# Unit 2 - Periodic Behavior and Ionic Bonding

**6.1 Organizing the Elements**

I. The Periodic Law
   A. The physical and chemical properties of the elements are periodic functions of their atomic numbers
   B. Elements on the table are arranged in order of increasing atomic number (number of protons)

II. Metals, Nonmetals, and Metalloids
   A. Metals
      1. Good conductors of heat and electricity
      2. Lustrous (shiny)
      3. Solids (except mercury)
      4. Ductile (can be drawn into wire)
      5. Malleable (can be hammered into thin sheets)
   B. Nonmetals
      1. Poor conductors of heat and electricity
      2. Most are gaseous
      3. Solids tend to be brittle
   C. Metalloids
      1. Some properties of metals, some of nonmetals

**6.2 Classifying the Elements**

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 meta**ls
         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

**6.3 Periodic Trends**

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. Trends in 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)

$$A + \text{energy} \rightarrow A^{+} + e^{-}$$

  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

$$Na + 496 \text{ kJ/mol} \rightarrow Na^{+} + e^{-}$$
$$Na^{+} + 4562 \text{ kJ/mol} \rightarrow Na^{++} + e^{-}$$
$$Na^{++} + 6912 \text{ kJ/mol} \rightarrow Na^{+++} + e^{-}$$

    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. Trends in Ionic Size
  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

IV. Trends in 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

**7.1 Ions**

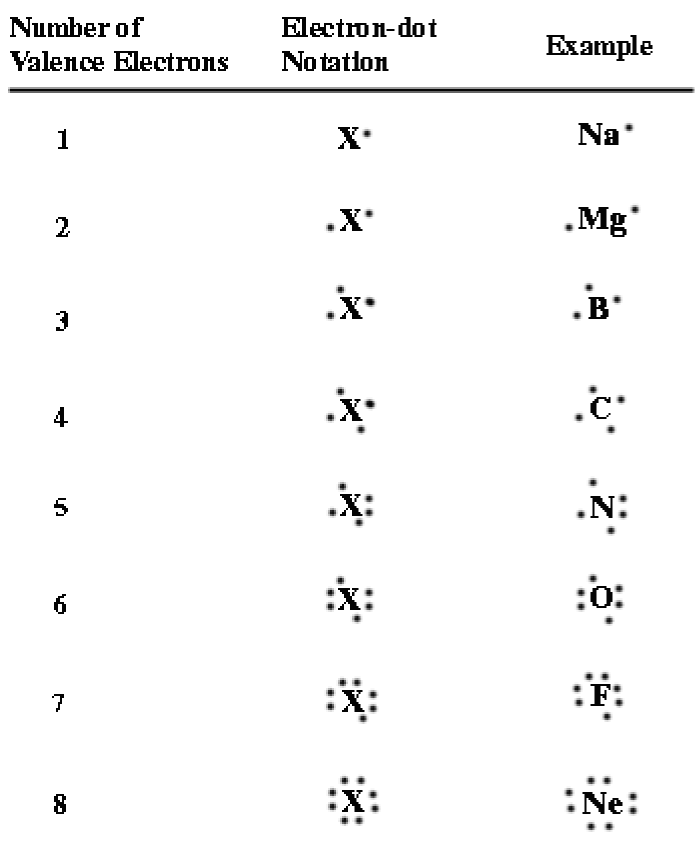
I. Valence Electrons – Outer energy level electrons

A. Electron-dot Notation

1. An electron-configuration notation in which only the valence electrons of an atom of a particular element are shown, indicated by dots placed around the element's symbol
2. Inner shell electrons are not shown
3. These are the electrons usually involved in the formation of covalent bonds

B. The Octet Rule – Ionic Compounds

1. Ionic compounds tend to form so that each atom, by gaining, or losing electrons, has an octet of electrons in its highest occupied energy level



| Number of Valence Electrons | Electron-dot Notation | Example |
|---|---|---|
| 1 | X | Na |
| 2 | X | Mg |
| 3 | X | B |
| 4 | X | C |
| 5 | X | N |
| 6 | X | O |
| 7 | X | F |
| 8 | X | Ne |

II. Formation of Ions

A. Electron Configuration Changes

1. Cations LOSE their valence electrons to attain a noble-gas configuration
2. Formation of a sodium ion

$$Na = 1s^2 2s^2 2p^6 3s^1 \rightarrow Na^+ = 1s^2 2s^2 2p^6$$

3. Anions GAIN electrons to complete their valence shell noble-gas configuration
4. Formation of chloride ion

$$Cl = 1s^2 2s^2 2p^6 3s^2 3p^5 \rightarrow Cl^- = 1s^2 2s^2 2p^6 3s^2 3p^6$$

# Common Ions and Their Charges

| Monatomic Cations | Name |
|---|---|
| $H^+$ | Hydrogen |
| $Li^+$ | Lithium |
| $Na^+$ | Sodium |
| $K^+$ | Potassium |
| $Mg^{2+}$ | Magnesium |
| $Ca^{2+}$ | Calcium |
| $Ba^{2+}$ | Barium |
| $Al^{3+}$ | Aluminum |

| Monatomic Anions | Name |
|---|---|
| $F^-$ | Fluoride |
| $Cl^-$ | Chloride |
| $Br^-$ | Bromide |
| $I^-$ | Iodide |
| $O^{2-}$ | Oxide |
| $S^{2-}$ | Sulfide |
| $N^{3-}$ | Nitride |
| $P^{3-}$ | Phosphide |

**7.2 Ionic Bonds and Ionic Compounds**

I. Introduction
  A. Ionic Compounds
    1. A compound composed of positive and negative ions that are combined so that the numbers of positive and negative charges are equal
      a. Most are crystalline solids
      b. Examples include NaCl, $MgBr_2$, $Na_2O$
  B. Formula Unit
    1. The simplest collection of atoms from which an ionic compound's formula can be established

II. Formation of Ionic Compounds
  A. Electron Configuration Changes
    1. Electrons are transferred from the highest energy level of one atom to the highest energy level of a second atom, creating noble gas configurations in all atoms involved
    2. Formation of sodium chloride
      a. Na = $3s^1$ Cl = $3s^2 3p^5$
      b. $Na^+ = 2s^2 2p^6$ $Cl^- = 3s^2 3p^6$

Oppositely charged ions come together in a ratio that produces a net charge = 0

$Na^+Cl^-$

III. Properties of Ionic Compounds

| | *Ionic Compounds* |
|---|---|
| ***Structure*** | Crystalline solids |
| ***Melting point*** | Generally high |
| ***Boiling Point*** | Generally high |
| ***Electrical Conductivity*** | Excellent conductors, molten and aqueous |
| ***Solubility in water*** | Generally soluble |

**7.3 Bonding in Metals**

I. The Metallic Bond Model
  A. Metallic Bonding
    1. The chemical bonding that results from the attraction between metal atoms and the surrounding sea of electrons
  B. Electron Delocalization in Metals
    1. Vacant *p* and *d* orbitals in metal's outer energy levels overlap, and allow outer electrons to move freely throughout the metal
    2. Valence electrons do not belong to any one atom

II. Metallic Properties

A. Metals are good conductors of heat and light
B. Metals are shiny
    1. Narrow range of energy differences between orbitals allows electrons to be easily excited, and emit light upon returning to a lower energy level
C. Metals are Malleable
    1. Can be hammered into thin sheets
D. Metals are ductile
    1. Ability to be drawn into wire
        a. Metallic bonding is the same in all directions, so metals tend not to be brittle
E. Metals atoms organized in compact, orderly crystalline patterns
F. Different metallic elements (and carbon) can be mixed to form alloys
    1. Sterling silver
        a. Ag = 92.5%, Cu = 7.5%
    2. Brass
        a. Cu = 60%, Zn = 40%