

# Chapter 7: Periodic Properties of the Elements

## 7.1 Development of the Periodic Table

-1869: Dmitri Mendeleev (Russia) and Lothar Meyer (Germany) publish nearly identical schemes for classifying the elements

-Both had arranged the elements in order of increasing atomic weights

-Mendeleev predicted the existence and properties of Germanium (Ge) and Gallium (Ga). When they were discovered, their properties were found to closely match those predicted by Mendeleev

-1913: Henry Moseley discovers concept of atomic numbers. Observed that frequencies of X-rays were different for each element. Was able to arrange these frequencies in order by assigning each element a unique whole number, which he called the *atomic number*

## 7.2 Electron Shells in Atoms

-Orbitals with the same value of  $n$  are referred to as a *shell*

-When looking at a radial electron-density graph, the maxima (peaks) designate the areas with the higher probabilities of finding electrons

-The 1s shell in Argon is closer to the nucleus than that of Helium. This is because both 1s shells are the only ones not shielded from the nucleus by other electrons. However, since Argon has a stronger nuclear charge, it can attract its electrons more easily

## 7.3 Sizes of Atoms

-**Atomic radius:** radius of an atom of a given element

1. Within each group, the atomic radius tends to increase going from top to bottom
2. Within each period, the atomic radius tends to decrease moving left to right

-Two factors determine the size of the outermost orbital

1. Its principal quantum number (as it increases, the size of the orbital increases)
2. The effective nuclear charge (reduces the size of the orbital by pulling electrons closer)

## 7.4 Ionization Energy

-**Ionization energy:** energy required to remove an electron from a gaseous atom when the atom is in its ground state

-The *first* ionization energy,  $I_1$ , is the energy needed to remove the first electron. The *second* ionization energy,  $I_2$ , removes the second electron, and so on. The HIGHER the ionization energy, the MORE DIFFICULT it is to remove an electron.

-Every element exhibits a large increase in ionization energy when electrons are removed from its noble-gas core. The inner electrons are too tightly bound to the nucleus to be lost from the atom or even shared with another atom

### Periodic Trends in Ionization Energies

1. Within each period,  $I_1$  generally increases with increasing atomic number. The alkali metals show the lowest ionization energy in each row, and the noble gases the highest

2. Within each group, ionization energy generally decreases with increasing atomic number.

-Energy needed to remove electrons from outer shell depends on nuclear charge and average distance between of the electron from the nucleus

-As we move across a period, there is both an increase in effective nuclear charge and a decrease in atomic radius, causing the ionization energy to increase

-As we move down a column, the atomic radius increases, while the effective charge remains essentially constant. Thus, the attraction between the nucleus and the electron decreases in this direction, causing the ionization energy to decrease.

## 7.5 Electron Affinities

-**Electron affinity**: energy change that occurs when an electron is added to a gaseous atom or ion

-For most neutral atoms and for all positively charged ions, energy is released when an electron is added

-The greater the attraction between the species and the added electron, the more exothermic the process

-On the other hand, adding an electron to anions and to some atoms makes them unstable.

-The general trend for electron affinity is to become increasingly negative (stronger binding of electrons) as we move across each period toward the halogens. The halogens have the most negative electron affinities since they are one electron short of noble-gas configuration

-Electron affinities do not change greatly as we go down a group

## 7.6 Metals, Nonmetals, and Metalloids

Metallic elements	Nonmetallic elements
Distinguishing luster	Nonlustrous; various colors
Malleable and ductile as solids	Solids are usually brittle; may be hard or soft
Good thermal and electrical conductivity	Poor conductors of heat and electricity
Most metallic oxides are basic, ionic solids	Most nonmetallic oxides are molecular, acidic compounds
Exist in aqueous solution mainly as cations	Exist in aqueous solution mainly as anions or oxyanions

### Metals

-See properties above. All are solids at room temperature except mercury, which is a liquid.

-Metals tend to have low ionization energies and are consequently oxidized (lose electrons) when they undergo chemical reaction.

-Compounds of metals with nonmetals tend to be ionic substances

-Most metals are **basic oxides** (oxide that either reacts with water to form a base or reacts with an acid to form a salt and water)

### Nonmetals

-Generally have lower melting points than those of metals.

-Seven nonmetals exist as diatomic molecules under ordinary conditions ( $H_2$ ,  $N_2$ ,  $O_2$ ,  $F_2$ ,  $Cl_2$ ,  $Br_2$ ,  $I_2$ ). Bromine is liquid; Iodine is a volatile solid, the rest are gases. The remaining nonmetals are solids.

-Nonmetals, in reacting with metals, tend to gain electrons and become anions

- Compound composed entirely of nonmetals are molecular substances
- Most nonmetals are **acidic oxides** (oxide that forms acid when added to water; soluble nonmetal oxides are acidic oxides)

### Metalloids

- Have properties intermediate between those of metals and nonmetals. May have some metallic properties but lack others.
- Several metalloids are electrical semiconductors

### Trends in Metallic and Nonmetallic Character

-**Metallic character:** the extent to which an element exhibits the physical and chemical properties characteristic of metals, for example, luster, malleability, ductility, and good thermal and electrical conductivity.

1. Metallic character is strongest for the elements in the leftmost part of the periodic table and tends to decrease as we move to the right in any period
2. Within any group of representative elements, the metallic character increases progressively from top to bottom

## **7.7 Group Trends: The Active Metals**

### Group 1A: The Alkali Metals

- Metallic properties; silvery, metallic luster, high thermal and electrical conductivity
- Low densities and melting points; increasing atomic radius and decreasing ionization energy as we move down the family
- Lowest  $I_1$  values of the elements, which reflects the ease with which their outer electrons can be removed. This makes the alkali metals very reactive, readily losing one electron.
- In the hydrides of the alkali metals (LiH, NaH, and so forth), hydrogen is present as  $H^-$ , called the **hydride ion**.
- The alkali metals react vigorously with water, producing hydrogen gas and solutions of alkali metal hydroxides (very exothermic).

### Group 2A: The Alkaline Earth Metals

- Solids, typical metallic properties, denser and harder than alkali metals (also melt at higher temperatures)
- Less reactive than alkali metals (have slightly higher ionization energy)
- Ease with which the elements lose their electrons increases as we move down the family

## **7.8 Group Trends: Selected Nonmetals**

### Hydrogen

- Hydrogen is a nonmetal, which occurs as a diatomic gas under normal conditions

- Its ionization energy is considerably higher (due to lack of shielding, and thus higher  $Z_{\text{eff}}$ ) than the metals and is more like the nonmetals
- Hydrogen generally reacts with other nonmetals to form molecular compounds (typically highly exothermic)
- Hydrogen reacts with active metals to form metal hydrides that contain the H<sup>-</sup> hydride ion
- Hydrogen can also lose an electron to yield the aqueous H<sup>+</sup> ion

### Group 6A: The Oxygen Family

- As we proceed down group 6A, the increase in metallic character is clearly evident
- Allotropes:** different forms of the same element in the same state (ie: O<sub>2</sub> and O<sub>3</sub>)
- Oxygen has a great tendency to attract electrons from other elements (“oxidize” them)
- Oxygen in combination with metals is almost always present as the oxide, O<sup>2-</sup> ion, which has noble-gas configuration and is particularly stable
- Sulfur is the second most important element in the 6A group. It also exists in several allotropic, the most common and stable being S<sub>8</sub>.
- Sulfur also has a great tendency to gain electrons from other elements to form sulfides
- Because Sulfur is below Oxygen, its tendency to form sulfide anions is not as great as that of Oxygen to form oxide ions

### Group 7A: The Halogens

- All halogens are typical nonmetals. Their melting and boiling points increase with increasing atomic numbers
- Fluorine and Chlorine are gases at room temperature, Bromine is a liquid, and Iodine is a solid
- The halogens have among the most negative electron affinities; they have a great tendency to gain electrons.
- Fluorine and Chlorine are the most reactive. Fluorine removes electrons from almost any substance, even water, and usually very exothermically.
- Halogens react with most metals to form ionic halides and with Hydrogen to form gaseous hydrogen halides.

### Group 8A: The Noble Gases

- All noble gases are monoatomic (single atom)
- All have large first ionization energies, with a decrease as we go down the group
- Until the early 1960's, they were called *inert gases* because they were thought to be incapable of forming chemical compounds.
- Only noble-gas compounds known today are: XeF<sub>2</sub>, XeF<sub>4</sub>, XeF<sub>6</sub>, and KrF<sub>2</sub>