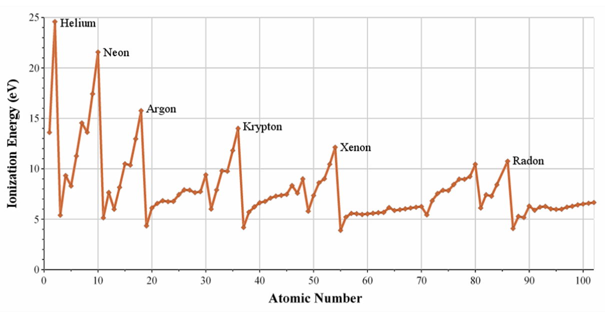
7.12 - Periodic Trends



Introduction to Periodic Trends

Now that we have looked at the ordering of electrons in the periodic table, it is meaningful to discuss elemental characteristics established by that electronic distribution.

The trends we shall be looking at are:
A. Atomic Radius - how big an atom is,
B. Ionization Energy - how much energy is needed to remove electrons,
C. Electronegativity - how much attraction an atom has towards its neighbor's electrons,
D. Electron Affinity - the energy change associated with the addition of an electron to a gaseous atom.

A. Atomic Radius Trends

The radius of an atom is defined as half the distance between the nuclei in a molecule consisting of neighboring atoms (ideally of the same element).
In cases where a variety of bonds play a part (for example carbon, where single, double, and triple bonds occur), an average is derived.
In metals, the distance between neighboring atoms in a metallic crtystal are obtained.
Due to electron cloud overlap, the expected size of atoms is less than the 90% electron density volumes of isolated atoms (the probability model).

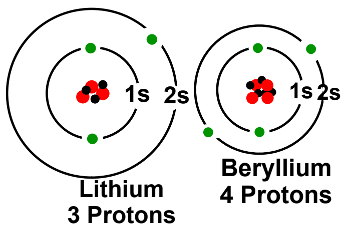
A. Atomic Radius Trends

The trend: left to right across a row, size *decreases*. Counterintuitive!

While there are blips as one goes from one sublevel to another, the overall trend is a general decrease in size.

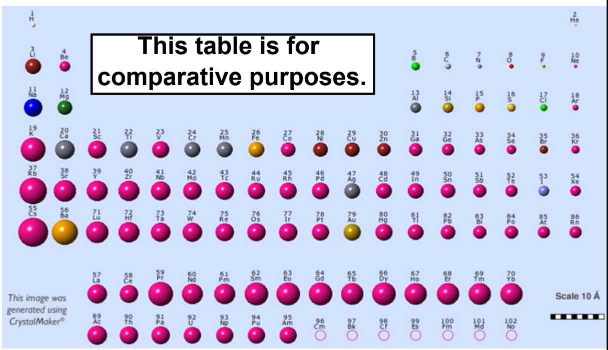
Explanation: the nucleus has greater pull on electrons going into the SAME energy level, and as one tracks along a row, quantum physics dictates that electrons go sequentially into available orbitals of the same (or lower) energy level.

Example:
Lithium vs. Beryllium.



Radius Trends

Radius increases down a group: adding electrons to higher energy levels requires more distant orbitals. Also, inner electrons repel outer ones.



Ionic Radius Trends

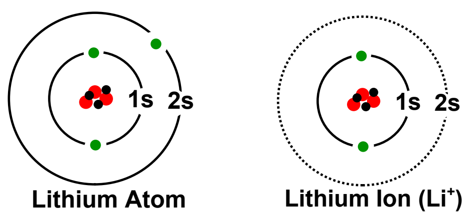
It is worth noting how sizes of atoms change as they become ions by gaining or losing electrons.

Atoms losing electrons get <u>smaller</u> for three reasons:

1. Valence electrons lost – orbitals become empty.

2. Electrostatic repulsion of electrons is less.

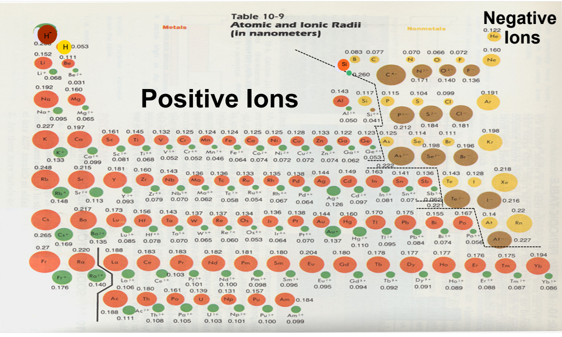
3. Nuclear pull is proportionally greater.



Ionic Radius Trends

Atoms gaining electrons become larger.
A. More repulsion between electrons.
B. Nucleus pulls electrons less (farther away).

Sizes of ions tend to get smaller across a period, and larger down a group (same as atomic radius trends). NOTE: as you go from cations to anions, there's a large jump in size, but the trend continues after that.



Ionic Trend Examples

1. How does a sulfide ion compare to a sulfur atom?
   Sulfide ion (S^{2-}) will be bigger that its atom because it has two more electrons, which repel each other.

2. How does a sodium ion compare to its atom?
   The sodium ion (Na^+) will be smaller that its atom because it lost its valence electron, thus clearing out its 3s orbital.

3. Rank the following ions in order of increasing size: K^+, Zn^{2+}, Ga^{3+}, Ca^{2+}.
   Ga^{3+}, Zn^{2+}, Ca^{2+}, K^+.

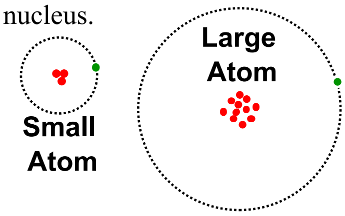
B. Ionization Energy

Energy required to remove an electron from a gaseous atom to make a *positive* ion.

$$\boxed{X_{(g)} \rightarrow X^+_{(g)} + e^-}$$

Trends: ionization energy increases from left to right across a period, and decreases down a group.

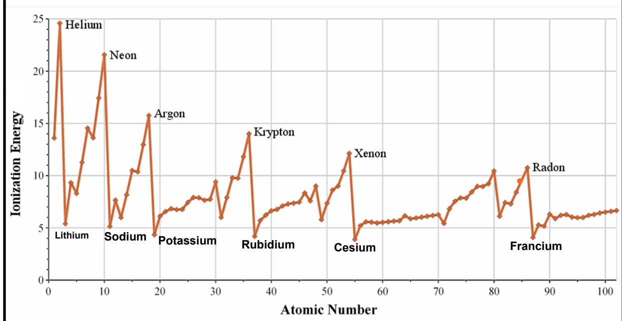
Bigger atoms have valence electrons farther away from the nucleus, so there is less attraction to overcome (even with more protons). Also, the electrons in lower energy levels repel those in outer levels, thus to some extent shielding the outer electrons from the nucleus.



Ionization Energy Chart

This chart plots first ionization energy vs. element.

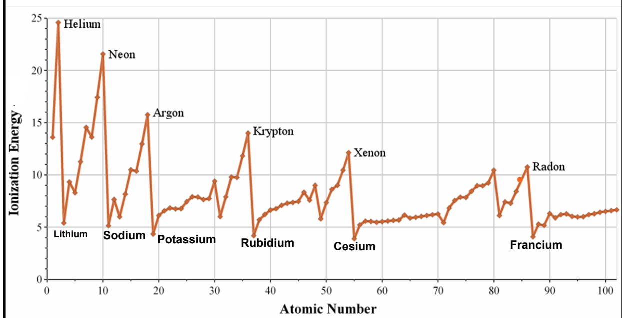
Note the large jumps present between noble gases and the next alkali metals.
Also note the blips as different sublevels are opened up, and as electrons get paired up in those sublevels.



Ionization Chart Interpretation

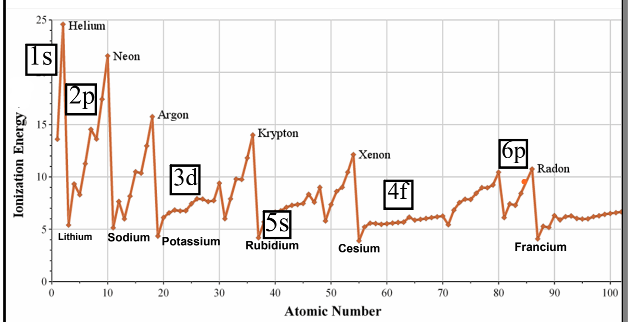
4. On your chart, circle and label the 1s, 2p, 3d, 4f, 5s, and 6p series of elements.

5. Use orbital diagrams to account for the jagged line of the second energy level elements (lithium through neon).



Ionization Chart Answers

4. Applied labels.



Ionization Chart Answers

5. The blip between beryllium and boron is coincident with the first electron being added to the 2p sublevel;

	1s	2s	2p
beryllium	↑↓	↑↓	[ ][ ][ ]
boron	↑↓	↑↓	[↑][ ][ ]

the blip between nitrogen and oxygen is where the first pairing of electrons happens in the 2p sublevel.

	1s	2s	2p
nitrogen	↑↓	↑↓	[↑][↑][↑]
oxygen	↑↓	↑↓	[↑↓][↑][↑]

Successive Ionization Energy

Additional electrons can be removed after the first one.

More energy is required to remove additional electrons - each electron removed makes the ion smaller.

There is a large jump in ionization energy as the last electron from the valence shell is removed, and inner shell electrons are exposed (and removed).

TABLE 7.2 Successive Values of Ionization Energies, I, for the Elements Sodium Through Argon (kJ/mol)

Element	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	496	4560			(Inner-shell electrons)		
Mg	738	1450	7730				
Al	578	1820	2750	11,600			
Si	786	1580	3230	4360	16,100		
P	1012	1900	2910	4960	6270	22,200	
S	1000	2250	3360	4560	7010	8500	27,100
Cl	1251	2300	3820	5160	6540	9460	11,000
Ar	1521	2670	3930	5770	7240	8780	12,000

C. Electronegativity

The relative ability an element has to attract electrons in a chemical bond. Think: blanket hog.

This will become important as we study bonding in the next unit.

Electronegativity trends are driven by atomic radius: the smaller an atom, the closer its nucleus is to the electrons involved in bonding, thus smaller atoms have greater electronegativity.
Trends: generally greater as you go from left to right across a period, and less as you go down a group.

Fluorine highest = 3.98 Paulings
Francium lowest = 0.70 Paulings

Electronegativity Chart

Here is a chart of accepted electronegativities.

Note: not all noble gases have a designated electronegativity: argon, neon, and helium have yet to be observed bonding with another element.

Electronegativity Chart

	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8	Group 9	Group 10	Group 11	Group 12	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
1	1 H 2.1																	2 He —
2	3 Li 1.0	4 Be 1.5											5 B 2.0	6 C 2.5	7 N 3.0	8 O 3.5	9 F 4.0	10 Ne —
3	11 Na 0.9	12 Mg 1.2											13 Al 1.5	14 Si 1.8	15 P 2.1	16 S 2.5	17 Cl 3.0	18 Ar —
4	19 K 0.8	20 Ca 1.0	21 Sc 1.3	22 Ti 1.5	23 V 1.6	24 Cr 1.6	25 Mn 1.5	26 Fe 1.8	27 Co 1.8	28 Ni 1.8	29 Cu 1.9	30 Zn 1.6	31 Ga 1.6	32 Ge 1.8	33 As 2.0	34 Se 2.4	35 Br 2.8	36 Kr 3.0
5	37 Rb 0.8	38 Sr 1.0	39 Y 1.2	40 Zr 1.4	41 Nb 1.6	42 Mo 1.8	43 Tc 1.9	44 Ru 2.2	45 Rh 2.2	46 Pd 2.2	47 Ag 1.9	48 Cd 1.7	49 In 1.7	50 Sn 1.8	51 Sb 1.9	52 Te 2.1	53 I 2.5	54 Xe 2.6
6	55 Cs 0.7	56 Ba 0.9	57 La 1.1	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.8	83 Bi 1.9	84 Po 2.0	85 At 2.2	86 Rn 2.4
7	87 Fr 0.7	88 Ra 0.9	89 Ac 1.1	104 Rf —	105 Db —	106 Sg —	107 Bh —	108 Hs —	109 Mt —									

Lanthanide series

58 Ce 1.1	59 Pr 1.1	60 Nd 1.1	61 Pm 1.1	62 Sm 1.2	63 Eu 1.1	64 Gd 1.2	65 Tb 1.1	66 Dy 1.2	67 Ho 1.2	68 Er 1.2	69 Tm 1.3	70 Yb 1.1	71 Lu 1.3
90 Th 1.3	91 Pa 1.5	92 U 1.4	93 Np 1.4	94 Pu 1.3	95 Am 1.3	96 Cm 1.3	97 Bk 1.3	98 Cf 1.3	99 Es 1.3	100 Fm 1.3	101 Md 1.3	102 No 1.3	103 Lr —

Quick Quiz!

6. Which element is larger, K or Ca?

7. Which element has a higher electronegativity, Si or Ba?

8. Which element has the greater ionic radius, Na^+ or Rb^+?

9. Which element has the greater ionic radius, Ca^{2+} or Se^{2-}?

10. Which is larger: Te or S?

11. Organize the following elements largest to smallest:

N Li F C
Be B O

Quick Quiz Answers

6. Which element is larger, K or Ca?

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8. Which element has the greater ionic radius, Na^+ or Rb^+?

9. Which element has the greater ionic radius, Ca^{2+} or Se^{2-}?

10. Which is larger: Te or S?

11. Largest to smallest:

Li Be B C N O F

D. Electron Affinity

Electron affinity is the energy change associated with the addition of an electron to a gaseous atom:

$$\boxed{X_{(g)} + e^{-} \rightarrow X^{-}_{(g)}}$$

Essentially it is the opposite of ionization energy.

Realize that non-metallic elements naturally take on extra electrons (becoming anions: ex. chloride). If this is a spontaneous process, energy is released (exothermic) and the sign of the process is negative.

In contrast, metallic atoms naturally lose electrons, so adding one requires energy, thus giving the process a positive sign.

Trends: electron affinity tends to become more negative (more exothermic) from left to right, and less negative going down a row. Please read pages 332-333 to explore exceptions and nuances.

Homework

Read 7.13 in your textbook.

7.12 Problems in your Booklet
Due: Next Class

Work on Unit 6.A Review Problems
Scanned: Tuesday 2/25