outer orbital (Noble Gas configuration)
- MAXIMUM for NOBLE GASES
  - Have full outer orbitals, so are especially stable

As you already know…

- Representative elements tend to gain or lose electrons until they become isoelectronic with the nearest Noble gas

(ie. They have the same electron configuration!)



So…

- Once an atom has become isoelectronic with a Noble Gas (has a full outer orbital) it is very hard to remove another electron

  ∴ ionization energy for that ion is VERY high

- Each successive ionization takes more energy
  - more net positive charge pulling on each electron, so they are harder to remove

Examples:

- Which atom has the highest ionization energy?

  He

- Which has the lowest ionization energy :
  - F or B ? B
  - Ba or Ba^{2+} ? Ba
  - Sr or P ? Sr

3) Electronegativity

- The ability of an atom to attract electrons in a chemical bond

- If an atom has HIGH electronegativity, it strongly attracts electrons from other atoms and might even remove the electron completely (forming ions)

Electronegativity – Left to Right

- nuclear charge
  - More protons, attract electrons more
    ∴ easier to remove electrons from other atoms
- "n" value
  - Stays the same in the same period
    ∴ no change
- Inter-electron repulsion
  - Adding more electrons would cause more inter-electron repulsion
    ∴ harder to remove electrons from other atoms

So what actually happens?

Overall: as you go from left to right along the periodic table…

The Electronegativity INCREASES

Electronegativity - Top to Bottom

- nuclear charge
  - More protons, but more shielding electrons
    ∴ not distinct change
- "n" value
  - Increases with each period you go down the table
  - electrons are farther away from protons
    ∴ harder to remove electrons from other atoms
- Inter-electron repulsion
  - More electrons, electrons push farther away
    ∴ harder to remove electrons from other atoms

So what actually happens?

Overall: as you go from the top to the bottom of the periodic table…

The Electronegativity DECREASES

Measuring Electronegativity:

- Pauling scale is most commonly used to quantify electronegativity
- Pauling assigned Fluorine (the most electronegative element) a value of 4.0
- Values range down to Cs and Fr, which are the least electronegative (value of 0.7)

Electronegativity Cont…

- VERY similar to ELECTRON AFFINITY (the energy change when an electron is accepted)
  - they follow the same trend
- <u>NOTE</u>: link to ionization energy
  - If the atom strongly attracts other electrons, it also strongly attracts its own valence electrons

∴ if an atom has a *<u>high electronegativity</u>*, it will also have a *<u>high ionization energy</u>*

TREND SUMMARY