

Chapter 7: Periodic Properties of the Elements

Section 7.1: Development of the Periodic Table

- 1869: Dmitri Mendeleev and Lothar Meyer published nearly identical classification schemes
 - Both scientist noted that when elements are arranged by increasing weight, chemical and physical properties repeated
 - Mendeleev's theory that elements of the same properties were in the same family, forced him to leave blanks in his table. These blanks predicted new elements.
- 1913: Henry Moseley developed concept of atomic numbers
 - Moseley arranged elements in increasing atomic numbers, instead of increasing atomic weight

Section 7.2: Electron Shells and the Sizes of Atoms

- Electron Shells in Atoms
 - ↓ a column on the periodic table, we change the **principle quantum number, n** , of the valence orbitals of the atoms.
 - Using the periodic table we can predict **radical electron density** → the probability of an electron at a specific distance from the nucleus
 - On a graph (Fig. 7.3 pg. 230 in old book), maxima represent highest probability of finding electrons
 - One maxima for each electron shell for element graphed
 - The larger the nucleus, the greater the positive charge, and the closer the maxima are to the nucleus
- Atomic Sizes
 - **Apparent radii**: the closest distance separating two nuclei of same element in a collision, also known as the **nonbonding radii**
 - **Bonding atomic radius**: distance separating two atoms when they are chemically bonded
 - Radius ↑, ↓ the periodic table
 - Radius ↓, → the periodic table
 - Atomic radius predicts bond length, the shorter bond length generally the stronger the bond

Section 7.3: Ionization Energy

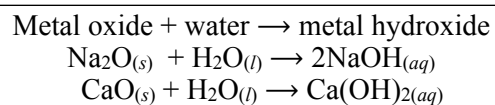
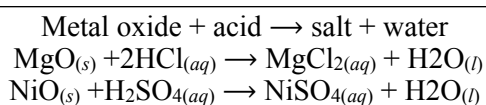
- **Ionization energy**: the minimum energy required to remove an electron from the ground state of an isolated gaseous atom or ion
 - *1st ionization energy, I_1* : the energy needed to remove the first electron from a neutral atom
 - $\text{Na}_{(g)} \rightarrow \text{Na}^+_{(g)} + e^-$
 - *2nd ionization energy, I_2* : the energy needed to remove the second electron
 - $\text{Na}^+_{(g)} \rightarrow \text{Na}^{2+}_{(g)} + e^-$
- The greater the ionization energy, the more difficult it is to remove an electron
 - The closer an atom is to having a full valence shell, the higher the ionization energy
 - Alkali metals have generally low ionization energies, noble gases have high ionization energies
 - The closer the valence shell is to the nucleus the higher the ionization energy, $\text{Ne } I_1 > \text{Ar } I_1$

Section 7.4: Electron Affinities

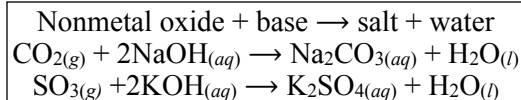
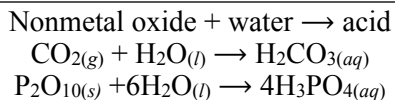
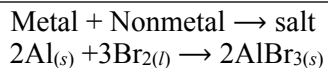
- **Electron affinity**: the energy change when an electron is added to a gaseous atom, generally energy is released
 - $\text{Cl}_{(g)} + e^- \rightarrow \text{Cl}^-_{(g)} \quad \Delta E = -349 \text{ kJ/mol}$
- The greater the attraction between an atom and an electron the more negative the electron affinity will be.
- Electron affinities change little ↓ a group, but become greater moving → the periodic table

Section 7.5: Metals, Nonmetals, and Metalloids

- **Metallic character**: the degree to which an element exhibits metallic behaviors, also *nonmetallic character* can be referred to.
 - Most metallic elements are shiny. Metals conduct heat and electricity. They are malleable and ductile. All except Hg are solids at room temp. Most have very high melting points.
 - Metals tend to have low ionization energies, oxidized in chemical reactions.
 - Metal-nonmetal compounds tend to be ionic ⇒ most metal oxides and halides are ionic solids.
 - Most metal oxides are **basic oxides** that dissolve in water to form metal hydroxides.
 - Metal oxides also react with acids to form salts and water



- Nonmetals when reacted w/ metals tend to gain electrons
 - Compounds formed completely of nonmetals are molecular substances. Most nonmetal oxides are **acidic oxides**



- Metalloids have properties intermediate between those of metals and nonmetals. Have some, lack others.
 - Many metalloids like Si are *semiconductors*, only conduct electricity in one direction.

Section 7.6: Group Trends for the Active Metals

- **Alkali metals:** group 1A (Li, Na, K, Rb, Cs, Fr)

Reactivity \uparrow, \downarrow the group

- Alkali metals are soft metallic solids
- Alkali metals are highly reactive forming 1+ ions $\Rightarrow \text{M} \rightarrow \text{M}^+ + \text{e}^-$

Flame Test Colors

Li - crimson red

Na - yellow

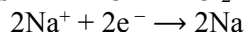
K - lilac

Calcium - brick red

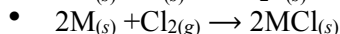
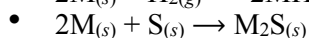
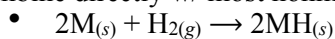
Strontium - crimson red

Barium - green

- They are only found in compounds in nature, *electrolysis* is used to obtain the metals
Electrolysis of salt, half equations $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$



- Combine directly w/ most nonmetals



In hydrides of alkali metals, hydrogen is present as the **hydride ion** H^-

- Alkali metals react violently w/ water: $2\text{M}_{(s)} + 2\text{H}_2\text{O}_{(l)} \rightarrow 2\text{MOH}_{(aq)} + \text{H}_{2(g)}$

- **Alkaline Earth Metals:** group 2A (Be, Mg, Ca, Sr, Ba, Ra)

- Compared w/ alkali metals more dense, melt at higher temperatures
- React less vigorously, and slower than alkali metals, Be does not react w/ water, Mg only reacts w/steam

Section 7.7: Group Trends for Select Nonmetals

- **Hydrogen:** does not truly belong to any family

- Nonmetal, found as diatomic gas
- Generally reacts w/ nonmetals to form molecular compounds, generally exothermic
 - $2\text{H}_{2(g)} + \text{O}_{2(g)} \rightarrow 2\text{H}_2\text{O}_{(l)} \quad \Delta H^\circ = -517.7 \text{ kJ}$

- **Group 6A: The Oxygen Group**

- \downarrow the group, metallic character increases
- Oxygen has two molecular forms, O_2 and O_3 ($\text{O}_3 = \text{ozone}$)
 - **Allotropes:** different forms of the same element in the same state

- **Group 7A: The Halogens**

- All halogens are typically nonmetals
- Each element consists of a diatomic atoms: $\text{F}_2, \text{Cl}_2, \text{Br}_2, \text{I}_2$
- Halogens have very negative electron affinities \therefore they accept electrons easily (only need one electron for full shell configuration)
 - $\text{X}_2 + 2\text{e}^- \rightarrow 2\text{X}^-$
- Chlorine is most industrially useful and is obtained through the electrolysis of brine (concentrated solution of NaCl)
 - $2\text{NaCl}_{(aq)} + \text{electricity} \rightarrow 2\text{NaOH}_{(aq)} + \text{H}_{2(g)} + \text{Cl}_{2(g)}$
- Halogens react directly w/ most metals to form ionic halides.
- Halogens react w/ hydrogen to form gaseous halide compounds
 - $\text{H}_{2(g)} + \text{X}_2 \rightarrow 2\text{HX}_{(g)}$

- **Group 8A: The Noble Gases**

- All are **monoatomic**, consisting of single atoms instead of molecules
- Their full valence shells mean that they are highly unreactive
 - Xe, the largest noble gas, reacts with PtF_6 to form $\text{XeF}_2, \text{XeF}_4$
 - Only stable compound with krypton is KrF_2
 - No known compounds have been formed w./ He, Ne, Ar as of yet