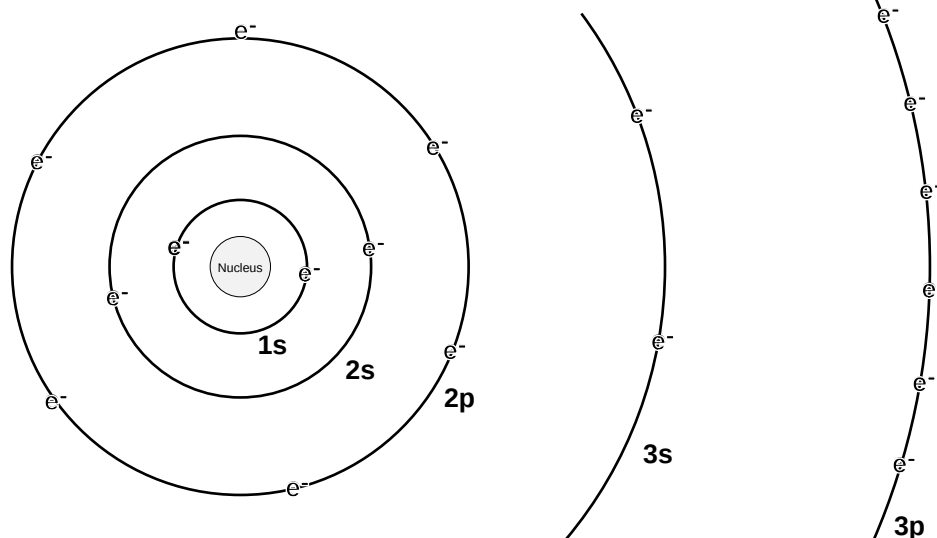


Shielding its underlying premises and how to derive periodic trends using it:

1. Understand that electrons add to atoms according to sublevels and by and large you don't proceed to the next sublevel until the current one is completely full.



2. The other idea in shielding is that opposite charges attract (electrostatic attraction) and the greater the charge of the nucleus that an outer electron sees, the greater the force of attraction between that electron and the nucleus.

Shielding says that electrons essentially only see the positive charge of the nucleus that is not blocked out by electrons occupying inner sublevels.

For example,

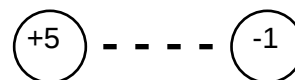
the outer electrons of Si ($[\text{Ne}]3s^23p^2$) essentially only see approximately +2 charge from the nucleus whereas the outer electrons of Cl ($[\text{Ne}]3s^23p^5$) essentially see a +5 charge from the nucleus.

force of attraction between



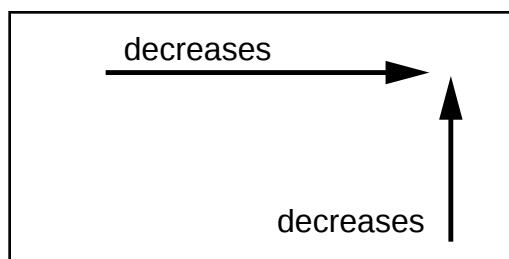
is less than

the force of attraction between



So the electrons of Cl are pulled in more tightly than that of Si. You can use this idea to predict trends in **Atomic Size** of atoms as you move across a period. The greater the charge difference between the shielded nucleus and the outer electrons, the greater attractive force and thus the smaller the atomic radius. As you go across the periods atomic radius decreases. (As you go down the groups you are simply adding energy levels and thus increasing atomic radius.)

**Atomic
Radius**



Ionic radius:

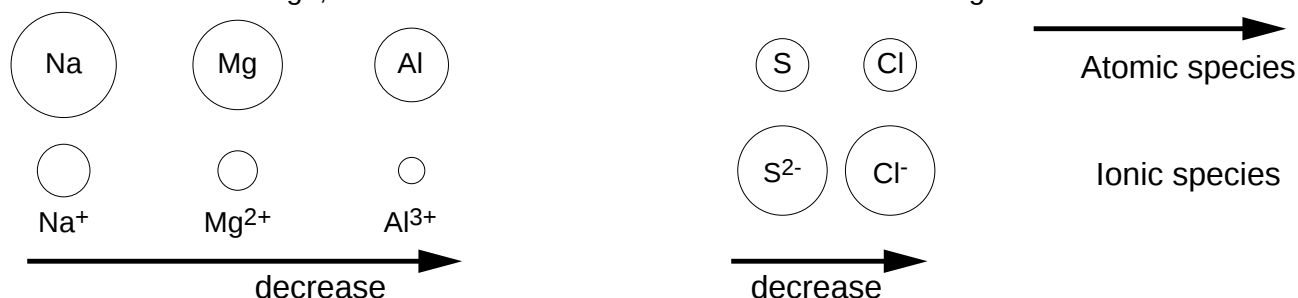
Shielding doesn't help much here but the same electrostatic force idea used to explain shielding does work. This is really tricky business because you have to look at what happens to atoms and their ions separately and then try to look at the trends in ionic radius for the periodic table as a whole.

Cations: get smaller than the atomic species because you lose outer electrons (often complete sublevels) and the net positive charge draws in the remaining electrons.

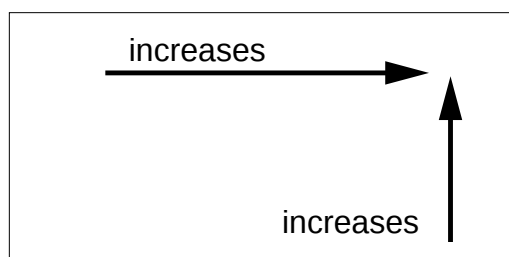
Anions: you gain electrons and they repulse each other. The ionic radius expands out to accommodate the repulsive forces.

For the Periodic Table

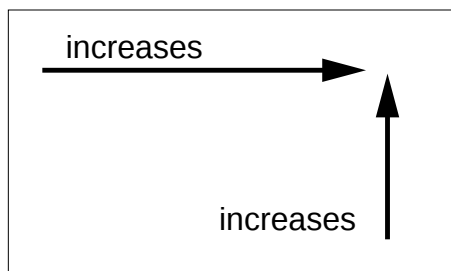
As with atoms, the size of ions increases as you go down the periodic table but cations (decrease in size) get smaller than anions (increase in size) and this contradicts the trend for the same atomic species. This behavior is a result of the nonuniformity of stable ionic charge across chemical groups. The alkali metals have stable ions of +1 charge, while the alkaline earth metals ions have +2 charges.



Ionization energy: This is the energy required to remove the outermost electron. The larger the unshielded charge on the nucleus the bigger the I.E. Watch out for jumps that occur because of orbital stability issues. Like when you lose a lone p-electron and then try to remove the next electron from a full s-sublevel. That will require a big jump in energy because the full s-sublevel is a stable configuration.



Electron Affinity/Electronegativity: These are both measures of how strongly an atom can attract an additional electron --> the greater the difference between the shielded nucleus and the outer electrons the higher the electron affinity and electronegativity. The two aren't the same thing, they just kind-of describe the same characteristics. Electron affinity is a physical characteristic whereas electronegativity was made up by Linus Pauling as a means of predicting bond polarity between atoms in molecules.



Student Packet: 1.5: Periodic Trends: AssessMe questions

- 1) Nitrogen has a higher first ionization energy than oxygen. This is principally the result of
- A) a nuclear charge effect.
 - B) greater penetration of the nitrogen p orbitals.
 - C) a crowding effect of the electrons.
 - D) the half-filled subshell of nitrogen.
 - E) I'm clueless
- 2) What is the correct order of increasing first ionization energies for the elements Be, B, and C?
- A) Be>B>C B) B>Be>C C) B>C>Be D) C>Be>B E) I'm clueless
- 3) Which of the following atoms would have the most negative electron affinity?
- A) Si B) P C) S D) Ge E) I'm clueless
- 4) Which of the following electronic configurations would represent an atom with the most negative electron affinity?
- A) ns^2np^1 B) ns^2np^2 C) ns^2 D) ns^2np^4 E) I'm clueless
- 5) In which of the following sets do the elements have the most similar atomic radii?
- A) Be, B, C, N
 - B) Ne, Ar, Kr, Xe
 - C) Mg, Ca, Sr, Ba
 - D) Cr, Mn, Fe, Co
 - E) I'm clueless
- 6) Atoms generally become smaller with increasing atomic number within a period because
- A) of additional electron repulsion.
 - B) of increased effective nuclear charge.
 - C) smaller electrons are used later in a period.
 - D) the nucleus absorbs some of the electrons.
 - E) I'm clueless

Student Packet: 1.5: Periodic Trends: ShowMe questions

1) Arrange the following groups of atoms or ions in order of increasing size:

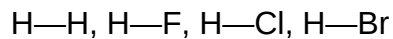
- a) Li, Na, K
- b) P, N, O
- c) Li^+ , Na^+ , F^-
- d) Na^+ , Mg^{2+}
- e) S^{2-} , Cl^-

2) The following table shows the ionization energies of sodium, magnesium, sulfur and aluminum, not necessarily in that order. Which set of ionization energies belongs to each element?

<u>Element #</u>	<u>Ionization Energies (kJ/mol)</u>						
	I_1	I_2	I_3	I_4	I_5	I_6	I_7
1	1005	2260	3375	4565	6950	8490	27000
2	735	1445	7730				
3	495	4560					
4	580	1815	2740	11600			

3) Arrange the alkali metals in terms of increasing first ionization energy.

4) Arrange the following bonds in order of increasing polarity:



Student Packet: 1.5: Periodic Trends: TestMe questions

- 1) To which element do the following ionization energies correspond? (All in kJ/mol)

$$I_1=780$$

$$I_2=1575$$

$$I_3=3220$$

$$I_4=4350$$

$$I_5=16,100$$

A) Si

B) F

C) Na

D) Mg

E) Se

- 2) Arrange the following elements and ions in order of increasingly exothermic electron affinity.

I. Cl

II. Cl⁻

III. I

IV. I⁻

A) I < II < III < IV

B) I < III < II < IV

C) II < I < III < IV

D) II < IV < I < III

E) IV < II < III < I

- 3) The first ionization energy of sodium is 495 kJ/mole. Which of the following reactions corresponds to the first ionization of sodium?

A) $\text{Na} + \text{e}^- \rightarrow \text{Na}^-$ B) $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$ C) $\text{Na}^{2+} + \text{e}^- \rightarrow \text{Na}^+$ D) $\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$ E) $\text{Na}^+ \rightarrow \text{Na}^{2+} + \text{e}^-$

- 4) Arrange the following elements and ions in order of increasing radius:

I. K⁺II. Ca²⁺III. S²⁻

A) I < II < III

B) I < III < II

C) II < I < III

D) II < III < I

E) III < II < I

- 5) Arrange the following elements in order of electronegativity:

I. F

II. P

III. O

IV. S

A) I < II < III < IV

B) I < III < IV < II

C) III < I < II < IV

D) IV < III < II < I

E) II < IV < III < I