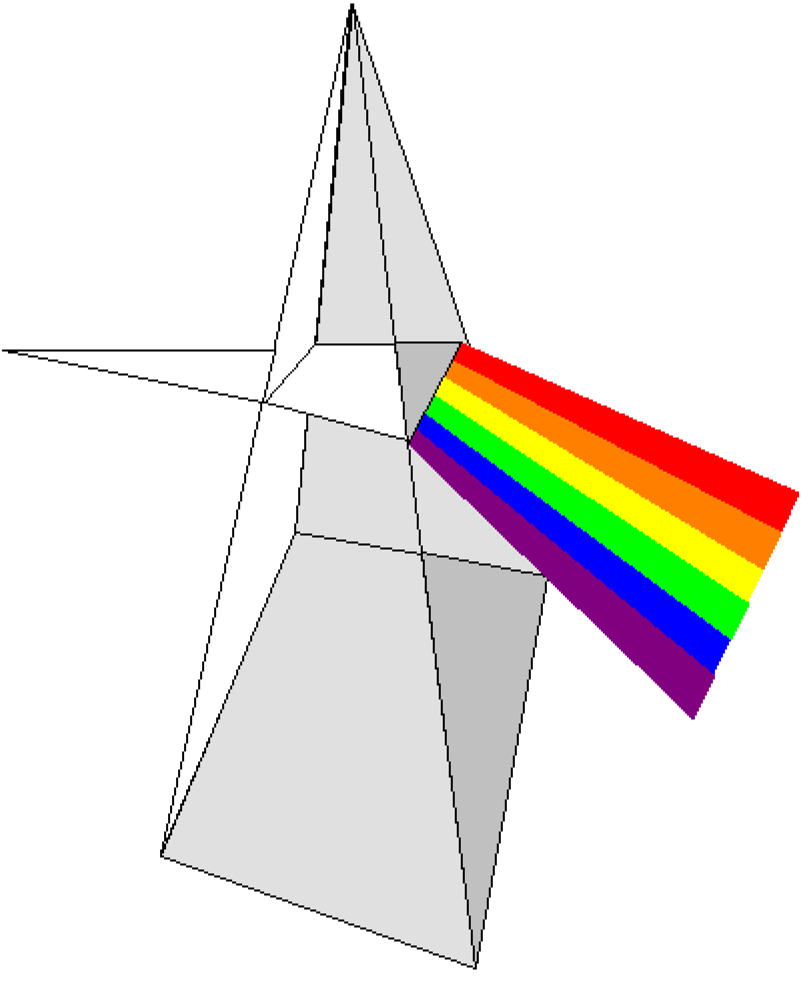


CH 110

ATOMIC STRUCTURE

&

PERIODICITY

5

Trends in Ionization Energy

- Ionization energy: the energy required to remove an electron from a gaseous atom or ion

  $X(g) \rightarrow X^+(g) + e^-$

- The highest energy electron is removed first.
- First ionization energy (I_1) is that required to remove the first electron.
- Second ionization energy (I_2) - the second electron etc. etc.

- Consider Al:

$$Al(g) \rightarrow Al^{+}(g) + e^{-} \qquad I_1 = 580\text{kJ/mol}$$

$$Al^{+}(g) \rightarrow Al^{2+}(g) + e^{-} \qquad I_2 = 1815\text{J/mol}$$

$$Al^{2+}(g) \rightarrow Al^{3+}(g) + e^{-} \qquad I_3 = 2740\text{kJ/mol}$$

$$Al^{3+}(g) \rightarrow Al^{4+}(g) + e^{-} \qquad I_4 = 11600\text{kJ/mol}$$

- Al Conf. is $[Ne]3s^2 3p^1$
- for Mg
  - I_1 = 735 kJ/mole
  - I_2 = 1445 kJ/mole
  - I_3 = 7730 kJ/mole

Trends in ionization energy

- Notice change in values from I_1 to I_4
- Why?
- The effective nuclear charge increases as you remove electrons.
- There is a high jump in IE after removing the valence electron(s)
- It takes much more energy to remove a core electron than a valence electron because there is less shielding

First, Second, Third, and Fourth Ionization Energies of Sodium, Magnesium, and Aluminum (kJ/mol)

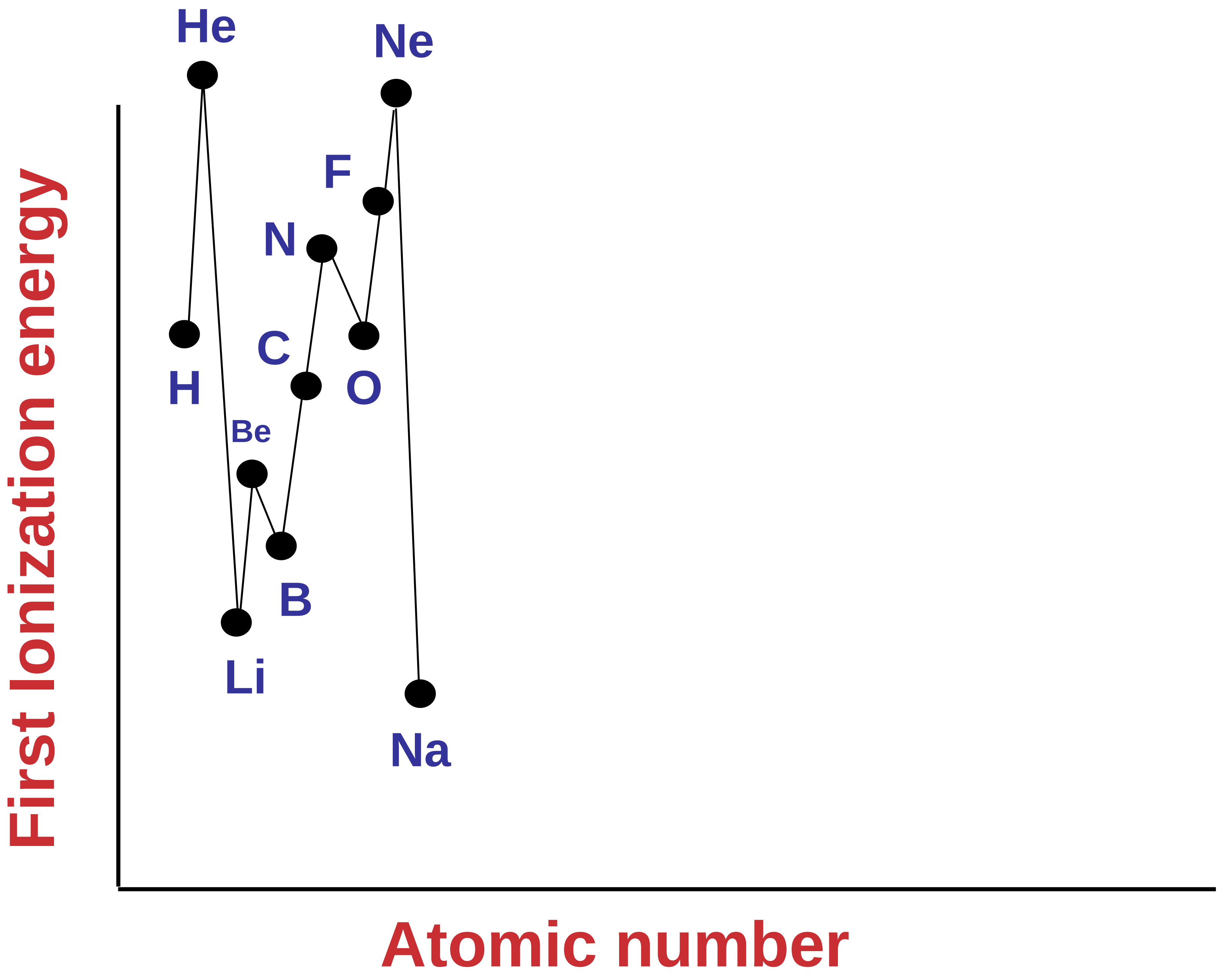
	1st IE	2nd IE	3rd IE	4th IE
Na	495.8	4562.4	6912	9543
Mg	737.7	1450.6	7732.6	10,540
Al	577.6	1816.6	2744.7	11,577

IE Trend Across a Period

- Generally from left to right, I_1 increases because
  - there is a greater nuclear charge with the same shielding.
- As you go down a group I_1 decreases because
  - electrons are further away and there is more shielding

IE Trend Across a Period

- Z_{eff} changes as you go across a period, so will I_1
- Half-filled and filled orbitals are harder to remove electrons from.
  - This brings variations within the period e.g Be to B & N to O.
- Here’s what it looks like


He
Ne
F
N
C
H
O
Be
B
Li
Na
First Ionization energy
Atomic number

IE energies for period 3 elements

Table 7.5 ▶ Successive Ionization Energies in Kilojoules per Mole for the Elements in Period 3

Element	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	495	4560					
Mg	735	1445	7730	Core electrons*			
Al	580	1815	2740	11,600			
Si	780	1575	3220	4350	16,100		
P	1060	1890	2905	4950	6270	21,200	
S	1005	2260	3375	4565	6950	8490	27,000
Cl	1255	2295	3850	5160	6560	9360	11,000
Ar	1527	2665	3945	5770	7230	8780	12,000

*Note the large jump in ionization energy in going from removal of valence electrons to removal of core electrons.

General decrease

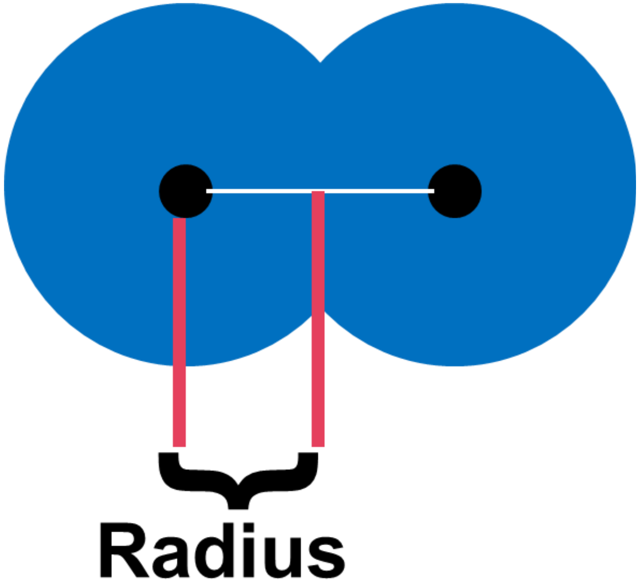
General increase

Try This

- Which atom in the following pairs has the greater first IE
  - Li or Be
  - Ca or Ba
  - Na or K
  - P or Ar
  - Cl or Si
  - Li or K

Trends in Atomic Size

- First problem is where do you start measuring.
- The electron cloud doesn't have a definite edge.
- We get around this by measuring more than 1 atom at a time.
- Atomic Radius = half the distance between two nuclei of a diatomic molecule

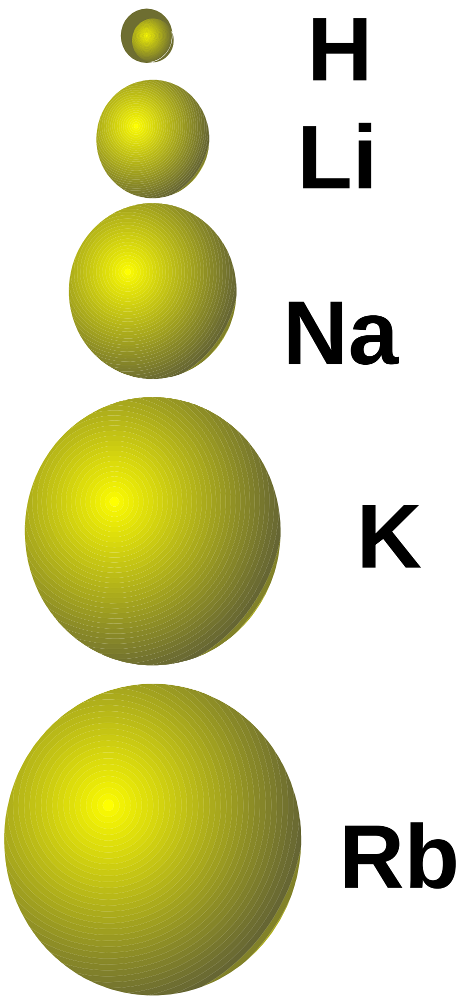


Trends in Atomic Size

- Atomic size is influenced by two factors:
  - Shielding
    - More shielding pushes electron further away
  - Charge on nucleus
    - More charge pulls electrons in closer

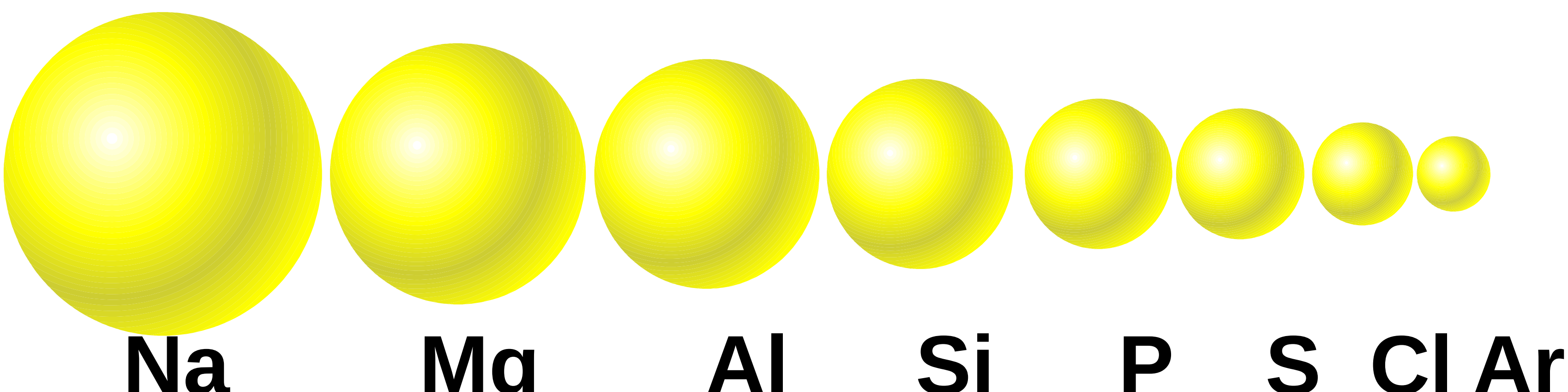
Group trends

- As we go down a group
  - Each atom has another energy level
  - So the atoms get bigger



Periodic Trends

- As you go across a period the radius gets smaller.
- Same energy level.
- But more nuclear charge.
- Outermost electrons are pulled closer




Atomic radius decreases
Atomic radius increases
1A
2A
3A
4A
5A
6A
7A
8A
H
37
He
32
Li
152
Be
113
B
88
C
77
N
70
O
66
F
64
Ne
69
Na
186
Mg
160
Al
143
Si
117
P
110
S
104
Cl
99
Ar
97
K
227
Ca
197
Ga
122
Ge
122
As
121
Se
117
Br
114
Kr
110
Rb
247
Sr
215
In
163
Sn
140
Sb
141
Te
143
I
133
Xe
130
Cs
265
Ba
217
Tl
170
Pb
175
Bi
155
Po
167
At
140
Rn
145

Try this

- Rank the following elements by increasing atomic radius
  - C, Al, O, K

Try this

- Rank the following elements by increasing atomic radius
  - O, C, Al, K
  - Xe, F, Rb, Sn, Sr

Try this

- Rank the following elements by increasing atomic radius
  - O, C, Al, K
  - F, Xe, Sn, Sr, Rb

Electron Affinity

- The energy change associated with adding an electron to a gaseous atom.

  $$X(g) + e^- \rightarrow X^-(g)$$

- High electron affinity gives you more negative energy
  - Exothermic
- EA increases (more -ve) from left to right
  - greater nuclear charge.
- EA decrease as we go down a group
  - More shielding

Ionic Size

- Cations are formed by losing electrons.
- Cations are smaller than the atom they come from.
- Metals form cations.
- Cations of representative elements have noble gas configuration.

Ionic size

- Anions are formed by gaining electrons.
- Anions are bigger than the atom they come from.
- Nonmetals form anions.
- Anions of representative elements have noble gas configuration.

Configuration of Ions

- Ions always have noble gas configuration
- Na is $1s^2 2s^2 2p^6 3s^1$
- Forms a +1 ion Na^+: $1s^2 2s^2 2p^6$
- Same configuration as Neon
- Metals form ions with the configuration of the noble gas before them - they lose electrons

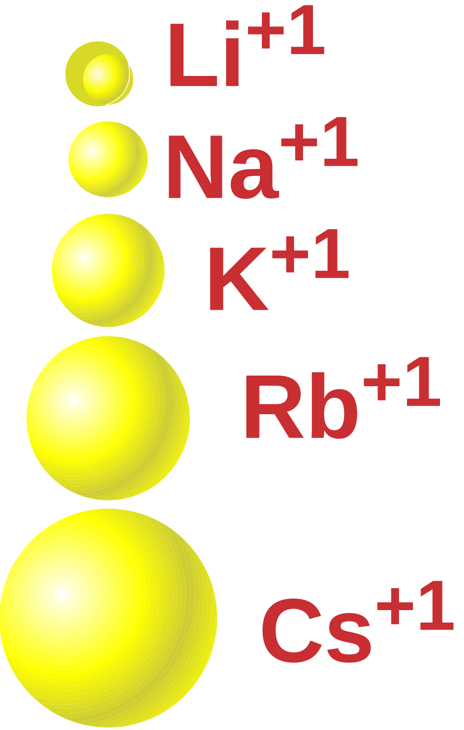
Configuration of Ions

- Non-metals form ions by gaining electrons to achieve noble gas configuration.

- They end up with the configuration of the noble gas after them.

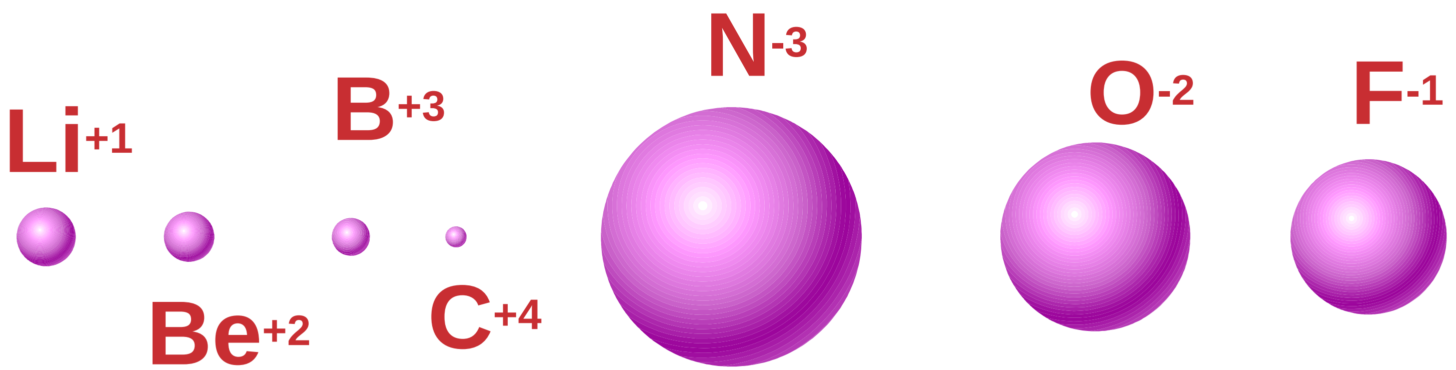
Group trends

- Adding energy level
- Ions get bigger as you go down



Periodic Trends

- Across the period, nuclear charge increases so they get smaller.
- Energy level changes between anions and cations.

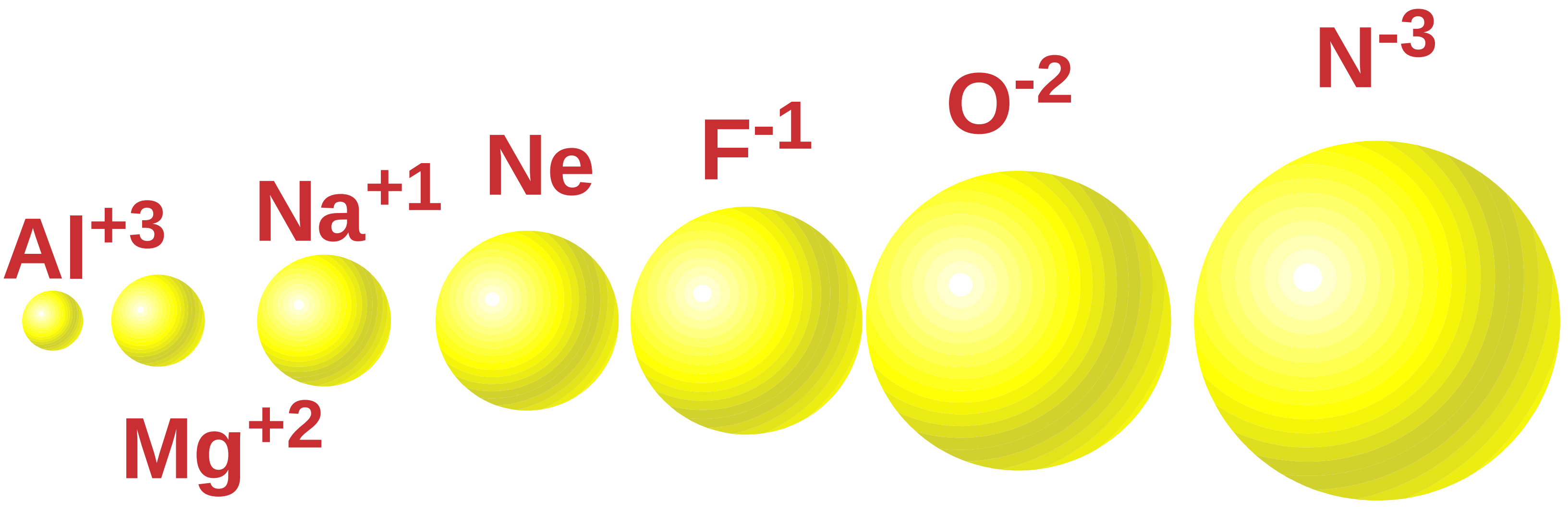


Size of Isoelectronic ions

- Iso - same
- Iso electronic ions have the same # of electrons
- Al^{+3} Mg^{+2} Na^{+1} Ne F^{-1} O^{-2} and N^{-3}
- all have 10 electrons
- all have the same configuration $1s^2 2s^2 2p^6$

Size of Isoelectronic ions

- Positive ions have more protons so they are smaller



Try This

- Which ion is larger in each of the pairs below:
  - Ca^{2+} and B^{3+}
  - K^{+} and P^{3-}
  - Li^{+} and Rb^{+}
  - Ca^{2+}
  - P^{3-}
  - Rb^{+}

Electronegativity

Electronegativity

- The tendency for an atom to attract electrons to itself when it is chemically combined with another element.
- How “greedy!”
- Big electronegativity means it pulls the electron toward itself.
- Atoms with large negative electron affinity have larger electronegativity.

Group Trend

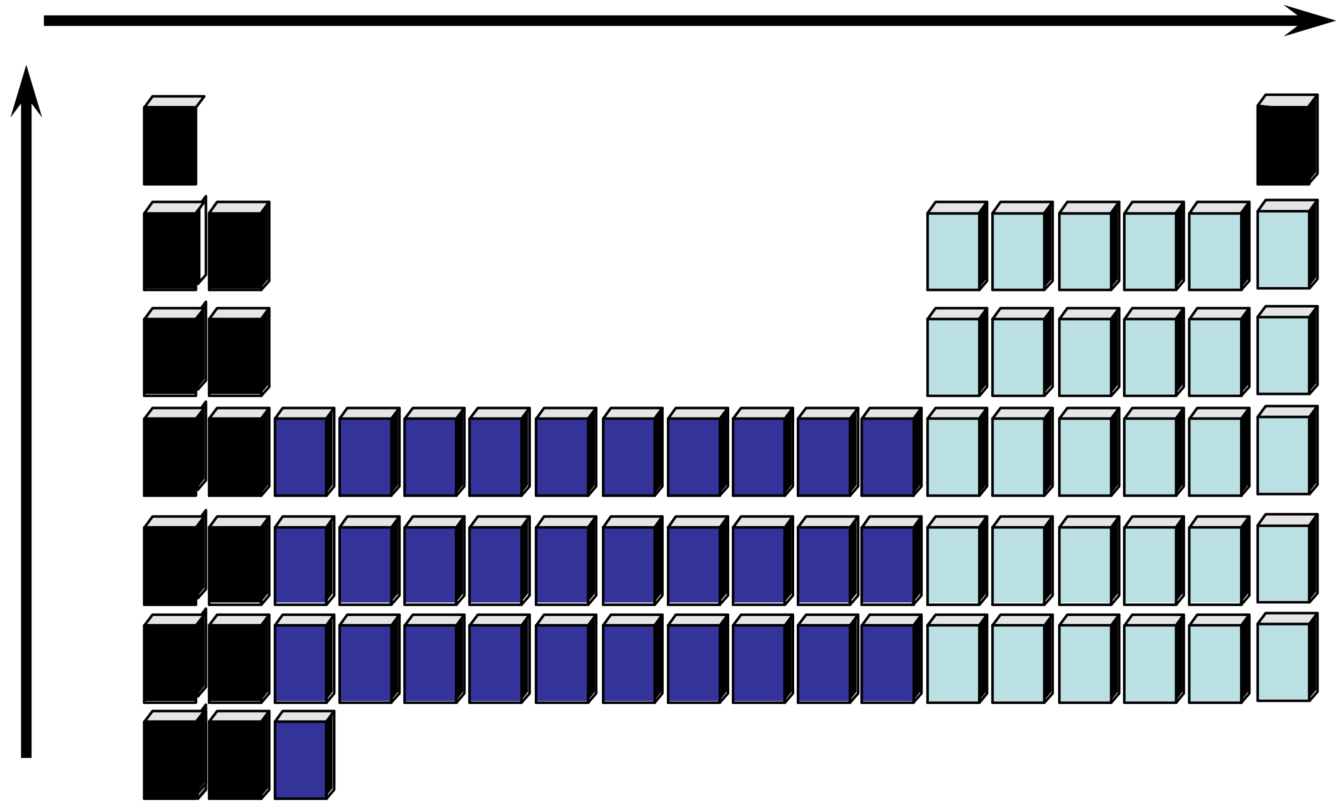
- The further down a group more shielding
- Less attracted (Z_{eff})
- Low electronegativity.

Periodic Trend

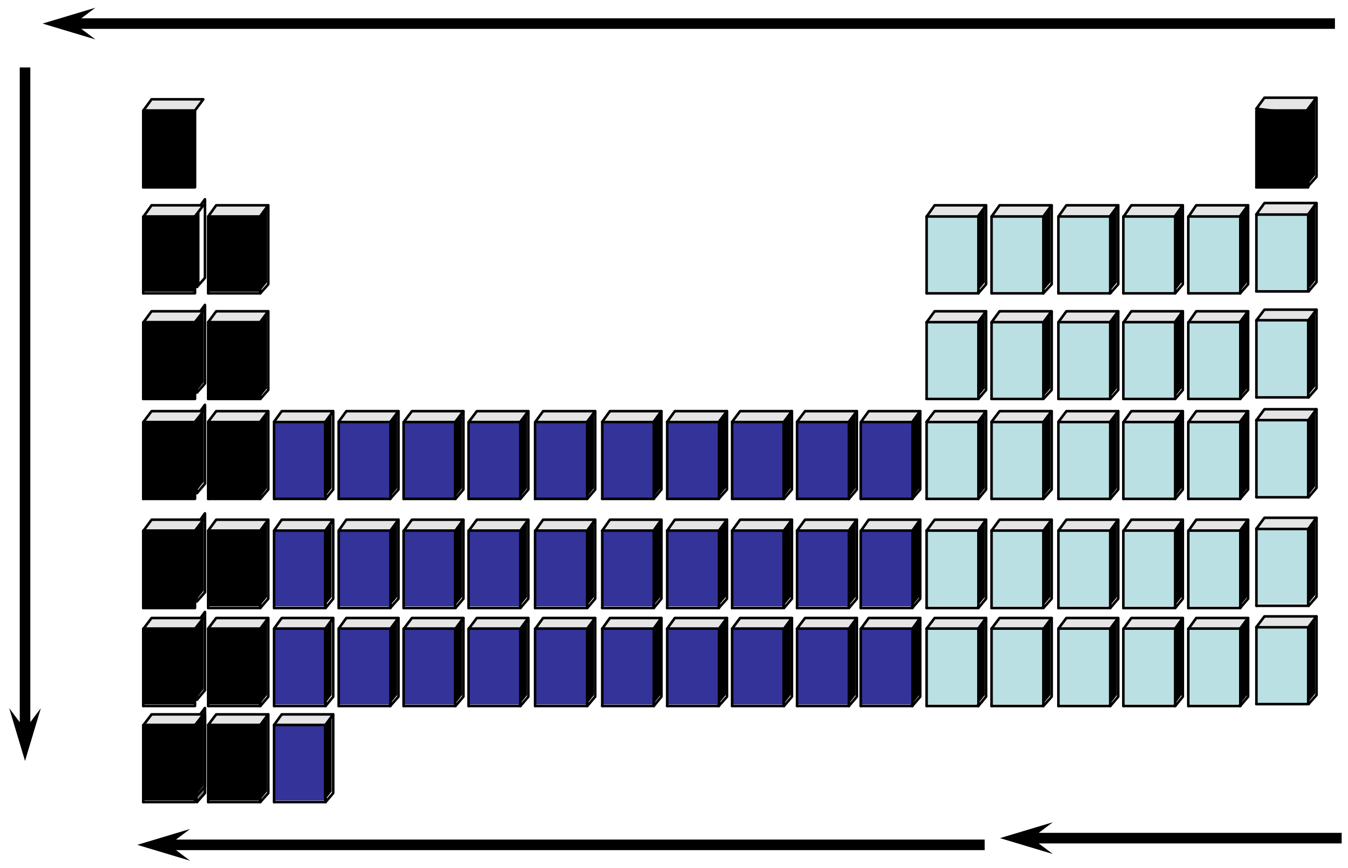
- Metals are at the left end.
  - Low ionization energy- low effective nuclear charge.
  - Low electronegativity.

- At the right end are the nonmetals
  - More negative electron affinity.
  - High electronegativity.
  - Except noble gases.

Ionization energy, electronegativity Electron affinity INCREASE



Atomic size increases,



Ionic size increases

- Which atom in each pair has greater electronegativity
  - Ca or Ga
  - Br or As
  - Li or O
  - Ba or Sr
  - Cl or S
  - O or S
- Rank the following elements by increasing electronegativity
  - S, O, Ne, Al

- Which atom in each pair has the greater electronegativity
  - Ca or Ga
  - Br or As
  - Li or O
  - Ba or Sr
  - Cl or S
  - O or S
- Rank the following elements by increasing electronegativity
  - Ne, Al, S, O

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  - Ca or Ga
  - Br or As
  - Li or O
  - Ba or Sr
  - Cl or S
  - O or S
- Rank the following elements by increasing electronegativity
  - Ne, Al, S, O
  - Fr, Rn, Cs, At

- Which atom in each pair has the greater electronegativity
  - Ca or Ga
  - Br or As
  - Li or O
  - Ba or Sr
  - Cl or S
  - O or S
- Rank the following elements by increasing electronegativity
  - Ne, Al, S, O
  - Rn, Fr, Cs, At