THE PERIODIC TABLE & TRENDS

UNIT LEARNING GOAL

Throughout this unit we will be working on **ALL** of our learning goals.

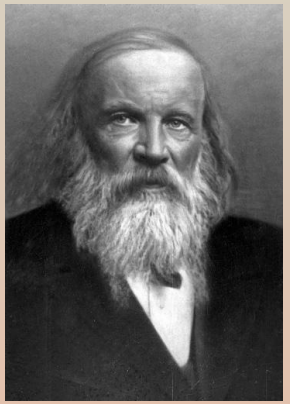
Our learning goal with respect to **<u>understanding concepts</u>** is:

We are learning to describe the periodic trends in the periodic table, and how elements combine to form chemical bonds.

TO SUPPORT THIS GOAL TODAY'S LEARNING GOALS ARE…

- We are learning to **explain** how the patterns in the electron arrangement and forces in atoms **result in periodic trends** in the periodic table **(understanding concepts).**
- We are learning how to use scientific terminology to explain these trends **(communication).**

THE PERIODIC TABLE



- Dmitri Mendeleev arranged elements in order of increasing **ATOMIC MASS**

- This showed a pattern described as Mendeleev's Periodic law:

*"If the elements are arranged according to their atomic mass, a **pattern** of similar properties can be seen"*

Mendeleev was not the first to attempt to find order within the elements, but it is his attempt that was so successful that it now forms the basis of the modern periodic table.

PERIODIC LAW

- **LATER REVISED:** When elements are arranged by **Atomic number** their chemical and physical properties recur periodically

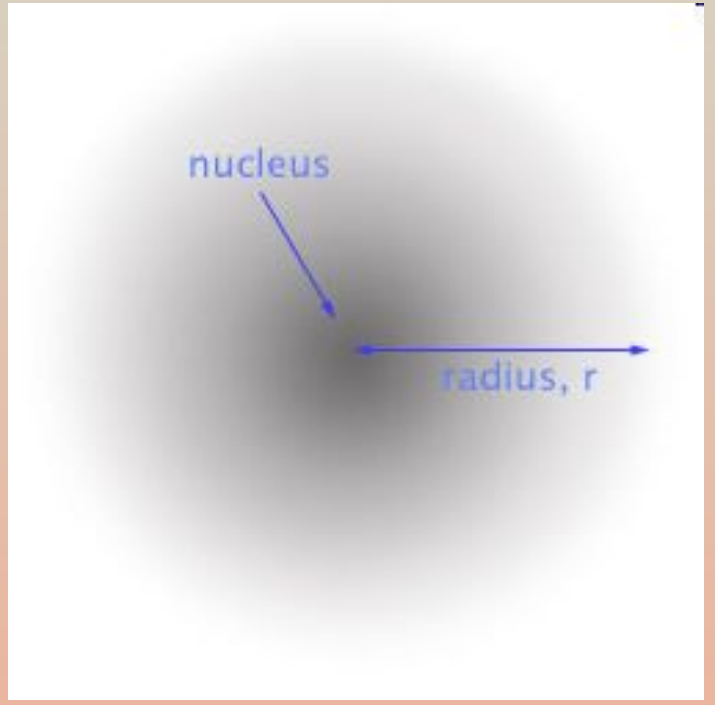
Periods: a row in the periodic table (left to right)

Groups: a column in the periodic table (top to bottom)

TRENDS

ATOMIC RADIUS

Distance between the centre of the nucleus and outer boundary of an atoms charge cloud.

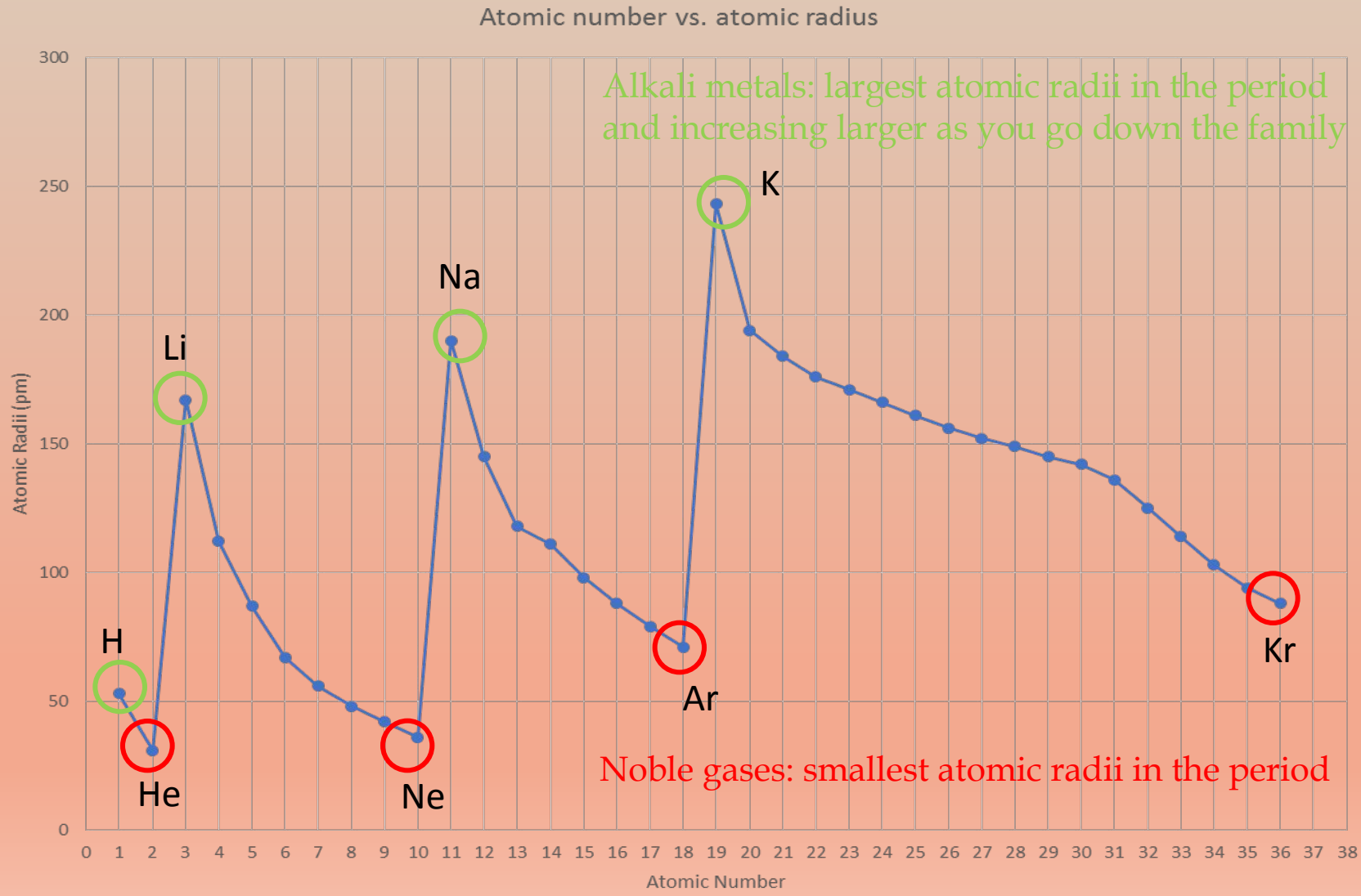


What influences size:

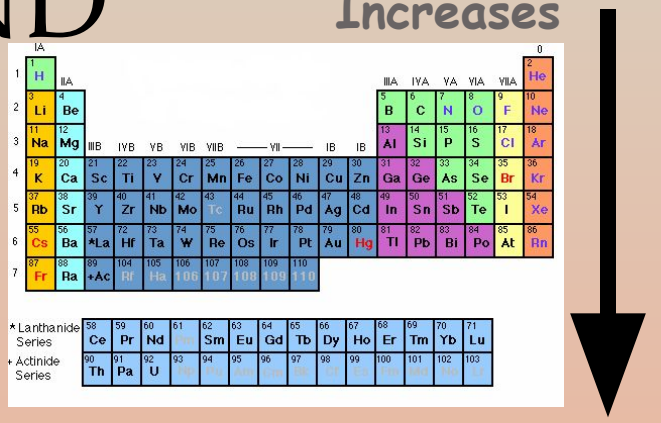
1. **Number of energy levels (n) present**
   - as 'n' increases, the probability that the outer electrons spend more time further from the nucleus increases and so the atoms become larger.

2. **The amount of nuclear charge felt by the outer electrons.**
   - as effective nuclear charge (Z_{eff}) increases, the outer electron cloud is pulled closer to the nucleus and the atom becomes smaller.

- As we move down a family or group the sizes generally increase.
- As we move across a period, the sizes of the atoms generally decrease.



ATOMIC RADIUS TREND



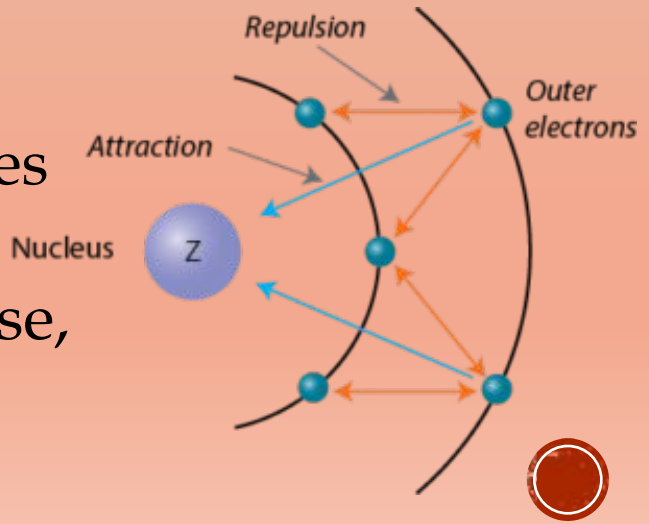
Top to bottom: Radius increases

WHY?????

Because there is an increase in the number of energy levels

- Even though the positive charge on the nucleus is increasing, each additional inner level of electrons effectively shields the outer cloud from the attractive force of that nucleus, and the electron cloud increases in size.

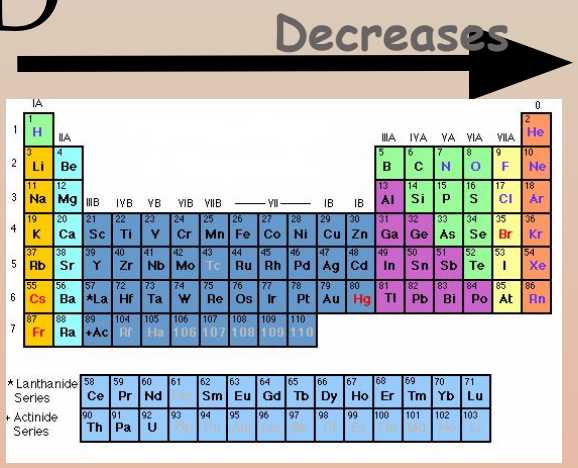
- The additional inner electron clouds also repel each other, which further contributes to the increasing size
- So, as the number of energy levels increase, the size of the atom increases.



ATOMIC RADIUS TREND

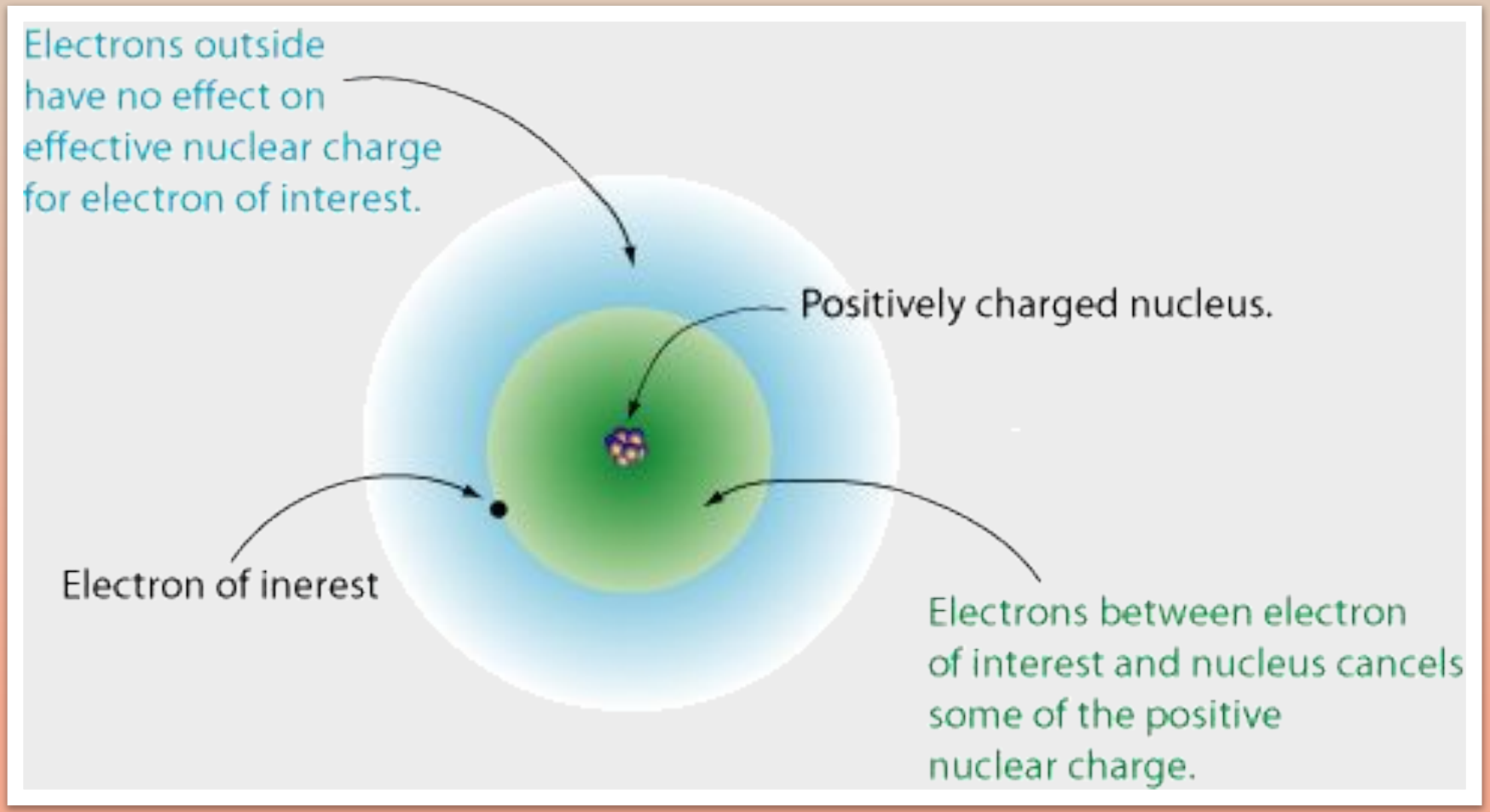
- **Left to Right: Radius decreases**

  WHY?????



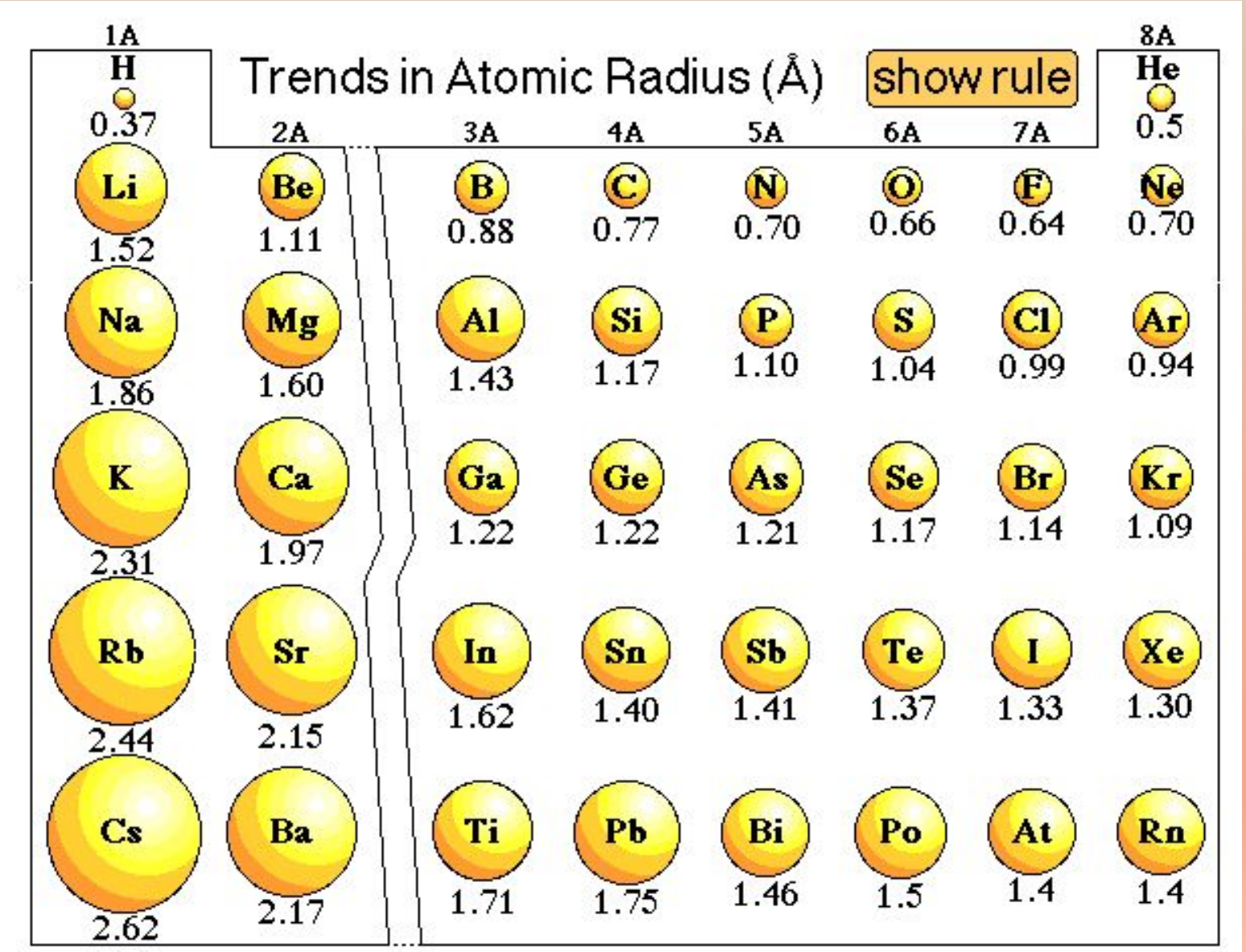
- The electrons are being added to the same outer level.
- This means that the level of shielding by the inner electrons remains the same.
- It also means that the added electrons are ineffective at shielding each other from the increasing positive charge on the nucleus.
- The result is that the Z_{eff} on those outer electrons increases and the charge cloud is pulled closer and closer to the nucleus, so the size of the atom decreases.

NUCLEAR CHARGE



This greatly effects period trends - shielding effect is not an issue across a period.

ATOMIC RADIUS TREND



EXCEPTIONS…

- These can be explained by considering the electron configurations of the atoms and the electron sublevels being filled in each case.

- Ex. Period 4

V: $[Ar]3d^3 4s^2$ vs Zn: $[Ar]3d^{10} 4s^2$

Even though zinc has 7 more protons, the extra shielding provided by its 10 electrons in the 3d sublevel is such that zinc's atomic radius is equal to vanadium's

WHAT ABOUT IONS???

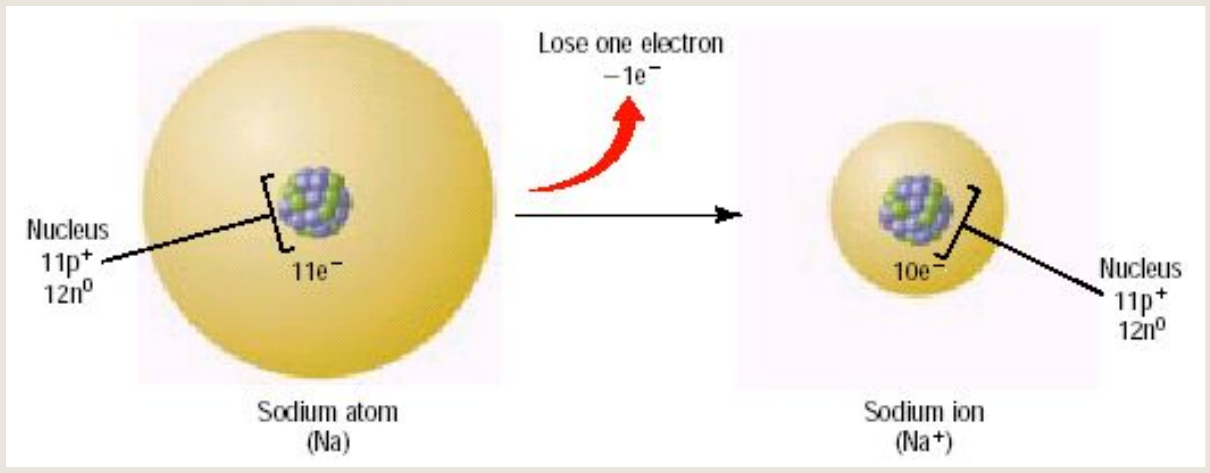
The same forces influence ions…

Consider: Al^{3+}, F^-, Mg^{2+}, N^{3-}, Na^+, Ne, O^{2-}

- All have 10 electrons therefore:
  - isoelectronic with Ne
  - same amount of shielding
  - So, attractive force from nucleus is only factor influencing size…great # protons = stronger attractive force on electron cloud = smaller atom or ion.

$$N^{3-} > O^{2-} > F^- > Ne > Na^+ > Mg^{2+} > Al^{3+}$$

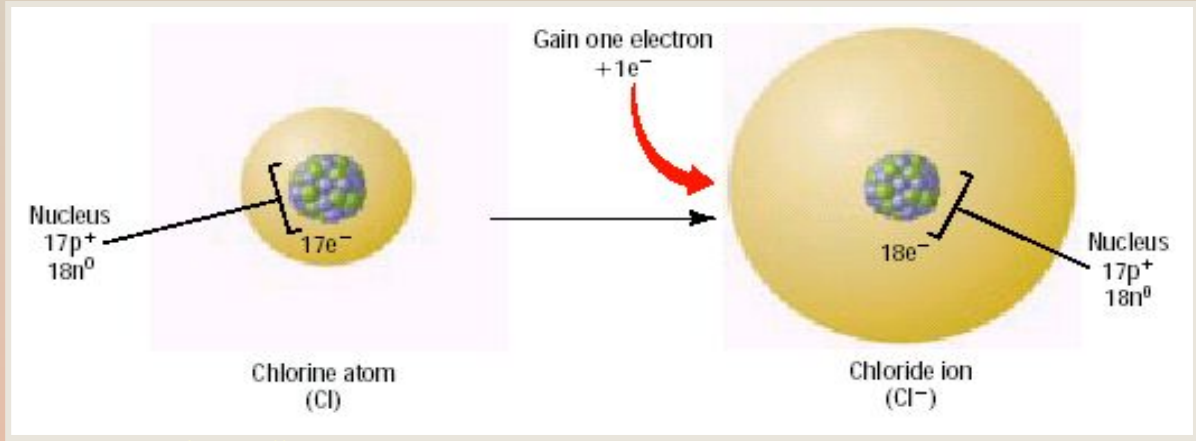
CATION



Removing electrons produces an ion that is <u>smaller</u> than the neutral atom

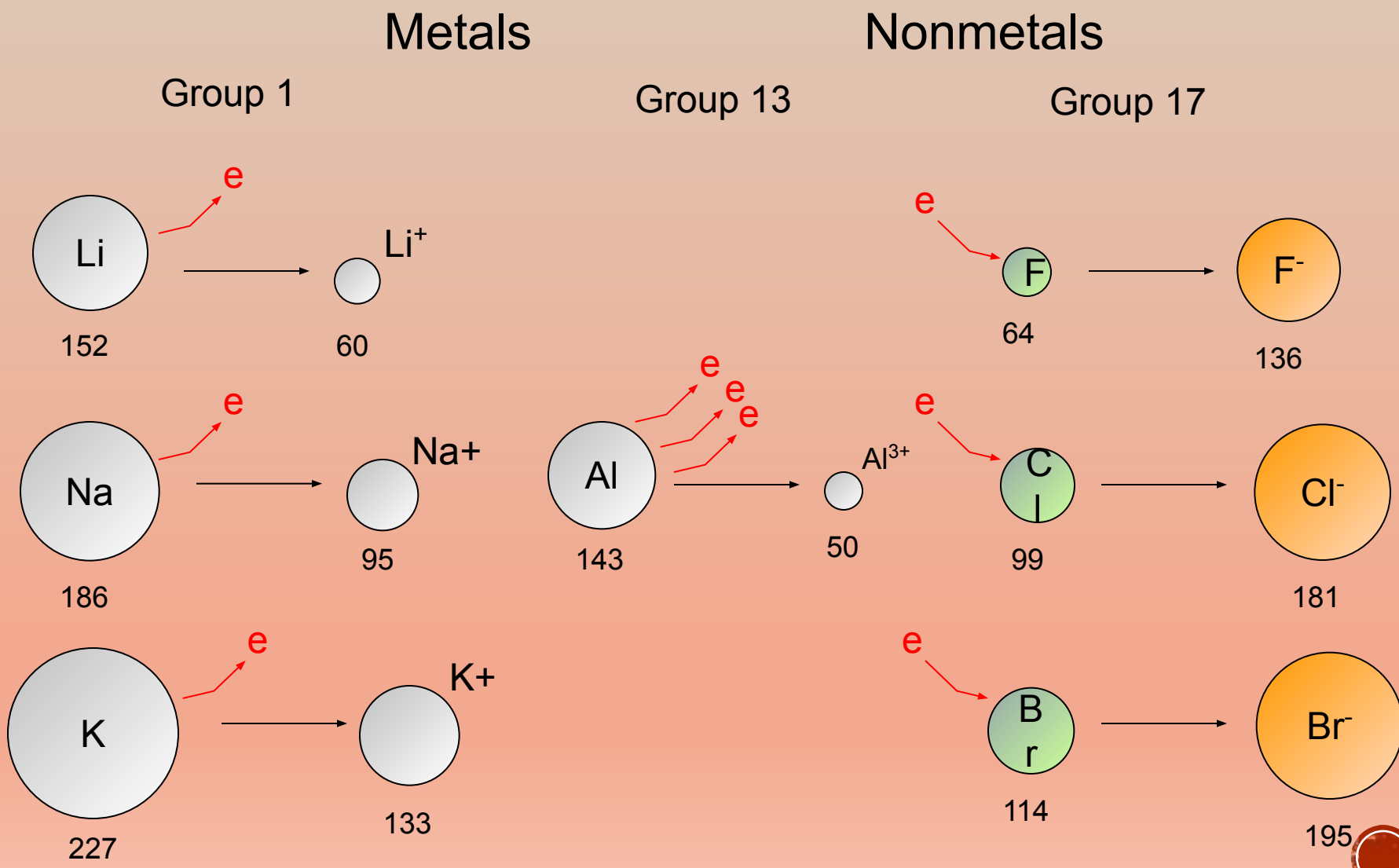
- less repulsion between the electrons.
- the number of shell decreases.
- Li+ has smaller radius than Li

ANION



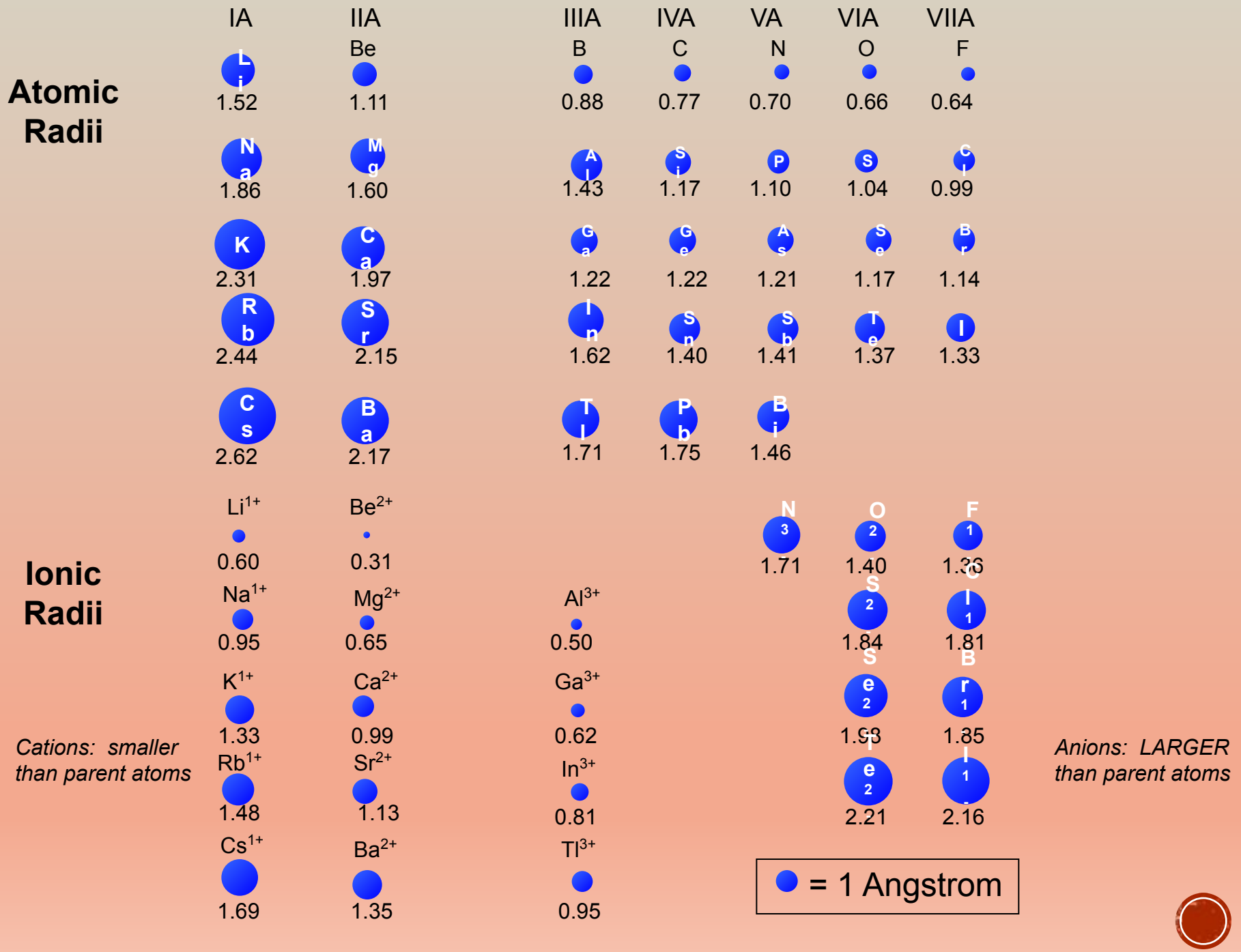
- By gaining electrons, there is an increase in electron repulsion while the number of shells stays the same
- For example, F- has a larger radius than F.

TRENDS IN ATOMIC AND IONIC SIZE



Cations are smaller than parent atoms

Anions are larger than parent atoms


Atomic Radii
IA
IIA
IIIA
IVA
VA
VIA
VIIA
Be
B
C
N
O
F
Li
1.52
1.11
0.88
0.77
0.70
0.66
0.64
Na
Mg
Al
Si
P
S
Cl
1.86
1.60
1.43
1.17
1.10
1.04
0.99
K
Ca
Ga
Ge
As
Se
Br
2.31
1.97
1.22
1.22
1.21
1.17
1.14
Rb
Sr
In
Sn
Sb
Te
I
2.44
2.15
1.62
1.40
1.41
1.37
1.33
Cs
Ba
Tl
Pb
Bi
2.62
2.17
1.71
1.75
1.46
Ionic Radii
Li1+
Be2+
N3-
O2-
F1-
0.60
0.31
1.71
1.40
1.36
Na1+
Mg2+
Al3+
S2-
Cl1-
0.95
0.65
0.50
1.84
1.81
K1+
Ca2+
Ga3+
Se2-
Br1-
1.33
0.99
0.62
1.93
1.85
Rb1+
Sr2+
In3+
Te2-
I1-
1.48
1.13
0.81
2.21
2.16
Cs1+
Ba2+
Tl3+
1.69
1.35
0.95
Cations: smaller than parent atoms
Anions: LARGER than parent atoms
= 1 Angstrom

IONIZATION ENERGY

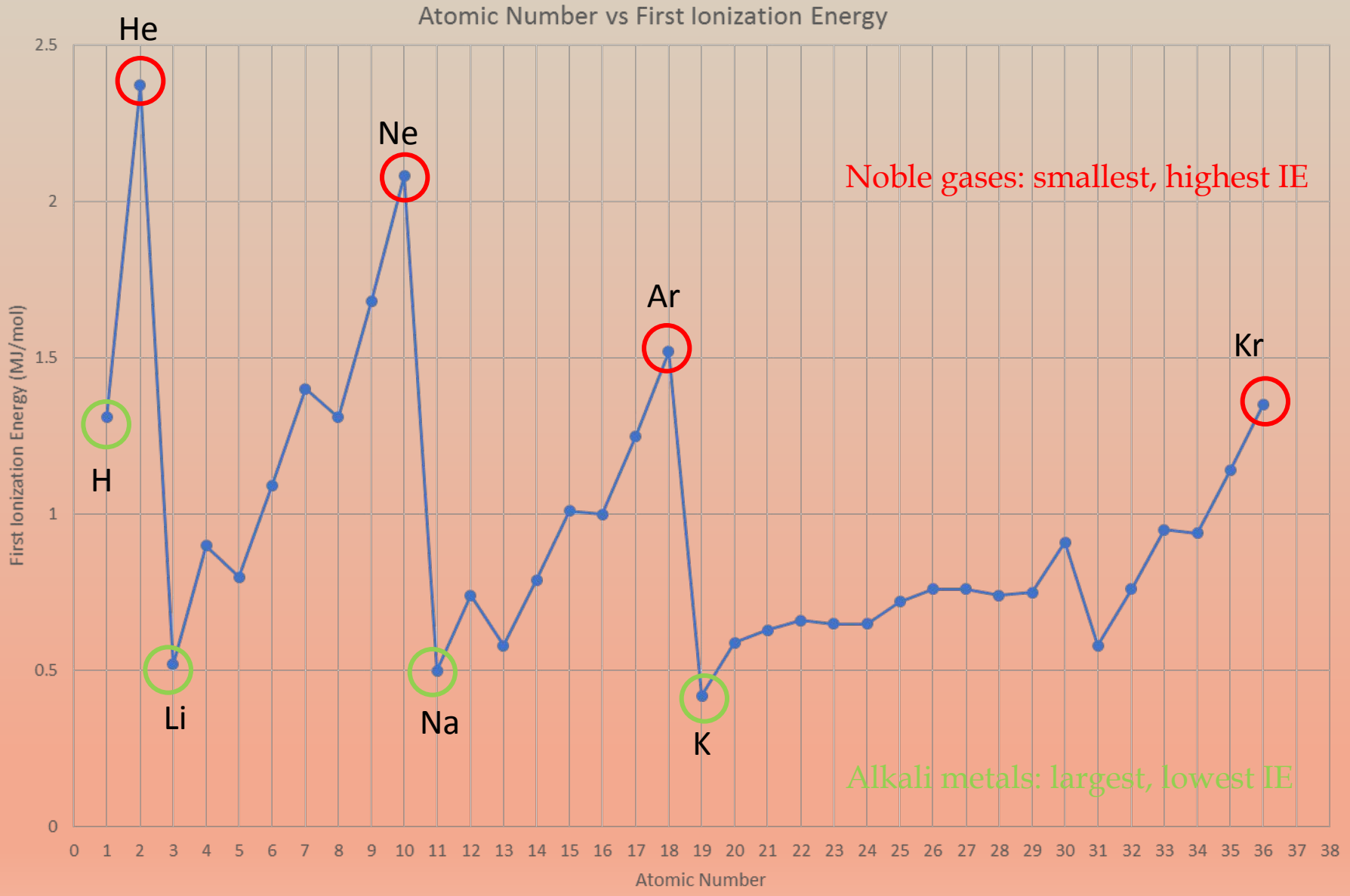
- The minimum energy required to **remove an electron** from a gaseous atom or ion.

- It often refers to the first ionization energy (IE_1) whereby a neutral atom becomes a 1^+ cation:

$$atom_{(g)} + IE_1 \rightarrow ion^+_{(g)} + e^-$$

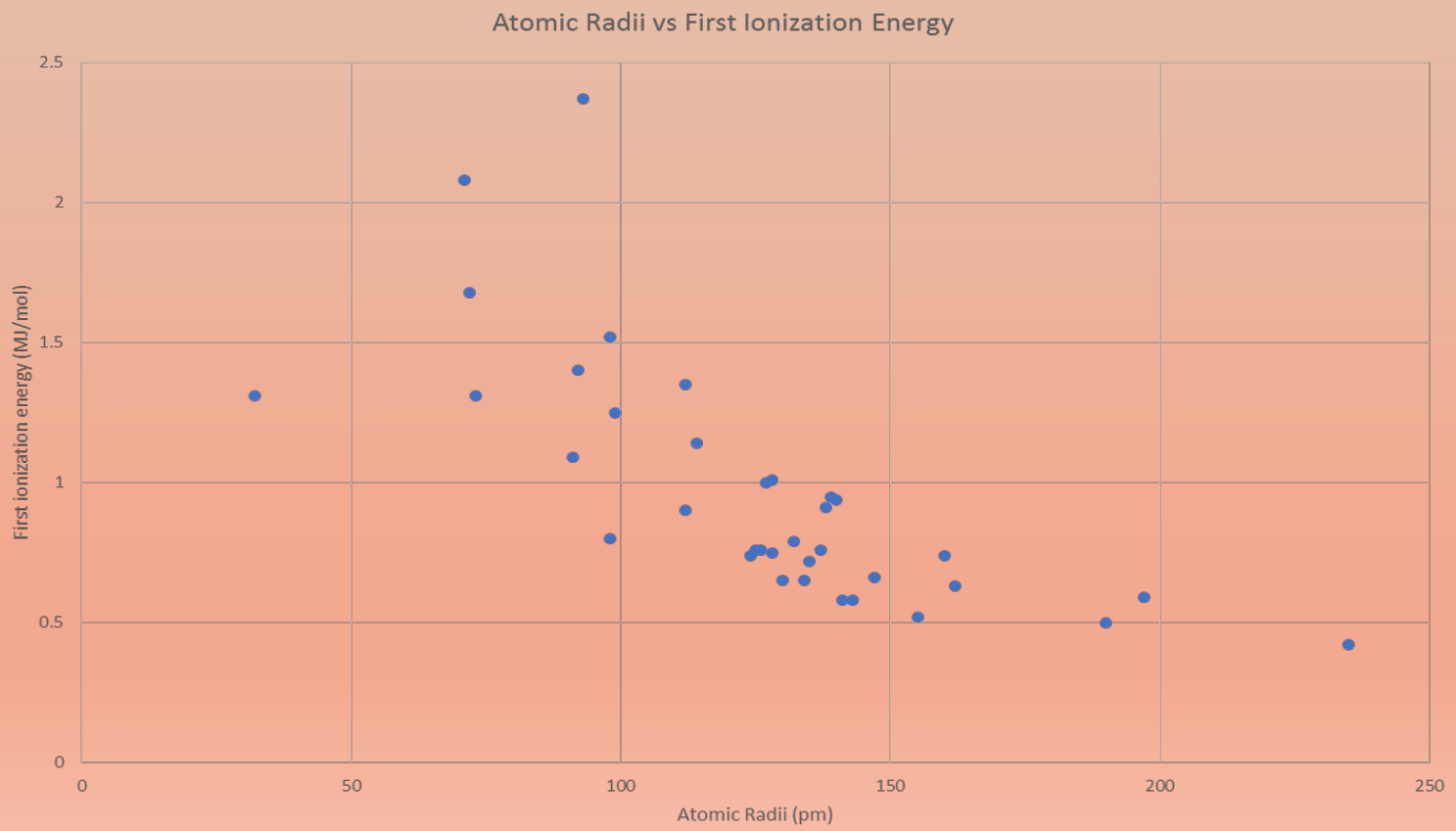
It tells us how strongly an atom holds onto its outermost electrons.

An element with a low IE will be more likely to lose electrons and form cations.


Atomic Number vs First Ionization Energy
First Ionization Energy (MJ/mol)
Atomic Number
He
Ne
Ar
Kr
H
Li
Na
K
Noble gases: smallest, highest IE
Alkali metals: largest, lowest IE

IONIZATION ENERGY

- Generally, a large atom, whose electrons are held less tightly, would have a lower IE, than a smaller atom whose outer electrons are held much more strongly.
- As atomic size decreases, ionization energy increases.

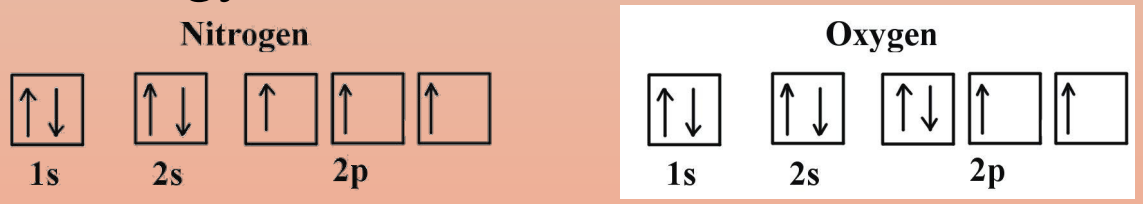


EXCEPTIONS…

- Again, exceptions can be explained by analyzing electron configurations.

N vs. O

Oxygen is smaller than Nitrogen, but oxygen has a lower first ionization energy.



Nitrogen has a single electron in each of its three 2p orbitals and a half-filled p sublevel that is quite stable.

Oxygen has a pair of electrons in one of its 2p orbitals. This increases electron-electron repulsion associated with that pairing making it easier for oxygen to lose one of those electrons.

- Atoms with more than one electron have more than one ionization energy.
  - energies correspond to the stepwise removal of electrons, one after another

- **First Ionization energy** is the energy needed to remove the first electron from an atom.
- **Second Ionization energy** is the energy needed to remove the second electron from an atom.
- **Third Ionization energy** is the energy needed to remove the third electron from an atom.

Ionization Energies in kJ/mol

	1	2	3	4	5	6	7
H	1312						
He	2372	5250					
Li	520	7297	11810				
Be	899	1757	14845	21000			
B	800	2426	3659	25020	32820		
C	1086	2352	4619	6221	37820	47260	
N	1402	2855	4576	7473	9442	53250	64340

After removing one electron from an atom, further ionization energies increase because each successive electron removed is being separated from an increasing positive ion.

The equation representing the first ionization energy of hydrogen... $H + 1312\ kJ \rightarrow H^+ + e^-$

SUCCESS CRITERIA

- **By the end of this lesson…**

- I can use the periodic trends **to predict the properties** of atomic radius, ionization energy, electron affinity and electronegativity **for elements on the periodic table (*undestanding concepts*).**
- I can **explain the periodic trends** of atomic radius, ioniation energy, electron affinity, and electronegativity based on **patterns in electron arrangement and forces in atoms (*communication*).**