Elements and Trends

p,n,e

- Atoms are made up of 3 particles: protons, neutrons, and electrons.
- Protons have a positive charge and are found in the nucleus.
- Neutrons have no charge and are found in the nucleus as well.
- Electrons have a negative charge and are found orbiting around the nucleus.
- Electrons do all the bonding, and protons hold them there.

Atomic Number and Mass Number

- The atomic number is always equal to the number of protons. Its what makes elements different from each other.
- Protons and neutrons each have a mass of 1 atomic mass unit.
- Electrons have a mass 1/2000 of a proton or neutron.
- The mass number is the number of protons plus the number of neutrons. Electrons don’t have enough mass to be counted.

Charge

- A charged particle is called an ion, a neutral particle is called an atom.
- A charge comes from having a different number of protons and electrons.
- Having more electrons than protons means you have more negative charges, and therefore a negative ion.
- Having fewer electrons than protons means you have more positive charge than negative, and therefore a positive ion.
- Atoms can only gain or lose electrons, never protons!

Symbol Writing

Mass Number

Charge

$$^{33}_{16}S^{-2}$$

Atomic Number

Isotopes

- Elements that have a different number of neutrons are called isotopes.
- This only changes their mass #, but that also changes the name.
- When writing an element name (especially when we do nuclear) you always write the name, followed by the mass #.
- Ex. Sulfur-32, or Sulfur-33

The Periodic Table

- Early chemists began to notice similarities in the properties of the elements so they looked to neatly classify them into a usable grouping.
- They made rows that represented an increase in the atomic mass, and they made columns based on similar chemical properties.
- Mendeleev got credit for the first periodic table because his was the first that could be used to predict the properties and existence of unknown elements.
- He was not the first, and his was not like ours now!

Metals vs. nonmetals

- The periodic table is broken into three major categories by their properties.
- Metals – left of staircase – elements give up electrons easily (+ ions), are conductive, malleable, shiny.
- Non-metals – right of staircase – elements gain electrons easily (- ions) , are insulators, brittle, dull.
- Semi-metals (metalloids) – on the staircase (B,Si,As,Ge,Te,Sb), and have varying properties of metals and non-metals.

Valence electrons

- The modern table puts the elements in order by atomic number (protons), and the columns are determined by the number of valence electrons (for now).
- Valence electrons are electrons in the outermost energy level (shell)
- This only applies to columns 1,2,13-18. We'll deal with the transition metals and rare earth metals later in the course.
- The number of valence electrons determines the chemical properties that the element will have.

Names of groups

- A row on the table is called a “period”. (Same energy level)
- A column is called a group or family (similar properties)
- Groups are given names to describe their reactions quicker.
  - Group 1 – Alkali metals
  - Group 2 –Alkaline Earth metals
  - Group 17 – Halogen
  - Group 18 – Noble gases
  - Groups 3-12 – Transition metals
  - Bottom 2 rows – Rare Earth Metals

Periodic Trends

- As you look at the properties of elements when you go in order of increasing atomic mass, trend become apparent.
- These trends are predictable and repeat as you go from row to row.
- They are:
  - Oxidation # = charge
  - Atomic Radius = size of atom
  - Electronegativity = attraction for other electrons
  - Ionization energy = energy need to remove electrons
  - Ionic radius = radius after electrons are gained or lost.

Oxidation

- Atoms want to have 8 valence electrons (called the octet). They can gain or lose electrons to get a perfect 8.
- If they have a few extra, they lose, and if they are a few short, they gain.
  - Group 1 = +1
  - Group 2 = +2
  - Group 13 = +3
  - Group 14 = +/-4
  - Group 15 = -3
  - Group 16 = -2
  - Group 17 = -1
  - Group 18 = 0

Atomic Radius

- Down a column, radius increases. Number of energy levels increases making your radius get much larger.
- Across a row – radius decreases. You are in the same energy level (same maximum size), but the number of protons is increasing.
- The attraction each electron has for the protons far outweighs the repulsion electrons have for each other.
- This causes the size to decrease as you add more protons to the nucleus.

Electronegativity

- As the size of the atom goes up, its ability to attract electrons goes down.
- Smaller size, higher electronegativity (other electrons can get closer to the nucleus and therefore be attracted.)
- Down a column – electronegativity goes down
- Across a row – electronegativity goes up.
- Noble gases – have zero electronegativity since they have no desire to gain electrons ever!

Ionization energy

- As the atom gets bigger, the amount of energy needed to remove an electron goes down.
- That electron is less attracted to the nucleus, so easier to pull away.
- Down a column – ionization energy goes down
- Across a row, ionization energy goes up.
- Noble gases have a very high ionization energy, they don’t want to give up electrons either!

Ionic Radius

- When atoms lose electrons, they will lose what is necessary to go to a lower energy level. (makes them smaller)
- When they gain electrons, they fill that energy level (makes them a little larger)
- Going down a column – ions get larger
- Going across a row, +ions are smaller than negative ions)
- + ions get smaller as you go right (more p+ for same # e^-)
- - ions get smaller as you go right (more p+)