7.1 – Development of the Periodic Table

Mendeleev and Mayer were founders of the periodic table and organized by atomic weight

Mosley developed the concept of atomic numbers

7.2 – Effective Nuclear Charge

**Effective Nuclear Charge (Z\text{eff})**: Strength of pull of electrons to nucleus

\[ Z_{\text{eff}} = Z - S \approx Z - \# \text{ of core electrons} \approx \# \text{ of valence electrons} \]

Increases down a group a little due to less shielding effect (more diffuse core electrons and increasing protons with no change in valence electrons) and across the group due to more positive protons in the nucleus

**Shielding Effect**: Electrons blocking positive nucleus’ pull

7.3 – Sizes of Atoms and Ions

**van der Waals Radius** – Nonbonding radius that is the shortest distance separating the two nuclei during a collision (twice the radii of the atoms) in solid state

**Bonding Atomic Radius/Covalent Radius** – Half of the nucleus-to-nucleus distance between two atoms of a covalent bond

**Metallic Diameter** – Distance between two nuclei in solid metal

Van der Waals Radius > Covalent Radius

Atomic radius increases down a group (due to higher quantum levels) and decreases across a group due to higher positive nucleus (increasing Z\text{eff})

Cations are smaller than their parent atoms and anions are larger

Anions are generally larger than cations

**Isoelectronic Series** – Ions that have the same number of electrons

7.4 – Ionization Energy
**Ionization Energy** – Amount of energy required to remove an electron from the ground state of an atom in the *gaseous state*

It requires more energy to remove each successive electron

There is more stability (higher ionization energy) when atoms have full or half-full orbitals

First ionization energy decreases down a group (less shielding effect) and increases across the row due to higher positive nucleus with protons (higher $Z_{\text{eff}}$)

To write an electronic configuration for a metal ion, write the configuration for the principal quantum numbers increasing and then remove electrons from right to left (d and f metals have close in energy valence energy levels that switch position for neutral atoms and cations, so this is why this must be done)

**7.5 – Electron Affinities**

**Electron Affinity** – Energy change that occurs when an electron is added to a gaseous atom

Electron affinities are positive for the noble gases

Electron affinities generally increase in magnitude of energy released across a period with exceptions for noble gases and elements that have stable p or s subshells

Electron affinities hardly change down a group

**7.6 – Metals, Nonmetals, and Metalloids**

Metal oxides are ionic solids that are *basic*

Metal oxide + Water $\rightarrow$ Metal hydroxide

Metal oxide + Acid $\rightarrow$ Salt + Water

Metals tend to form cations while nonmetals tend to form anions

Nonmetal oxides are molecular substances that form *acidic* solutions

Nonmetal Oxide + Water $\rightarrow$ Acid

(The nonmetal has the same oxidation number in the oxide and the acid: no REDOX)

Nonmetal Oxide + Base $\rightarrow$ Salt + Water

Metallic character increases down a group and decreases across the period
Mercury is the only liquid metal at room temperature

 Metals have low ionization energies

 Of the diatomic molecules, bromine is liquid and iodine is a volatile solid

 Nonmetals have very negative electron affinities (easily accept)

 Metalloids are shiny but brittle solids and can act as semiconductors

 **Metalloid oxides are weakly acidic and behave like analogs of the above group**

### 7.7 – Trends for Group 1A and Group 2A Metals

**Alkali Metals:**

Alkali metals are soft, metallic solids

Francium is an unstable alkali metal

Alkali metals have **low melting points** (decrease down a group), **low densities** (increase down a group), large atomic radii (increase down a group), and low ionization energies (decrease down a group)

Alkali metals are reactive with water and reactivity increases **down the group**

**The reaction with lithium and oxygen is the only “expected” one**

- *Sodium and oxygen gives a sodium peroxide*
  - *Potassium, rubidium, and cesium produce superoxides with an O\(_2^-\) anion*

Lithium (Bright Red), Sodium (Bright yellow), Potassium/Rubidium/Cesium (Pale Violet-Blue) flame tests

They are found only as compounds in nature

Alkali metals react with hydrogen gas to form metal hydrides

**Alkali and heavy alkaline earth metals (Li, Na, K, Rb, Cs, Ca, Sr, Ba) react with water to form strong bases (soluble metal hydroxides) and hydrogen gas.**

**Alkaline Earth Metals:**

**Have higher densities and melting points** than alkali metals

Have low ionization energies but not as low as group 1

Less reactive than alkali metals with reactivity still increasing down the group
Be has no reaction with water, Mg has hot steam, Ca/Sr/Ba have slow reactions

Strontium (Bright Red), Barium (Bright Green) flame tests

**Group 6A:**

**Allotrope** – A different form of elemental substance

Oxygen has two allotropes – O$_2$ and O$_3$

Oxygen forms two compounds with hydrogen – water and hydrogen peroxide

S$_8$ is the most stable allotrope of sulfur

Se, Te, and Po are silvery gray while sulfur is yellow

Group 6A elements are reduced when in contact with metals to form -2 anions

**React with hydrogen gas to form H$_2$X gases** (except oxygen which makes water)

**Are oxidized with oxygen gas to form dioxides**

**Halogens:**

Melting points and boiling points increase down the group unlike alkali metals

Fluorine and chlorine gas are pale, yellow gases

Bromine is a fuming brown liquid

Iodine is a fuming dark-violet solid

Halogenes are toxic in elemental poisons

Can serve as disinfectants in low concentrations

Dissolved chlorine reacts with water reversibly: Cl$_2$ (g) + H$_2$O (l) $\leftrightarrow$ HCl (aq) + HOCl (aq)

**Noble Gases:**

Heavy noble gases (Ar, Kr, Xe) can form a few covalent compounds with other nonmetals and are unstable as well as highly reactive: XeF$_2$, XeF$_4$, XeF$_6$, KrF$_2$, HArF