History of the Periodic Table

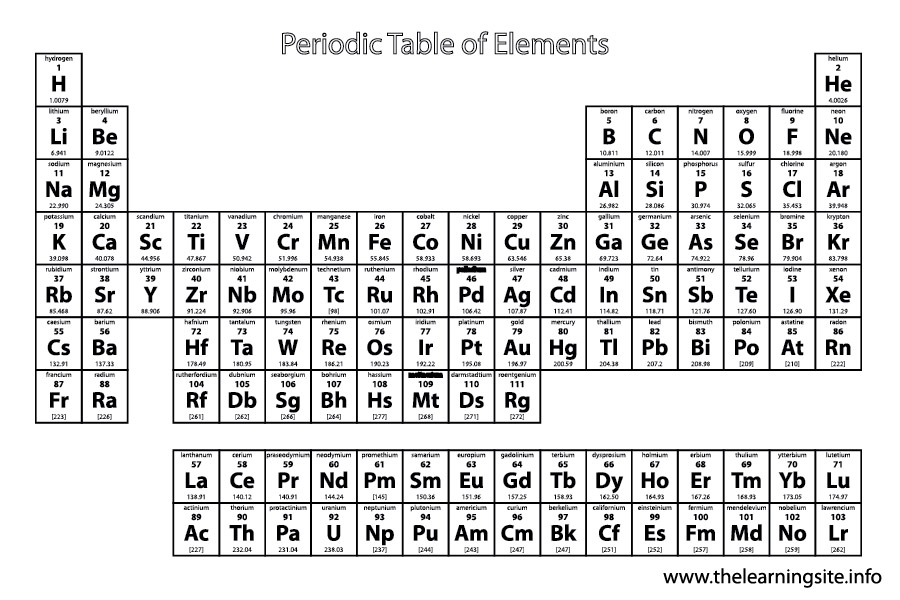
* Antoine Lavoisier
  + 1700’s
  + French scientist
  + Compiled a list of all elements that were known at the time.
    - Contained 33 elements, that were organized into 4 categories
* John Newland
  + 1864
  + English chemist
  + Arranged elements by increasing atomic mass
  + Noticed the repetition of properties every 8th element
  + Created the law of octaves (after the musical octave in which notes repeat every 8th tone). He showed that the properties of elements repeat in a periodic way.
* Lothar Meyer
  + 1869
  + German chemist 🡪 Worked with Mendeleev
  + Demonstrated a connection between atomic mass and elements’ properties.
  + Arranged the elements in order of increasing atomic mass.
* Dmitri Mendeleev
  + 1869
  + Russian chemist 🡪 Worked with Meyer
  + Demonstrated a connection between atomic mass and elements’ properties
  + Arranged the elements in order of increasing atomic mass.
  + Predicted the existence and properties of undiscovered elements.
* Henry Moseley
  + 1913
  + English chemist
  + Discovered that atoms contain a unique number of protons called the atomic number.
  + Arranged elements in order of increasing atomic number, which resulted in a periodic pattern of properties.

Periodic Law – States that when the elements are arranged by increasing atomic number, there is a periodic repetition of their properties.

Groups – A vertical column of elements in the periodic table arranged in order of increasing atomic number.

* Also called a family

Periods – A horizontal row of elements in the modern periodic table



**Representative Elements** – Elements from group 1, 2, and 13-18 in the modern periodic table, possessing a wide range of chemical and physical properties.

**Transition Elements** – Elements in groups 3-12 of the modern periodic table and are further divided into transition metals and inner transition metals.

**Alkali Metals** – Group 1 elements, except for hydrogen, they are reactive and usually exist as compounds with other elements.

**Alkaline Earth Metals** – Group2 elements in the modern periodic table and are highly reactive.

**Transition Metals** – An elements in groups 3-12 that is contained in the d-block of the periodic table and, with some exceptions, is characterized by a filled outermost s orbital of energy level n, and filled or partially filled d orbitals or energy level n-1.

**Inner Transition Metals** – A type of group B element that is contained in the f-block of the periodic table and is characterized by a filled outermost orbital, and filled or partially filled 4f and 5f orbitals.

* **Lanthanide Series** – In the periodic table, the f-block elements from period 6 that follow the element lanthanum.
* **Actinide Series** – In the periodic table, the f-block elements from period 7 that follow the element actinium.

**Halogens** – A highly reactive group 17 element.

**Noble Gases** – An extremely unreactive group 18 element.

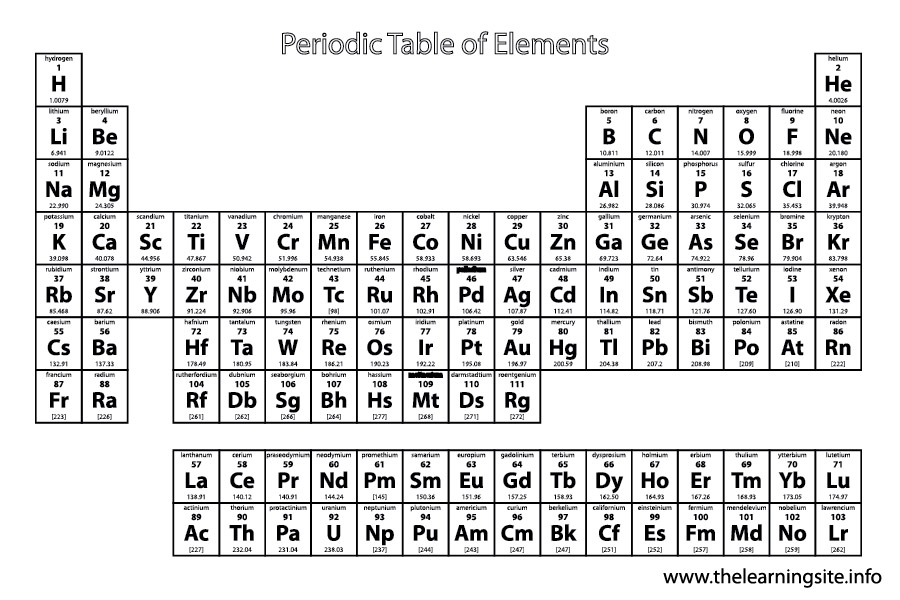
**Metals**

**Metalloids** – An element that has physical and chemical properties of both metals and nonmetals.

**Nonmetals**

**Valence Electrons On the Periodic Table**

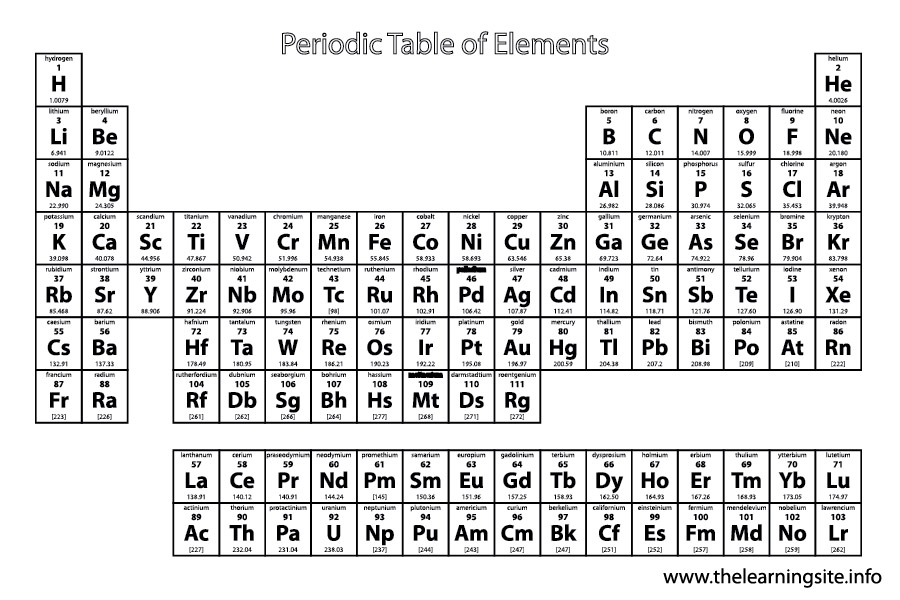
|  |  |
| --- | --- |
| **Alkali Metals** |  |
| **Alkaline Earth Metals** |  |
| **Transition Metals** |  |
| **Group 13** |  |
| **Group 14** |  |
| **Group 15** |  |
| **Group 16** |  |
| **Halogens** |  |
| **Noble Gases** |  |



**Periodic Table Trends**

Atomic Radius – The size of the atom

Trend



Ion – An atom or a bonded group of atoms that has a positive or negative charge.

* Cation
  + A positively charged ion
    - It’s positive because it has lost an electron. So the positively charged protons now outnumber the negatively charged electrons, giving the ion a positive charge.
  + Smaller in size than the original atom
* Anion
  + A negatively charged ion
    - It’s negative because it has gained an electron. So the negatively charged electrons now outnumber the positively charged protons, giving the ion a negative charge.
  + Larger in size than the original atom

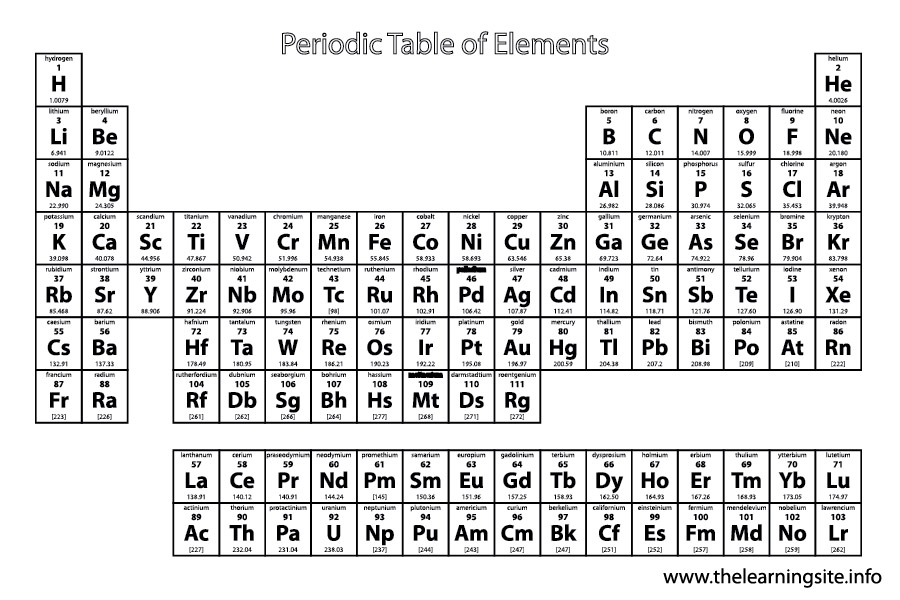
Element Charges (Oxidation Number)

|  |  |
| --- | --- |
| **Alkali Metals** |  |
| **Alkaline Earth Metals** |  |
| **Transition Metals** |  |
| **Group 13** |  |
| **Group 14** |  |
| **Group 15** |  |
| **Group 16** |  |
| **Halogens** |  |
| **Noble Gases** |  |

Ionization Energy – The energy required to remove an electron from a gaseous atom.

* Think of ionization energy as an indication of how strongly an atom’s nucleus holds on to its electrons.
* A high ionization energy indicates the atom has a strong hold on its electrons & vice versa.

Trend



The Octet Rule – States that atoms tend to gain, lose, or share electrons in order to acquire full set of eight valence electrons.

* *Example*: Sodium Atom vs. Sodium Ion

Electronegativity – The electronegativity of an element indicates the relative ability of its atoms to attract electrons in a chemical bond.

* Values are on p.194
* Trend (Below).

