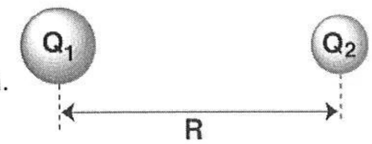
Periodic Trends

Periodic Trends are specific patterns that are present in the periodic table that illustrate different aspects of a certain element.

Almost all of the properties that are asked about in exam questions rely on Coulombic (electrostatic) attraction between outer electrons and the nucleus. Hence, you should refer to Coulomb's law in your answers!

$$E \propto \frac{Q_1 Q_2}{r}$$



E = ionization energy, the energy needed to remove the outermost electron.
Q_1 = charge of an electron, -1.
Q_2 = effective nuclear charge (Zeff) of protons in nucleus.
r = distance between charged particles, which can be approximated by the period (n, energy level).

In short, the **energy of attraction or repulsion** between charged particles is:

- Directly proportional to the **magnitude** (size) of the charges
- Inversely proportional to the **distance** between the charges

All Periodic Trends can be understood in terms of Coulomb's Law(!)

1. Electrons are attracted to the protons in the nucleus of an atom.
   a. The closer an electron is to the nucleus, the **MORE** strongly it is attracted.
      i. Distance between electrons and the nucleus can be approximated by n (main energy level)
   b. The more protons in a nucleus, the **MORE** strongly an electron is attracted.
2. Electrons are repelled by other electrons in an atom.
   a. If other electrons are between a valence electron and the nucleus, the valence electron will be **LESS** strongly attracted to the nucleus (this is known as e^- shielding).

Effective Nuclear Charge (Zeff): net positive charge experienced by electrons

→ Attraction of a valence electron to a proton is partially shielded by the inner shell (core) electrons, and so the farther away an electron is from the nucleus, the less + charge it feels from the nucleus!

→ Thus, following Coulomb's Law:

- The nucleus of atoms with a **higher Z_{eff}** (at the same energy level, n) will be more attractive to their valence electrons
- The nucleus of atoms with a **lower Z_{eff}** (at the same energy level, n) will be less attractive to their valence electrons

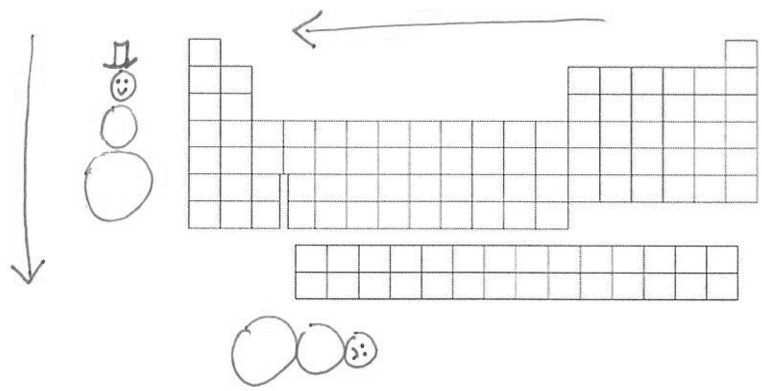
→ At the same energy level, Z_{eff} is directly proportional to the number of p^+!

NOTE: Energy level trumps Z_{eff}!
If you're comparing elements that have different n AND different Z_{eff}, n matters more!!

Trendy Trends: Periodic Trends to Know

Atomic radius (size of atom): distance between the nucleus and valence electrons.

Hint: Snowman!



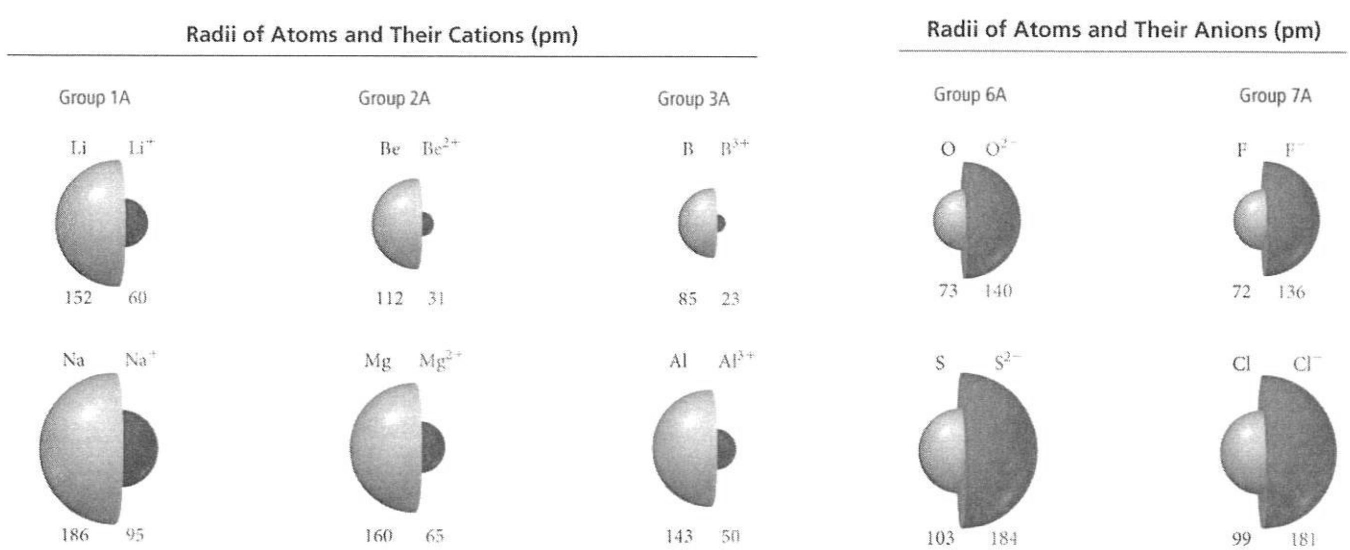
Ionic radius: distance from the nucleus to valence electrons in a charged ion.

Most comparisons between an atom and their ion or ions of the same atom can be explained by **e^-/e^- repulsion**:

- Positively charged cations are **SMALLER** than the neutral atom because fewer electrons in the outermost shell results in less e^-/e^- repulsion, thus valence electrons are closer to the nucleus.
- Negatively charged anions are **LARGER** than the neutral atom because more electrons in the outermost shell results in more e^-/e^- repulsion, thus valence electrons are farther from the nucleus.

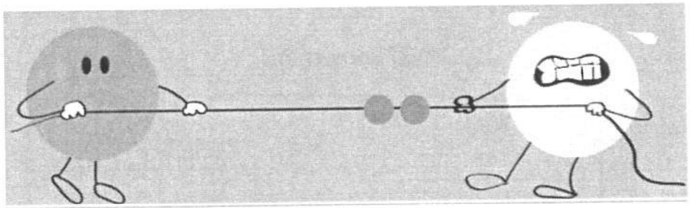
Only one type of comparison between an atom and their ion(s) should **NOT** utilize the e^-/e^- repulsion argument:

- Metals cations which have lost sufficient electrons such that their valence electrons are now in a different energy level (n). Examples: Sr vs Sr^{2+}, Na vs Na^+, Al^{2+} vs Al^{3+}, etc.
- Why? Because **n matters more than Z_{eff}!!!** If a species has their outermost electrons on a lower energy level (n), their valence electrons are closer to and thus more attracted to the nucleus.



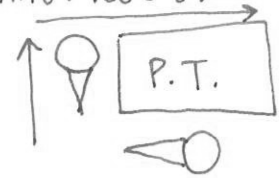
Electronegativity: attraction of an atom for pair of valence level electrons in a covalent bond with another atom.

Think of the atoms as playing "tug of war" with their Valence shell electrons!



Ionization Energy + Electron Affinity

hint: ice cream!



Electronegativity of an atom is determined by combining the following two trends:

Ionization Energy vs Electron Affinity

Ionization Energy	Electron Affinity
Energy required to remove an electron	Energy change when electron is added to an atom
$X(g) + energy \rightarrow X^{+}(g) + e^{-}$	$X(g) + e^{-} \rightarrow X^{-}(g) + energy$
endothermic (+kJ/mol)	first electron added is always exothermic (–kJ/mol)

Higher attraction between nucleus and electron = ↑ ionization energy and ↑ electron affinity!

Metallic vs Non-Metallic Character

Metallic Character	Non-Metallic Character
Metals react by losing electrons	Non-metals react by gaining electrons
How easy it is to remove an electron from an atom or ion	How hard is it to remove an electron from an atom or ion
↓ Ionization Energy = ↑ Metallic Character	↑ Ionization Energy = ↑ Non-Metallic Character

Let's Practice!

O Oxygen	F Fluorine
S Sulfur	Cl Chlorine

Consider the elements shown above. Which element has the:

a) Highest ionization energy? F

b) Highest electron affinity? F

c) Highest Metallic Character? S

d) Highest Non-Metallic Character? F

Successive Ionization Energies

Removing one electron, and then another electron, and then another electron...

- 1st Ionization Energy: (IE_1) energy required to remove the first (highest energy) electron
- 2nd Ionization Energy: (IE_2) energy required to remove the second electron (second highest energy)
- Each additional electron requires **MORE** energy to remove than the previous one:

$$IE_1 < IE_2 < IE_3 \text{ (etc.)}$$

- Valence (outer) electrons require much **LESS** energy to remove than core (inner) electrons

$$IE_{valence} << IE_{core}$$

TABLE 8.1 Successive Values of Ionization Energies for the Elements Sodium through Argon (kJ/mol)

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7
Na	496	4560					
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

Core electrons

You can identify an element from the pattern of successive ionization energies!

FR Practice

huge jump!

IE_1	IE_2	IE_3	IE_4	IE_5
801 kJ/mol	2,426 kJ/mol	3,660 kJ/mol	24,682 kJ/mol	32,508 kJ/mol

1. Which element from Period 2 does the table of successive ionizations energies above represent? Explain.

This is boron, b/c its 4th e^- requires significantly more energy to remove than the first 3 e^-, so it must have 3 valence e^- and the 4th + higher e^- are closer to the nucleus. (core e^-)

F: IE_2 IE_3

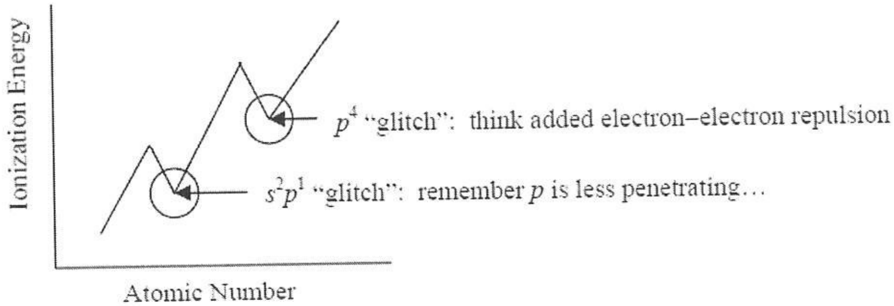
2. Would it require more energy to remove an electron from an F^+ ion or an F^{2+} ion? Justify your answer using Coulomb's Law.

It would require more energy to remove an e^- from F^{2+}, b/c removing the first e^- from F^+ to make F^{2+} decreases e^-/e^- repulsion, so the e^- cloud shrinks, so the next e^- removed is closer to (and more attracted to) the nucleus, + takes more energy to remove.

Ionization Energy Exceptions:

You MUST know these!

*these exceptions only matter when both elements are found on the same period (energy level).

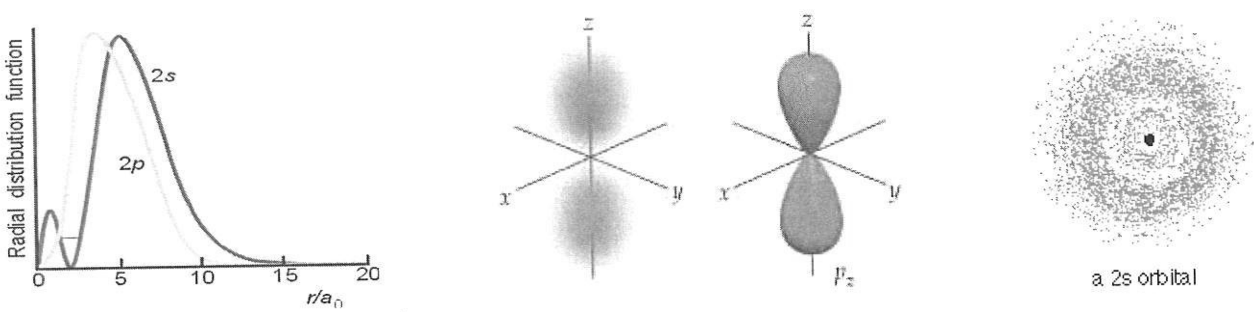


1. **Group 5 vs Group 6: "p^4 glitch"** (Example: oxygen vs nitrogen): It's easier to remove an electron from oxygen ($2p^4$) because of the repulsion of the paired electrons (whereas the $2p^3$ electrons in nitrogen (Group 5) are all unpaired).

N: 1 1 1 (2p) O: 1L 1 1 (2p)

Do not state that having a ½-filled sublevel is more stable : nope! Instead, the p^4-filled sublevel is less stable than the general trend.

2. **Group 2 vs Group 3: "s^2p^1 glitch"** (Example: beryllium vs boron): It's easier to remove an electron from boron ($2s^22p^1$) because it's being removed from the p orbital, and p orbitals do not penetrate the nuclear region as greatly as the s orbitals, so electrons in a p orbital are not as tightly held.

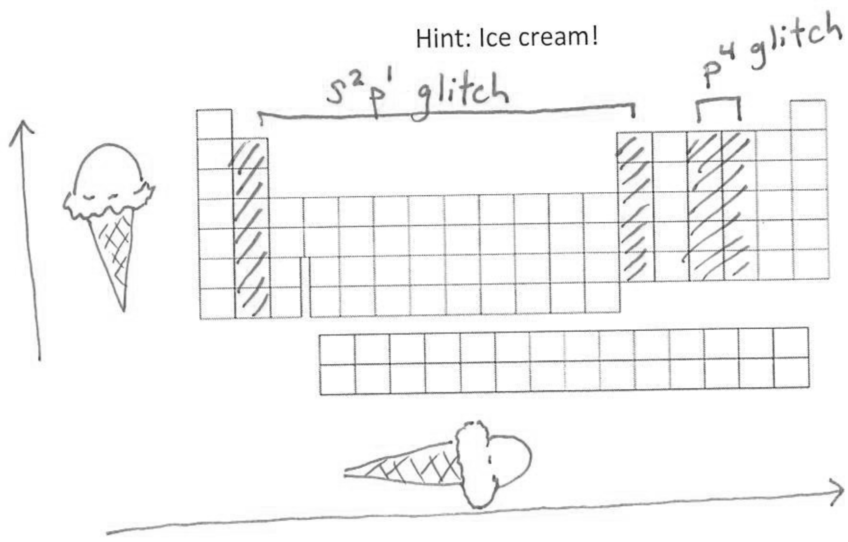


s orbital: higher nuclear penetration more electron density in greater proximity to nucleus

BE CAREFUL! This is NOT about distance between electron and nucleus - on average, electrons in a 2s orbital and 2p orbital are the same distance from the nucleus.

Ionization Energy Trend

Hint: Ice cream!



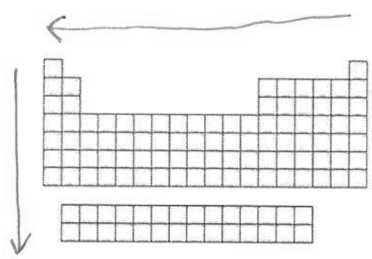
One last trend:

- **Reactivity** depends on whether the element reacts by losing electrons (metals) or gaining electrons (nonmetals).
  - Metals are MORE reactive as you move down a column: because metals lose electrons as they react, **LESS** attraction between valence electrons and the nucleus will result in a more reactive metal.
  - Non-metals are LESS reactive as you move down a column: because non-metals gain electrons as they react, **LESS** attraction between valence electrons and the nucleus will result in a less reactive non-metal.

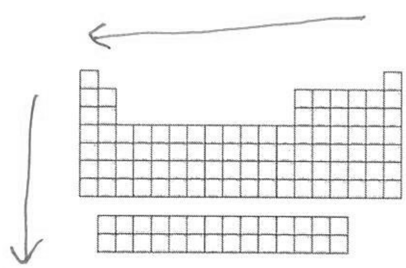
In Summary

Draw arrows in the direction that each trend increases!

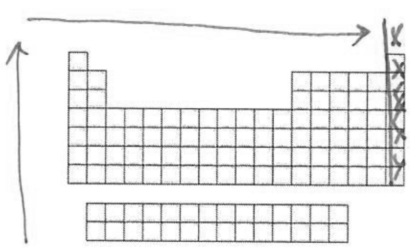
Atomic Radius Trend



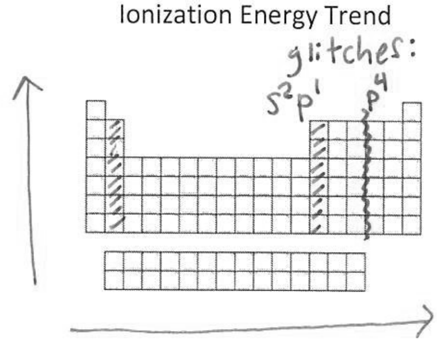
Metallic Character Trend



Electronegativity Trend



Ionization Energy Trend



Reactivity Trend