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
    - $Na_{(g)} \longrightarrow Na^{+}_{(g)} + e^{-}$
  - *2nd ionization energy, I_2*: the energy needed to remove the second electron
    - $Na^{+}_{(g)} \longrightarrow Na^{2+}_{(g)} + e^{-}$
- 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, Ne I_1 > 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
  - $Cl_{(g)} + e^{-} \longrightarrow Cl^{-}_{(g)}$ $\Delta E = -349$ 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

Metal oxide + acid ⟶ salt + water	Metal oxide + water ⟶ metal hydroxide
$MgO_{(s)} + 2HCl_{(aq)} \longrightarrow MgCl_{2(aq)} + H2O_{(l)}$	$Na_2O_{(s)} + H_2O_{(l)} \longrightarrow 2NaOH_{(aq)}$
$NiO_{(s)} + H_2SO_{4(aq)} \longrightarrow NiSO_{4(aq)} + H2O_{(l)}$	$CaO_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2(aq)}$

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

Metal + Nonmetal ⟶ salt	Nonmetal oxide + water ⟶ acid	Nonmetal oxide + base ⟶ salt + water
$2Al_{(s)} + 3Br_{2(l)} \longrightarrow 2AlBr_{3(s)}$	$CO_{2(g)} + H_2O_{(l)} \longrightarrow H_2CO_{3(aq)}$ $P_2O_{10(s)} + 6H_2O_{(l)} \longrightarrow 4H_3PO_{4(aq)}$	$CO_{2(g)} + 2NaOH_{(aq)} \longrightarrow Na_2CO_{3(aq)} + H_2O_{(l)}$ $SO_{3(g)} + 2KOH_{(aq)} \longrightarrow K_2SO_{4(aq)} + H_2O_{(l)}$

- 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 ↑, ↓ the group**
  - Alkali metals are soft metallic solids
  - Alkali metals are highly reactive forming 1+ ions ⇒ $M \longrightarrow M^+ + e^-$
    - They are only found in compounds in nature, *electrolysis* is used to obtain the metals
      Electrolysis of salt, half equations $2Cl^- \longrightarrow Cl_2 + 2e^-$
      $2Na^+ + 2e^- \longrightarrow 2Na$
    - Combine directly w/ most nonmetals
      - $2M_{(s)} + H_{2(g)} \longrightarrow 2MH_{(s)}$
      - $2M_{(s)} + S_{(s)} \longrightarrow M_2S_{(s)}$
      - $2M_{(s)} + Cl_{2(g)} \longrightarrow 2MCl_{(s)}$

      In hydrides of alkali metals, hydrogen is present as the **hydride ion** H^-
    - Alkali metals react violently w/ water: $2M_{(s)} + 2H_2O_{(l)} \longrightarrow 2MOH_{(aq)} + H_{2(g)}$

Flame Test Colors
Li - crimson red
Na - yellow
K - lilac
Calcium - brick red
Strontium - crimson red
Barium -green

- **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
    - $2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(l)}$ $\Delta H^o = -517.7$ kJ
- **Group 6A: The Oxygen Group**
  - ↓ the group, metallic character increases
  - Oxygen has two molecular forms, O_2 and O_3 (O_3 = **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: F_2, Cl_2, Br_2, I_2
  - Halogens have very negative electron affinities ∵ they accept electrons easily (only need one electron for full shell configuration)
    - $X_2 + 2e^- \longrightarrow 2X^-$
  - Chlorine is most industrially useful and is obtained through the electrolysis of brine (concentrated solution of NaCl)
    - $2NaCl_{(aq)} + \text{electricity} \longrightarrow 2NaOH_{(aq)} + H_{2(g)} + Cl_{2(g)}$
  - Halogens react directly w/ most metals to form ionic halides.
  - Halogens react w/ hydrogen to form gaseous halide compounds
    - $H_{2(g)} + X_2 \longrightarrow 2HX_{(g)}$
- **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 XeF_2, XeF_4
    - Only stable compound with krypton is KrF_2
    - No known compounds have been formed w./ He, Ne, Ar as of yet