Trends of the Periodic Table Basics

Trends are patterns that atoms on the periodic table of elements follow. Trends hold true "most" of the time, but there are exceptions, or "blips", where the trend seems to do the wrong thing. It is important when deciding a particular trend that you examine at least 4 atoms in a group—or period to see what the numbers do. Choosing just a pair of atoms might show you the exception to a trend rather than the trend itself.

The 7 trends we study in class are these:

1. atomic radius (relative size, measured in picometers, pm)
2. average weighted atomic mass (measured in amu)
3. net nuclear charge (how positive is the nucleus, related to # of protons)
4. ion size (cations or anions)
5. electro negativity (relates to bonding)
6. 1st Ionization Energy (energy required to change a mole of atoms → a mole of +1 cations)
7. metal property or non-metal property (how metallic or non-metallic is this atom?)

Group Trends: the trend that the atoms follow going <u>down</u> any particular group

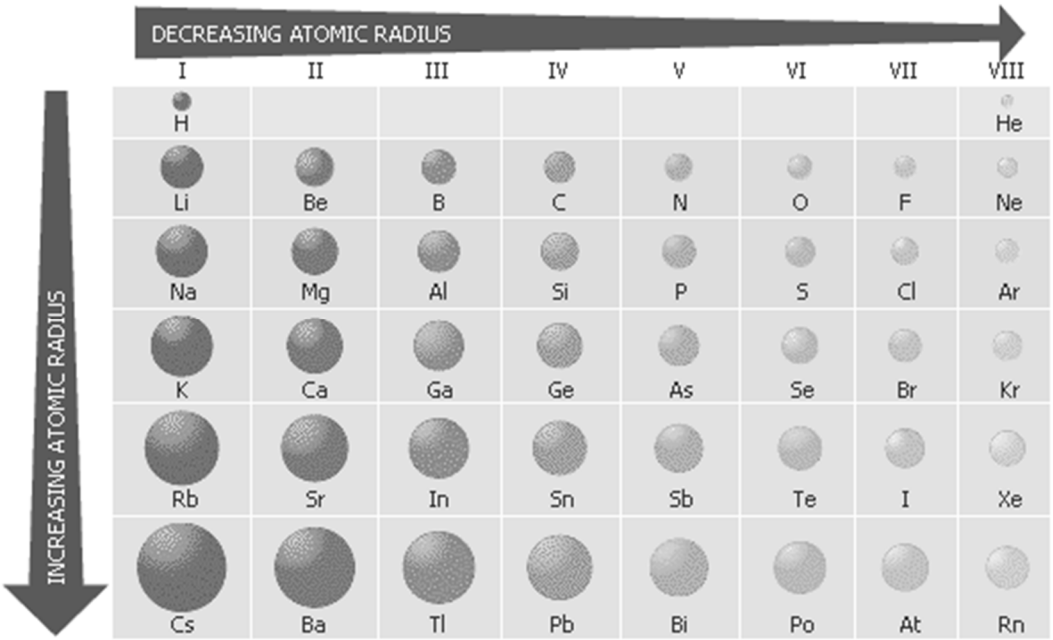
Period Trends: the trend that the atoms follow going <u>across</u> any particular period

Trend #1: Atomic Size

Reference table S shows us atomic radius, which is the measure of distance from nucleus to outer most electron orbital. The measurement is in picometers (1×10^{-12} meters which is trillionths of a meter).

<u>The group trend for atomic size is INCREASING</u>. That is because period 1 atoms (H and He only) have only one orbital. Period 2 atoms have two orbitals—which make them bigger. Period 3 atoms have three orbitals, bigger still, etc. Period number = number of orbitals.
The group trend for atomic radius is increasing, without exception.

<u>The period trend for atomic size is DECREASING</u>. As you go across a period each atom has the SAME number of orbitals, but each box adds another positive proton in the nucleus. These positive protons "pull" the negative electrons closer and closer. Usually the smallest atom in the period would be the noble gas, which has the MOST protons in the period.



Trend #2: Atomic Mass

Atomic mass is measured in amu, atomic mass units. The average weighted atomic mass for each atom is listed on the Periodic Table of Elements. Generally speaking the smallest atoms are those with the lowest atomic numbers, and they get heavier as this number increases.

Atomic mass is a measure of the number of protons and neutrons in a nucleus, we accept that the mass of an electrons is so small we disregard it in high school. Atomic mass is how many particles are in the nucleus.

All atoms have isotopes, chemically identical atoms with different masses because they have different numbers of neutrons. Neutrons are neutral, they don't really affect the chemical properties, so all isotopes for a given atom react the same way.

<u>The group trend for atomic mass is INCREASING</u>. Check this with the Periodic Table. Going down each group on the table each box adds many protons and many neutrons.

<u>The period trend for atomic mass is also INCREASING</u>, as we move to the right we are always adding at least one proton per box, and usually neutrons as well. There are some exceptions (see cobalt → nickel). Exceptions like this are due to the numbers of neutrons in isotopes of certain atoms.

Trend #3: Net Nuclear Charge

The subatomic particles: electrons, protons, & neutrons all have particular charges. Electrons are negative (-1) and are all located outside the nucleus. Neutrons are neutral (Ø) and even though they are in the nucleus, add NO CHARGE to the nucleus. The protons of the nucleus are positively charged (+1) and these protons alone are the measure of net nuclear charge.

This trend is a measure of <u>how positive</u> the nucleus of the atom is. It is measured by how many protons, each with a +1 charge, are in a nucleus of an atom. Since each atom has a certain number of protons (the ATOMIC NUMBER), it's easy enough to count the net nuclear charges.

Examples: He has 2 protons and 2 neutrons in the nucleus, this adds to a +2 net nuclear charge

Ar has 18 protons and 22 neutrons in the nucleus, this adds to a +18 net nuclear charge.

<u>The group trend for net nuclear charge is INCREASING</u>.

<u>The period trend for net nuclear charge is INCREASING</u>.

There are NO exceptions to these trends. This would require a + sign to be true. Helium has 2 protons, but the net nuclear charge for helium is +2. <u>2 is not the same as +2.</u>

Trend #4: Cation and Anion Sizes

Ions come in 2 varieties, cations are atoms that have lost electrons and become positively charged, they are always metals. Anions are atoms that have gained electrons and become net negatively charged, these are always non-metals.

Ions form by gaining or losing enough electrons to get that "special" stable, noble gas electron configuration. When an ion forms, it obtains a noble gas electron configuration, which is called being ISOELECTRIC to a noble gas. These ions are not noble gases, they obtain the same electron configuration as a noble gas.

When an atom becomes a cation it usually loses ALL of its valence, or outermost electrons. Group one atoms all lose only one electron. Group 2 atoms all lose 2 electrons as they become +2 cations. Metals always lose all of their valence electrons, to become isoelectric to a noble gas.

Because cations lose their WHOLE VALENCE ORBITAL when they form,
Cations are always smaller than the atoms they started out as.
The sodium cation is smaller than the sodium atom. The calcium cation is smaller than the calcium atom. The aluminum cation is smaller than the aluminum atom.

Cations are always quite a bit smaller than their atoms—a whole orbital smaller.

When a non-metal atom becomes an anion it gains enough electrons to obtain a full outer orbital, so it can be ISOELECTRIC to a noble gas. Non-metals can gain one, two, or even three electrons to fill the outer valence orbital. When the atoms gain electrons in the valence orbitals this valence orbital must stretch a bit to accommodate this influx of negative charge. The electrons all repel each other, and the extra electron will force all the electrons in that orbital a bit further away from each other.

Anions are always slightly larger than the atoms that they formed from.

Atomic size increases going down a group because each atom lower on the table has more orbitals. Same for cations, although a cations are smaller than the atoms they formed from, each successive cation has more orbitals than the cation above it.

Cations get smaller going across a period as each cation has the same number of electrons as all the other cations in the period, but they gain electrons across the period, so more protons are pulling on the same number of electrons. Cations Na^{+1} , Mg^{+2}, and Al^{+3} all have 10 electrons in a 2-8 configuration but they will have 11, 12, and 13 protons respectively. Sodium is larger than magnesium because Mg has that extra proton pulling the outer orbital in tighter. The aluminum +3 ion is smallest because it has yet another proton pulling on the same number of electrons in the same number of orbitals, 13 positive protons pulls "harder" than 12.

The Group Trend for Cation size is INCREASING.

The Period Trend for Cation size is DECREASING.

Anions are all slightly larger than the atoms they formed from because they add extra electrons into the outer orbital, where it can fit, but this forces all the negatively charged electrons a bit further away from themselves. Looking at group 16, oxygen sulfur and selenium, all get larger as atoms moving down the table. Each anion is larger than it's atom, and the anions get progressively larger as well. Each anion below has one more orbital than the ion above it.

For the period trend of anion size let us look at period 3: phosphorous, sulfur and chlorine.
The atoms get progressively smaller going across the period (more protons pulling on the same number of electron orbitals).

The anions for these three form as P^{-3}, S^{-2}, and Cl^{-1} respectively. Each has TEN electrons in the same 2-8 configuration, but since there are more and more protons in these ions moving across the period, the anions get smaller moving across the period.

The Group Trend for Anion size is INCREASING.

The Period Trend for Anion size is DECREASING.

Trend #5: Electronegativity

Electronegativity is the tendency of an atom to attract electrons to itself when it makes a bond.

When atoms make covalent bonds, they "share" electrons. Each atom puts up one electron, and when a pair of electrons is shared, a single covalent bond forms.

How well they share is measured using the difference in their electronegativity values. The values were set by Linus Pauling (Nobel prize for this work). The higher the Electronegativity value, the harder the atoms pulls electrons to itself. The lower the value, the less attraction the atom has for these bonding electrons.

If two atoms in the bond have the SAME electronegativity values, there is NO difference in EN Value, which means the atoms share the electrons evenly, or, the bond is NONPOLAR.

If the atoms in a bond have different EN Values, the atom with the higher EN Value "pulls" harder, getting the bonding electrons more of the time, making that side of the bond NEGATIVE. This leaves the atom with the lower EN Value get the electrons in the bond less often, which leaves that side of the bond mostly POSITIVE.

The greater the difference the less fairly the atoms share the electrons, which makes the bonds MORE POLAR

Examples of NONPOLAR BONDS include: H_2, Cl_2, and Br_2.
Each shares a single pair of electrons, forming a nonpolar covalent bond. Since both of the H atoms have the same 2.2 EN Value, the difference between these values is zero. That makes this bond nonpolar.

With HCl, hydrogen monochloride forms, the H atom (2.2 EN Value) and the Cl atom (3.2 EN Value) have a big difference in EN Values— 1.0 difference. This bond is a POLAR BOND.

When H_2O forms, there are 2 H atoms that bond to ONE oxygen atom. Each bond is individual, they both get examined separately. The H atom (2.2 EN Value) and the O atom (3.4 EN Value) make a 1.2 EN Value difference. This bond is polar. So is the OTHER H—O bond in this same molecule. Water makes two polar covalent bonds inside the one molecule.

Comparing H—Cl and the H—O bonds, we compare the differences in EN Values. The H—O bond is MORE POLAR tha the H—Cl bond. We can use the EN Value differences to "rank" bond polarity.

In Regents Chem bonds are NONpolar with a zero EN Value difference, or they are POLAR because there is an EN Value difference. In college you may see that sometimes a very small EN Value difference is called nonpolar, but not in high school chem.

Linus Pauling created this concept of electronegativity and determined that fluorine has the greatest tendency to gain electrons in a bond compared to ALL OTHER ATOMS. Since he created this scale, he could do what he wanted, and he did.

Dr. Pauling <u>decided</u> that F would has an EN value of 4.0, the highest value on the table. All other atoms would be compared or ranked to Fluorine. Since all atoms are compared to a single "standard" atom, this makes electronegativity a relative scale. All atoms are measured, in relation to, or compared to a standard that Pauling decided upon. Electro negativity is <u>an example</u> of a relative scale.

Electronegativity does not have units.

He choose the numbers 0 to 4.0 for no particular reason. The numbers don't "mean" anything, they are just numbers that "rank" the atoms. Electronegativity is a relative scale, and it is ARBITRARY as well.

Table S shows all the EN values for the elements. Some atoms, the smaller noble gases have no EN values. These atoms do not make bonds ever, they have NO TENDENCY to gain (or lose) electrons.

The Group Trend for electro negativity is decreasing.

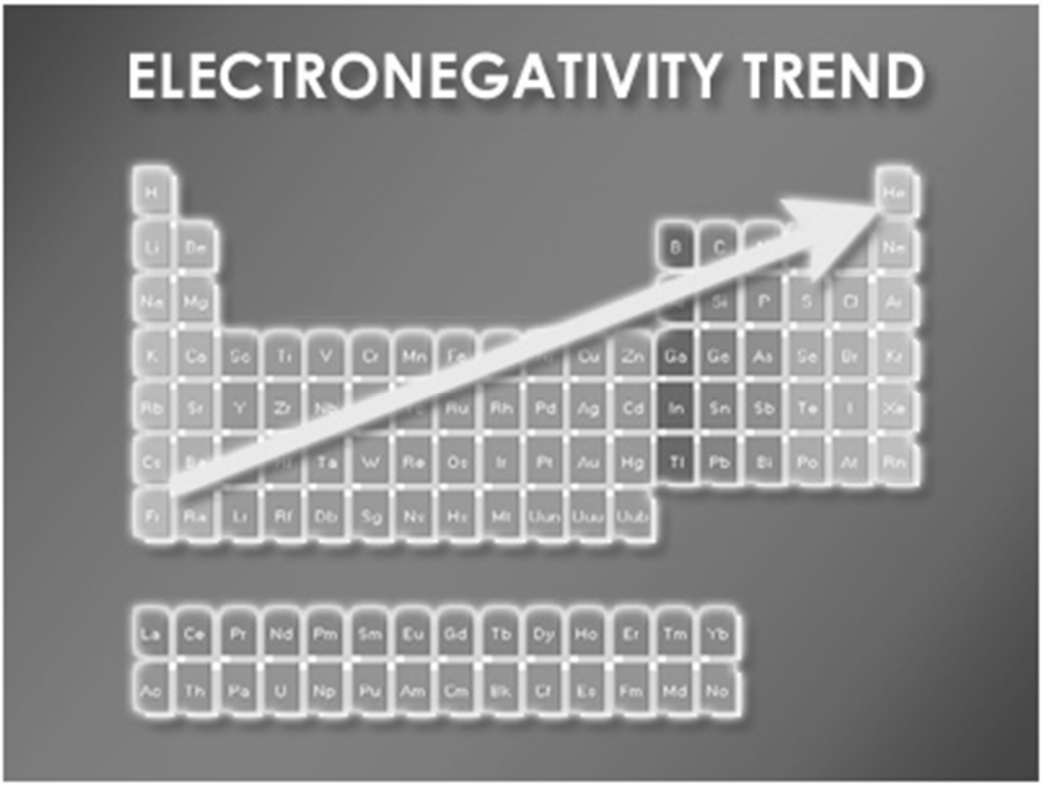
The Period Trend for electro negativity is increasing.

The general table trend is the closer you are to F, the higher the EN value (except for noble gases).

The noble gas XENON does have an EN value, and under some unusual conditions (very high pressure and catalysts) it can be forced into some bonds. This is an exception, noble gases tend to be relatively INERT, or non-reactive.

The period trend for electronegativity is increasing, because in any period you have one orbital, and more and more protons pulling tighter and tighter. There is a greater inward attraction to gain electrons as you move across the table. Until you get to group 18, which already has full orbitals.

The group trend for electronegativity is decreasing. This seems odd at first, more and more protons are in the atoms going down any group, but this increased positive charge in the nucleus pulling inwards is offset by the distance the outermost valence orbital is to the nucleus. More protons helps pull inward, distance hurts more than the increased number of protons helps.



Trend #6: 1st Ionization Energy

When atoms of group 1 become cations and "lose" an electron, even though they "want" to do this to gain the noble gas electron configuration, it requires some energy. The electrons do not FALL OFF of the atoms. 1st Ionization energy is the energy required to pull a mole of electrons off of a mole of atoms.

The amount of energy required to take a mole of atoms and make them a mole of +1 cations is called the 1st IONIZATION ENERGY. The unit is kJ/mole or kilo-joules per mole.

The metals have lower first ionization energy requirements for several reasons. Metals will tend to lose electrons easier than nonmetals. Metals form cations. Nonmetals gain electrons to form anions.

The largest atoms in any period are in group 1, and these atoms have the lowest 1st ionization energy. Moving across any period, you have the same number of orbitals, but more and more protons, pulling the atoms smaller. For the same reason that the atoms get smaller (greater inward attraction) the more difficult it becomes to pull these electrons off. <u>The period trend for first ionization energy is increasing</u>.

Going down any group, since the atoms are getting larger and larger, even though there are more protons, the inward attraction the nucleus on these electrons weakens over this greater atomic radius.
<u>The group trend for 1st ionization energy is decreasing</u>.

A mole of atoms can be converted into a mole of +1 ions by applying the 1st ionization energy. Anything (almost) is possible. Making fluorine a +1 cation is possible, but hard. It requires a lot of energy compared to lithium. Noble gases can be forced into +1 cations too. That is really hard to do—but possible.

Atoms like Mg and Ca make +2 ions. Al makes a +3. To convert a mole of Mg into a mole of +2 ions requires the application of the 1st Ionization energy PLUS the application of the 2nd Ionization energy (the energy required to remove a second electron from a mole of +1 ions to make them form into +2 ions.)

There is also a 3rd Ionization energy. (We concern ourselves only with the 1st level energy, not the others. I add it here because you are smart enough to understand it, and also smart enough to let it go).

Noble gases tend to be "perfect" atoms, and it's very hard to remove their electrons, as the Table S shows us with very high 1st ionization energies. The atom with the highest 1st ionization energy is helium, only two electrons but super small. It's so hard to remove those electrons!

Trend #7: Metallic Property and Non-Metallic Property

Metals are on the left side of the Periodic Table of Elements. They have many properties that make them "metallic", such as: luster, electrical conductivity, heat conductivity, they're malleable, ductile, they have low specific heat capacity constants, higher density, higher melting point, and only form cations, etc.

If each property could somehow be ranked, if we "measure" all metals against each other, the MOST METALLIC METAL would be Francium, Fr.

In fact, the closer to Fr a metal is on the table (LITERALLY, in inches) the more metallic it is.
Example: polonium, lead, silver and zirconium are all metals. Zr is the "closest" to Fr on the table, therefore, Zr is the most metallic of these four metals.

Non-metals are on the right side of the table (plus Hydrogen). They have pretty much the opposite properties of metals. Nonmetal solids are brittle, not able to change shape. They don't conduct heat or electricity, they form only anions, and are dull colored. The nonmetal gases are clearly not metallic.

If these nonmetals were ranked (which is the most nonmetallic of them all?) Helium would come up as the MOST NON-METALLIC nonmetal. The closer an atom is to He on the table, the more non-metallic it is.

Example, C, Cl and Ne are all non-metals and neon is the closest on the table to helium, so Ne is the most non-metallic of these three non-metals.

Sometimes kids think of crazy questions, like which is more metallic, Cesium or radium, or which is more non-metallic, F or Ne. These questions cannot be answered by our simple "proximity" rule, so don't worry, no one will ever ask those questions of you.

Parts of the Periodic Table of Elements

Even the name of the table is important. When the atoms are arranged in increasing order of atomic number there is a periodic repetition of their chemical properties, in the groups. That is the PERIODIC LAW and why it's the PERIODIC table of the elements.

Group 1 = the alkali metals
Group 2 = the alkaline Earth metals
Group 3 → 12 <u>plus</u> some more atoms under the staircase = Transitional metals
Group 17 = the halogens
Group 18 = the Noble Gases
The bottom 2 rows of the table are detached. These are the Inner Transitional metals.
They all fit into GROUP 3, in periods 6 + 7

Hydrogen is an exception, it's a non-metal that *sometimes* acts like a group 1 metal when bonding

7 Metalloids touch the staircase line from group 13 down into group 17, with 2 exceptions: Al & Po.

All atomic masses are based against carbon-12, an atom with 6 protons and 6 neutrons. It's said to have the exact mass of 12 amu. One AMU is one twelfth the mass of one C-12 atom.

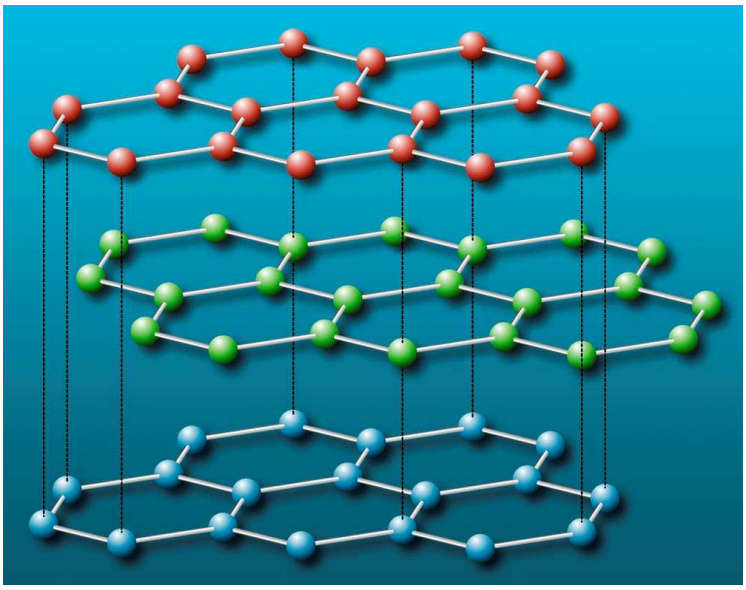
At STP, all metals are solids, except for Hg, which is liquid.
At STP most non-metals are gases, but Br is a liquid and B, C, Si, P, S, As, Se, Te and I are solids.

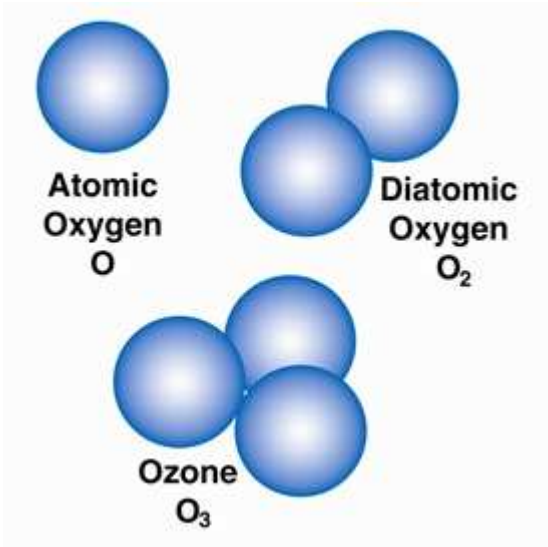
The modern periodic table was first devised by Dimitri Mendeleev, a Russian scientist. It was a tremendous achievement of figuring out a pattern for 62 plus elements into a table that had no particular shape. (remember the puzzle lab!)

ALLOTROPES

An allotrope is a pure form of an element, but it is bonded together in a different way, so it has some different properties. Examples include CARBON: in the form of graphite, diamond, and "bucky balls".
Another example is oxygen as O_2 we breathe and O_3 called ozone. Oxygen and ozone are allotropes of oxygen.

3 allotropes of carbon, All forms of the pure element,
but in different structure, with different properties.





O_2 (oxygen) + O_3 (ozone) are
allotropes of oxygen.

"O" atoms do not normally occur on Earth
except under very high temperatures.