




Chemical Periodicity

Periodic Table




Classification of the Elements

- OBJECTIVES:
 - Explain why you can infer the properties of an element based on those of other elements in the periodic table.
- 



Classification of the Elements

- OBJECTIVES:
 - Use electron configurations to classify elements as noble gases, representative elements, transition metals, or inner transition metals.
- 



Periodic Table Revisited


- Russian scientist Dmitri Mendeleev taught chemistry in terms of properties.
- Mid 1800's - molar masses of elements were known.
- Wrote down the elements in order of increasing mass.
- Found a pattern of repeating properties.

Mendeleev's Table

- Grouped elements in columns by similar properties in order of increasing atomic mass.
- Found some inconsistencies - felt that the properties were more important than the mass, so switched order.
- Also found some gaps.
- Must be undiscovered elements.
- Predicted their properties before they were found.

The modern table

- Elements are still grouped by properties.
- Similar properties are in the same column.
- Order is by increasing atomic number.
- Added a column of elements Mendeleev didn't know about.
- The noble gases weren't found because they didn't react with anything.

- 
- Horizontal rows are called periods
 - There are 7 periods



Vertical columns called groups

Elements are placed in columns by similar properties

Also called families



1A

- The elements in the A groups are called the representative elements

8A
0

2A outer s or p filling

3A 4A 5A 6A 7A



1

- Modern numbering of groups is 1 to 18 group 18 is also called zero group

18

2


13 14 15 16 17

3 4 5 6 7 8 9 10 11 12





Why?

- The part of the atom another atom sees is the electron cloud.
 - More importantly the outside orbitals.
 - The orbitals fill up in a regular pattern.
 - The outside orbital electron configuration repeats.
 - The properties of atoms repeat.
- 

H
1 $1s^1$

Li
3 $1s^2 2s^1$

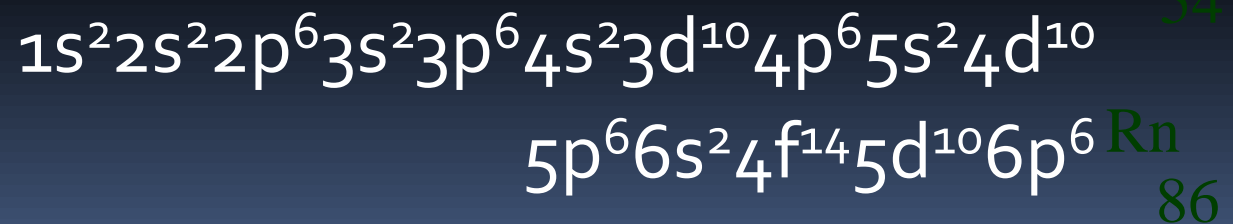
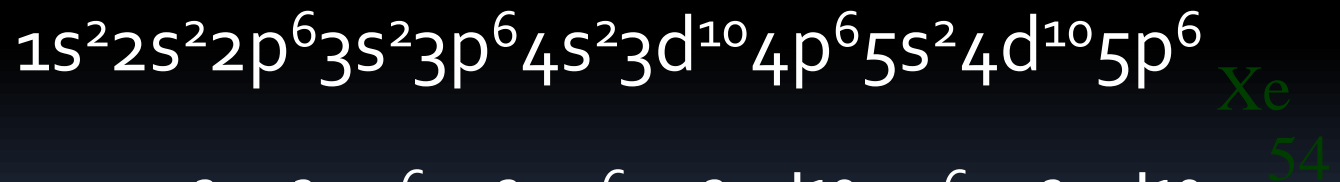
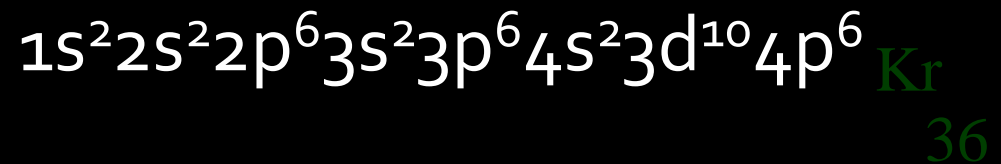
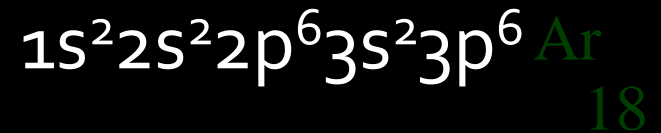
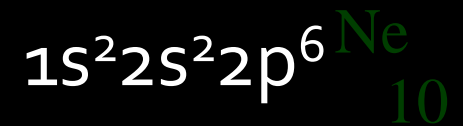
Na
11 $1s^2 2s^2 2p^6 3s^1$

K
19 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

Rb
37 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$

Cs
55 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$

Fr
87 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^1$



S- block

s^1

s^2

- Alkali metals all end in s^1
- Alkaline earth metals all end in s^2
- really should include He, but it fits better later.
- He has the properties of the noble gases.

Transition Metals -d block

d^1 d^2 d^3 $s^1 d^5$ d^5 d^6 d^7 d^8 $s^1 d^{10}$ d^{10}



The P-block

p^1

p^2

p^3

p^4

p^5

p^6





F - block

- inner transition elements

f^1 f^2 f^3 f^4 f^5 f^6 f^7 f^8 f^9 f^{10} f^{11} f^{12} f^{13} f^{14}





Writing electron configurations the easy way

Yes there is a shorthand



Electron Configurations repeat

- The shape of the periodic table is a representation of this repetition.
- When we get to the end of the column the outermost energy level is full.
- This is the basis for our shorthand.

The Shorthand


- Write symbol of the noble gas before the element, in [].
- Then, the rest of the electrons.
- Aluminum's full configuration:
 $1s^2 2s^2 2p^6 3s^2 3p^1$
- previous noble gas Ne is: $1s^2 2s^2 2p^6$
- so, Al is: $[\text{Ne}] 3s^2 3p^1$

More examples

- $\text{Ge} = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$
 - Thus, $\text{Ge} = [\text{Ar}] 4s^2 3d^{10} 4p^2$
- $\text{Hf} = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^2$
 - Thus, $\text{Hf} = [\text{Xe}] 6s^2 4f^{14} 5d^2$




Periodic Trends

- OBJECTIVES:
 - Interpret group trends in atomic radii, ionic radii, ionization energies, and electronegativities.
- 



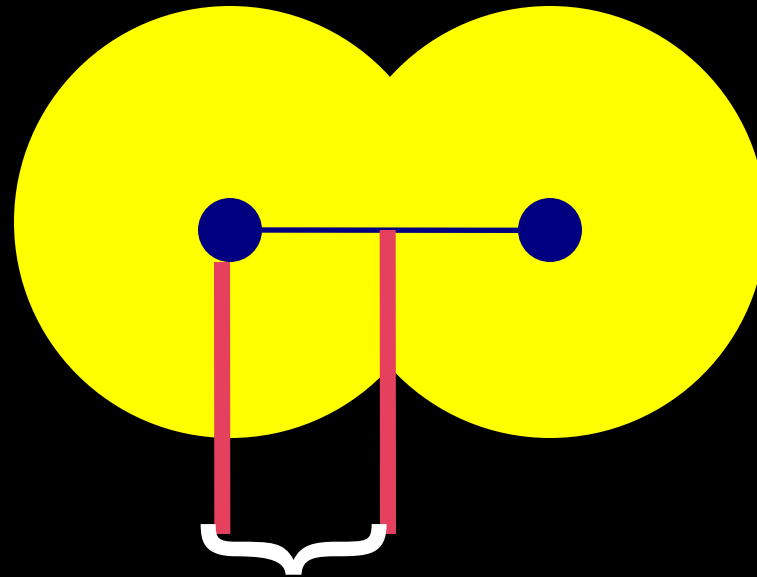
Periodic Trends

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- 

Trends in Atomic Size

- First problem: Where do you start measuring from?
- The electron cloud doesn't have a definite edge.
- They get around this by measuring more than 1 atom at a time.

Atomic Size



Radius

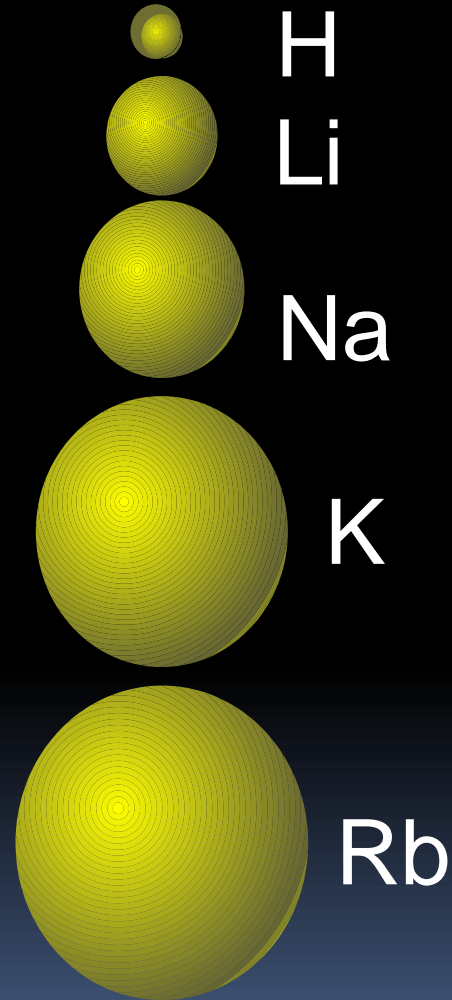
- Atomic Radius = half the distance between two nuclei of a diatomic molecule.

Trends in Atomic Size

- Influenced by three factors:
 1. Energy Level
 - Higher energy level is further away.
 2. Charge on nucleus
 - More charge pulls electrons in closer.
 - 3. Shielding effect
 - (blocking effect?)

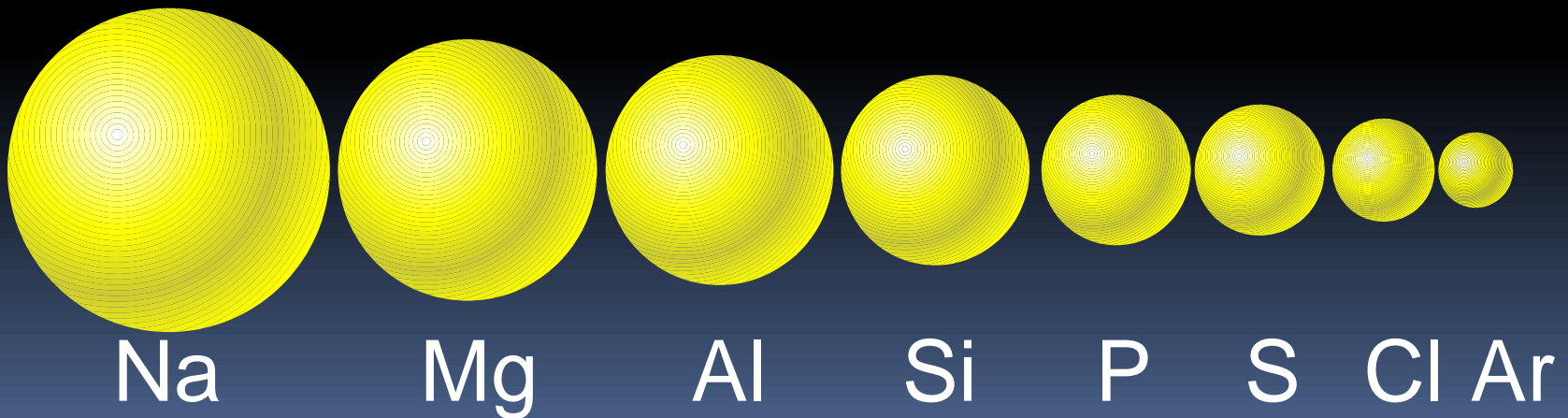
Group trends

- As we go down a group...
- each atom has another energy level,
- so the atoms get bigger.



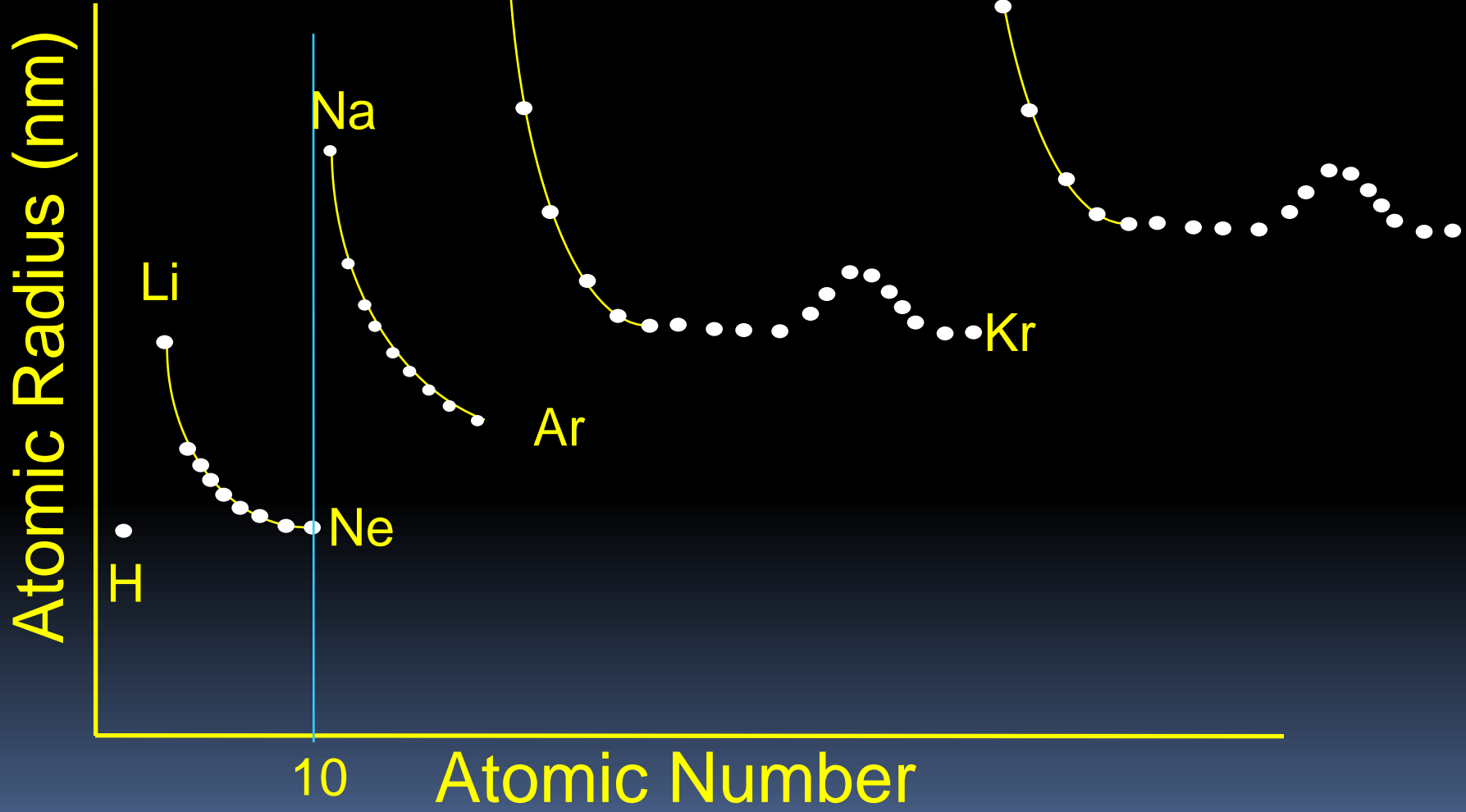
Periodic Trends

- As you go across a period, the radius gets smaller.
- Electrons are in same energy level.
- More nuclear charge.
- Outermost electrons are closer.





Overall



Trends in Ionization Energy

- The amount of energy required to completely remove an electron from a gaseous atom.
- Removing one electron makes a $1+$ ion.
- The energy required to remove the first electron is called the first ionization energy.

Ionization Energy

- The second ionization energy is the energy required to remove the second electron.
- Always greater than first IE.
- The third IE is the energy required to remove a third electron.
- Greater than 1st or 2nd IE.

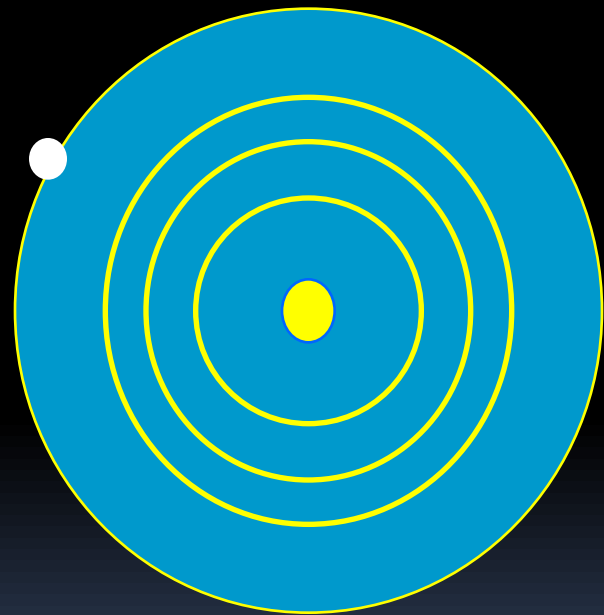
Symbol	First	Second	Third
H	1312		
He	2731	5247	
Li	520	7297	11810
Be	900	1757	14840
B	800	2430	3569
C	1086	2352	4619
N	1402	2857	4577
O	1314	3391	5301
F	1681	3375	6045
Ne	2080	3963	6276

What determines IE

- The greater the nuclear charge, the greater IE.
- Greater distance from nucleus decreases IE
- Filled and half-filled orbitals have lower energy, so achieving them is easier, higher IE.
- Shielding effect


Shielding

- The electron on the outermost energy level has to look through all the other energy levels to see the nucleus.
- Second electron has same shielding, if it is in the same period





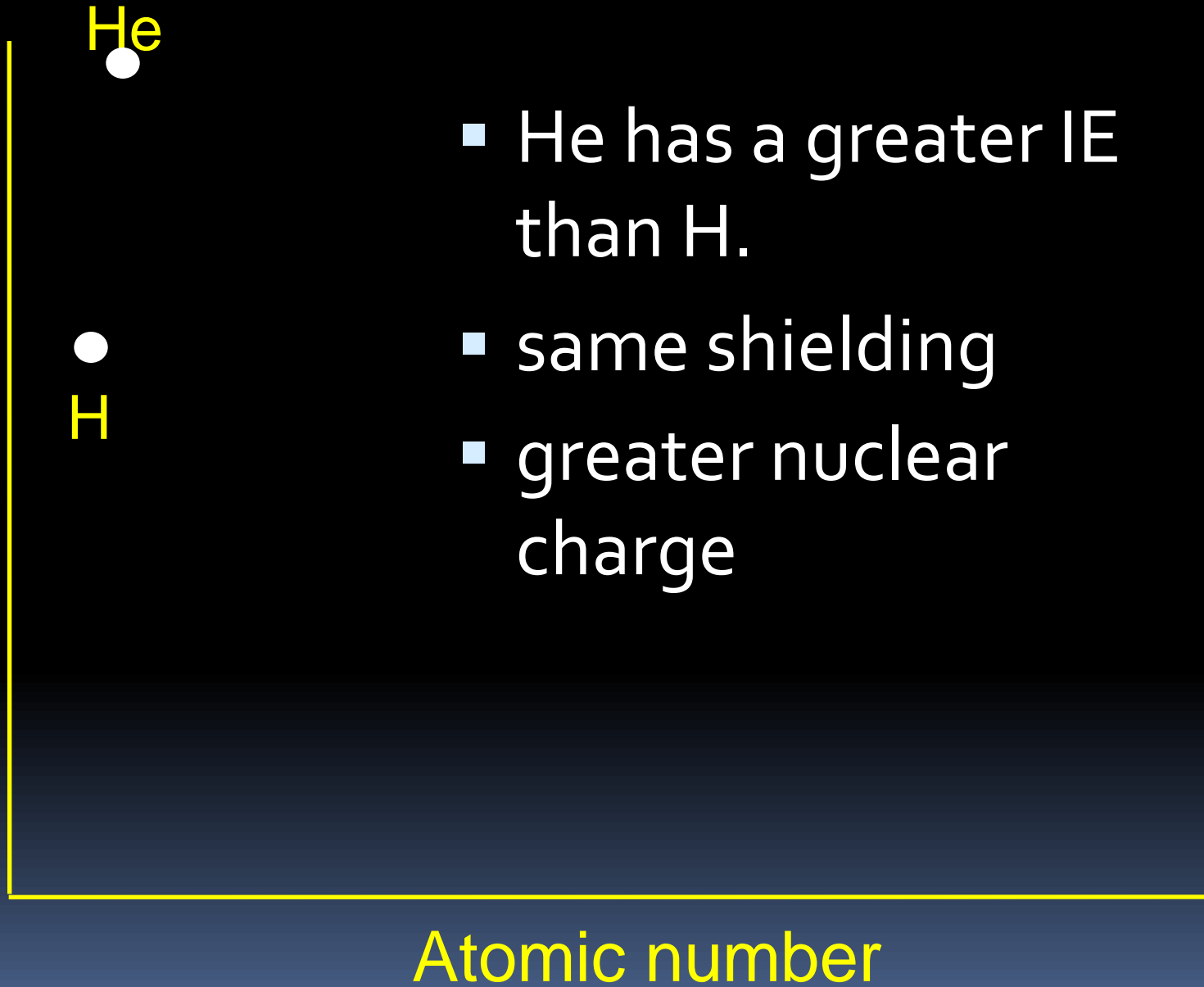
Group trends

- As you go down a group, first IE decreases because...
 - The electron is further away.
 - More shielding.
- 

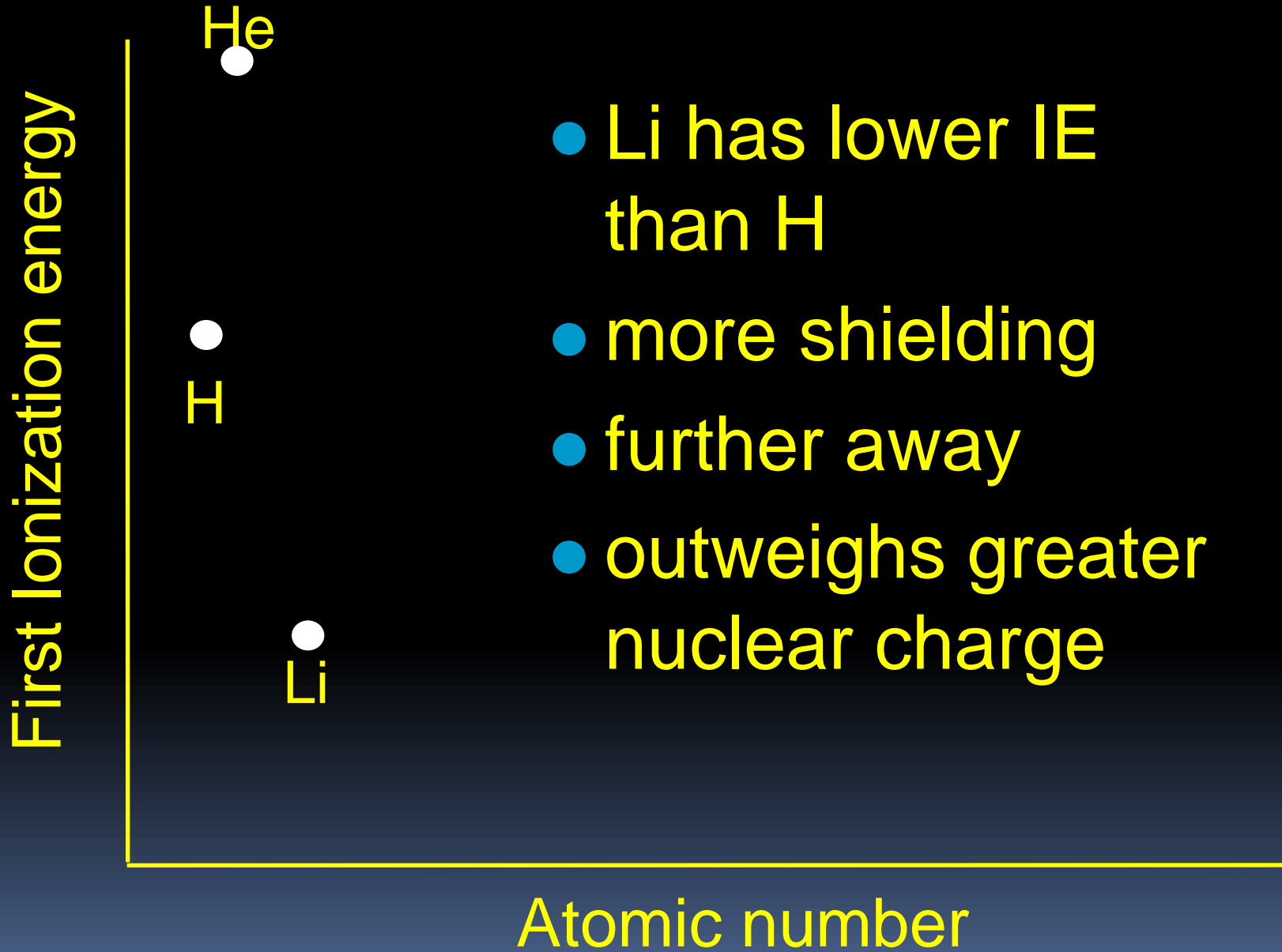
Periodic trends

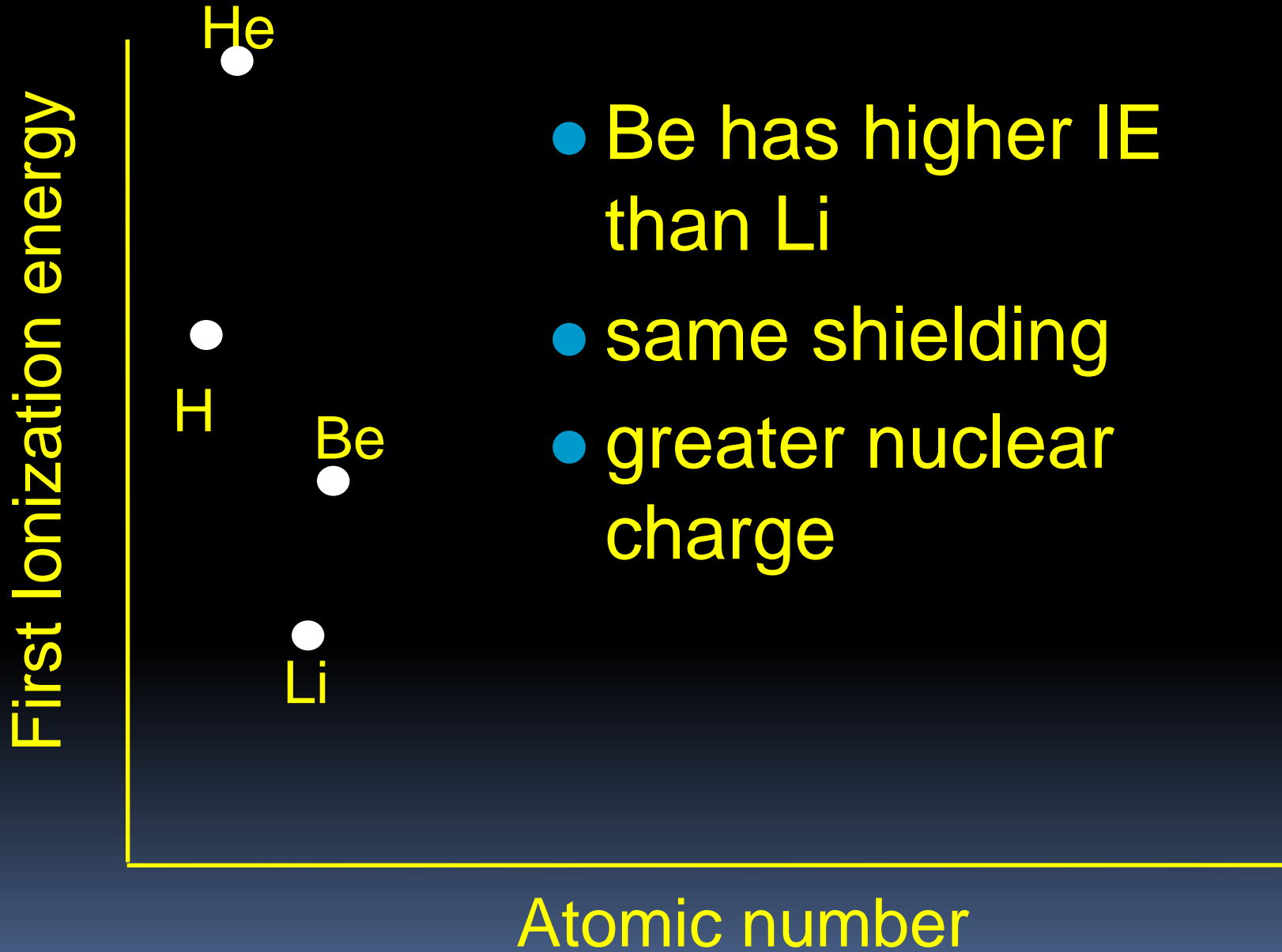
- All the atoms in the same period have the same energy level.
- Same shielding.
- But, increasing nuclear charge
- So IE generally increases from left to right.
- Exceptions at full and $1/2$ full orbitals.

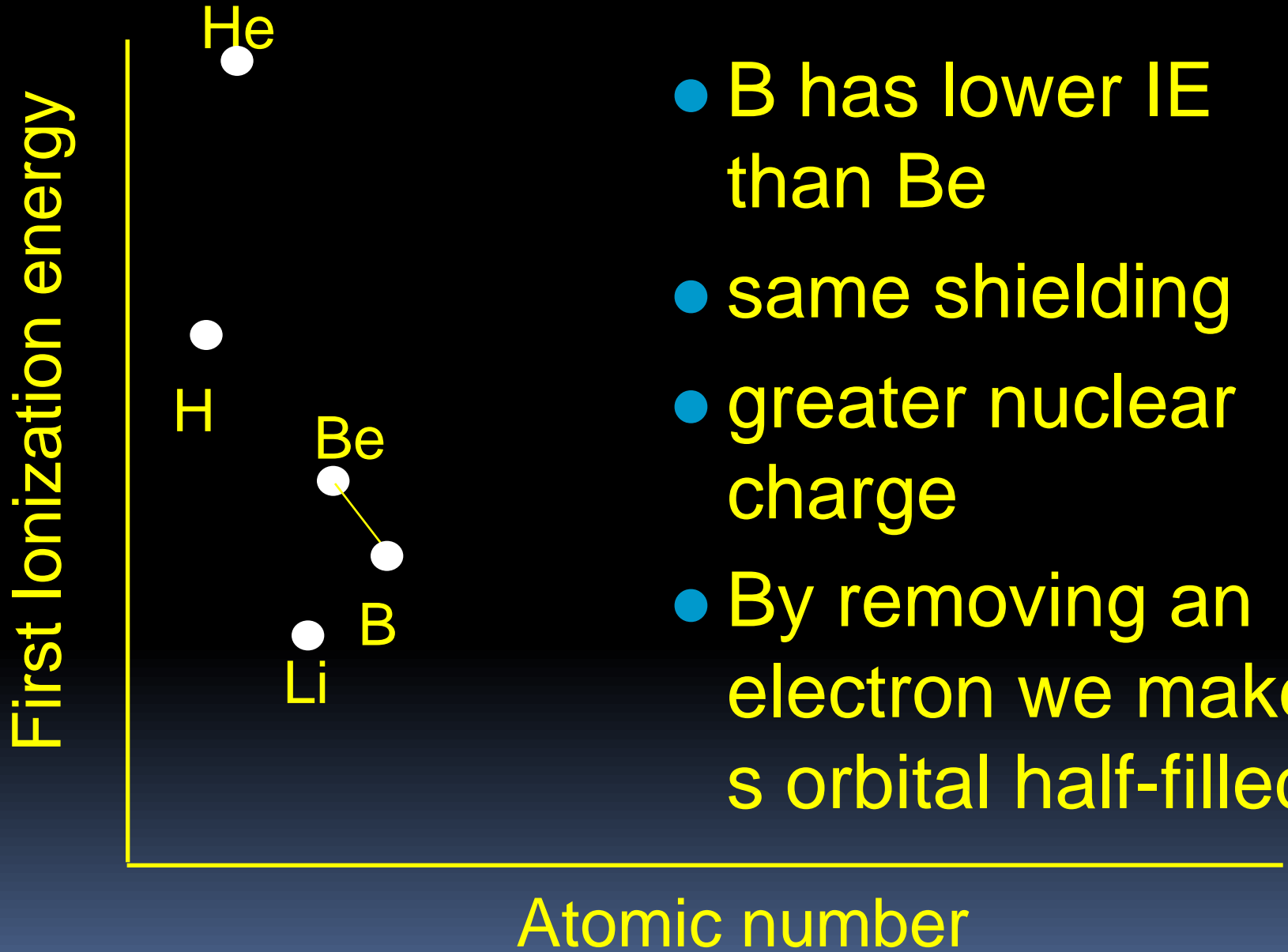
First ionization energy



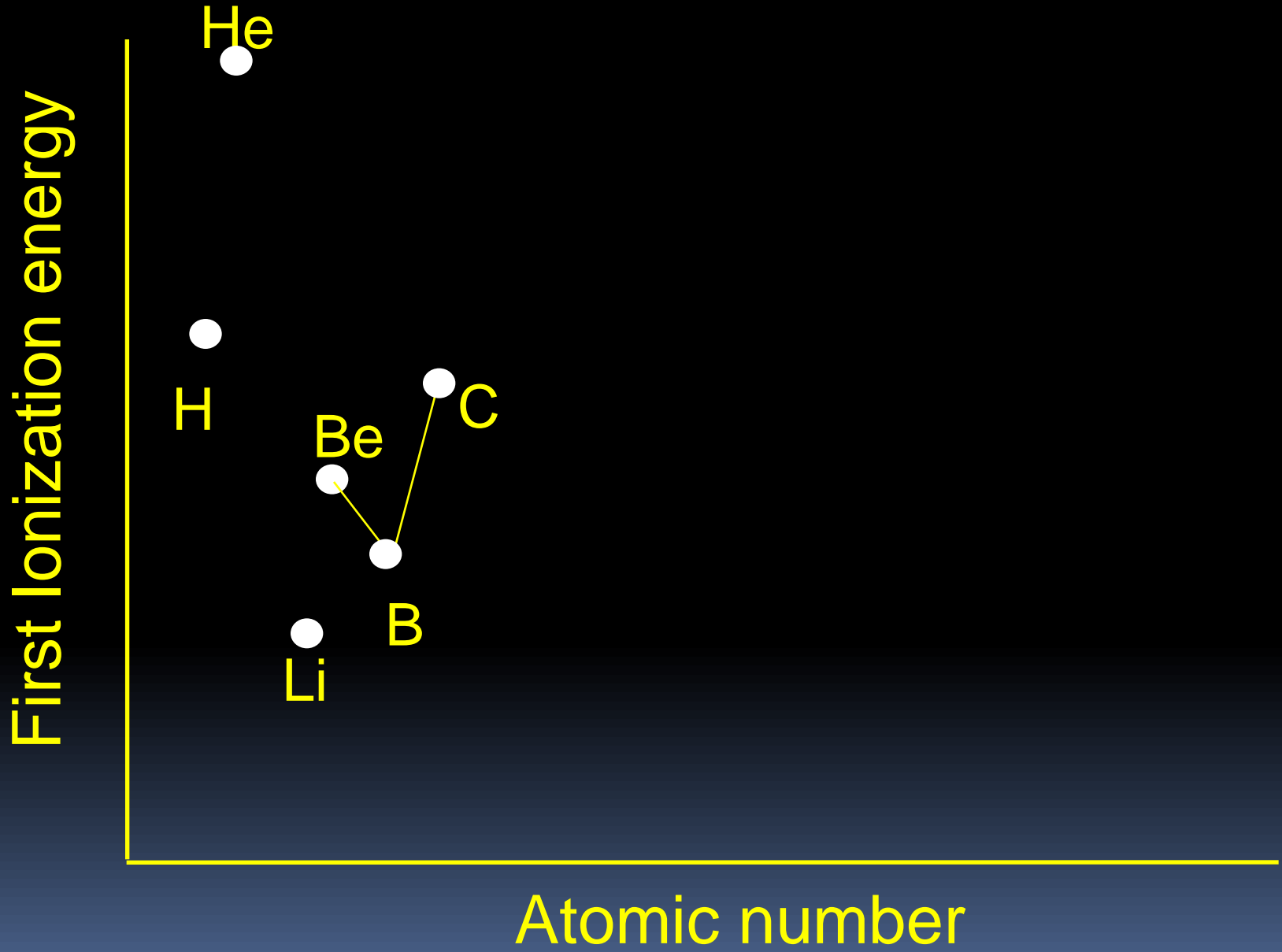
- He has a greater IE than H.
- same shielding
- greater nuclear charge



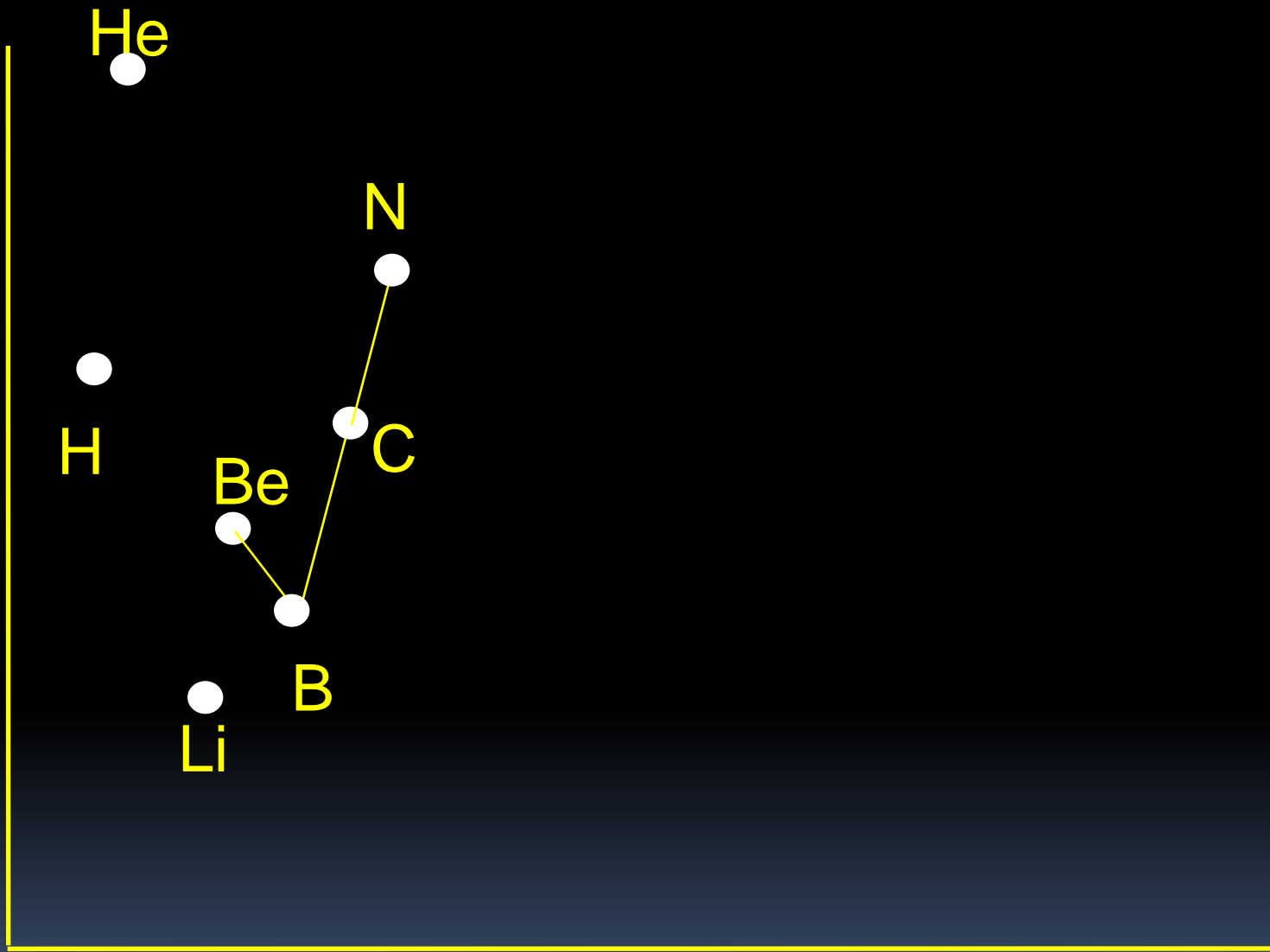




- B has lower IE than Be
- same shielding
- greater nuclear charge
- By removing an electron we make s orbital half-filled

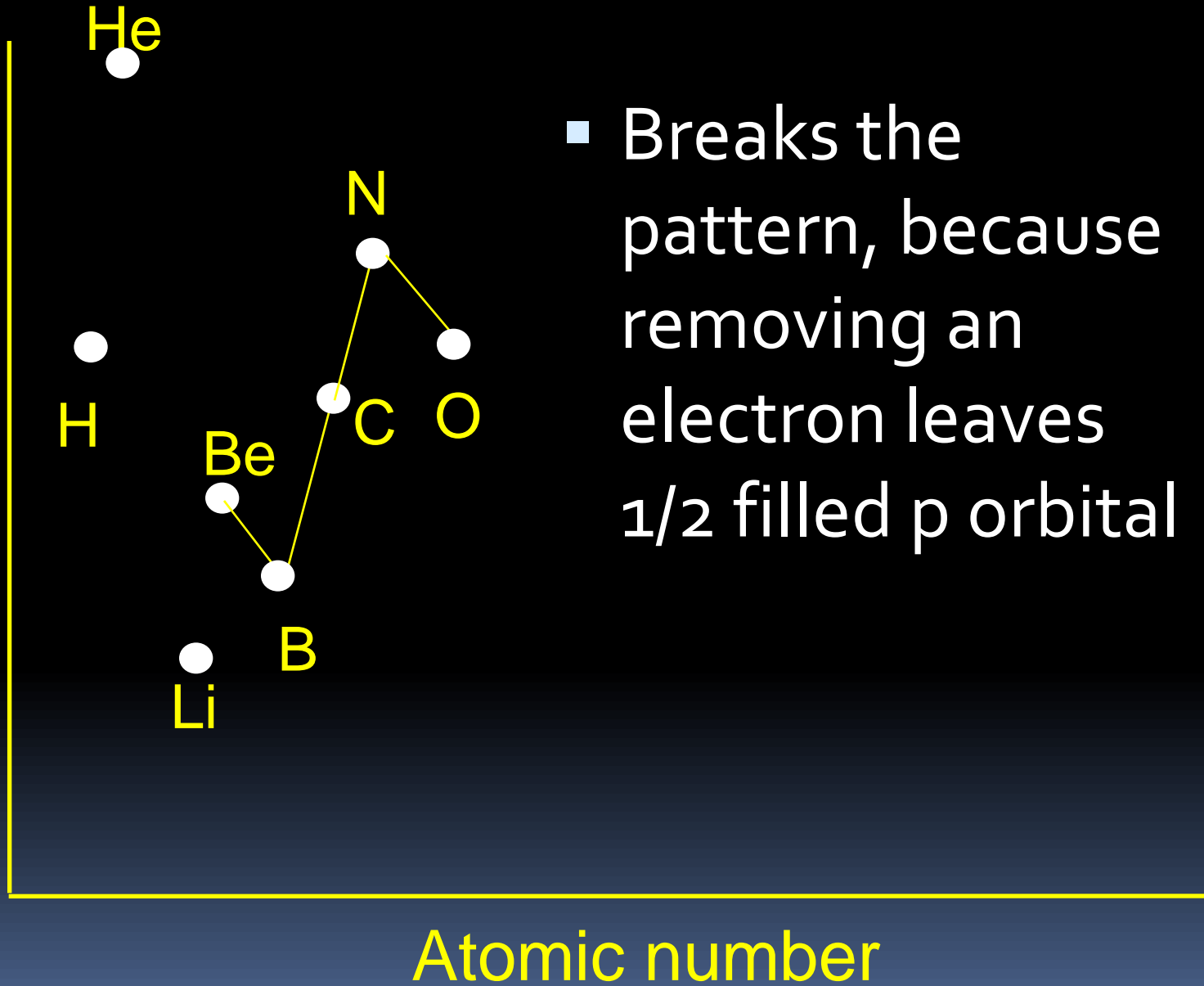


First ionization energy

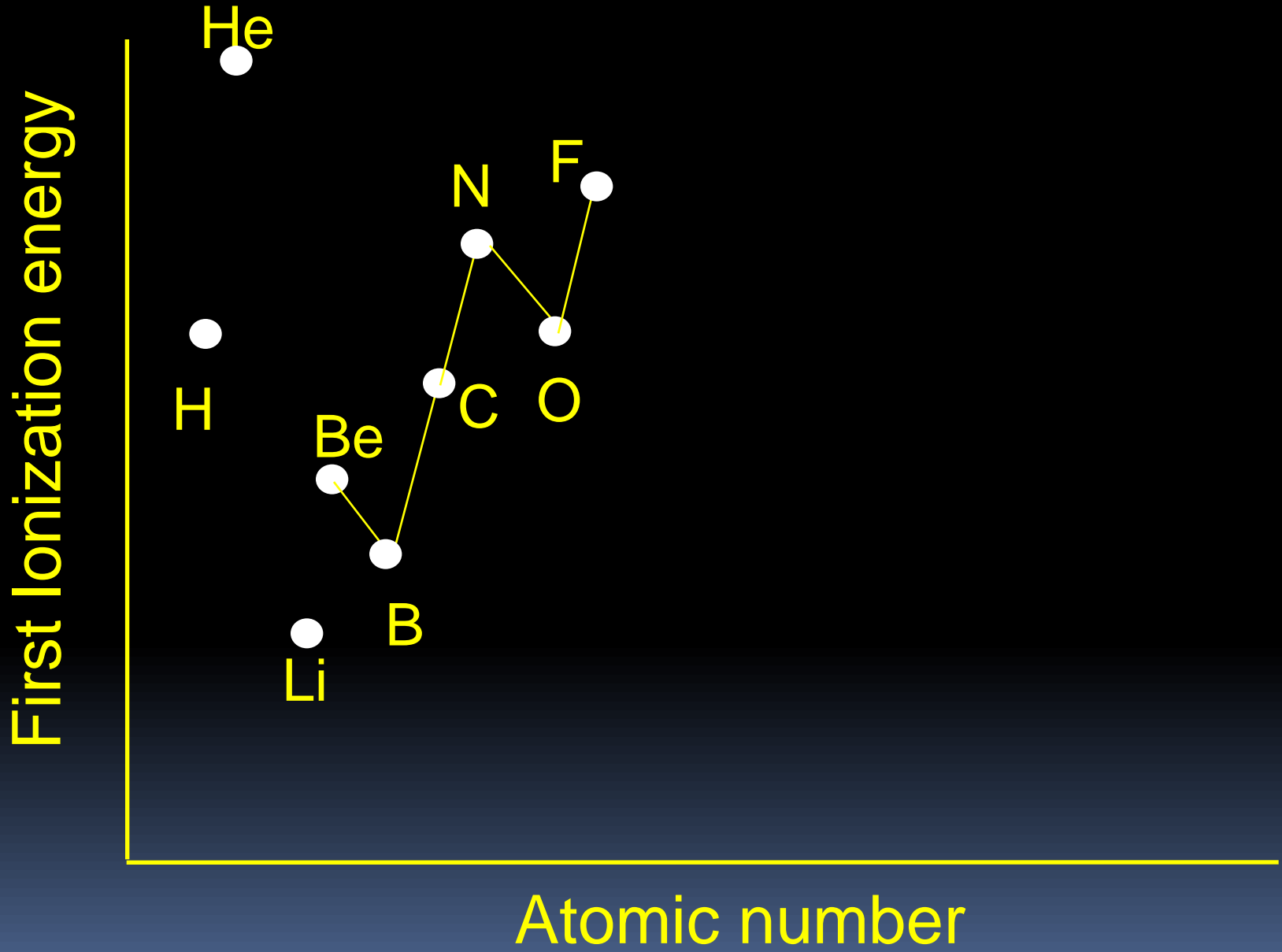


Atomic number

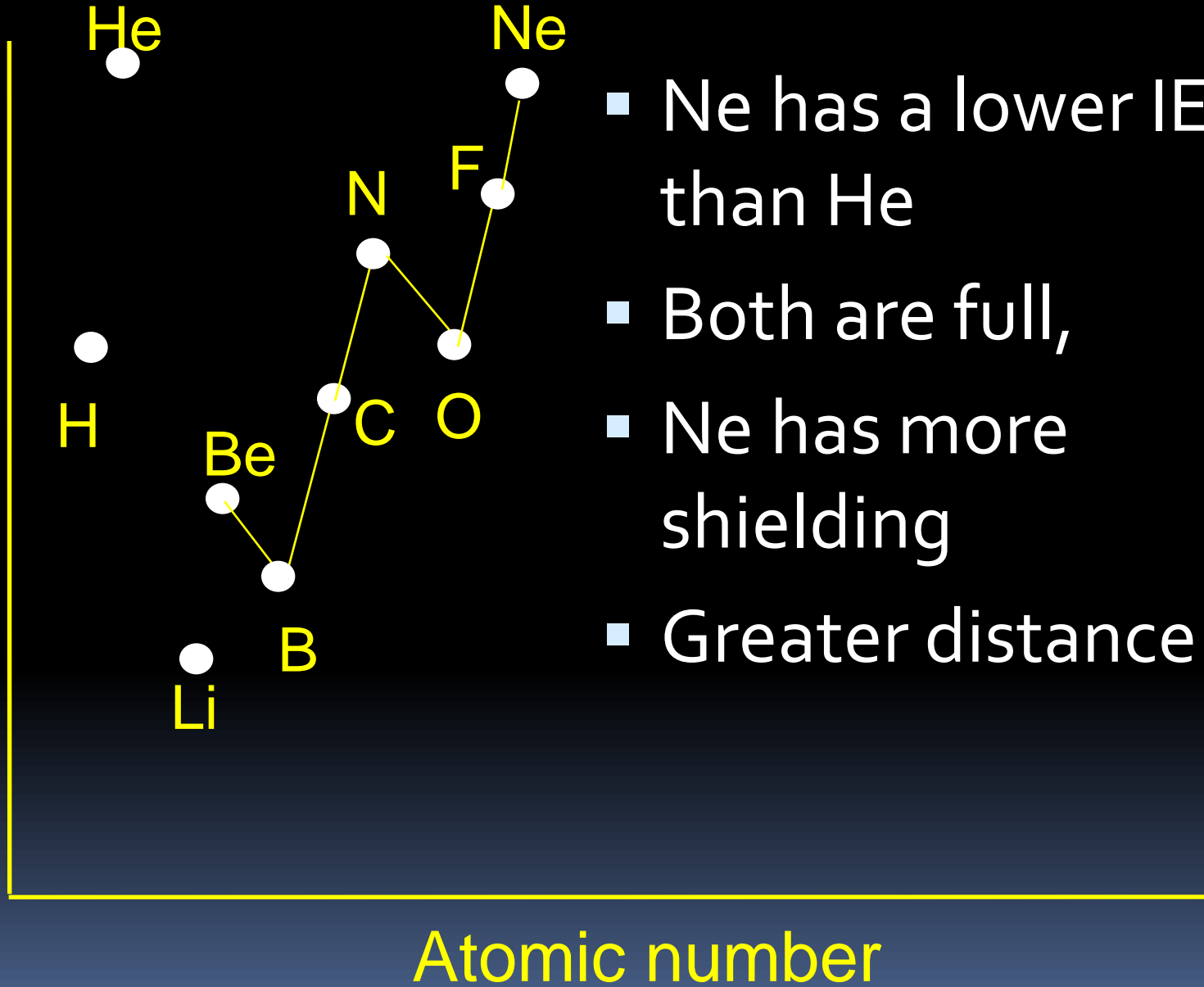
First ionization energy



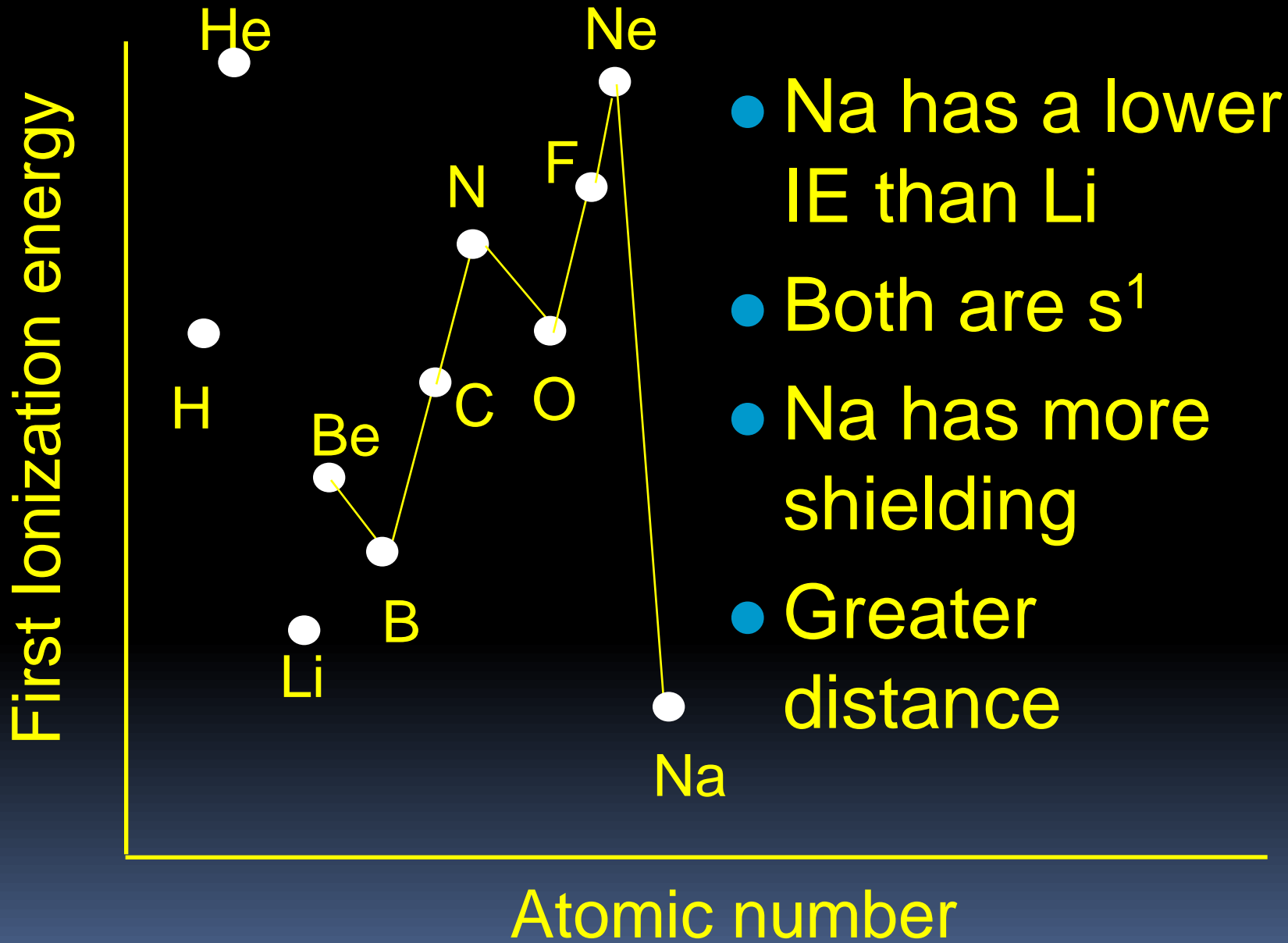
- Breaks the pattern, because removing an electron leaves $1/2$ filled p orbital



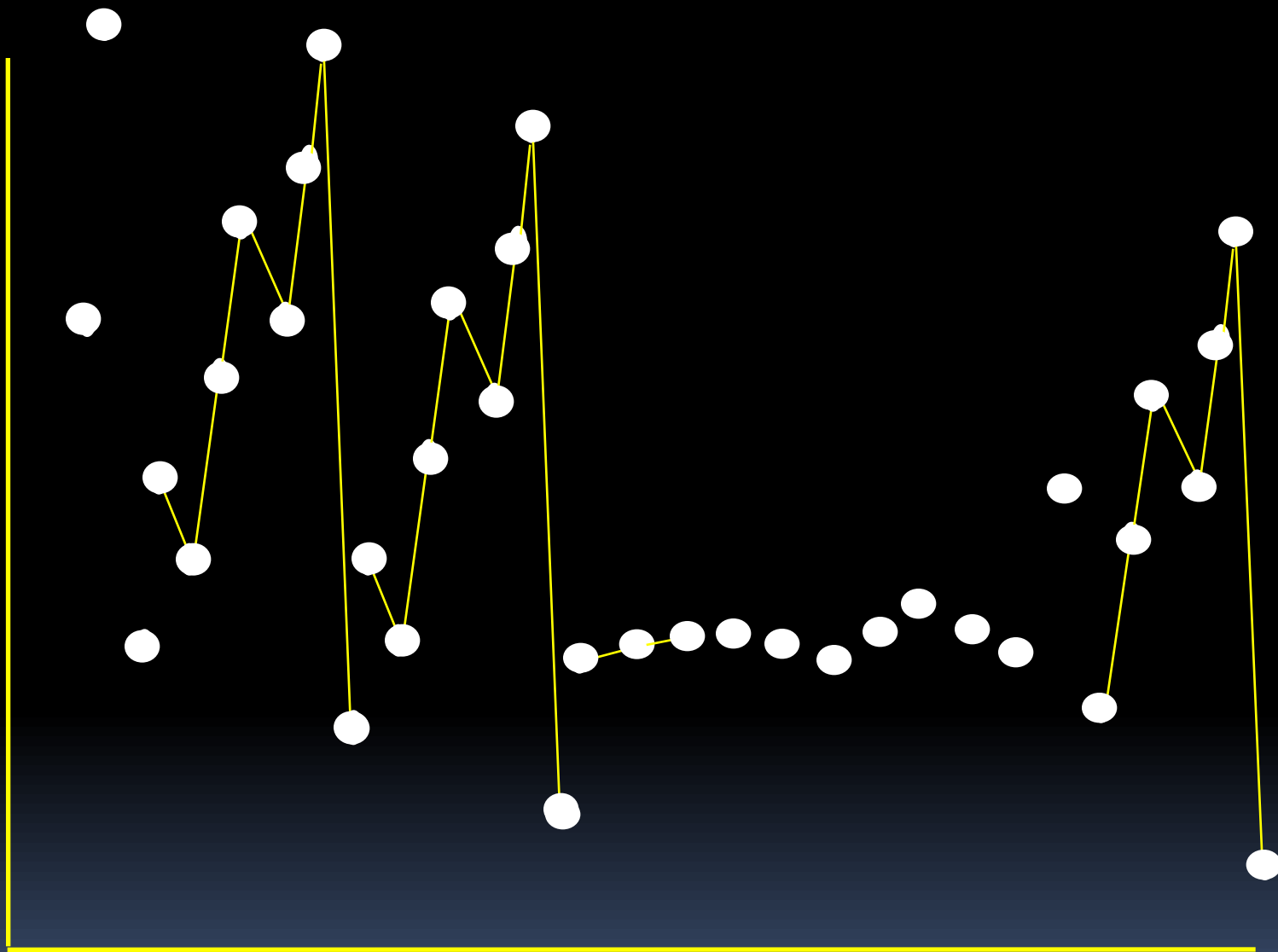
First ionization energy



- Ne has a lower IE than He
- Both are full,
- Ne has more shielding
- Greater distance




First ionization energy



Atomic number



Driving Force

- Full Energy Levels require lots of energy to remove their electrons.
 - Noble Gases have full orbitals.
 - Atoms behave in ways to achieve noble gas configuration.
- 

2nd Ionization Energy

- For elements that reach a filled or half-filled orbital by removing 2 electrons, 2nd IE is lower than expected.
- True for s^2
- Alkaline earth metals form 2+ ions.

3rd IE

- Using the same logic s^2p^1 atoms have a low 3rd IE.
- Atoms in the aluminum family form $3+$ ions.
- 2nd IE and 3rd IE are always higher than 1st IE!!!

Trends in Electron Affinity

- The energy change associated with adding an electron to a gaseous atom.
- Easiest to add to group 7A.
- Gets them to full energy level.
- Increase from left to right: atoms become smaller, with greater nuclear charge.
- Decrease as we go down a group.

Trends in Ionic Size

- Cations form 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 form 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

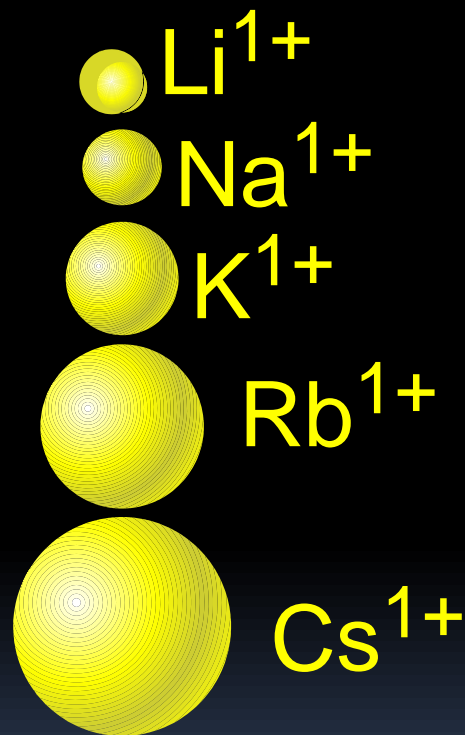
- Ions always have noble gas configuration.
- Na is: $1s^2 2s^2 2p^6 3s^1$
- Forms a 1+ ion: $1s^2 2s^2 2p^6$
- Same configuration as neon.
- Metals form ions with the configuration of the noble gas before them - they lose electrons.

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.

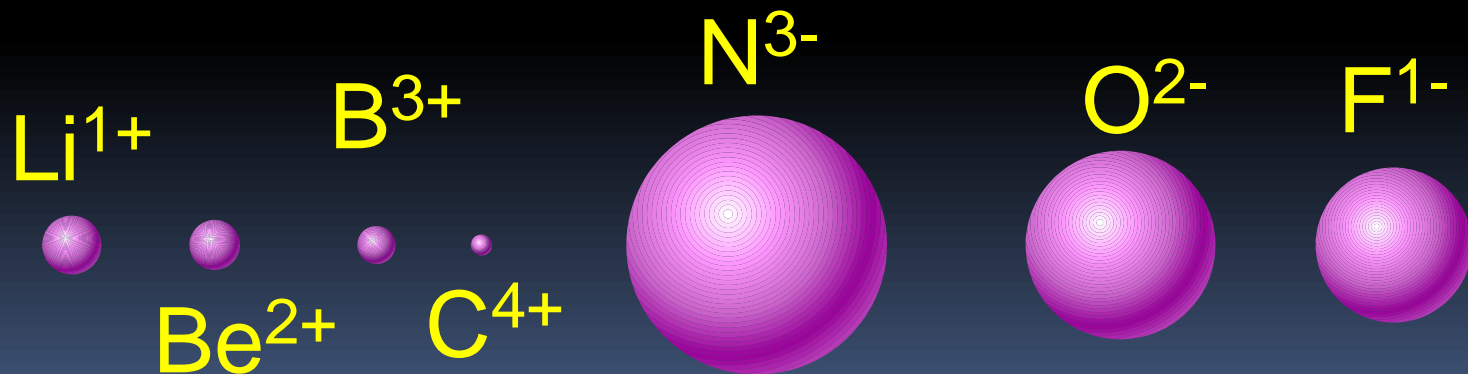
Group trends

- 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.

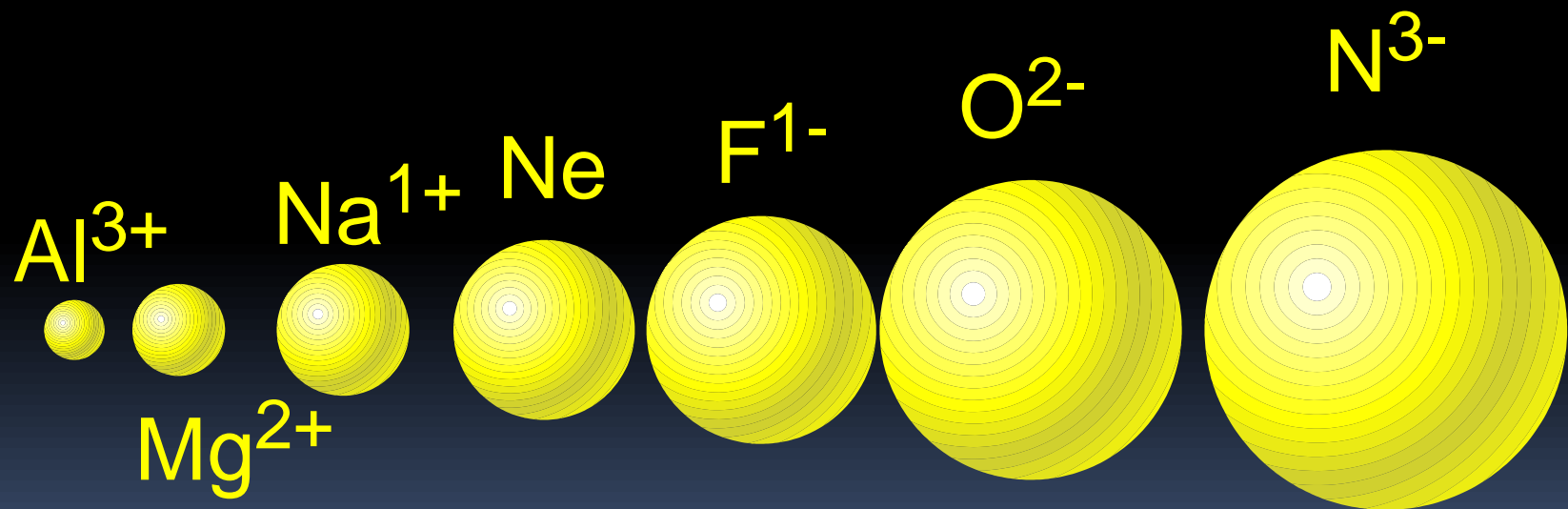


Size of Isoelectronic ions

- Iso- means the 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 configuration: $1s^2 2s^2 2p^6$

Size of Isoelectronic ions


- Positive ions that have more protons would be smaller.






Electronegativity

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



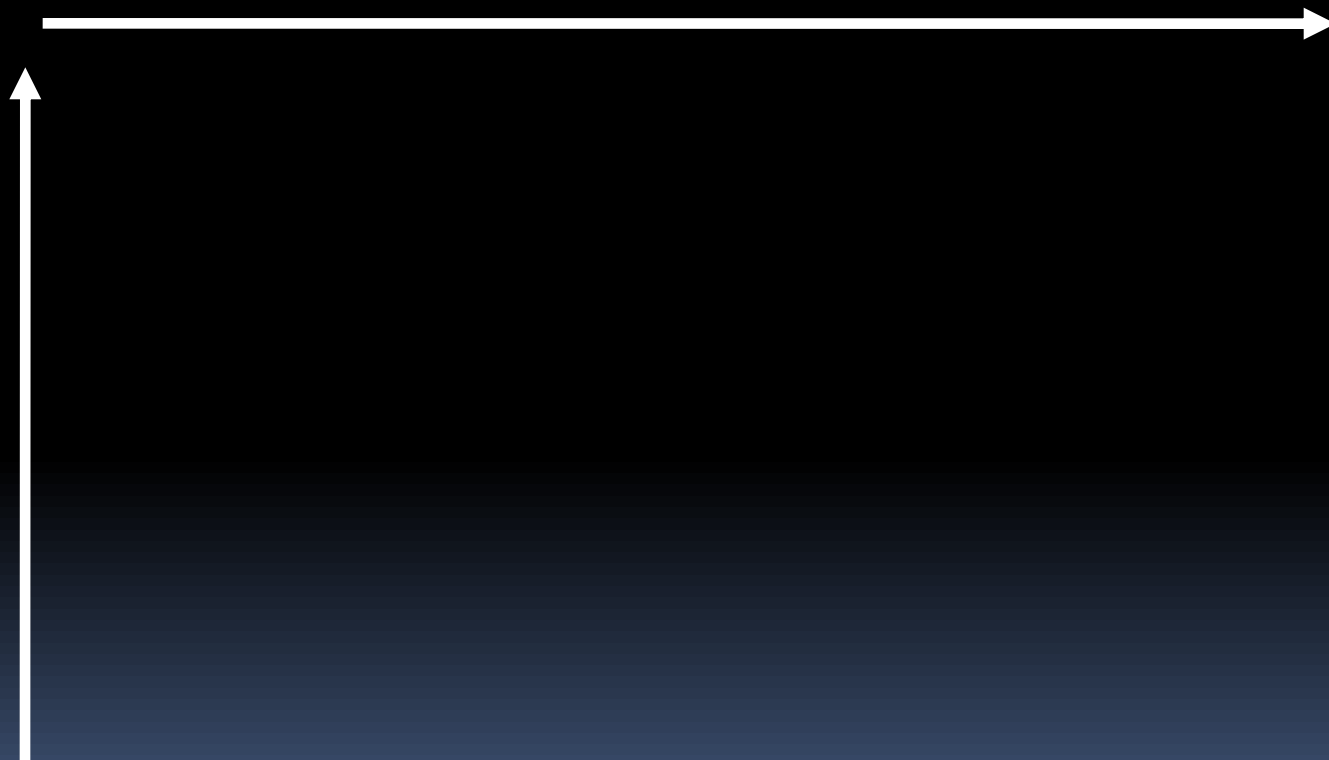
Group Trend

- The further down a group, the farther the electron is away, and the more electrons an atom has.
 - More willing to share.
 - Low electronegativity.
- 

Periodic Trend

- Metals are at the left of the table.
- They let their electrons go easily
- Low electronegativity
- At the right end are the nonmetals.
- They want more electrons.
- Try to take them away from others
- High electronegativity.

Ionization energy, Electronegativity,
and Electron Affinity INCREASE



Atomic size increases,
shielding constant



Ionic size increases