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 investigating a particular trend that you examine at least four atoms in a group or period and see what the trend numbers are doing. Choosing just one pair of atoms might show you the exception to a trend rather than the trend itself.

The seven 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 down any particular group

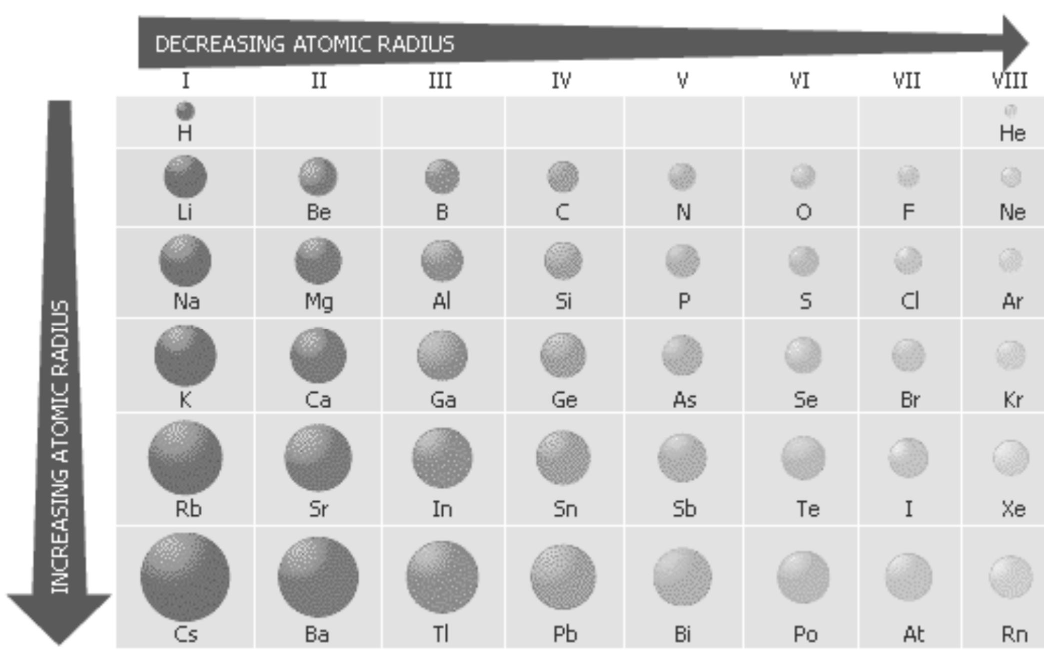
Period Trends: the trend that the atoms follow going across any particular period

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).

The group trend for atomic size is INCREASING. That is because each atom that follows going down a group has <u>one more orbital</u> than the atom above it. Three orbitals are larger than two orbitals, four orbitals are larger than three.

The period trend for atomic size is DECREASING. As you go across a period you always keep the <u>same number of electron orbitals</u> but you are adding with each atom an extra proton. The more protons pulling on the same number
of electron orbitals pulls the atoms smaller and smaller as you go to the right on the table. The smallest atom in any period is the noble gas, that's the atom with the MOST number of protons with that particular number of orbitals.



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, as we accept that the mass of electrons is so small that 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 chemistry or properties (other than mass), so all isotopes for a given atom react the same way.

The group trend for atomic mass is INCREASING. Check this with the Periodic Table.

The period trend for atomic mass is also INCREASING, but there are some exceptions (see cobalt and nickel). Exceptions like this are due to the numbers of neutrons in isotopes of certain atoms.

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 are the measure of net nuclear charge.

This trend is a measure of how much positive charge is in the nucleus of the atom, which 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.

The group trend for net nuclear charge is INCREASING.

The period trend for net nuclear charge is INCREASING.

There are NO exceptions to this trend. 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>

Ion Size

Ions come in two varieties, cations are atoms that have lost electrons and become positively charged, and are always metals. Anions are atoms that have gained electrons and become net negatively charged, and 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 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 a whole outer (valence) orbital, they 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 the atoms

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 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 cation is smaller than its atom, each successive cation has more orbitals than the previous.

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 e^- in a 2-8 configuration but have 11, 12, and 13 protons respectively. Sodium is larger than magnesium because Mg has that extra proton pulling the outer orbital in. 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.

The Group Trend for Cations is INCREASING.
The Period Trend for Cations is DECREASING.

Anions are all larger than their atoms because they squeeze an extra electron into the outer orbital, where it can fit, but 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.

For a period trend, look at period 3, phosphorous, sulfur and chlorine. The atoms get progressively smaller going across the period. Anions for these three form as P^{-3}, S^{-2}, and Cl^{-1} respectively. 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 Anions is INCREASING.
The Period Trend for Anions is DECREASING.

Electro negativity

Electronegativity is the tendency of an atom to attract electrons to itself in 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. Examples include: H_2, Cl_2, and Br_2. Each shares a single pair of electrons, each makes a single covalent bond. Since these atoms are identical to each other, and they have the same tendency to attract an electron to itself (they have the same electronegativity values), the bond is a single NON-POLAR covalent bond, neither atom "gets" those electrons more than the other atom does.

When O_2 forms, each oxygen has 6 valence electrons, and each "needs" to borrow 2 electrons from the other atom in the bond, they form a DOUBLE covalent bond, sharing 2 pairs of electrons. The both oxygen atoms have the same electronegativity value, so it's a double NON-POLAR bond.

N_2 has two nitrogen atoms, each with 5 valence electrons, each "needing" to borrow 3 more electrons to fill the outermost valence orbital. To do this, they both "share" three electrons, sharing these three pairs of electrons is called a TRIPLE NON-POLAR bond. It's non-polar because again, the two atoms are identical, neither "pulls" electrons stronger than the other.

Electronegativity truly comes into play when atoms with different electronegativity values bond. A common example is the H-Cl bond in hydrogen monochloride . Since the electronegativity values for H and Cl are 2.2 + 3.2 respectively, there is a difference in electronegativity. This shows the chlorine gets that electron from hydrogen more than hydrogen gets it from the chorine. This bond is not "balanced", it's POLAR. It makes a single, polar covalent bond.

When CO_2 forms, carbon and each oxygen form double bonds, but since the electronegativity values are 2.6 and 3.2, the oxygen atoms pull these electrons in the bond more so to the oxygen side of the bond.
This bond is a double polar covalent.

When you compare bonds, such as H-Br, H-Cl, and H-F, the greater the difference in electronegativity, the greater the bond polarity. The E.N. differences here are 0.8, 1.0 and 1.8 respectively. So, the H-F bond is the most polar, the H-Br bond is least polar.

ANY difference in electronegativity means the bond is polar. If the difference is ZERO, the bond is non-polar.

Linus Pauling created this concept of electronegativity and measured that fluorine has the greatest tendency to gain electrons in a bond. Since he created this scale, he could do what he wanted, and he did just that. Dr. Pauling <u>decided</u> that F would have an EN value of 4.0, the highest value on the table. All other atoms would be compared to Fluorine. Since all atoms are compared to a single "standard" atom, this is a relative scale. All atoms are measured, relative or compared to a standard that HE 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 to "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.

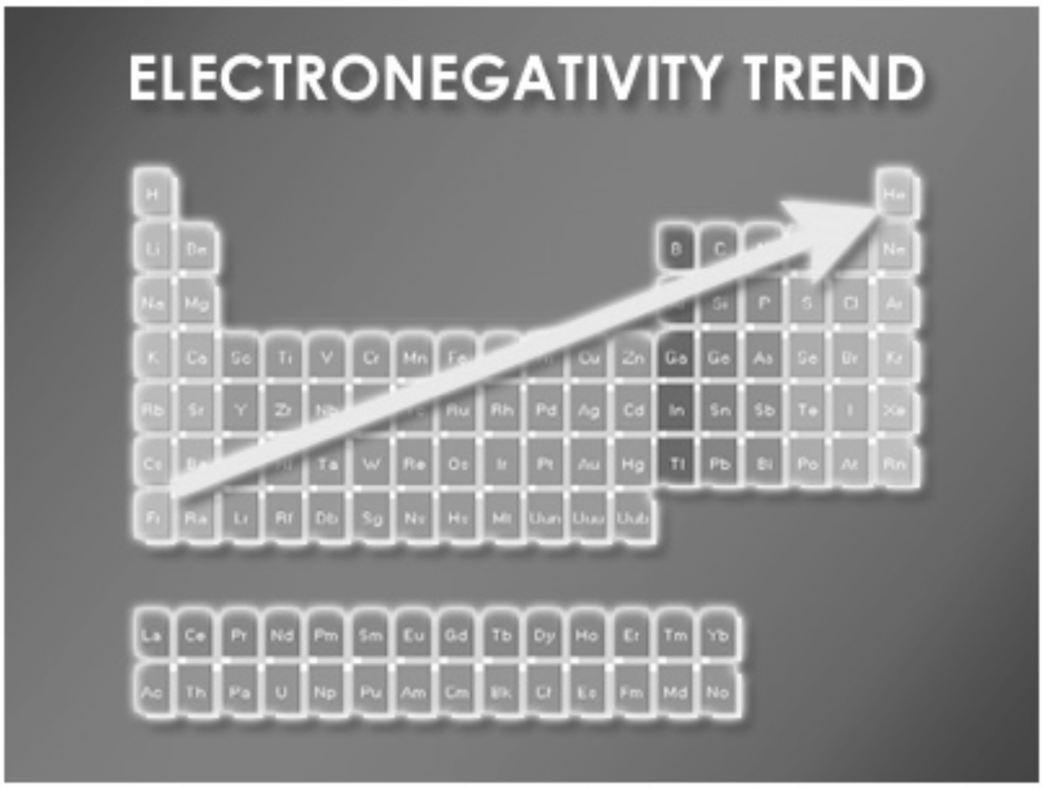
The Group Trend for electro negativity is decreasing.
The Period Trend for electro negativity is increasing.

The 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 it can be forced into a bond. 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.



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 whole mole of electrons off of a whole 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. The group trend for first ionization energy is increasing.

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 distance. The group trend for 1st ionization energy is decreasing.

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!

Metallic Property and Non-Metallic Property

Metals are on the left side of the Periodic Table of Elements. They have a variety of properties that make them "metallic", such as: luster, electric conductivity, electric conductivity, malleable, ductile, low specific heat capacity, higher density, higher melting point, 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, the more metallic it is. Use this idea to rate or rank groups of metals. 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. The gases are weirder. They also tend to form anions, or no ions, they don't conduct heat, don't conduct electricity, and are dull rather than lustrous.

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

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 tests, 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. The properties of the elements periodically repeat themselves, so that is why it's the PERIODIC table.

Group 1 = alkali metals
Group 2 = alkaline Earth metals

Group 17 = halogens
Group 18 = Noble Gases

hydrogen is the exception, it's a non-metal that acts like a group 1 metal in bonding

Transitional metals stretch the middle of the table, plus some under groups 13 to 16.
Inner Transitionals are at the bottom of the chart, and all fit into GROUP 3, and periods 6 + 7
Under Y-39 fits 57-71. Under that comes 89-103. All of these are group 3 metals

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 some are solids as well.

The modern periodic table was first devised by Dimitri Mendeleev, a Russian chemist. It was a tremendous achievement of figuring out a pattern for 70 plus elements into a table that had no particular shape. (remember the puzzle lab!)

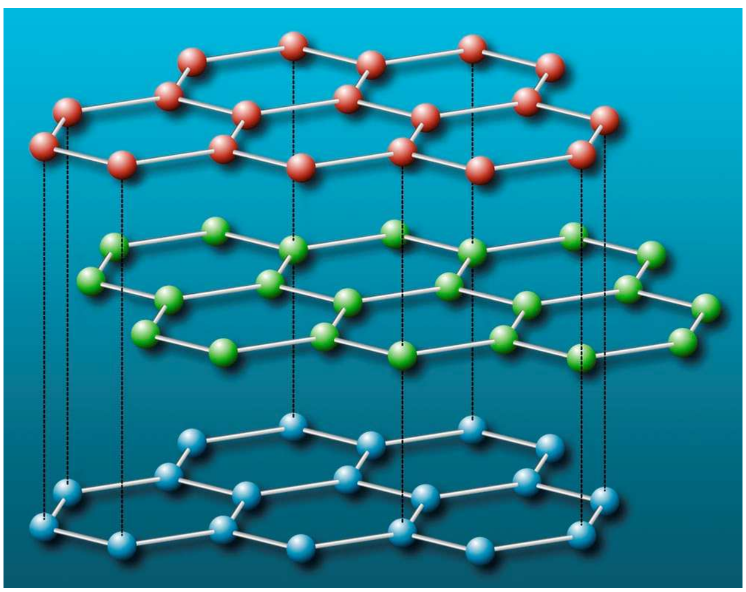
The Periodic Law

When the elements are arranged in order of increasing atomic number, there is a periodic repetition of their chemical and physical characteristics in the groups.

ALLOTROPES

An allotrope is a pure form of an element, but it is bonded together in a different way, so it has 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 + O_3 (ozone) are also allotropes of oxygen.