

Chapter 12: Chemical Periodicity

The first elements that were discovered were shiny metals that melted out of rocks heated in fires. The first elements discovered were copper, silver and gold. This occurred approximately eight thousand years ago. Lead, tin, iron, mercury, carbon and sulfur were discovered next.

As people began working with metals, the Alchemists searched for a method of turning lead (or other ordinary metals) into gold. They were not successful, but they did discover other elements in the process. Scientists began to predict the existence of other elements. Experimentation continued and improved. By 1940 all 90 naturally occurring elements had been discovered. Since then work in nuclear science has led to the discovery of radioactive elements.

Only about one quarter of the elements occur in the free (or elemental) state. The others are found in compounds - two or more elements chemically bonded together.

Metals - More than three quarters of the known elements are metals. Metals share similar properties:

- good conductors of heat and electricity
- luster
- malleable and ductile
- high densities
- high boiling points and melting points
- resist stretching and twisting
- solids at room temperature

Nonmetals

- no luster
- poor conductors
- may be solid, liquid or gas
- low densities
- low melting and boiling points

Metalloids (or Semimetals) - have properties of both metals and nonmetals

The Development of the Periodic Table

As more elements were discovered, some patterns of similarities in elements were recognized. Scientists began trying to organize the elements. Dmitri Mendeleev organized the elements in a table. He placed the elements in order of increasing atomic mass. He used columns of different lengths so that elements with similar properties were side by side. He left blank spaces where no known element seemed to fit. He predicted that these blank spaces represented elements which had not been discovered yet. Based on the location of the blanks, he predicted the properties of undiscovered elements. When these elements were discovered, their actual properties were very close to Mendeleev's predictions.

There were a few discrepancies with Mendeleev's table. When the elements were arranged according to increasing atomic mass, some of the elements were out of order. Mendeleev assumed that their masses were not calculated accurately enough. He believed that when more accurate methods were available, that these discrepancies would disappear.

The Modern Periodic Table: With the discovery of atomic number (the number of protons in the nucleus) the periodic table was rearranged. The elements are now arranged in order of increasing atomic number. This corrected the discrepancies in Mendeleev's table.

The Modern Periodic Law - When the elements are arranged in order of increasing atomic number, there is a periodic repetition of their physical and chemical properties.

Organization of the Periodic Table

Periods: - horizontal rows

- properties change from left to right

Groups (or Families):

- vertical columns
- share similar properties

Group 1: Alkali Metals

- low densities and melting points
- good electrical conductivity
- the most reactive metals
- react with oxygen and moisture in the air
- they will react violently with water and are stored under oil or kerosene
- due to their reactivity, they are not found uncombined in nature

Group 2: Alkaline Earth Metals

- similar to the alkali metals, but less reactive, not required to store under oil
- harder than the alkali metals

Group 17: Halogens

- nonmetals
- do not exist in nature in the uncombined state, but their compounds are common
- form salts with group 1 or 2 metals

Group 18: Noble Gases

- gases at room temperature
- occur in the atmosphere in very small amounts
- extremely unreactive

Representative or Main Group Elements

- groups 1,2 and 13 - 18

Transition Metals: Groups 3 - 12

- typical metals - shiny, and good conductors

Inner Transition Metals: Lanthanides: elements 57 - 70 Actinides: elements 89 - 102

- typical metals

Trends in the Periodic Table

Atomic Size: The size of an atom is generally given in terms of the atomic radius of the atom.

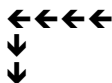
Group Trends: Atomic size generally increases as we move down a group of the periodic table. As we descend, electrons are added to successively higher energy levels. The nuclear charge also increases because of additional protons. However, the full nuclear charge is shielded from the outer electrons because of the inner electrons and their distance from the nucleus.

Period Trends: Atomic size generally decreases as we move from left to right across a period. As we move across a period, electrons are added to the same energy level. Protons are also added to the nucleus. The effect of the increasing nuclear charge on the outermost electrons is to pull them closer to the nucleus. Atomic size therefore decreases.

In general:



or:



Ionization Energy - the amount of energy required to remove an electron from an atom. When the outermost electron is located close to the nucleus, there is a greater attraction between the positive protons (in the nucleus) and the negative electrons. Because of this, more energy is required to remove an electron. When the outermost electrons are located further from the nucleus, the attractive force is less, and it is easier to remove an electron. Therefore, the smaller the atom, the higher the ionization energy.

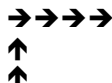
Group Trends: Ionization energy decreases as we go down a group.

Period Trends: Ionization energy increases as we go across a period.

In general:



or



Electron Affinity: The amount of energy released when an electron is added to an atom. When electrons are added close to the nucleus, there is a greater attraction between the negative electron and the positive nucleus and electron affinity is higher. Therefore, the smaller the atom, the higher the electron affinity.

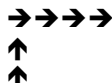
Group Trends: Electron affinity generally decreases as we go down a group.

Period Trends: Electron affinity generally increases as we go across a period.

In general:



or



Exceptions: Noble gases have full s and p orbitals at their highest energy levels. Because of this stability, they do not gain or lose electrons easily. Noble gases are not included in this general trend.

Electronegativity: Describes an atoms ability to attract electrons. This trend is similar to electron affinity.

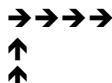
Group Trends: Electronegativity generally decreases as we go down a group.

Period Trends: Electronegativity generally increases as we go across a period.

In general:



or



Exceptions: Noble gases are not included in this general trend.