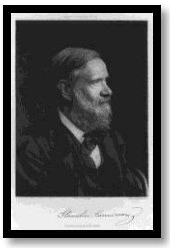
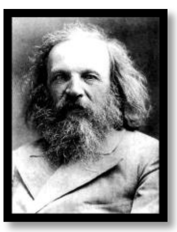
# *Topic 4: Periodic Table & Trends*

## The Periodic Law (Chapter 5 in Modern Chemistry)

**The History of the Periodic Table**



**Stanislao Cannizzaro** discovered a method of accurately *measuring atomic masses*.

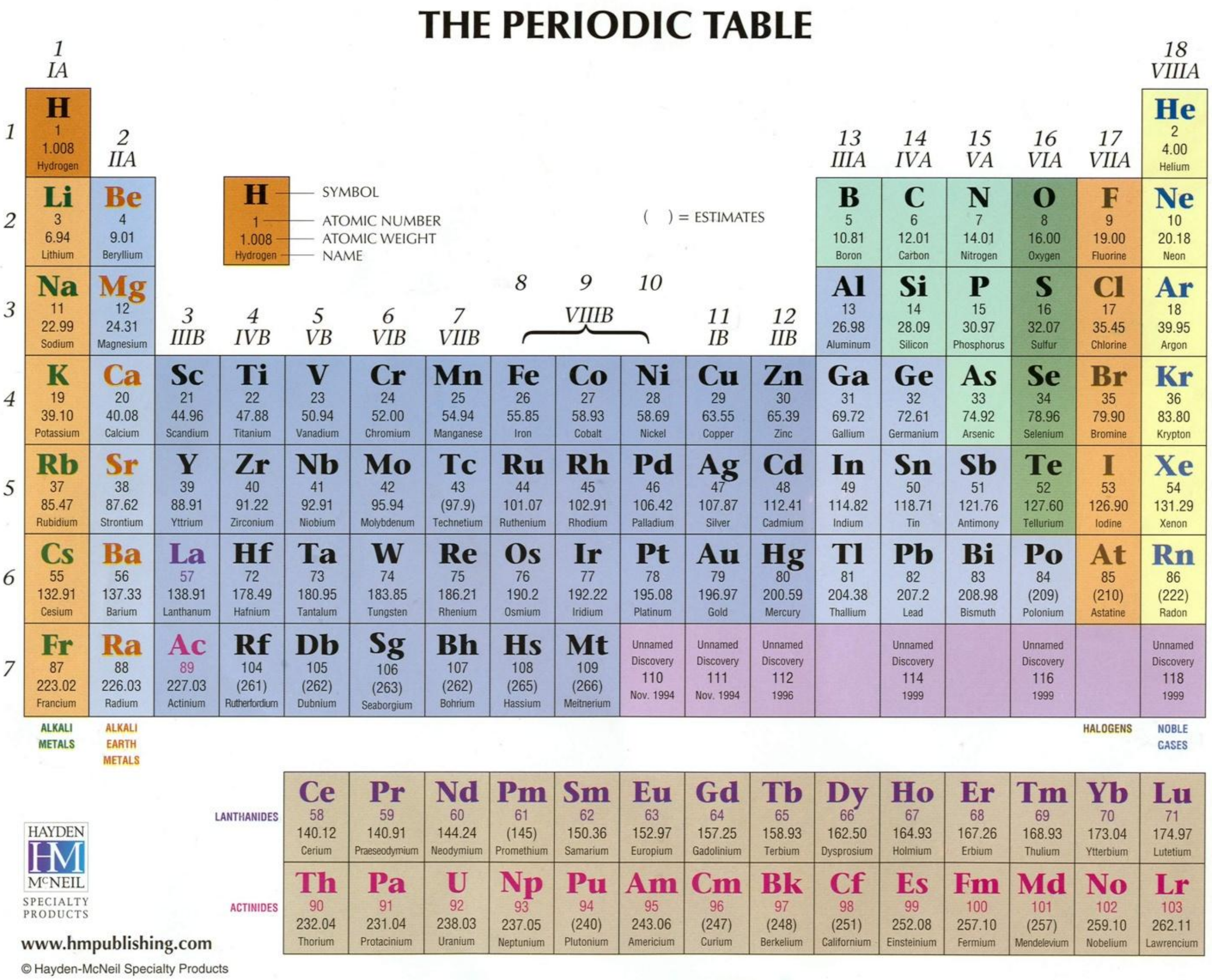


**Dmitri Mendeleev** is credited with *organizing the first periodic table based on atomic masses.(1869)* He noticed the when the elements were arranged in order of increasing atomic mass, certain similarities in their chemical properties appeared at regular intervals. These repeating patterns are referred to as periodic. Mendeleev did find some discrepancies, he placed iodine after tellurium even though based on atomic masses they should be reversed. Tellurium acted more like O, S, and Se. Iodine acted more like F, Cl, and Br. He knew there was some problems with using atomic masses for ordering the table. He also left blanks in his periodic table. He boldly predicted the existence and properties of the elements that would fill three of these blanks based on the properties of elements that were similar in his table. Eventually, all three of these elements were discovered, Sc, Ga, and Ge. Their properties were strikingly similar to those predicted by Mendeleev.

**Henri Mosely** *discovered that the elements in the periodic table fit into patterns better when they were arranged in increasing order according to nuclear charge, or number of protons (atomic number). (1911)* This corrected the discrepancies in Mendeleev's table.

**Periodic Law** *state the physical and chemical properties of the elements are periodic functins of their atomic numbers.*

The **Modern Periodic Table** *is an arrangement of the elements in order of their atomic numbers so that elements with similar properties fall in the same column, or group.* Remember that the numbers across the top represent the group, while the numbers down the side represent the period.




**The Element Song** by Tom Lehrer http://www.privatehand.com/flash/elements.html

### Noble gases (Group 18)

All atoms of the noble gases have their outer s and p orbitals filled.
We will see later that these atoms require very large amounts of energy to form ions, so much in fact, that they are difficult to alter chemically and as such are inert (unreactive) and do not tend form ions.

### Alkali metals (Group 1) & Alkaline Earth metals (Group 2) METALS

**Group 1** atoms have an electronic structure [Noble gas] ns1. This means that they tend to lose the s electron when they from an ion, leaving behind an inert noble gas type structure. This explains why Group 1 elements tend to only form $1^+$ ions.
**Group 2** atoms have an electronic structure [Noble gas] ns2. This means that they tend to lose the two s electrons when they from an ion, leaving behind an inert noble gas type structure. This explains why Group 2 elements tend to only form $2^+$ ions.

A similar argument can be applied to group 3 atoms and their simple ions.

**Groups 16 & 17 (Chalcogens & Halogens) NON-METALS**

**Group 16** atoms have an electronic structure [Noble gas] ns2 np4. This means that they tend to gain two p electrons when they from an ion, to reach an inert noble gas type structure with a charge of $2^-$. Check out this animation: http://www.privatehand.com/flash/oxygen.html . It is about oxygen. Very cute!

**Group 17** atoms have an electronic structure [Noble gas] ns2 np5. This means that they tend to gain one p electron when they from an ion, to reach an inert noble gas type structure with a charge of 1-.

### Task 4a

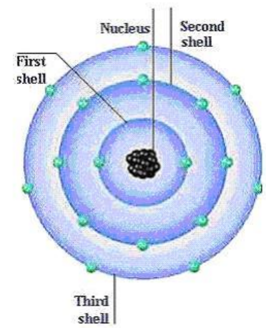
1. What name is given to each of the following groups of elements in the periodic table?
   a. Group 1
   b. Group 2
   c. Groups 3-12
   d. Group 17
   e. Group 18

2. Based on what you know about their electron configurations, which groups do you think are the most active? Why? The least active? Why?

**Periodic Properties**

As you have learned earlier, the elements in the same group have similar electron configurations, therefore they also have similar physical and chemical properties because their valence electrons are the same. You need to be able to identify the properties and how they relate to each other based on their placement on the periodic table.

Before we discuss the periodic property trends, we need to discuss two reasons that properties change based on the periodic table.

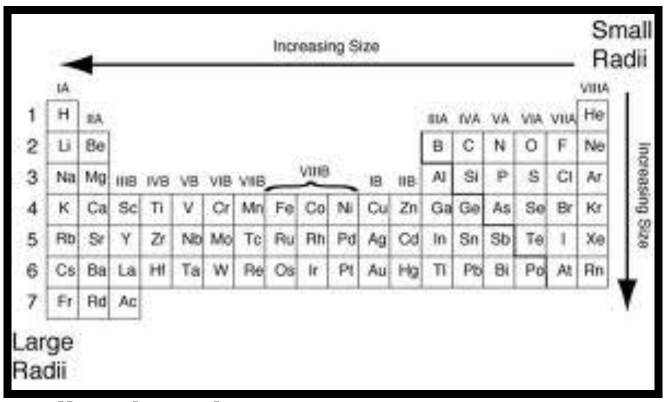


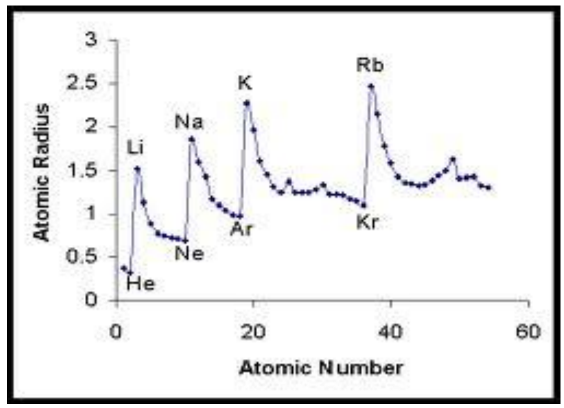

The **shielding effect** causes properties to change that are in the same group (column). *The shielding effect occurs because the inner electrons shield the nucleus from the outer electrons.* The more inner electrons there are between the valence electrons and the nucleus then the smaller the attraction of the nucleus on the outer electrons. Even though the valence electrons are being attracted to the nucleus, they are also being repelled by all the inner electrons.

The **effective nuclear charge** causes properties to change that are in the same period (row),. *This means that the positive charge of the nucleus is increasing while the number of energy levels is the same.* Since the electrons are negatively charged, they are attracted to nuclei, which are positively charged. Many of the properties of atoms depend not only on their electron configurations but also on how strongly their outer electrons are attracted to the nucleus. The distance from the nucleus is not changing but the charge of the nucleus is.

**Atomic radii**

**Atomic radii** refers to *the size of an atom*. It is actually defined as *one-half the distance between the nuclei of identical atoms that are bonded together.* Atomic radii increases down a group and decreases across a period. *It increases toward Fr.* The size of the atom naturally increases down a group because the volume of the electron cloud gets bigger as the number of energy levels increase. The size of the atom gets smaller as the atomic number increases because the charge of the positive nucleus is getting larger and attracts the electrons more and more. It is true that the number of electrons is also increasing, but they are all placed in the same energy level, meaning that they are the same distance from the nucleus. The higher the nuclear charge the stronger the




pull on those electrons.

**Task 4b**

1. Referring to the periodic table, arrange the following atoms in order of increasing size: P, S, As, Se.
2. Of cesium, Cs, hafnium, Hf, and gold, Au, which element has the smallest atomic radius? Explain your answer in terms of trends in the periodic table.

**Ionic Radii**

*A positive ion is known as a* **cation**. The formation of a cation by the loss of one or more electrons always leads to a decrease in atomic radius because as electrons are removed the nucleus has a greater attraction for the electrons that remain. The positive ion has lost a whole layer of electrons.

*A negative ion is known as an* **anion**. The formation of an anion by the addition of one or more electrons always leads to an increase in atomic radius. The extra electrons cause the electron cloud to spread out due to greater electron repulsion while the attraction from the protons remains the same.

Positive ions are smaller than the atom from which they came, while negative ions are larger than the atom from which they came.


### Task 4c

1. Distinguish between a cation and an anion.
2. Which of the following cations is least likely to form: $Sr^{2+}$, $Al^{3+}$, $K^{2+}$?
3. Which of the following anions is least likely to form: $I^-$, $Cl^{2-}$, $O^{2-}$?
4. From each set, determine which atom or ion is the largest.
   a. Ca, $Ca^{2+}$
   b. Cl, $Cl^-$
   c. S, $S^-$, $S^{2-}$
   d. Sr, $Sr^+$, $Sr^{2+}$
   e. $K^+$, $Ca^{2+}$
   f. $N^{3-}$, $O^{2-}$, $F^-$
   g. $Ca^{2+}$, $Sr^{2+}$, $Ba^{2+}$
   h. $O^{2-}$, $S^{2-}$, $Se^{2-}$

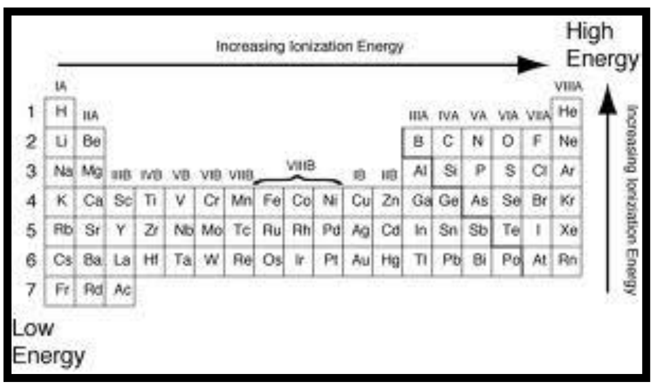
**Ionization Energy**

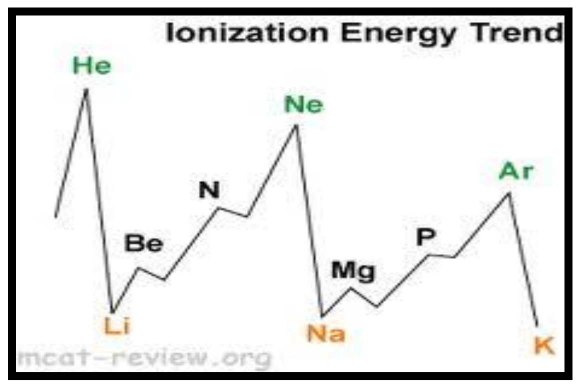
An electron can be removed from an atom if enough energy is supplied. *Any process that results in the formation of an ion is referred to as* **ionization**. Suppose A is any element.

$$A + \text{energy} \rightarrow A^{+} + e^{-}$$

$$A^{+} + \text{energy} \rightarrow A^{2+} + e^{-}$$

**Ionization Energy** *is the amount of energy required to remove one electron from a neutral atom of an element to make an ion*. Specifically, this is the first ionization energy or $IE_1$. Ionization energy *increases toward F.*




Ionization energy decreases down a group because the farther the valence electron is from the nucleus the lower the attraction between them. Remember that as you move down a group the period number increases, meaning there is another level of electrons being added. Therefore, there will be more shielding electrons between the nucleus and the valence electrons causing repulsion. So, the farther down the element is in the group the lower the ionization energy.

Ionization energy increases across a period because the effective nuclear charge (the positive charge) is getting larger while the shielding effect (levels of electrons) remains the same. This means there will be a stronger attraction between the valence electrons and the nucleus, so it will take more energy to remove an electron.

The **second ionization energy** *is the amount of energy required to remove the second electron from an ion.* The **third ionization energy** *is the amount of energy required to remove the third electron from an ion, etc.* Successive ionizations require more energy. The more electrons that are removed the harder it is to remove the next. Each successive electron removed from an ion feels an increasingly stronger effective nuclear charge.

***Successive molar ionization energies in kJ/mol***
**(96.485 kJ/mol = 1 eV/particle)**

| Element | First | Second | Third | Fourth | Fifth | Sixth | Seventh |
|---|---|---|---|---|---|---|---|
| Na | 496 | 4,560 | | | | | |
| Mg | 738 | 1,450 | 7,730 | | | | |
| Al | 577 | 1,816 | 2,881 | 11,600 | | | |
| Si | 786 | 1,577 | 3,228 | 4,354 | 16,100 | | |
| P | 1,060 | 1,890 | 2,905 | 4,950 | 6,270 | 21,200 | |
| S | 999.6 | 2,260 | 3,375 | 4,565 | 6,950 | 8,490 | 27,107 |
| Cl | 1,256 | 2,295 | 3,850 | 5,160 | 6,560 | 9,360 | 11,000 |
| Ar | 1,520 | 2,665 | 3,945 | 5,770 | 7,230 | 8,780 | 12,000 |

Notice that there is a big jump in the $IE_1$ to $IE_2$ of sodium. Sodium has only one valence electron. It is relatively easy for it to lose that electron when bonding. It will then have a stable configuration with 8 valence electrons. It will be much more difficult to remove the second electron from sodium because this will make the ion become unstable so it takes a lot more energy to do this. Basically, it is easier to remove valence electrons than it is to remove the electron immediately after the last valence electron.

### Task 4d

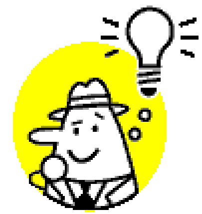
1. Referring to the periodic table, arrange the following atoms in order of increasing first ionization energy: Ne, Na, P, Ar, K.

2. In general ionization energy increases toward F. Refer to the graph on ionization energy trends. Considering electron configurations, why do you think B has a lower $IE_1$ than Be? O has a lower $IE_1$ than N?

3. Write the equations that show the process for the following.

   a. The first ionization energy for tin

   b. The second ionization energy for the tin(I) ion



4. Explain each of the following.

   a. Why does Li have a larger first ionization energy than Na?

   b. The difference between the third and fourth ionization energies of scandium is much larger than the difference between the third and fourth ionization energies of titanium. Why?

   c. Why does Li have a much larger second ionization energy than Be?

5. Here are the ionization energies for an element in period 2: 900, 1757, 14849, 21007. Which element is represented by these energies?

**Electron Affinity**

Neutral atoms can also acquire electrons. *The energy change that occurs when an electron is acquired by a neutral atom is called the atom's* **electron affinity**. Most atoms release energy when they acquire an electron. Since energy is given off it will have a negative sign (exothermic).

$$A + e^- \rightarrow A^- + \text{energy}$$

*Electron affinity increases toward F.*
It decreases down a group because there is an increase in atomic radius down a group, which decreases electron affinities. Therefore the nucleus is not as likely to gain an electron.

Noble gases have a 0 electron affinities. They have a full valence level; therefore they do not gain electrons.

Across the periods, the effective nuclear charge is increasing. The shielding effect remains essentially the same. The larger nuclei tend to be more likely to attract electrons other than its own than nuclei that are smaller.

## Task 4e

1. Order the atoms in each of the following sets from the least electron affinity to the most.

   a. O, S

   b. F, Cl, Br, I

   c. N, O, F

   d. B, C, N

## Electronegativity

**Electronegativity** *is a measure of the ability of an atom in a chemical compound to attract electrons from another atom in the compound*. Fluorine is the most electronegative element. It is arbitrarily assigned a value of four. The other values are calculated in relation to this value.

*Electronegativities increase toward F.*

| | Group | | | | | | | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| Period | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 |
| 1 | H 2.1 | | | | | | | | | | | | | | | | | He 0 |
| 2 | Li 0.98 | Be 1.57 | | | | | | | | | | | B 2.04 | C 2.55 | N 3.04 | O 3.44 | F 3.98 | Ne 0 |
| 3 | Na 0.93 | Mg 1.31 | | | | | | | | | | | Al 1.61 | Si 1.9 | P 2.19 | S 2.58 | Cl 3.16 | Ar 0 |
| 4 | K 0.82 | Ca 1 | Sc 1.36 | Ti 1.54 | V 1.63 | Cr 1.66 | Mn 1.55 | Fe 1.83 | Co 1.88 | Ni 1.91 | Cu 1.9 | Zn 1.65 | Ga 1.81 | Ge 2.01 | As 2.18 | Se 2.55 | Br 2.96 | Kr 0 |
| 5 | Rb 0.82 | Sr 0.95 | Y 1.22 | Zr 1.33 | Nb 1.6 | Mo 2.16 | Tc 1.9 | Ru 2.2 | Rh 2.28 | Pd 2.2 | Ag 1.93 | Cd 1.69 | In 1.78 | Sn 1.96 | Sb 2.05 | Te 2.1 | I 2.66 | Xe 2.6 |
| 6 | Cs 0.79 | Ba 0.89 | La 1.1 | Hf 1.3 | Ta 1.5 | W 2.36 | Re 1.9 | Os 2.2 | Ir 2.2 | Pt 2.28 | Au 2.54 | Hg 2 | Tl 2.04 | Pb 2.33 | Bi 2.02 | Po 2 | At 2.2 | Rn 0 |
| 7 | Fr 0.7 | Ra 0.89 | Ac 1.1 | Rf | Db | Sg | Bh | Hs | Mt | Uun | Uuu | Uub | | | | | | |

Note that the noble gases have a 0 electronegativity. Remember that they are very stable and do not normally form compounds so they do not pull additional electrons toward them.

The alkali and alkaline-earth metal are the least electronegative. In compounds, their atoms have a low attraction for electrons.

Nitrogen, oxygen, and the halogens are the most electronegative elements. Their atoms attract electrons strongly in compounds.

Electronegativities tend to either decrease down a group or stay the same.

### Task 4f

1. Using the periodic table, place the following element in order of increasing electronegativity: O, Fe, Ge, Sr, S, Zr

## Other Periodic Properties

### Valence electrons

The number of valence electrons can be determined by the group on the periodic table. Notice that the number of valence electrons is related to the group number.

| Group | # of valence electrons |
|---|---|
| 1 | 1 |
| 2-12 | 2 (vary) |
| 13 | 3 |
| 14 | 4 |
| 15 | 5 |
| 16 | 6 |
| 17 | 7 |
| 18 | 8 (except He, 2) |

### Oxidation Numbers

Oxidation numbers also vary in a regular pattern on the periodic table. Notice that groups 1-13 tend to lose electrons to become positive. Group 14 can either lose or gain 4 electrons. Groups 15-17 gain electrons to become negative. Group 18 is stable, so it neither loses nor gains electrons.

| Group | # of valence electrons |
|---|---|
| 1 | $1^+$ |
| 2-12 | $2^+$ (vary with RN) |
| 13 | $3^+$ |
| 14 | $4^+$,$4^-$ |
| 15 | $3^-$ |
| 16 | $2^-$ |
| 17 | $1^-$ |
| 18 | 0 |