Periodic trends Summary

Atomic radius- half the distance between 2 nuclei between 2 atoms sharing electrons.

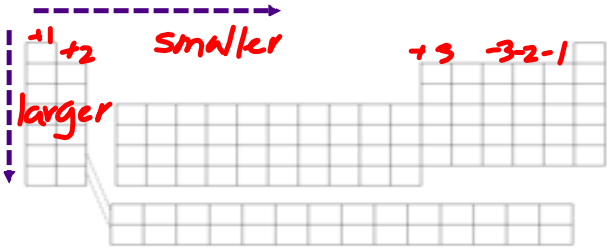
Down a group- atomic radius increases, valence electrons are farther away from attraction to nucleus.

Across a period- atomic radius decreases. More protons as you move across period, electrons have attraction to nucleus.

Ionic radius distance from center of nucleus to valence electrons

Cation vs atom- atom radius larger as cation would have lost electron(s)

Anion vs. atom ion radius larger than atom due to gain of electrons.



Ionization Energy 173 1→10

Electroneg. 1→8

C. IONIZATION ENERGY

1. Define ionization energy.

energy required to remove valence e⁻'s.

2. Is it easier to form a positive ion with an element that has a high ionization energy or an element that has a low ionization energy?

- low ionization energy = easier to remove e⁻.

3. Na^+ and Mg^{2+} ions each have ten electrons surrounding their nuclei. Which ion would you expect to have the larger radius? Explain your answer.

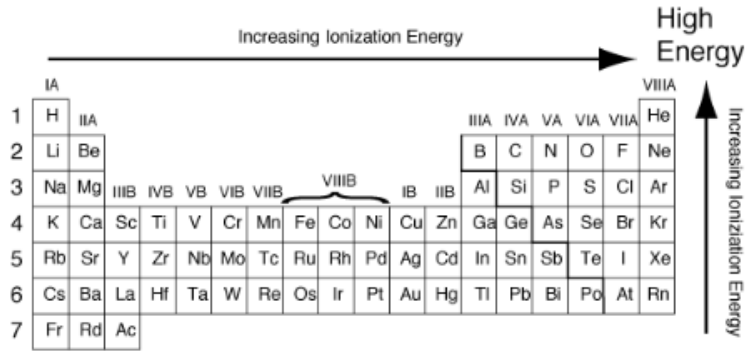
Na^+ as ions get smaller across period.

4. a. Explain why it is harder to remove an inner shell electron than a valence electron from an atom.

- valence electrons are farther from nucleus so less pull/attraction to nucleus.

b. Explain why sodium forms a +1 ion (Na^+) but magnesium forms a +2 ion (Mg^{2+}).

Na has one valence e^- to lose to be stable, but Mg has to lose 2.



5. What trend in ionization energy do you see as you go **down a group/family** on the periodic table?

- ionization energy decreases.

6. What causes this trend?

- easier to remove e^- as it is farther away from nucleus.

7. What trend in ionization energy do you see as you go **across a period/row** on the periodic table?

- ionization energy increases.

8. What causes this trend?

- harder to remove electron as attracted to nucleus.

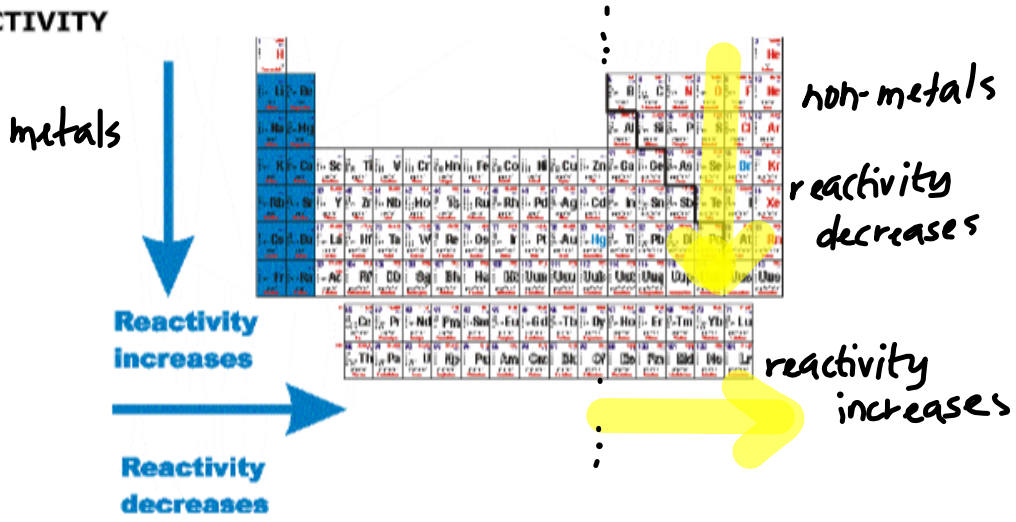
9. Circle the atom in each pair that has the **greater ionization energy**.

a) **Li** **Be** b) **Na** **K** c) **Cl** **Si** d) **Ca** **Ba** e) **P** **Ar** f) **Li** **K**

10. Why does fluorine have a higher ionization energy than iodine?

- electrons for florine are closer to nucleus; more attraction harder to remove.

C. REACTIVITY



1. What trend in reactivity do you see as you go **down a group/family** on the periodic table? metals: increasing reactivity; non-metals: decreasing reactivity

2. What causes this trend?

metals

easier to give up e^- if farther away from nucleus.

non-metals

electronegativity higher at top of group as elements are electron greedy ∴ will react more readily.

3. What trend in reactivity do you see as you go **across a period/row** on the periodic table?

metals decrease

non metals increase

4. What causes this trend?

metals - still need to lose electrons, but more of them so more energy required.

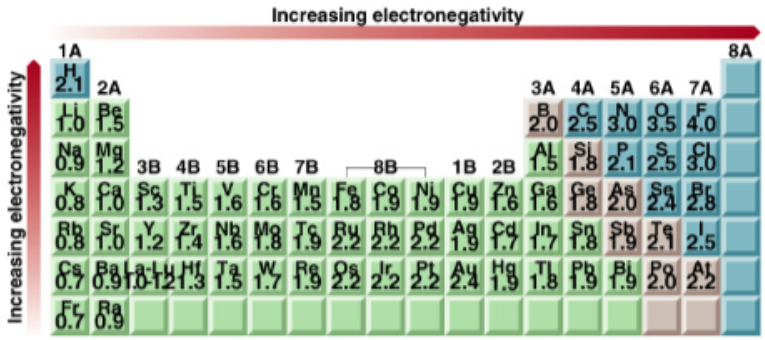
5. Circle the atom in each pair that has the **greater reactivity?**

a) **Li** **Be** b) **Na** **K** c) **Cl** **Si** d) **Ca** **Ba** e) **P** **Ar** f) **Li** **K**

6. How are ionization energy and reactivity related?

non-metals - closer to having valence level filled with electrons react to be stable.

D. ELECTRONEGATIVITY



1. Define electronegativity. attraction for an atom to electrons involved in bonding.
ELECTRON GREEDINESS :)

3. What trend in electronegativity do you see as you go down a group/family on the periodic table?

decreases

4. What causes this trend?

atoms lose valence electrons easily as they are farther away from nucleus.

5. What trend in electronegativity do you see as you go across a period/row on the periodic table?

-increases.

6. What causes this trend?

-moving from metals to non-metal section; non-metals are more electron greedy.

7. Circle the atom in each pair that has the **greater electronegativity**.

a) **Ca** **Ga** b) **Li** **O** c) **Cl** **S** d) **Br** **As** e) **Ba** **Sr** f) **O** **S**

8. Rank the following elements by increasing electronegativity:
sulfur, oxygen, neon, aluminum.

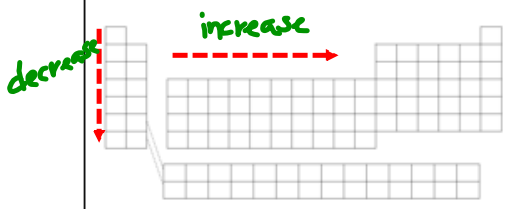
Ne, Al, S, O
0 1.5 2.5 3.5

Ionization Energy

energy required to remove a valence electron.

Down a group- ionization energy decreases, easie to remove electron as it is farther away from nucleus.

Across a period ionization energy increases, harder to remove electron due to attraction to nucleus.



Electronegativity-	an atoms ability to attract bonding electrons. Electron greediness... metals have low electroneg and non-metals high electroneg.
Down a group-	electronegativity decreases as electrons are farther away from nucleus.
Across a period-	electronegativity increases as electons have greater attration to nucleus.