PERIODIC TRENDS

notes 3

Mr. Yeung

PERIODIC TRENDS

- Based on how the elements are situated on the Periodic table, there are 3 trends that can be observed.

- 1. Electronegativity
- 2. Atomic radius
- 3. Ionization Energy

Memorize 1 fact and
Use logic to deduce the other

ELECTRONEGATIVITY

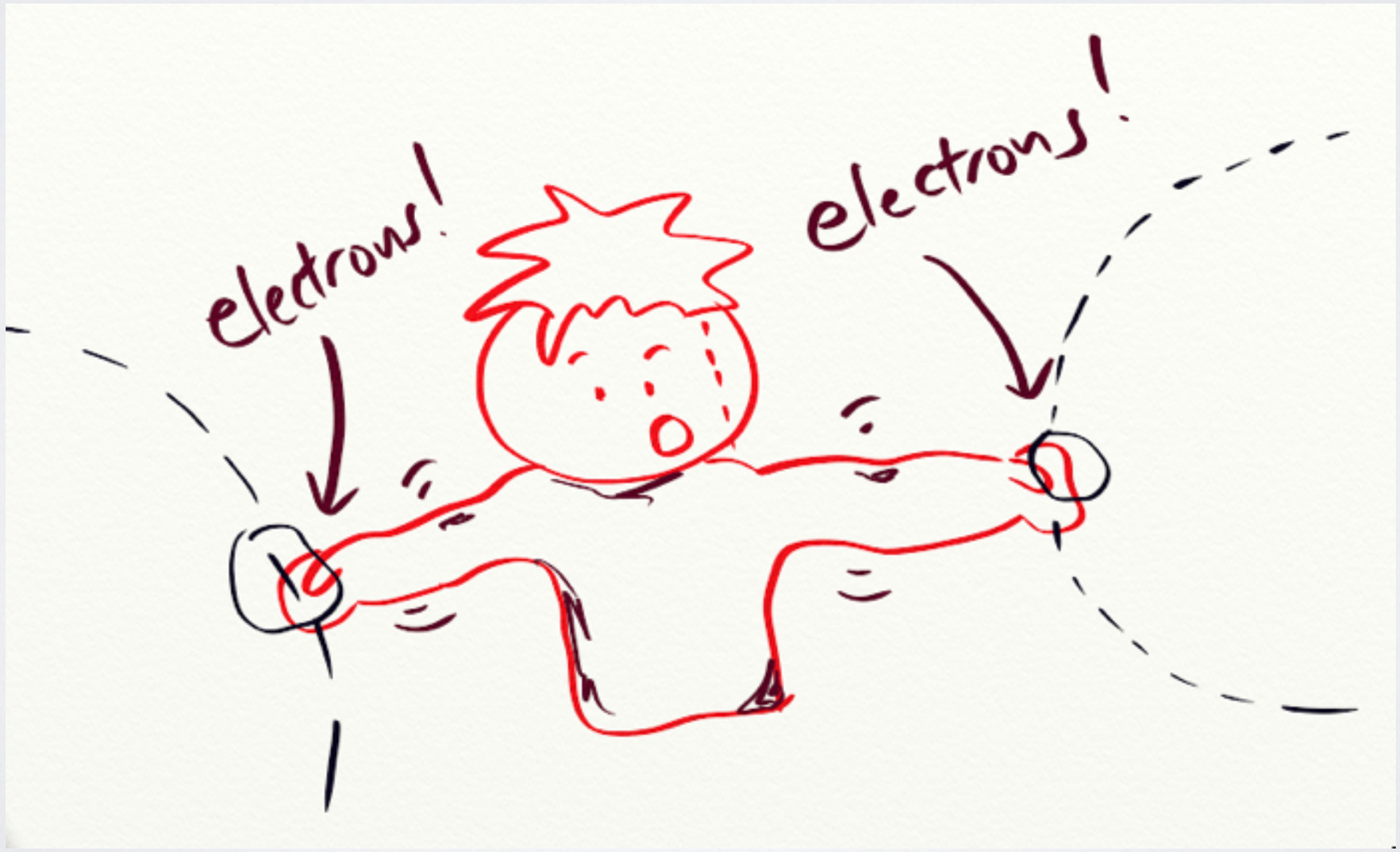
Memorize this one!

- What is it?
  - Designated as **δ** (delta)
  - It is the ability for an atom to **attract** electrons to a bond.
  - Fluorine is the **MOST** electronegative element

Just think of this…

Electronegativity is the ability of an atom to **PULL** electrons to itself

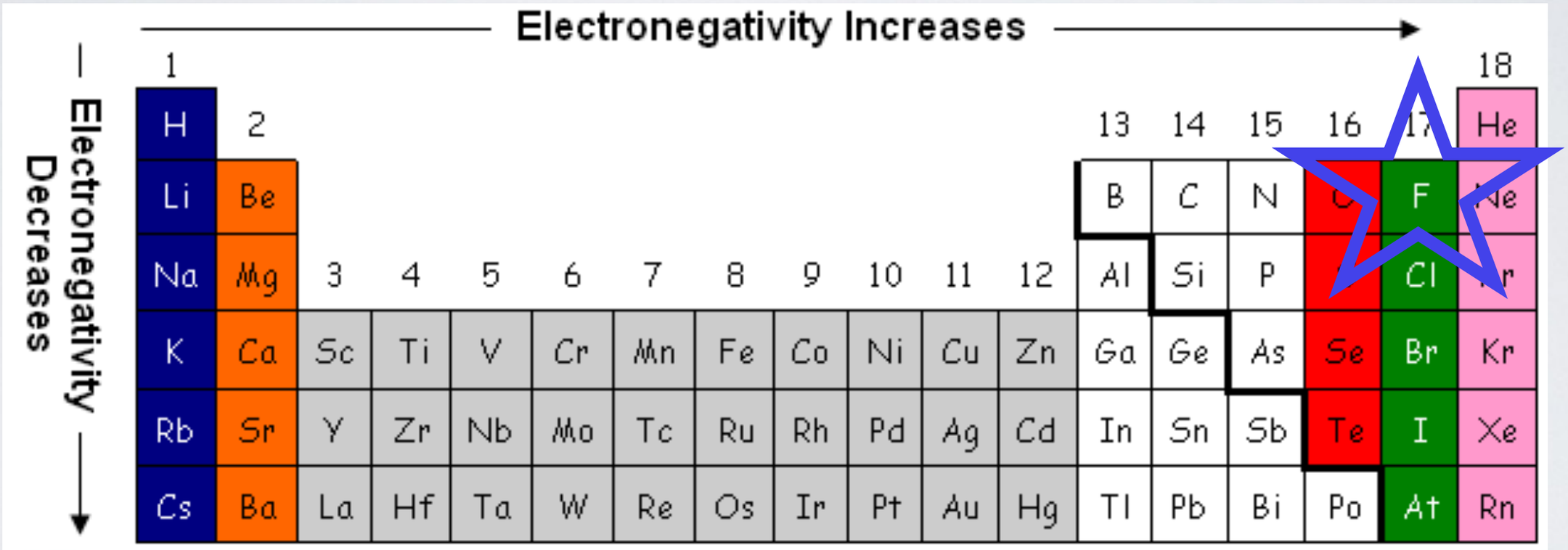
Higher electronegativity = stronger pull to itself



Since Fluorine has the highest electronegativity value, it has the strongest pull.

Increases across

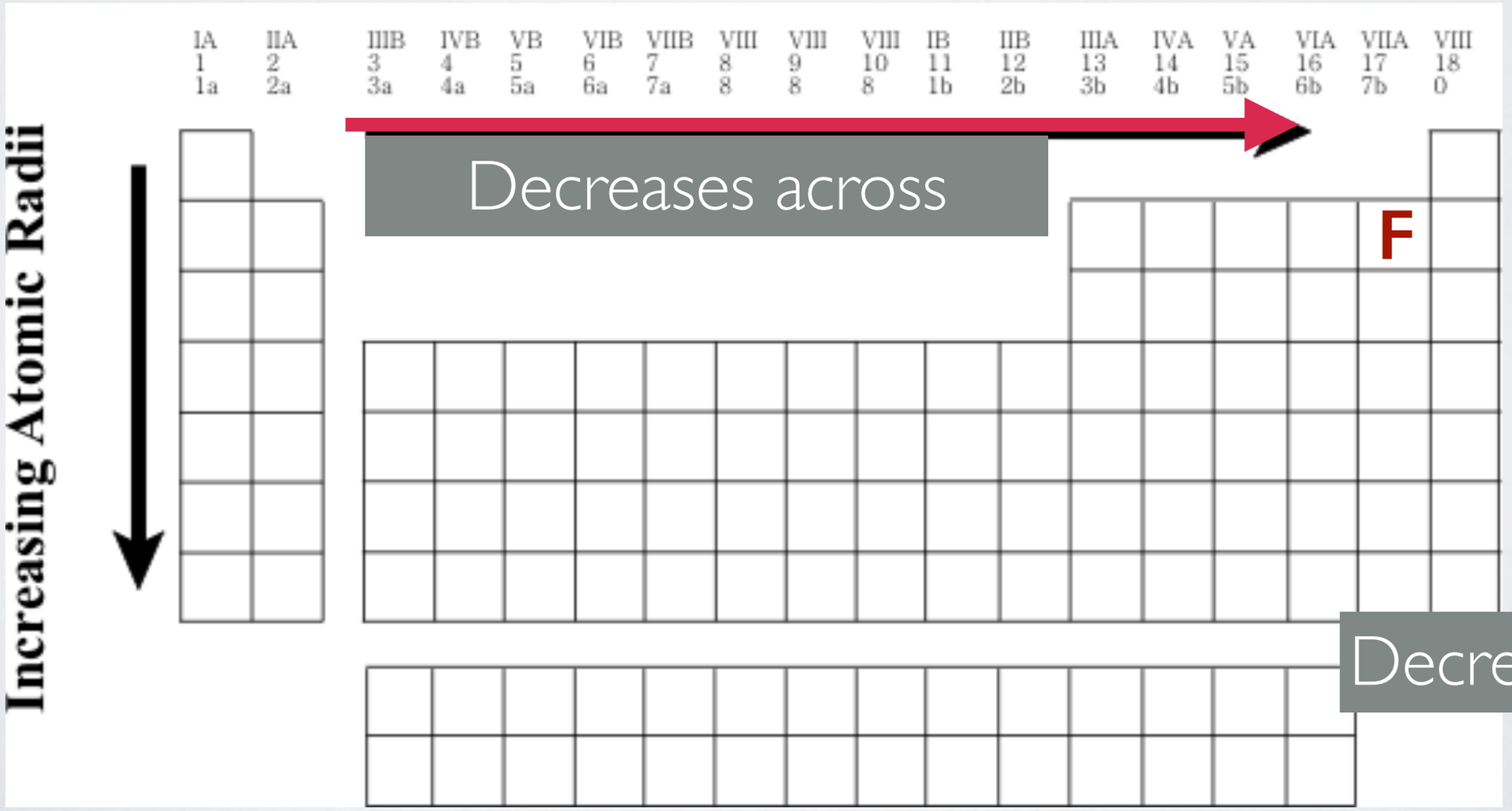
Highest

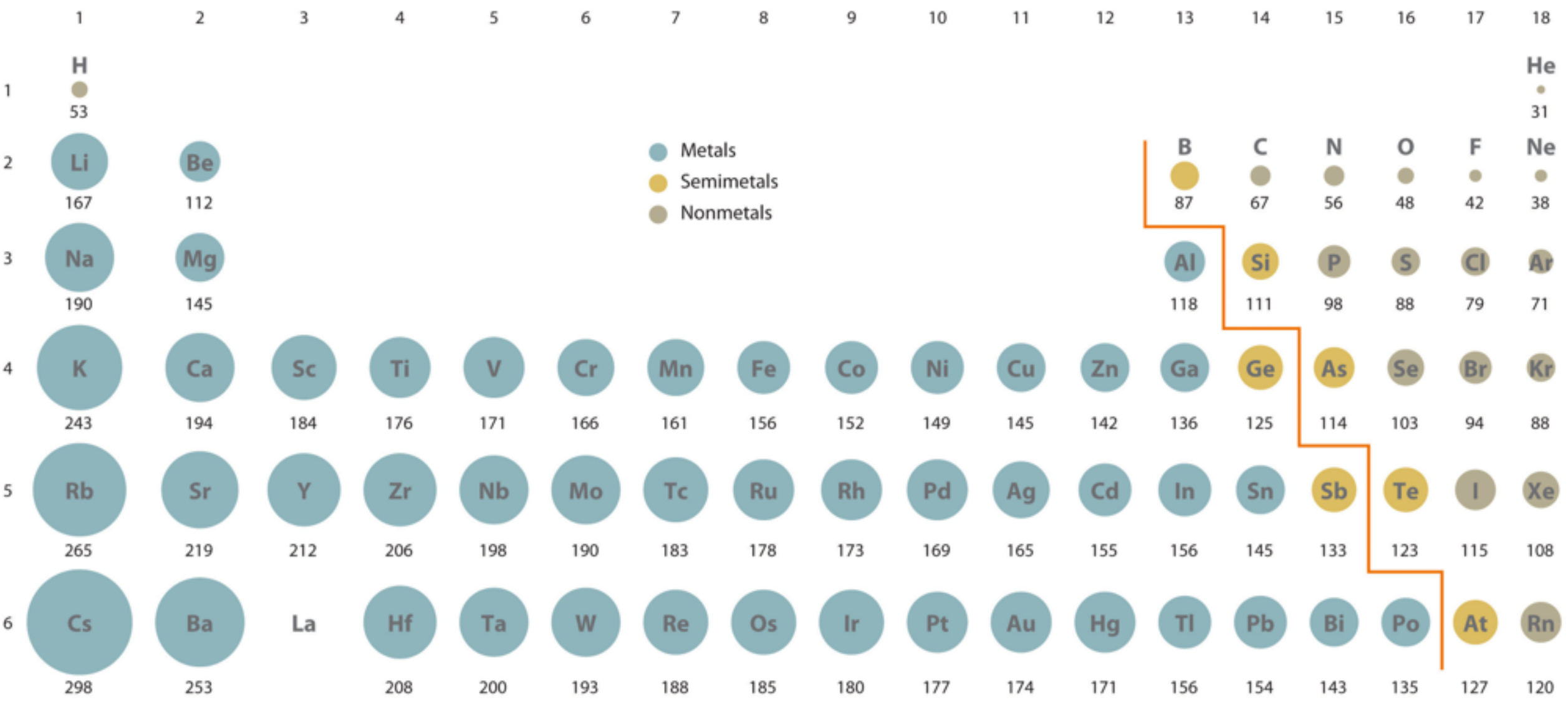


Increases up

2ND TREND: ATOMIC RADIUS

- If we know Fluorine has the highest electronegativity value, then it ***must*** hold the electrons closest to the nucleus…
- Fluorine would have a relatively ***small*** *atomic radius*
- *We can reason that the trend would* ***decrease*** *across the Periodic table (****opposite*** *of the electronegativity trend)*

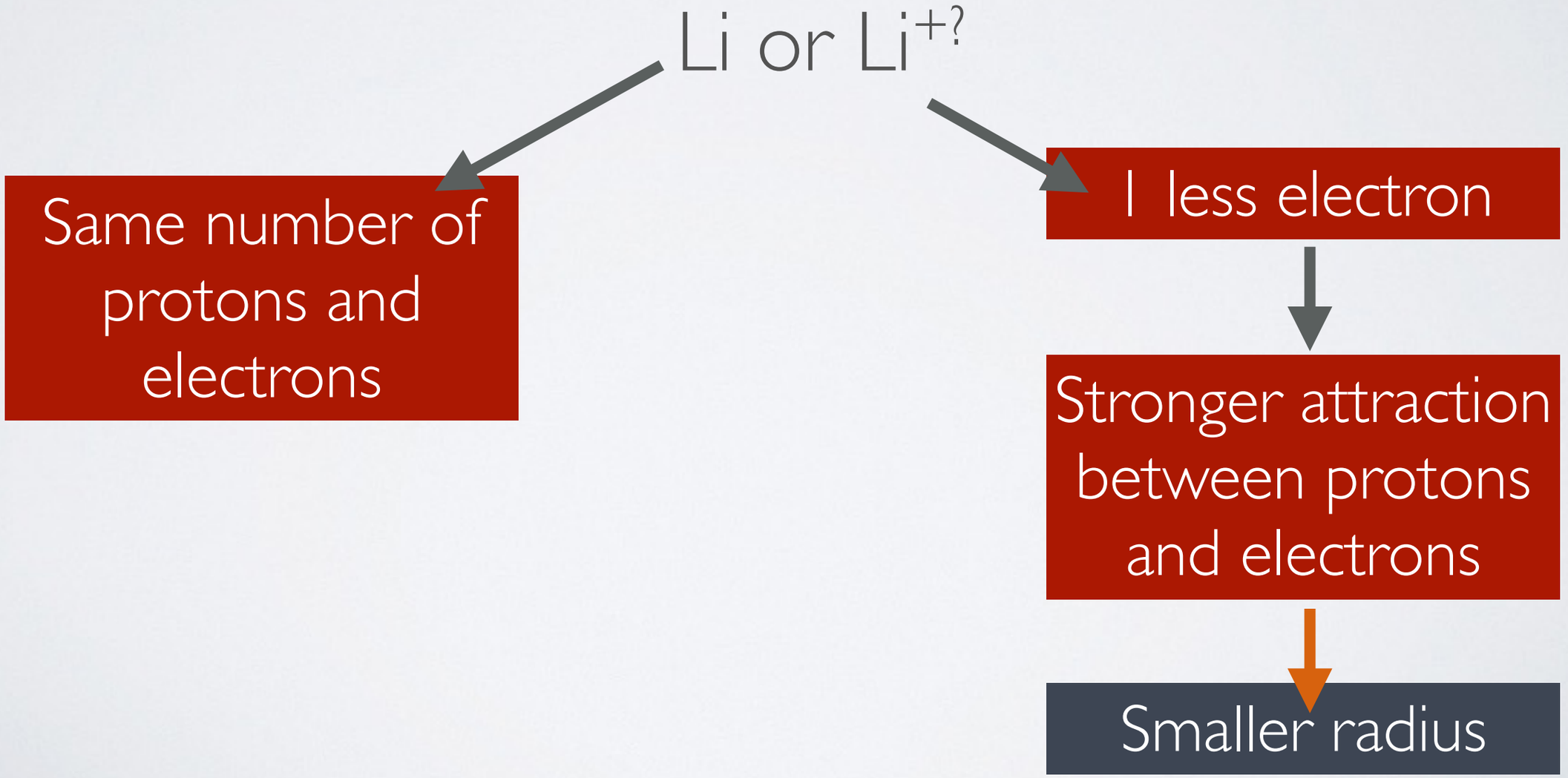




Relate the atomic radius with electronegativity!

WHAT ABOUT THE ATOMIC RADIUS OF IONS?

- Which one would have a larger atomic radius?



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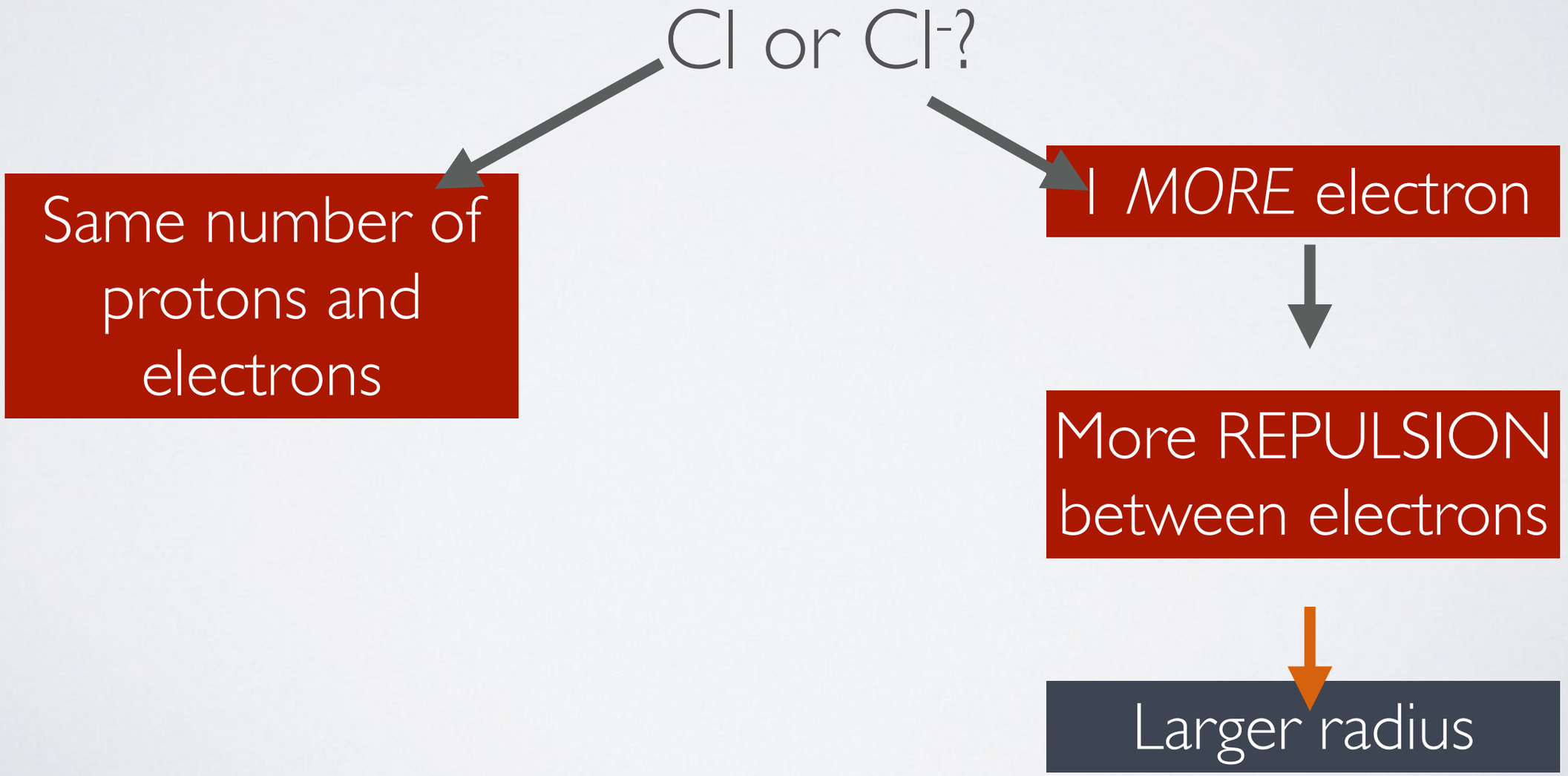
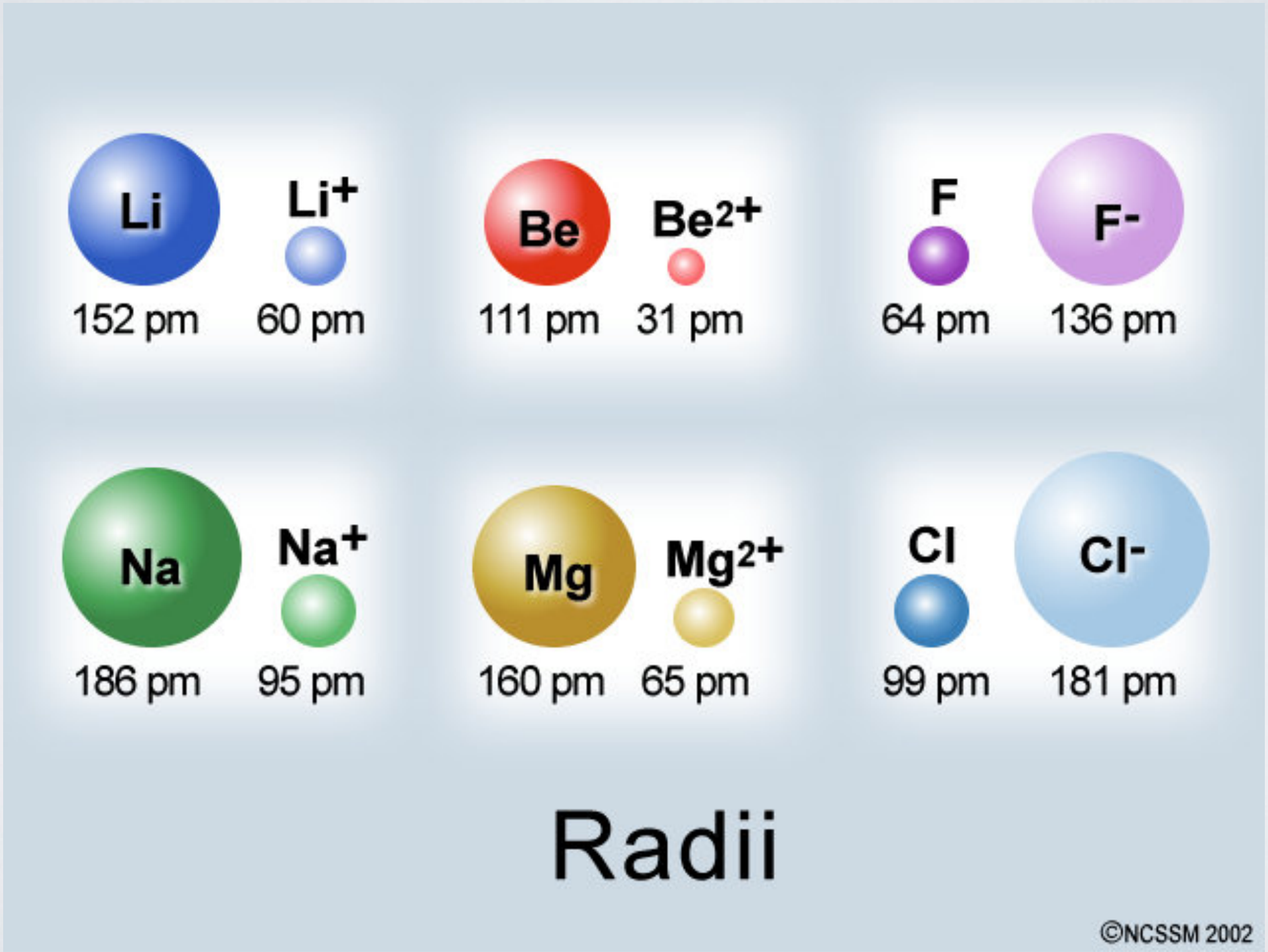


Try to make sense than to memorize



3RD TREND: IONIZATION ENERGY

- What is it?
  - Ionization energy is the amount of energy required to ***REMOVE*** *an electron*.

Can we use logic to **deduce** the trend of ionization energy?

If fluorine has the highest electronegativity

Has the strongest hold of electrons

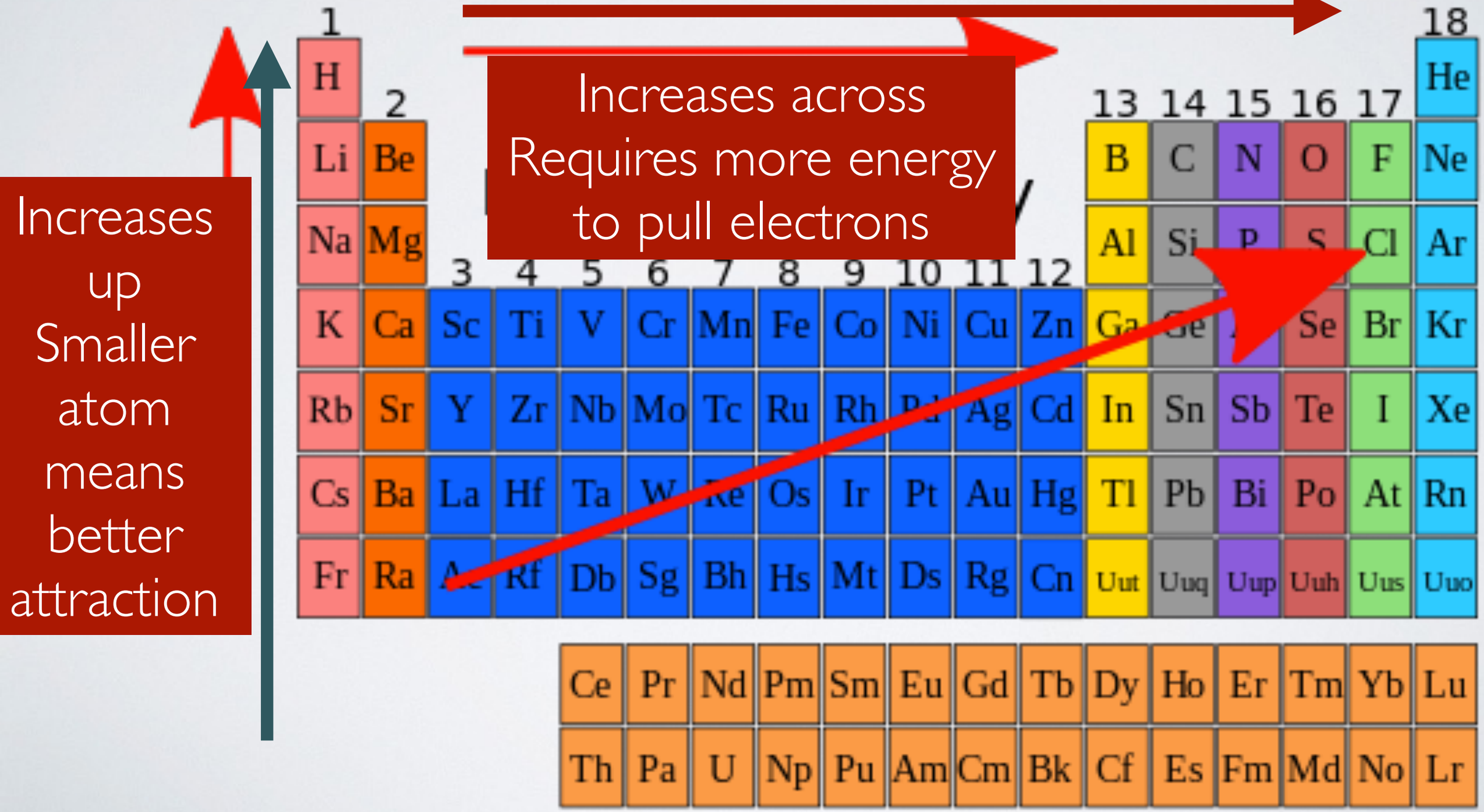
Electrons must be difficult to remove

High ionization energy

Make sense?

Increases across
Requires more energy
to pull electrons

Increases up
Smaller atom means better attraction



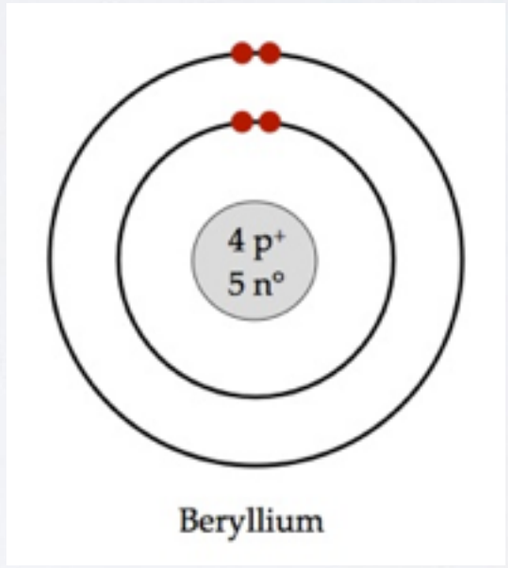
More to Ionization Energy

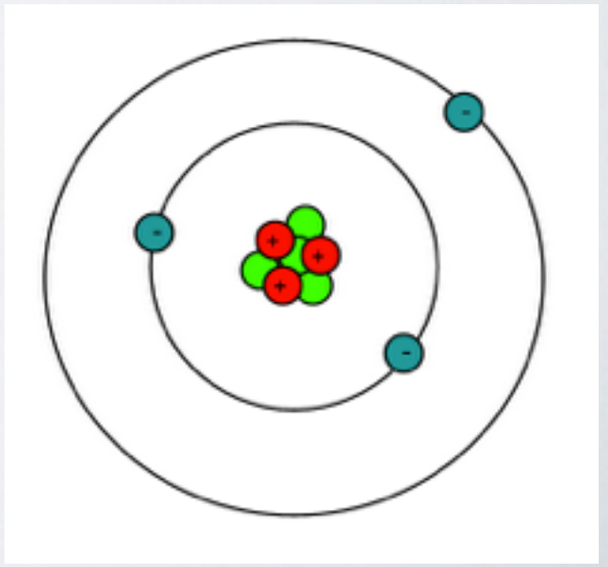
- First Ionization Energy
    - The energy required to remove 1st electron (lowest energy) from the element

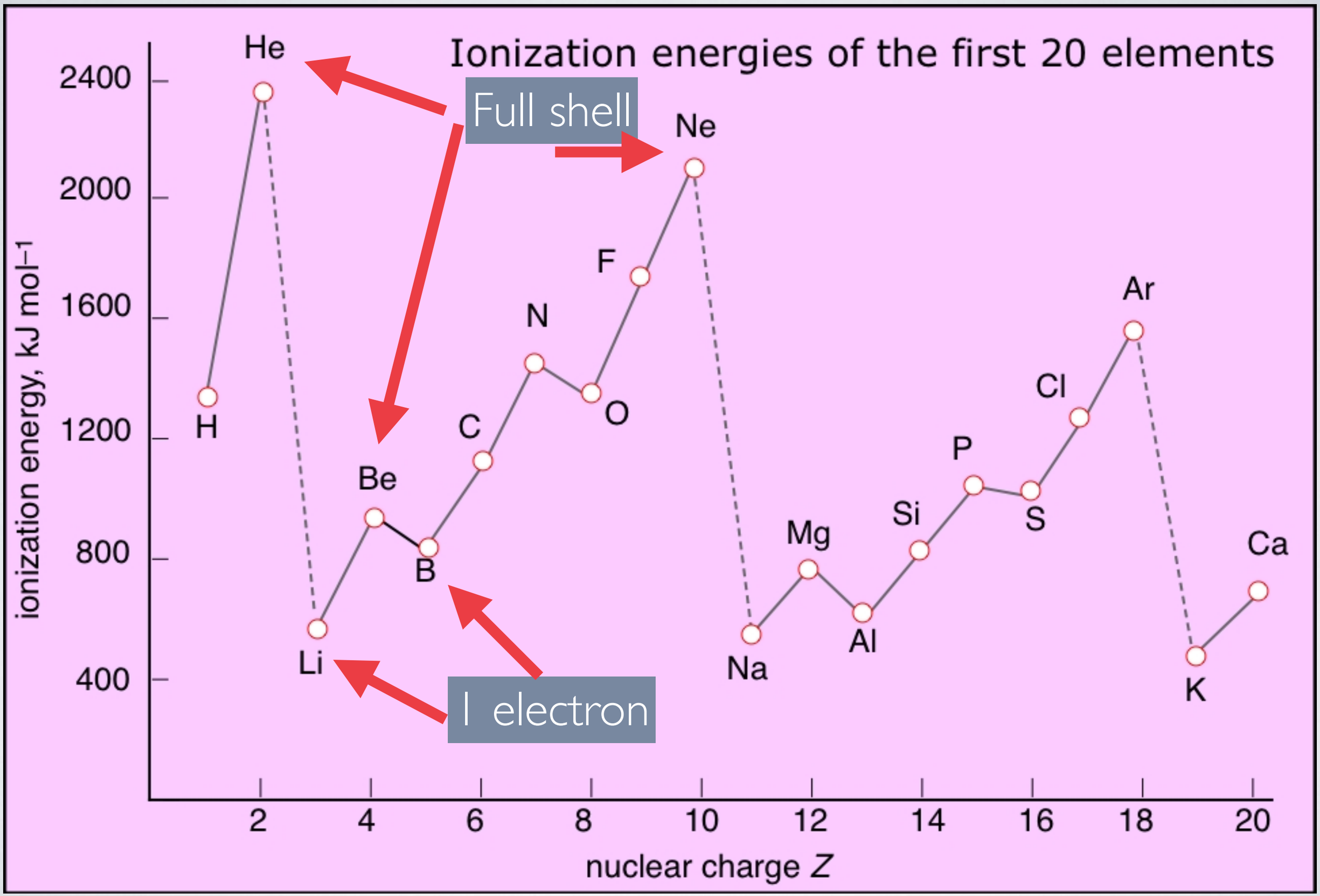
- 1st ionization energy of complete outer orbitals are higher than orbitals with just 1 electron

Example:

Requires more energy to remove an electron from Beryllium than Lithium






Ionization energies of the first 20 elements
Full shell
1 electron
ionization energy, kJ mol–1
2400
2000
1600
1200
800
400
He
Ne
Ar
H
C
N
O
F
Cl
P
S
Be
B
Mg
Si
Ca
Li
Na
Al
K
2
4
6
8
10
12
14
16
18
20
nuclear charge Z

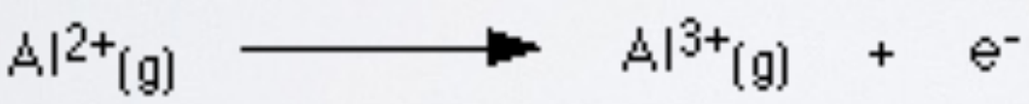
Would you expect the ionization energy to increase or decrease in *successive* electrons (the next electrons).

Ionization energy for successive electrons (2nd or 3rd electron)

Ionization energies of Aluminum

1st I.E: 577 kJ/mol

2nd I.E: 1820 kJ/mol



3rd I.E: 2740 kJ/mol

4th I.E: 11600 kJ/mol

Why the big jump for the 4th one?

Aluminum's electron configuration

ELECTRONEGATIVITY CAN DETERMINE THE TYPE OF BOND

- Consider a bond between two atoms, A and B

- A and B both have the same electronegativity value, that means they both pull the electrons the same strength.

- This is similar to H_2 or Cl_2 (Diatomic molecules)

- A non-polar covalent bond would result.

- But if B becomes more electronegativity than A,

- B atom would have more electron density making it slightly negative.

- Resulting in a *Polar bond*.

To determine the type of bond

Ex: KCl

Take the higher electronegativity value of one element

Subtract

The other element

Cl: 3.0

subtract

K: 0.8

= 2.2 Ionic Bond

Electronegativity Difference	Character of Bond	Percent Ionic Character
less than 0.4	non-polar covalent	0% – 5%
0.4–1.9	polar covalent	5% – 60%
greater than 1.9	ionic	>60%

END OF
UNIT 1: ATOMIC STRUCTURE