

## Chapter 5 - The Periodic Law

### **5-1 History of the Periodic Table**

#### I. Mendeleev's Periodic Table

##### A. Organization

1. Vertical columns in atomic weight order
  - a. Mendeleev made some exceptions to place elements in rows with similar properties
    - (1) Tellurium and iodine's places were switched
2. Horizontal rows have similar chemical properties

##### B. Missing Elements

1. Gaps existed in Mendeleev's table
  - a. Mendeleev predicted the properties of the "yet to be discovered" elements
    - (1) Scandium, germanium and gallium agreed with predictions

##### C. Unanswered Questions

1. Why didn't some elements fit in order of increasing atomic mass?
2. Why did elements exhibit periodic behavior?

#### II. Moseley and the Periodic Table (1911)

##### A. Protons and Atomic Number

1. Xray experiments revealed a way to determine the number of protons in the nucleus of an atom
2. The periodic table was found to be in atomic number order, not atomic mass order
  - a. The tellurium-iodine anomaly was explained

##### B. The Periodic Law

1. The physical and chemical properties of the elements are periodic functions of their atomic numbers

\*\*\*Moseley was killed in battle in 1915, during WWI. He was 28 years old

#### III. The Modern Periodic Table

##### A. Discovery of the Noble Gases

1. 1868 - Helium discovered as a component of the sun, based on the emission spectrum of sunlight
2. 1894 - William Ramsay discovers argon
3. 1895 - Ramsay finds helium on Earth
4. 1898 - Ramsay discovers krypton and xenon
5. 1900 - Freidrich Dorn discovers radon

##### B. The Lanthanides

1. Early 1900's the elements from cerium (#58) to lutetium (#71) are separated and identified

##### C. The Actinides

1. Discovery (or synthesis) of elements 90 to 103

##### D. Periodicity

1. Elements with similar properties are found at regular intervals within the "periodic" table

## 5-2 Electron Configuration and the Periodic Table

### I. Periods and the Blocks of the Periodic Table

#### A. Periods

1. Horizontal rows on the periodic table
2. Period number corresponds to the highest principal quantum number of the elements in the period

#### B. Sublevel Blocks

1. Periodic table can be broken into blocks corresponding to *s*, *p*, *d*, *f* sublevels

### II. Blocks and Groups

#### A. s-Block, Groups 1 and 2

1. Group 1 - The alkali metals
  - a. One *s* electron in outer shell
  - b. Soft, silvery metals of low density and low melting points
  - c. Highly reactive, never found pure in nature
2. Group 2 - The alkaline earth metals
  - a. Two *s* electrons in outer shell
  - b. Denser, harder, stronger, less reactive than Group 1
  - c. Too reactive to be found pure in nature

#### B. d-Block, Groups 3 - 12

1. Metals with typical metallic properties
2. Referred to as "transition" metals
3. Group number = sum of outermost *s* and *d* electrons

#### C. p-Block elements, Groups 13 - 18

1. Properties vary greatly
  - a. Metals
    - (1) softer and less dense than d-block metals
    - (2) harder and more dense than s-block metals
  - b. Metalloids
    - (1) Brittle solids with some metallic and some nonmetallic properties
    - (2) Semiconductors
  - c. Nonmetals
    - (1) Halogens (Group 17) are most reactive of the nonmetals

#### D. f-Block, Lanthanides and Actinides

1. Lanthanides are shiny metals similar in reactivity to the Group 2 metals
2. Actinides
  - a. All are radioactive
  - b. Plutonium (94) through Lawrencium (103) are man-made

## 5-3 Electron Configuration and Periodic Properties

## I. Atomic Radii

### A. Atomic Radius

1. One half the distance between nuclei of identical atoms that are bonded together

### B. Trends

1. Atomic radius tends to decrease across a period due to increasing positive nuclear charge
2. Atomic radii tend to increase down a group due to increasing number energy levels (outer electrons are farther from the nucleus)

## II. Ionization Energy

### A. Ion

1. An atom or a group of atoms that has a positive or negative charge

### B. Ionization

1. Any process that results in the formation of an ion

### C. Ionization Energy

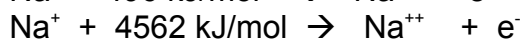
1. The energy required to remove one electron from a neutral atom of an element, measured in kilojoules/mole (kJ/mol)



### D. Trends

1. Ionization energy of main-group elements tends to increase across each period
  - a. Atoms are getting smaller, electrons are closer to the nucleus
2. Ionization energy of main-group elements tends to decrease as atomic number increases in a group
  - a. Atoms are getting larger, electrons are farther from the nucleus
  - b. Outer electrons become increasingly more shielded from the nucleus by inner electrons
3. Metals have a characteristic low ionization energy
4. Nonmetals have a high ionization energy
5. Noble gases have a very high ionization energy

### E. Removing Additional Electrons

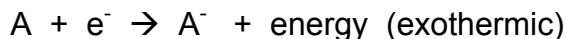


1. Ionization energy increases for each successive electron
2. Each electron removed experiences a stronger effective nuclear charge
3. The greatest increase in ionization energy comes when trying to remove an electron from a stable, noble gas configuration

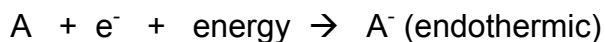
### III. Electron Affinity

#### A. Electron Affinity

1. The energy change that occurs when an electron is acquired by a neutral atom, measured in kJ/mol
  - a. Most atoms release energy when they acquire an electron



- b. Some atoms must be forced to gain an electron



#### B. Trends

1. Halogens have the highest electron affinities
2. Metals have characteristically low electron affinities
3. Affinity tends to increase across a period
  - a. Irregularities are due to the extra stability of half-filled and filled sublevels
4. Electron affinity tends to decrease down a group
5. Second electron affinities are always positive (endothermic)

### IV. Ionic Radii

#### A. Cations

1. Positive ions
2. Smaller than the corresponding atom
  - a. Protons outnumber electrons
  - b. Less shielding of electrons

#### B. Anions

1. Negative ions
2. Larger than the corresponding atoms
  - a. Electrons outnumber protons
  - b. Greater electron-electron repulsion

#### C. Trends

1. Ion size tends to increase downward within a group

### V. Valence Electrons

#### A. Valence Electrons

1. The electrons available to be lost, gained, or shared in the formation of chemical compounds
2. Main group element valence electrons are outermost energy level *s* and *p* sublevels

Group #	1	2	13	14	15	16	17	18
Number of valence Electrons	1	2	3	4	5	6	7	8

## VI. Electronegativity

### A. Electronegativity

1. A measure of the ability of an atom in a chemical compound to attract electrons
2. Elements that do not form compounds are not assigned electronegativities

### B. Trends

1. Nonmetals have characteristically high electronegativity
  - a. Highest in the upper right corner
2. Metals have characteristically low electronegativity
  - a. Lowest in the lower left corner of the table
3. Electronegativity tends to increase across a period
4. Electronegativity tends to decrease down a group of main-group elements

## VII. Periodic Properties of the *d*- and *f*- Block Elements

### A. Atomic Radii

1. Smaller decrease in radius across a period within the *d*- Block than within the main-group elements
  - a. Added electrons are partially shielded from the increasing positive nuclear charge
  - b. Slight increase at the end of the *d*-Block is due to electron-electron repulsion
2. Little change occurs in radius across an *f*-block of elements

### B. Ionization Energy

1. Tends to increase across *d*- and *f*-Blocks

### C. Ion Formation and Ionic Radii

1. Electrons are removed from the outermost energy level *s*-sublevel first
  - a. Most *d*-block elements form 2+ ions (losing 2 *s* electrons)
2. Ions of *d*- and *f*-Blocks are cations, smaller than the corresponding atoms

### D. Electronegativity

1. Characteristically low electronegativity of metals
2. Electronegativity increases as atomic radius decreases