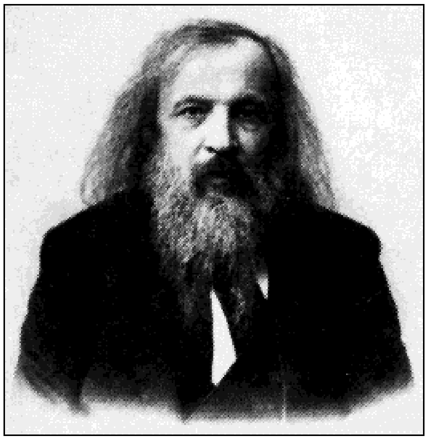
Chapter 5 Trends of the Periodic Table Diary

A Brief History of the Periodic Table

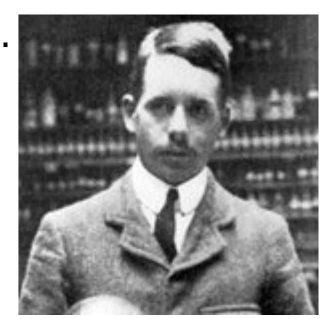


"I began to look about and write down the elements with their atomic weights and typical properties, analogous elements and like atomic weights on separate cards, and this soon convinced me that the properties of elements are in periodic dependence upon their atomic weights."
--*Mendeleev*, Principles of Chemistry, *1905, Vol. II*

The modern periodic table was first devised by Dmitri 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. Using the table, he predicted the properties of elements yet to be discovered. Mendeleev's original work covered a wide range, from questions in applied chemistry to the most general problems of chemical and physical theory. His name is best known for his work on the Periodic Law.

Reihen	Gruppe I. — R^2O	Gruppe II. — RO	Gruppe III. — R^2O^3	Gruppe IV. RH^4 RO^2	Gruppe V. RH^3 R^2O^5	Gruppe VI. RH^2 RO^3	Gruppe VII. RH R^2O^7	Gruppe VIII. — RO^4
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	—=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	—=100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

Between 1868 and 1870, in the process of writing his book, *The Principles of Chemistry,* Mendeleev created a table or chart that listed the known elements according to increasing order of atomic weights. When he organized the table into horizontal rows, a pattern became apparent--but only if he left blanks in the table. If he did so, elements with similar chemical properties appeared at regular intervals--periodically--in vertical columns on the table.



Henry Moseley 1887-1915) subjected known elements to x-rays. He was able to derive the relationship between x-ray frequency and number of protons. When Moseley arranged the elements according to increasing atomic numbers and not atomic masses, some of the inconsistencies associated with Mendeleev's table were eliminated. The modern periodic table is based on Moseley's Periodic Law (atomic numbers). At age 28, Moseley was killed in action during World War I and as a direct result Britain adopted the policy of exempting scientists from fighting in wars. Shown below is a periodic table from 1930:

Group 0	I a b	II a b	III a b	IV a b	V a b	VI a b	VII a b	VIII
	H 1							
He 2	Li 3	Be 4	B 5	C 6	N 7	O 8	F 9	
Ne 10	Na 11	Mg 12	Al 13	Si 14	P 15	S 16	Cl 17	
Ar 18	K 19 / Cu 29	Ca 20 / Zn 30	Sc 21 / Ga 31	Ti 22 / Ge 32	V 23 / As 33	Cr 24 / Se 34	Mn 25 / Br 35	Fe 26, Co 27, Ni 28
Kr 36	Rb 37 / Ag 47	Sr 38 / Cd 48	Y 39 / In 49	Zr 40 / Sn 50	Nb 41 / Sb 51	Mo 42 / Te 52	– / I 53	Ru 44, Rh 45, Pd 46
Xe 54	Cs 55 / Au 79	Ba 56 / Hg 80	57-71* / Tl 81	Hf 72 / Pb 82	Ta 73 / Bi 83	W 74 / Po 84	Re 75 / –	Os 76, Ir 77, Pt 78
Rn 86	–	Ra 88	Ac 89	Th 90	Pa 91	U 92		

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 Transitional are at the bottom of the chart, and all fit into GROUP 3.
Under Y-39 fits 57-71. Under that first 89-103. All of these are group 3 metals

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.
At STP most non-metals are gases, but Br is a liquid and some are solids as well.

You can use the melting and boiling points listed on table S to determine the state of matter.

Periodic Trends

Trends are patterns of behaviors 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 their numbers are doing. Choosing just a pair of atoms might show you the exception to a trend rather than the trend itself.

The seven trends we study in class are:

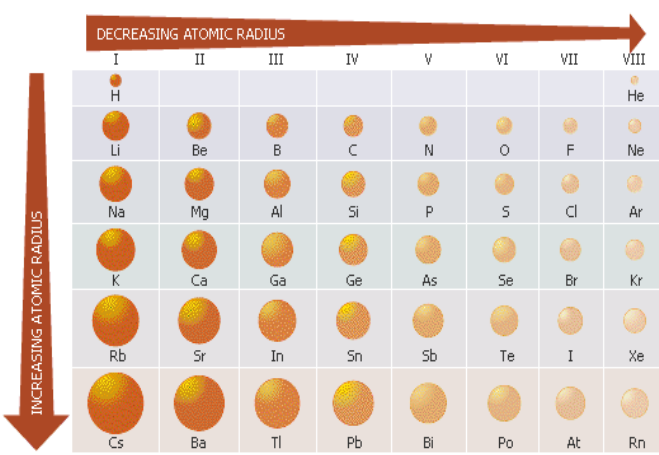
1. atomic size (measured in picometers, pm, as atomic radius)
2. atomic mass (measured in amu)
3. net nuclear charge (how many protons are in the nucleus)
4. ion size (cations or anions)
5. electro negativity (a relative scale)
6. 1st Ionization Energy (removing a mole's worth of electrons from a mole of atoms)
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 one more orbital 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 keep the same number of electron orbitals but you are adding with each atom an extra pair of opposite charges. The increasing pairs of electrons/protons pulling on each other pulls the atoms smaller and smaller as you go to the right on the table.

Atomic Mass

Atomic mass is measured in amu, atomic mass units. The average 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. For our purposes. atomic mass is how many particles are in the nucleus.
Most 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 isotopes of the atoms.

Net Nuclear Charge

The three subatomic particles, electrons, protons, and 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 many protons, each with a +1 charge, are in a nucleus of an atom. The modern periodic table is arranged in order of increasing atomic number. Since each atom has a certain number of protons (the ATOMIC NUMBER), it's easy enough to count the net nuclear charges.

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

Argon (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.**

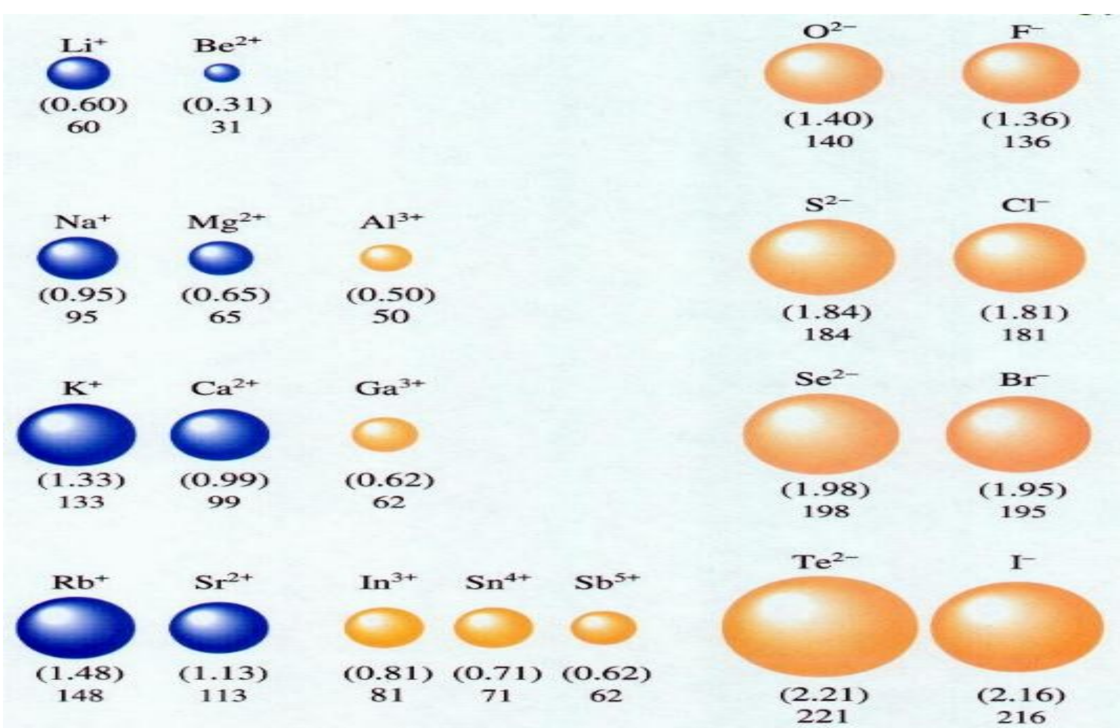
Ion Size

Ions come in two varieties, cations are atoms that have lost electrons and become net 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 to get those atoms perfectly filled electron orbitals, so their electron orbitals can match the NOBLE GAS electron configurations. Atoms like to be neutral (they all are) but they love to have full outer orbitals (like noble gases).

When an atom becomes a cation it loses ALL of its valence electrons. For group one atoms this means one electron. Group 2 loses 2 electrons and they become +2 cations. They lose the electrons to become cations, but they always lose enough electrons that they lose that whole outer orbital as well. **SO, cations are always smaller than their atoms.**

When an atom becomes an anion it gains electrons to create a full outer orbital, so it can be like 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 the orbitals 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. **SO, anions are always 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.



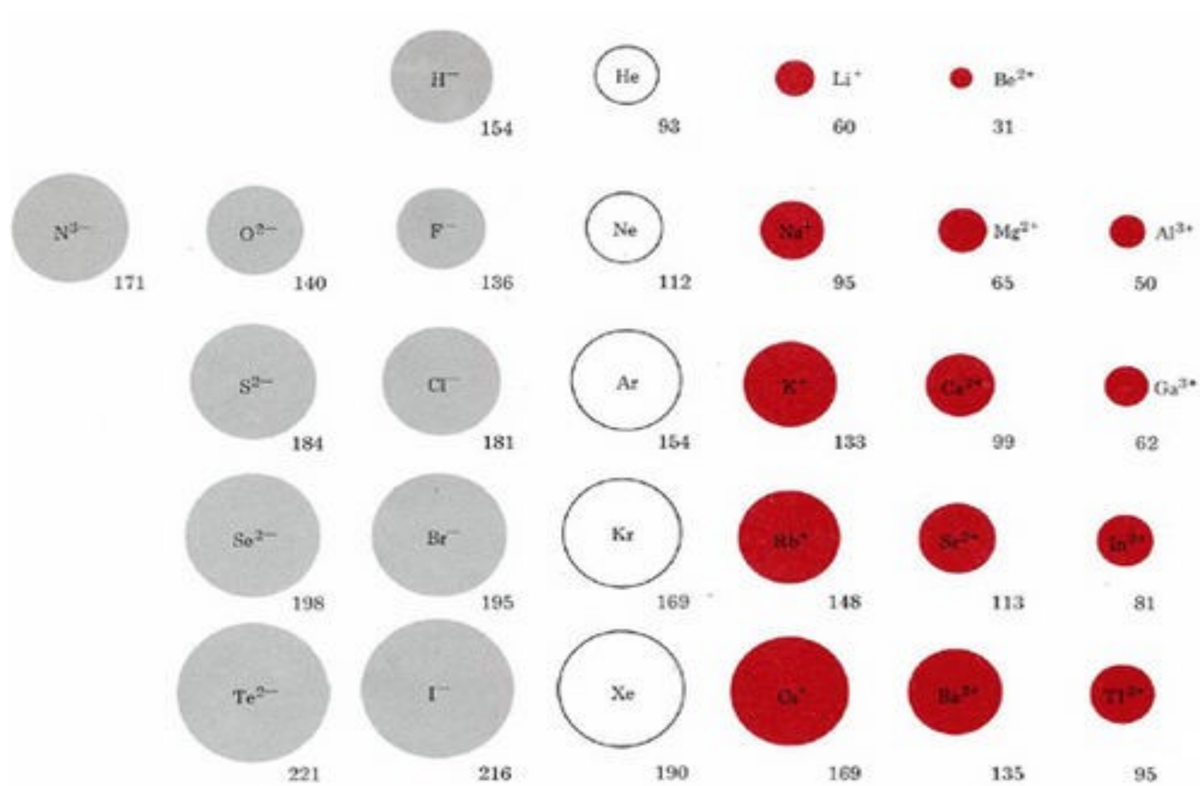
Cations get smaller going across a period as each cation has the same number of electrons as all the other cations in the period, but they gain electrons across the period, so more protons are pulling on the same number of electrons. Cations Na^{+1} , Mg^{+2}, and Al^{+3} all have ten electrons in a 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 tend to get smaller moving across the period.

The Group Trend for **Anions is INCREASING.**
The Period Trend for **Anions is DECREASING.**



Electronegativity

Electronegativity is defined as the ability of an atom to attract electrons to itself in a chemical reactions/bond.

Linus Pauling created this concept and decided 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 decided that F would have an EN value of 4.0, the highest on the table. All other atoms would be compared to F, relative or compared to a standard that HE ARBITRARILY decided upon.

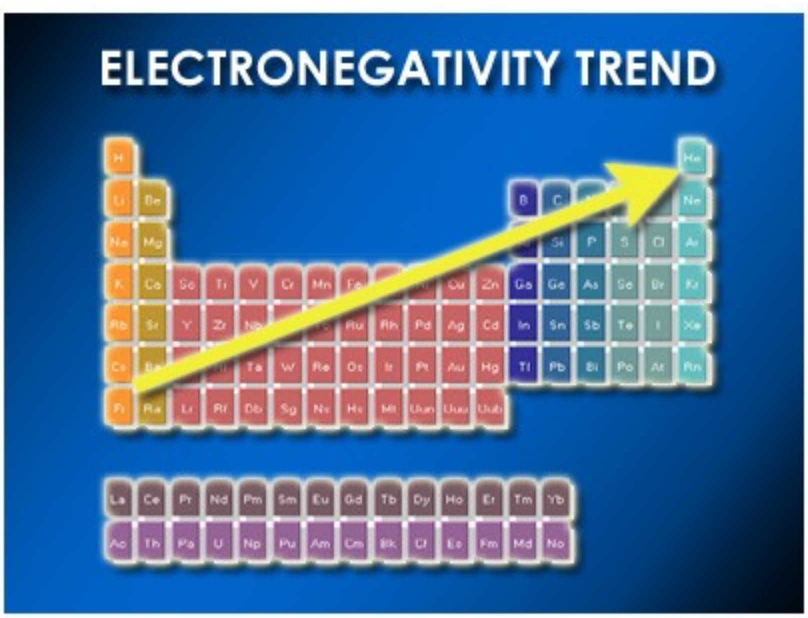
Electro negativity is <u>an example</u> of a relative scale. A relative scale can have units, or not. Electronegativity does not have units.

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

The Group Trend for **electronegativity is decreasing**.
The Period Trend for **electronegativity is increasing.**
The table trend is the closer you are to F, the higher the EN value (except for noble gases).

Krypton does have an EN value, and it can make some bonds under some unusual conditions. It’s an exception to the noble gases being “inert”, it clearly is not.

Group 1 atoms have a much lower EN value than the non-metals, since these atoms want to lose an electron to become +1 cations. It’s relatively easy to give away an electron, as they have little tendency to gain electrons.



Group 16 and 17 non-metals have much higher EN values. They want to gain electrons to become anions.

The lower you go on the table (down a group) the lower the EN value is. The further away you are from the nucleus, and the positive charges, the less pull that the nucleus has to take in more electrons.

1st Ionization Energy

When atoms of group 1 become cations and "lose" an electron, even though they "want" to do this to gain that coveted noble gas electron configuration, it requires some energy. 1st Ionization energy is just that. The amount of energy required to remove the outermost electron in the valence shell is called the FIRST IONIZATION ENERGY. Group 1 atoms have a lower level than the non-metals because they are much easier to convert into +1 cations. The further down a group you go, the energy required to do this decreases, because the electrons that have to be removed are further from the nucleus, and that distance makes it easier to pull the outer electron away.

The group trend for **1st Ionization energy is decreasing.**
The period trend for **1st Ionization energy is increasing.**

Going across a period, especially when you move to the non-metal side of the table, atoms want to gain electrons to get a complete outer orbital. Losing electrons is opposite of what is "normal" for these atoms. It's possible, however, the energy required to do so increases dramatically.

Atoms like Mg and Ca make +2 ions. Aluminum makes a +3. To convert an atom of Mg into a +2 ion 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 the +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.

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.

Atomic Number	Symbol	Name	First Ionization Energy (kJ/mol)	Electro-negativity	Melting Point (K)	Boiling* Point (K)	Density** (g/cm^3)	Atomic Radius (pm)
1	H	hydrogen	1312	2.1	14	20	0.00009	37
2	He	helium	2372	—	1	4	0.000179	32
3	Li	lithium	520	1.0	454	1620	0.534	155
4	Be	beryllium	900	1.6	1551	3243	1.8477	112
5	B	boron	801	2.0	2573	3931	2.340	98

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, and 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. We can 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. If they were ranked as well, **Helium, HE , 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. Example, C, N, and Cl non-metals, and Chlorine is the closest on the table to Helium, so Cl is the most non-metallic of these three non-metals.