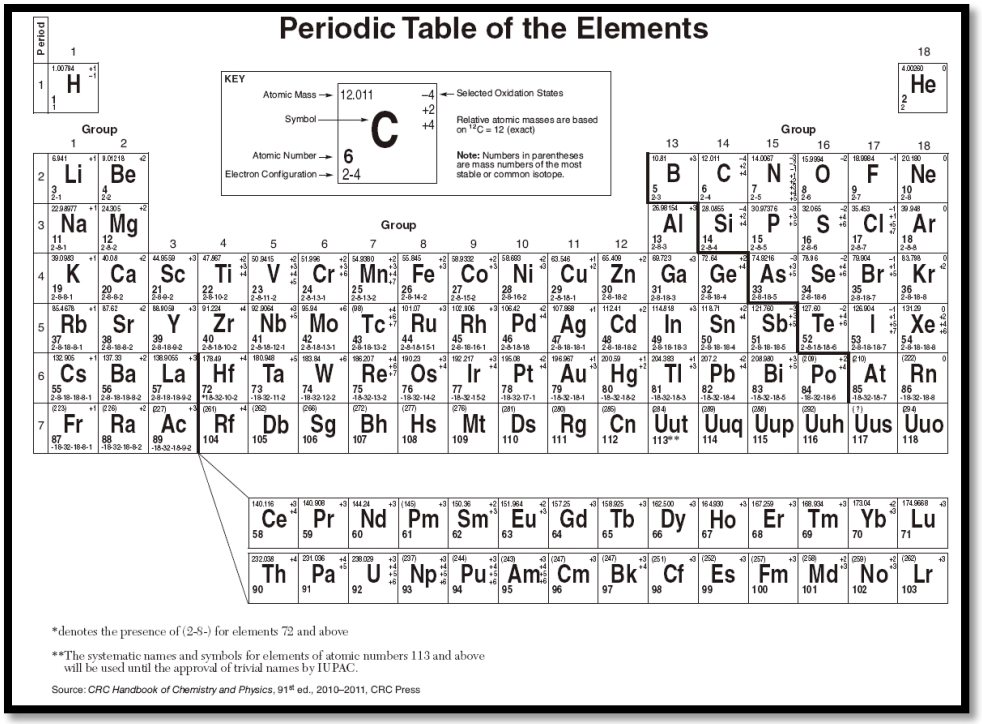
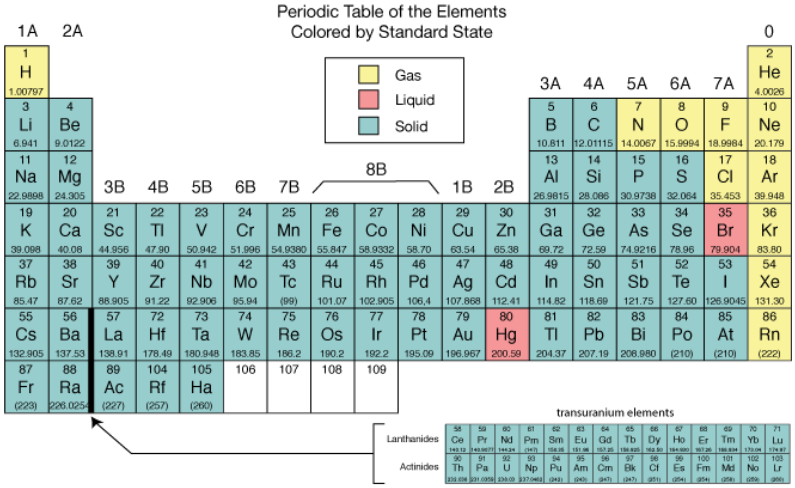
Unit 3 (Periodic Table) Test Review Sheet



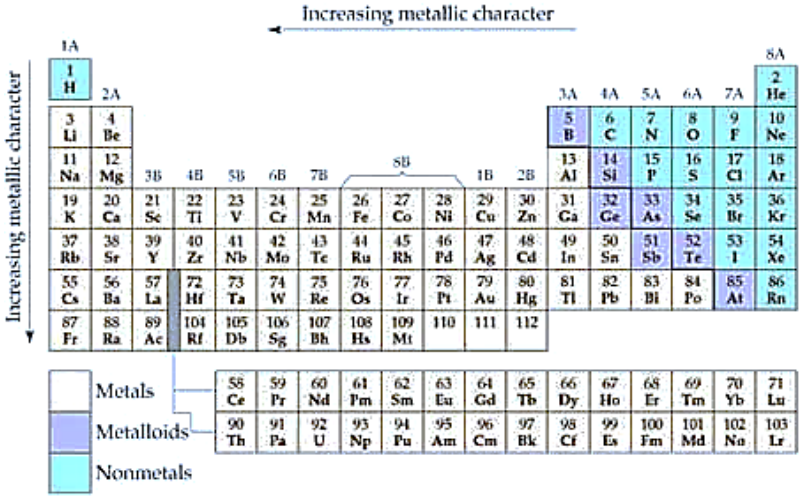
Basic Organization of Periodic Table

- Periodic Table is organized by increasing atomic number (equal to number of protons)
- Periodic Law: There is repetition of chemical properties when elements are arranged by increasing atomic number
- Elements in a column are referred to being in the same Group (18 Groups)
- Elements in a row are referred to as being in the same Period (7 Periods)
- All elements are classified as metals (left side), nonmetals (right side, but not including Group 18 elements) or metalloids. Look for the **BOLD** staircase in your Periodic Table as the "boundary between metals and nonmetals. Elements with boundaries on the staircase are the metalloids, except aluminum ("Al").
- Hydrogen is the only element on the left side of the **BOLD** "staircase" boundary that is a nonmetal
- Elements in the same Period have the same number of energy shells. For example, all the elements in Period 5 have 5 energy shells and their valence electrons are in the 5th energy shell (highest energy shell).
- Generally, elements in same Group have same # of valence electrons.
- Elements in the same <u>Group</u> have similar chemical properties because they have the same number of valence electrons

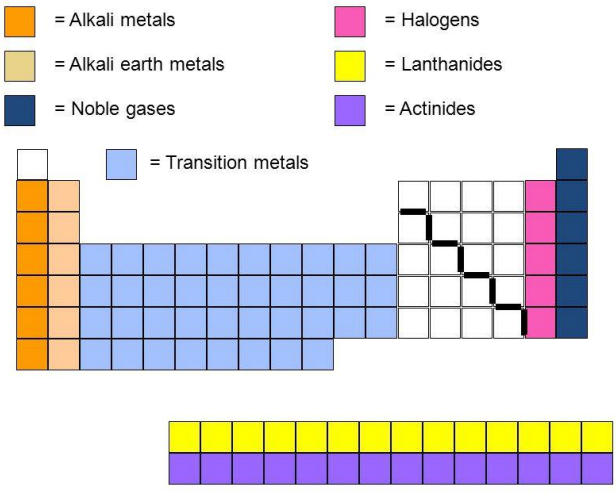
Phases of the Elements at STP



Metals, Nonmetals, Metalloids (Semi-Metals)



Important Groups of the Periodic Table



Properties of the Elements

Physical Properties of Elements

There are several physical properties that can be used to describe and identify the elements.
The following is a list of these physical properties and their definitions.

Concept Facts! Study to remember these properties.

Malleable describes a solid that is easily hammered and flattened into a thin sheet. (Ex. Aluminum, Al)

Ductile describes a solid that is easily drawn into a thin wire. (Ex. Copper, Cu)

Brittle describes a solid that is easily broken or shattered into pieces when struck. (Ex. Sulfur, S)

Luster describes the shininess of a substance. (Ex. Gold, Au)

Conductivity describes the ability of heat or electricity to flow through a substance. (Ex. Copper, Cu)

Electronegativity describes an atom's ability to attract electrons from another atom during bonding.

Ionization energy describes an atom's ability to lose its most loosely bound valence electrons.

Density describes the mass to volume ratio of an element.

Atomic radius describes the size of the atom of an element.

See Table S

Ionic radius describes the size of the element after it has lost or gained electrons to become an ion.

Use *Reference Table S* to find and compare electronegativity, ionization energy, atomic radius, and density values of the elements.

Metals

Metallic elements are located to the left of the periodic table.

All elements in Groups 1 to 12 (except hydrogen) are classified as metals.
The rest of the metallic elements are located near the bottom of Groups 13 through 17.
The majority (about 75%) of the elements are metals.

Iron (Fe)

Below are some general properties (characteristics) of metals.

Concept Facts! Study to remember these properties.

- Almost all metals are solids at room temperature. The exception is mercury (Hg), which is a liquid metal.
- Solid metals are malleable and ductile. Many have high luster.
- Metals tend to have high heat (thermal) and electrical conductivity due to their *mobile valence electrons*
- Metals tend to have low electronegativity values (because they do not attract electrons easily)
- Metals tend to have low ionization energy values (which is why metals lose their electrons easily)
- Metals lose electrons and form positive ions
- Radius (size) of a metal atom decreases as it loses electrons and form a positive ion
- The size of a positive (+) ion is always smaller than the size of the neutral atom

Metals, Nonmetals, and Metalloids

Metalloids

Metalloids are the seven elements located between the metals and nonmetals.
Metalloid elements are located on the periodic table along the thick zigzag line.

Below are some generally properties (characteristics) of metalloids.

Concept Facts! Study to remember these properties.

- Metalloids tend to have properties of both metals and nonmetals
- Metalloids properties are more like those of metals and less like those of nonmetals
- Metalloids exist only as solids at room temperature.

Tellurium (Te)

Nonmetals

Nonmetallic elements are located to the right of the periodic table.

All elements in Groups 17 and 18 (except At) are classified as nonmetals. The rest of the nonmetals are located near the top of Groups 14, 15 and 16. Hydrogen (in Group 1) is also a nonmetal.

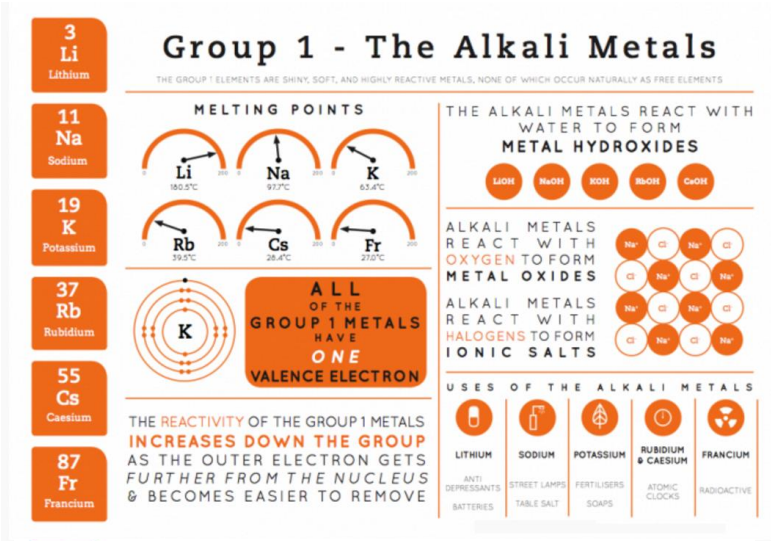
Below are some general properties (characteristics) of nonmetals.

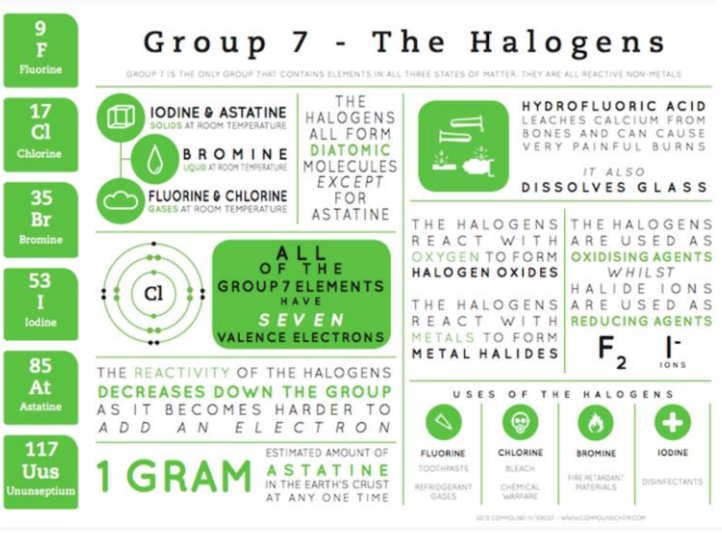
Sulfur (S)

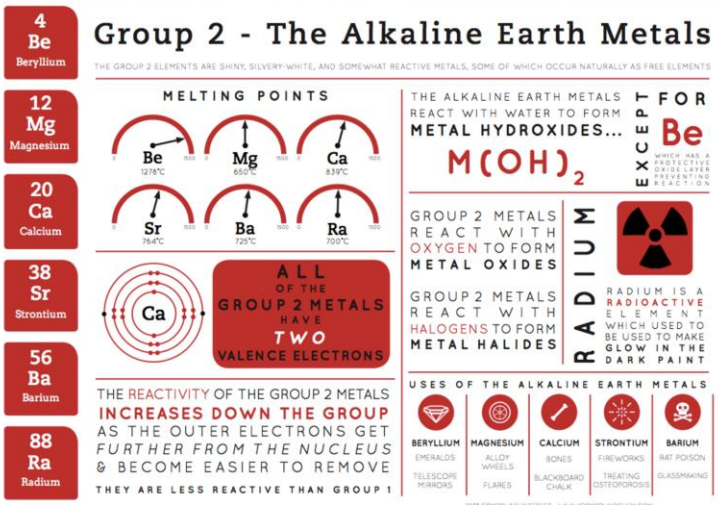
Concept Facts! Study to remember these properties.

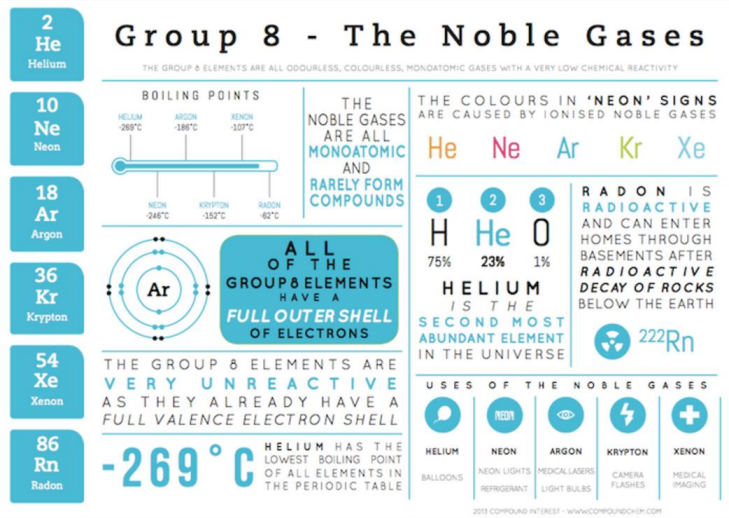
- Nonmetals are found in all three phases: solid, liquid and gas.
- Most nonmetals exist as molecular gases and solids. Bromine is the only liquid nonmetal.
- Solid nonmetals are generally brittle and dull (lack luster, not shiny)
- Nonmetals have low or poor electrical and heat (thermal) conductivity
- Nonmetals tend to have high electronegativity values (because they attract and gain electrons easily)
- Nonmetals tend to have high ionization energy (which is why nonmetals do not lose electrons easily)
- Nonmetals gain electrons and form negative ions
- Radius of a nonmetal atom increases as it gains electrons and forms a negative ion
- The size of a negative (-) ion is always bigger than that of the neutral atom

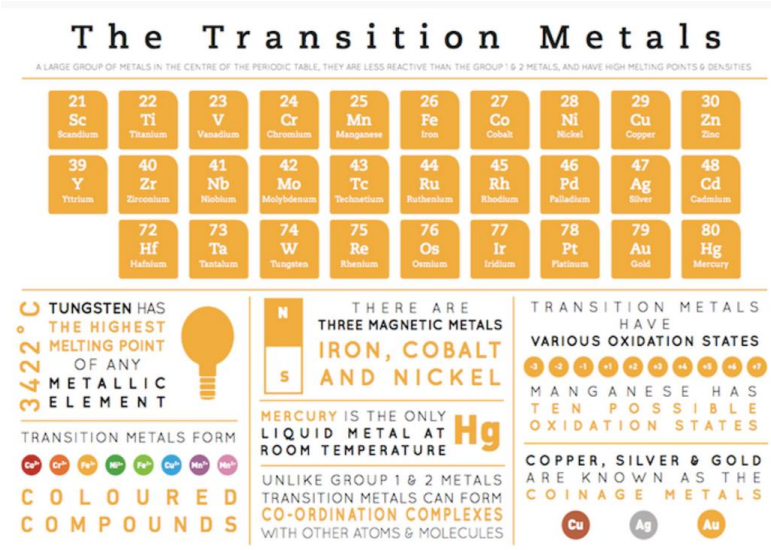
Important Groups of the Periodic Table





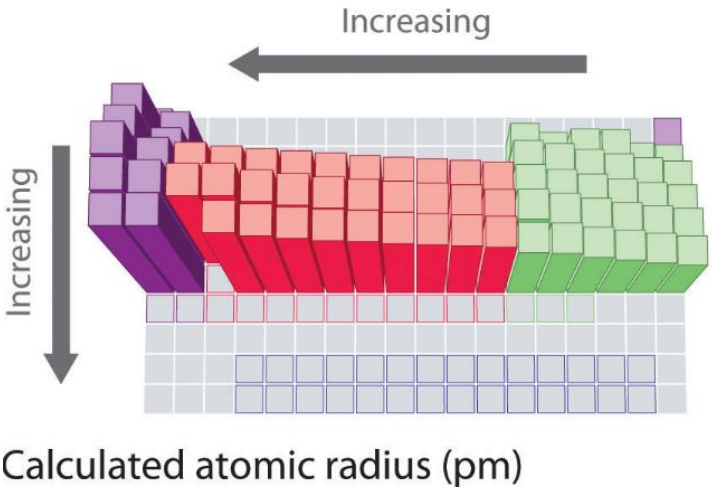






Trends of the Periodic Table

- **Atomic Radius (Size)**



- Atomic radius decreases along a Period as the atomic number increases (left to right due to the larger positive charge of the nucleus pulling closer valence electrons that are all in the same energy level for all the elements in the same Period)
- Atomic radius increases down a Group as the valence electrons are in higher energy shell which allows the valence electrons to be farther away from the nucleus
- Atomic radius for an element is provided in Table S

- **Ionic Radius vs. Atomic Radius**

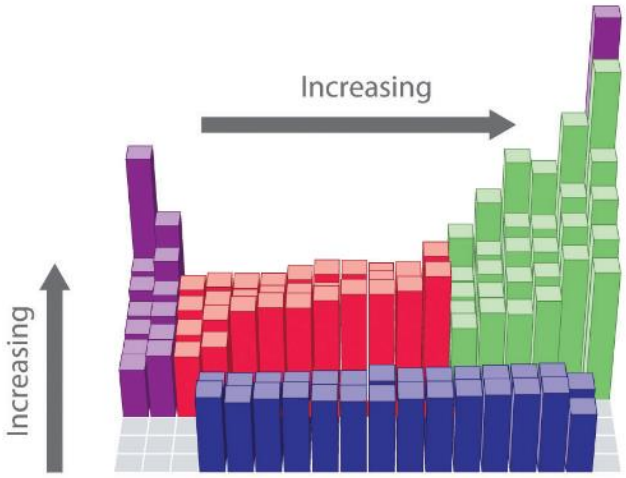
Positive ions have **smaller** atomic radius than neutral atom

For example: K^{1+} ion has a smaller radius that the neutral K atom. We say the ionic radius is smaller. This is due to the valence electron being lost and the ion having one less energy shell, making the ion smaller.

Negative ions have **larger** atomic radius than positive atom

For example: Cl^{1-} has a larger radius than the neutral Cl atom. We say the ionic radius is larger. The reason is more difficult to understand since the number of energy shells does not change. The reason the ionic radius is larger is because the extra valence electrons gained cause the electrons to repel each other and force the other valence electrons to move further out to create more space. Thus, the radius of the ion gets larger.

- **Ionization Energy**



Ionization energy is the energy required to remove an electron from the outermost shell of an atom

-Ionization energy generally increases along a Period due to increase in the positive charge of the nucleus (see atomic size)

-Ionization energy decreases down a Group as valence electrons are in higher Principal Energy levels. If the valence electron has more energy, then it is farther away from the nucleus and it is easier to remove the electron from the atom.

-Ionization energy for an element is provided in Table S

Since metals tend to lose electrons, then the lower the ionization energy, the easier it is for the metal to lose electrons. This makes the metal more reactive.

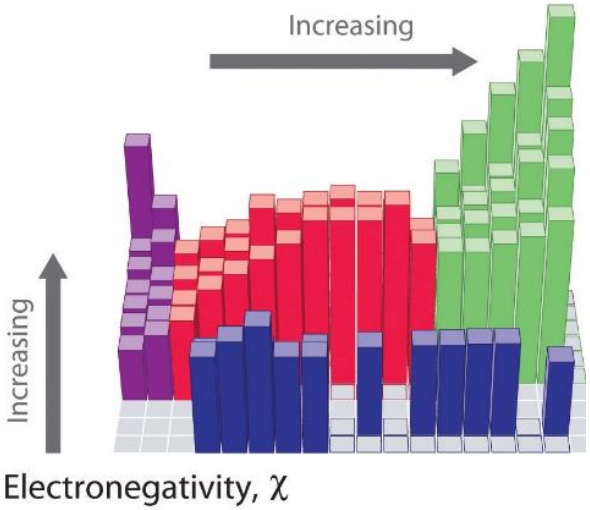
- **Electronegativity**

Electronegativity is the ability of an atom to attract electrons

-Nonmetals (except for noble gases) have higher electronegativities than metals. So the higher the electronegativity of a nonmetal, the more reactive it is.

-Scale of electronegativity is from 0 (low) to 4 (high)

-Electronegativity values for an element are provided in Table S. You can compare the electronegativities of any two elements to determine which one will attract electrons more. If the electronegativity is the same, then both have the same ability to attract electrons



*Noble Gases have an unusually **high** ionization energy and **little (or no)** electronegativity because they have a completely filled outer shell of electrons. They do not want to lose the electrons they have and they do not want any more electrons-because this would destabilize them.