The Periodic Table Class Notes

Development of the Periodic Table:

The modern periodic table of elements is mostly the work of Russian chemist Dmitri Mendeleev (1834-1907). It was set up in such a way so group elements in columns according to their properties, but also put them in ascending order according to their atomic number.

- Developed table to illustrate periodic trends in the properties of the then-known elements.
- Predicted some properties of then-unknown elements that would be expected to fill gaps in this table.
- Most of his predictions were proved correct when the elements in question were subsequently discovered.

Periodic Law: In 1869, **Dmitri Mendeleev** and **Lothar Meyer** individually came up with their own periodic law That said:

When elements R arranged in order of increasing atomic #, there is a periodic repetition of their physical + chemical properties

The periodic table was eventually reorganized according to Atomic Number (# of protons)

Basic info found on the periodic table:

EXAMPLE: Arsenic

Symbol ~ As

e- dot diagram~ ·As·

Protons ~ 33

Electrons ~ 33

Neutrons ~ 42

atomic # ~ 33

atomic mass ~ 74.9216amu

e- configuration ~ 2-8-18-5

Valence Shell /valence PEL ~ $4s^2 4p^3$

Valence electrons ~ 5

Total # of P E L' s ~ 4

2. Periodic Table General Information:

- ***Identifying phases*** of elements at room temperature
  Liquids ~ Hg (metal) & Br (non metal)

  Gases ~ O N Cl F H Noble Gases

  All others are ***solids*** …

  *identified by color

- Diatomic elements~ Br_2 I_2 N_2 Cl_2 H_2 O_2 F_2

- monatomic elements ~ all others

- Radioactive Elements (unstable nucleus)
  all elements with atomic number ABOVE 82!

3. Organization of the Periodic Table:

The vertical columns are called groups
same # of PEL's occupied, tells the valence shell

The horizontal rows are called periods.

similar properties & reactivity due to same number of valence electrons

Representative elements: groups of elements designated with an A (1A – 8A) and exhibit the entire range of chemical properties

Transition elements: several series of elements in which inner orbitals (d or f orbitals) are being filled

3 classifications of the elements: 1. Metals 2. Non-metals 3. Metalloids

1. METALS – Characteristics of metals:
   - Left of Staircase!
   a. good conductors of heat + electricity
   b. Form cations (lose electrons) + ions
   c. solid (except for mercury)
   d. Shiny when smooth + clean
   e. malleable - can be made into flat sheets
   f. ductile - can be made into thin wires

   a. alkali metals: Group 1A: Li, Na, K, Rb, Cs, Fr. Does not include hydrogen.
      - very reactive because have 1 valence electron
      - Not found as free elements
      - s^1 sublevel

   b. alkaline earth metals: Group 2A: Be, Mg, Ca, Sr, Ba, Ra
      - not as reactive as alkali metals, but still reactive + bonds easily
      - 2 valence electrons
      - s^2 sublevel
      - Not found as free elements

   c. transition metals:
      - have a special electron arrangement where 2 of their outer levels are not full.
      - d sublevel
      - used in construction, coins, jewelry * colored ions

   d. inner-transition: f-block, lanthanides and actinides;
      Actinides are radioactive.

2. NONMETALS → Right of Staircase
   - Poor conductors of heat + electricity
   - mostly gas, when solid are dull + brittle. (Br only liquid)
   - Forms anions (gain e-)

a. halogens: has 7 valence electrons (almost full)
- Often Bond with alkali metals to form octet (8 valence e-)
- F, Cl, Br, I, At

b. noble gases (inert gases): 8A group: Include: He, Ne, Ar, Kr, Xe, Rn.
- has 8 valence electrons
- do not gain/lose electrons because already have 8 valence electrons

3. METALLOIDS (semiconductors) –
- Partially conduct
- has properties of both metals + non-metals
- used in the computer chip industry.

*As you go across the periodic table, the number of valence electrons is equal to the group number for the s- and p-block ← valence electrons determine chemical properties!

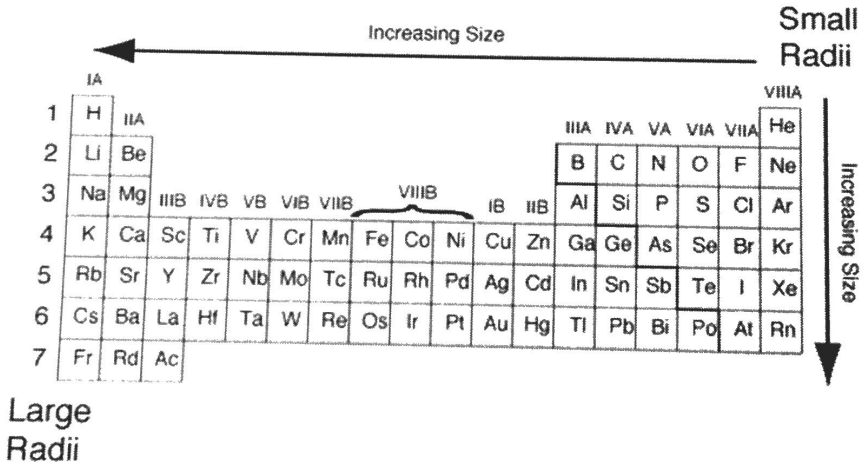
*Atoms in the same group have similar chemical properties because they have the same number of valence electrons

Periodic Trends

nuclear charge - charge on nucleus

Shielding -

(1.) Atomic Radii: size of atom



Group Trends: size increases as you go down a group. b/c of additional e- added

Periodic Trends: decreases across a period. Greater charge to nucleus (nuclear hold)

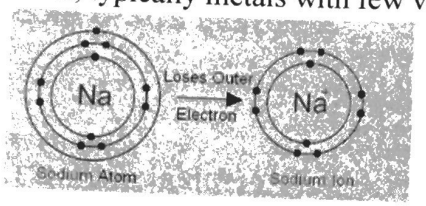
(2.) Ionization Energy: energy needed to REMOVE valence e-

- What is an oxidation state? The charge an atom will become (form an ion) during the process of bonding indicating the gain or loss of electrons.

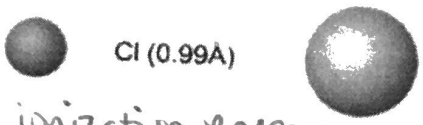
Ionic bond: attraction between a positive ion and a negative ion

ion:

cation: atom will lose electrons; typically metals with few valence electrons; ion is now smaller than original atom



anion: atom will gain electrons; typically non metals with many valence electrons; ion is now larger than original atom



- Noble gases have highest ionization energy

octet rule: atoms tend to gain, lose, or share electrons in order to acquire a full set of 8 valence electrons (atoms want to be like the noble gases)

- Predicting group locations from oxidation states
  - Check out an elements number of valence electrons!
    - 3 or less → tend to lose valence electrons
    - 5 or more → tend to gain additional valence electrons
    - 4 electrons → it depends on the bonding situation!

Group: down group IE↓
E further + easier to remove

3. Electronegativity: tendency to attract e-

Period: Across AR dec +

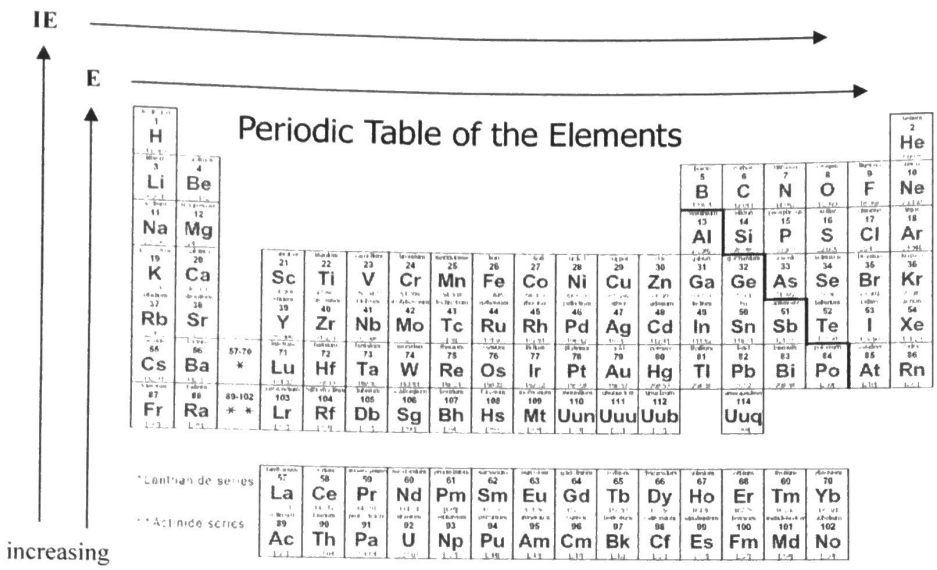
Group Trends: ↓ b/c greater energy levels + shielding Size↓ so greater nuclear hold

Periodic Trends: ↑ b/c greater # of VE + greater nuclear hold

- more likely to lose e-

- Fluorine is the most electronegative element

E = electronegativity IE = ionization energy



4. **Metallic Character:** tendency of atom to lose e⁻

5. **Non-metallic character:** tendency of atom to gain e⁻

In SUMMARY:

ALL Changes in periodic trends are due to

- ***Nuclear hold*** ~ the hold of the positively charged nucleus on the electrons

- ***Shielding effect*** ~ effective nuclear charge on the electron cloud in any atom with more than one electron shell. Greater number of PEL's = greater shielding effect

Examining the trends:

#of occupied PEL

down a group	INCREASE	More PEL's
Across a period	REMAIN THE SAME	Same PEL