Chapter 7: Periodic Properties of the Elements

- **valence orbitals** –outer shell orbitals

7.1 Development of the Periodic Table

- 1869 Dmitri Medeleev in Russian, Lothar Meyer in Germany
  - similar chemical and physical properties are periodic when elements arranged with increasing atomic weight
- Henry Moseley (1887-1915)
  - Developed concept of atomic numbers
  - Determined frequencies of X-rays emitted from elements after being bombarded with high-energy electrons
  - Frequency increased with increasing atomic mass

7.2 Electron Shells and the Sizes of Atoms

7.2.1 Electron Shells in Atoms

- **radial electron density** – probability of finding the electron at a particular distance from the nucleus
- as nuclear charge increases, 1s shells pulled closer to nucleus

7.2.2 Atomic Sizes

- **atomic radius** – estimation of radius of atoms
- estimate by assuming that atoms are spheres that touch each other
- atomic radius increases down a group, decrease down row
- atomic radius affected by principal quantum number and effective nuclear charge
- increase principal quantum number, increases size of orbital
- increase effective nuclear charge, reduces size or orbital

7.3 Ionization Energy

- **ionization energy** – minimum energy required to remove an electron from the ground state
- **first ionization energy** – energy needed to remove the first electron
- **second ionization energy** – energy needed to remove second electron
- the greater ionization energy, harder it is to remove electrons
- $I_1 < I_2 < I_3$
- Positive nuclear charge remains same, number electrons decreases → effective nuclear charge increases
- Greater effective nuclear charge, greater energy required to remove electron
- Sharp increase in ionization energy when inner-shell electron removed
- Only outer most electrons involved in sharing and transfer of electrons in bonding and reactions

7.3.1 Periodic Trends in Ionization Energies

- I_1 increases with increasing atomic number
  - Alkali metals lowest ionization energy, noble gases greatest ionization energy
  - Ionization energy decreases down each group
  - Ionization energy of transition metals increase slowly across row
  - Ionization energy depends of effective nuclear charge and average distance of the electrons from the nucleus
  - Increase attraction of electrons to nucleus – increase effective nuclear charge or decrease distance from nucleus
  - Across period – increase in effective nuclear charge and decrease in atomic radius, increase in ionization energy
  - Down group – atomic radius increases, effective nuclear charge remains constant, ionization energy decreases

7.4 Electron Affinities

- **electron affinity** – energy change when electron added to atom
- greater attraction between atom and electron, the more negative electron affinity

- **across row** – becomes more negative
- electron affinity of noble gases is positive
- **down group** – does not change much

7.5 Metals, Nonmetals, and Metalloids

Metals	Nonmetals
Shiny luster, various colors, most silvery Solids are malleable and ductile Good conductors of heat and electricity Metal oxides are ionic solids that are basic Tend to form cations in a aqueous solution	Do not have luster; various colors Solids are brittle; some hard and soft Poor conductors of heat and electricity Nonmetallic oxides are molecular substances that are acidic For anions or oxyanions in aqueous solution

7.5.1 Metals
- low ionization energy
- **transition metal –** form more than one positive ion
- compounds of metals and nonmetals are ionic
- metal oxides dissolve in water to form metal hydroxides
- metal oxides react with acids to form salts and water

7.5.2 Nonmetals
- seven nonmetals exist under ordinary conditions as diatomic molecules
- H_2, O_2, N_2, F_2, Cl_2, $Br_{2(l)}$, $I_{2(s)}$
- Nonmetals gain electrons
- Compounds composed of all nonmetals are molecular substances
- Nonmetal oxides – acidic oxides

7.5.3 Metalloids
- properties of both metals and nonmetals

7.6 Group Trends for the Active Metals

7.6.1 Group 1A: The Alkali Metals
- soft metallic solids
- characteristics of metals
- low densities and melting points
- down group – increasing atomic radius, decreasing first ionization energy
- lowest first ionization energy
- most active metals
- combine directly with nonmetals
- react with hydrogen to form hydrides, with sulfur to form sulfides, with chlorine to form chlorides
- hydride ion – H^{-1}
- react vigorously with water – produce hydrogen gas and solutions of alkali metal hydroxides

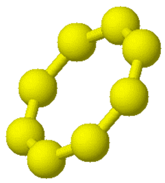
7.6.2 Group 2A: The Alkaline Earth Metals
- solids, harder, more dense, higher melting point that alkali metals
- tend to lose 2 electrons to form 2+ ions

7.7 Group Trends for Selected Nonmetals

7.7.1 Hydrogen
- ionization energy higher than active metals
- reacts with nonmetals to form molecular compounds
- reacts with active metals for form hydrides

7.7.2 Group 6A: The Oxygen Group

- oxygen has two forms – O_2 and O_3
- **allotropes** – different forms of the same element in the same state
- 21 percent dry air is O_2
- oxygen attracts electrons from other elements
- oxygen with other metals from oxides
- peroxide O_2^{2-} , superoxide O_2^{-}
    - compounds with these ions react with themselves to produce an oxide and O_2
- sulfur – several allotropic forms
- most common and most stable – S_8, usually written as S(s)
- sulfur and its compounds burned in oxygen to form sulfur dioxide



7.7.3 Group 7A: The Halogens

- salt formers
- halogens are typical nonmetals
- melting and boiling points increase with increasing atomic number
- exist as diatomic molecules
- have highly negative electron affinities
- gain electrons to form halide ions
- Fluorine removes electrons from almost any substance

7.7.4 Group 8A: The Noble Gases

- nonmetals, gas at room temperature
- monoatomic
- completely filled s and p subshells
- 1962 Neil Bartlett
    - first noble-gas compound by reacting Xe with F_2(g) to form XeF_2 and XeF_4 and XeF_6
- KrF_2, only single stable compound