The Periodic Table

The how and why

History

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

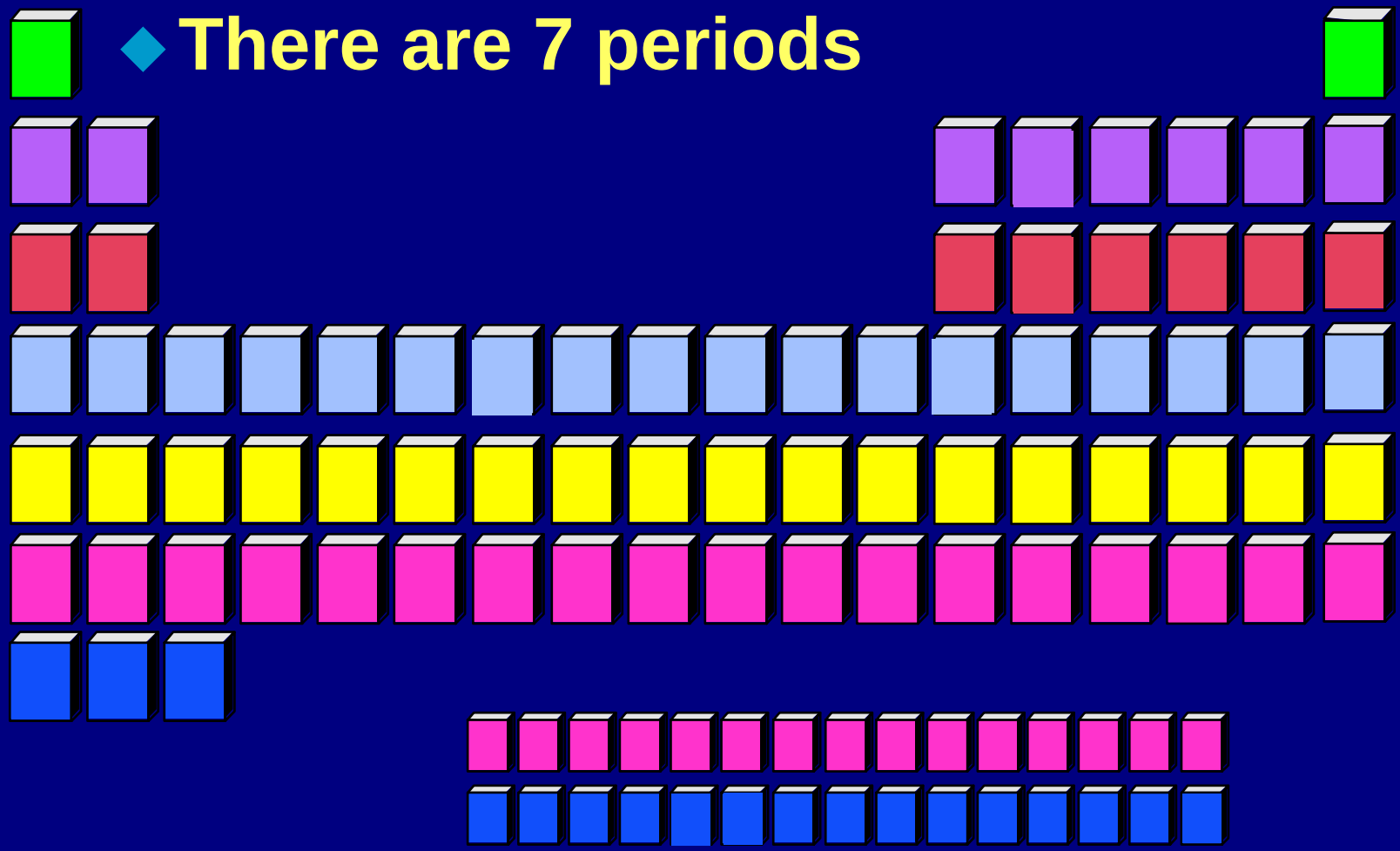
Mendeleev's Table

- **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.**
- **Found some gaps.**
- **Must be undiscovered elements.**
- **Predicted their properties before they were found.**

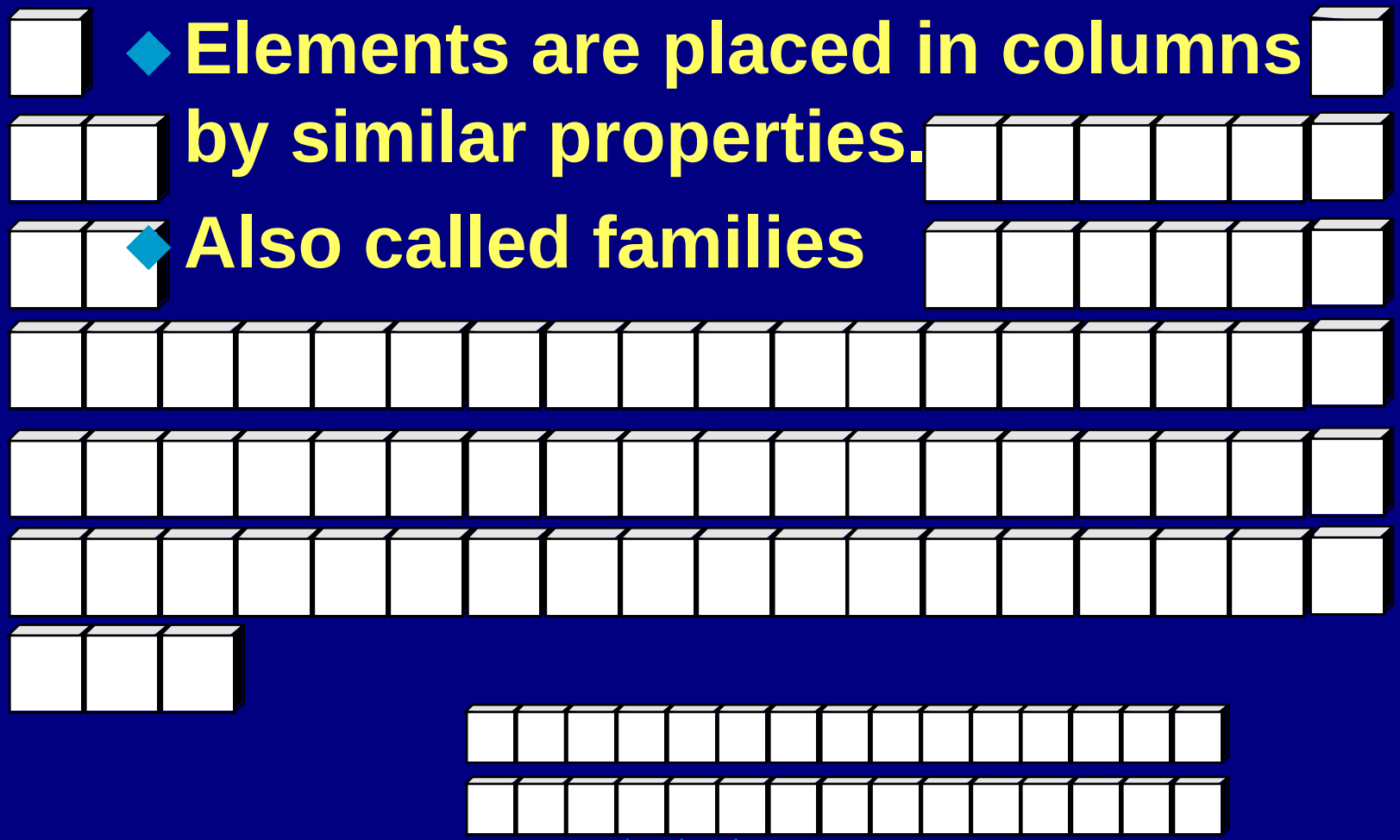
The modern table

- **Elements are still grouped by properties.**
- **Similar properties are in the same column.**
- **Order is in 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.**

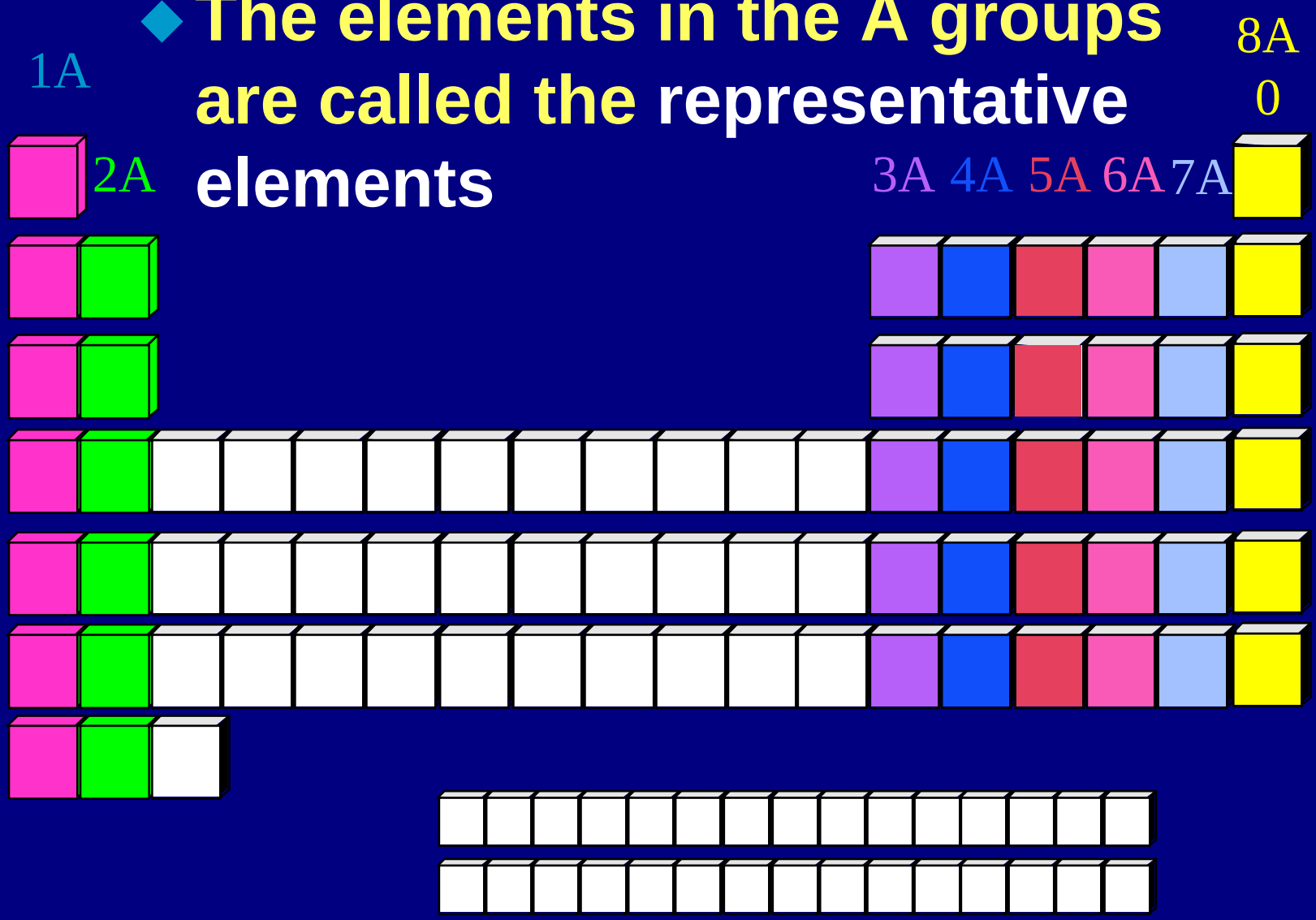
- Horizontal rows are called periods
- There are 7 periods



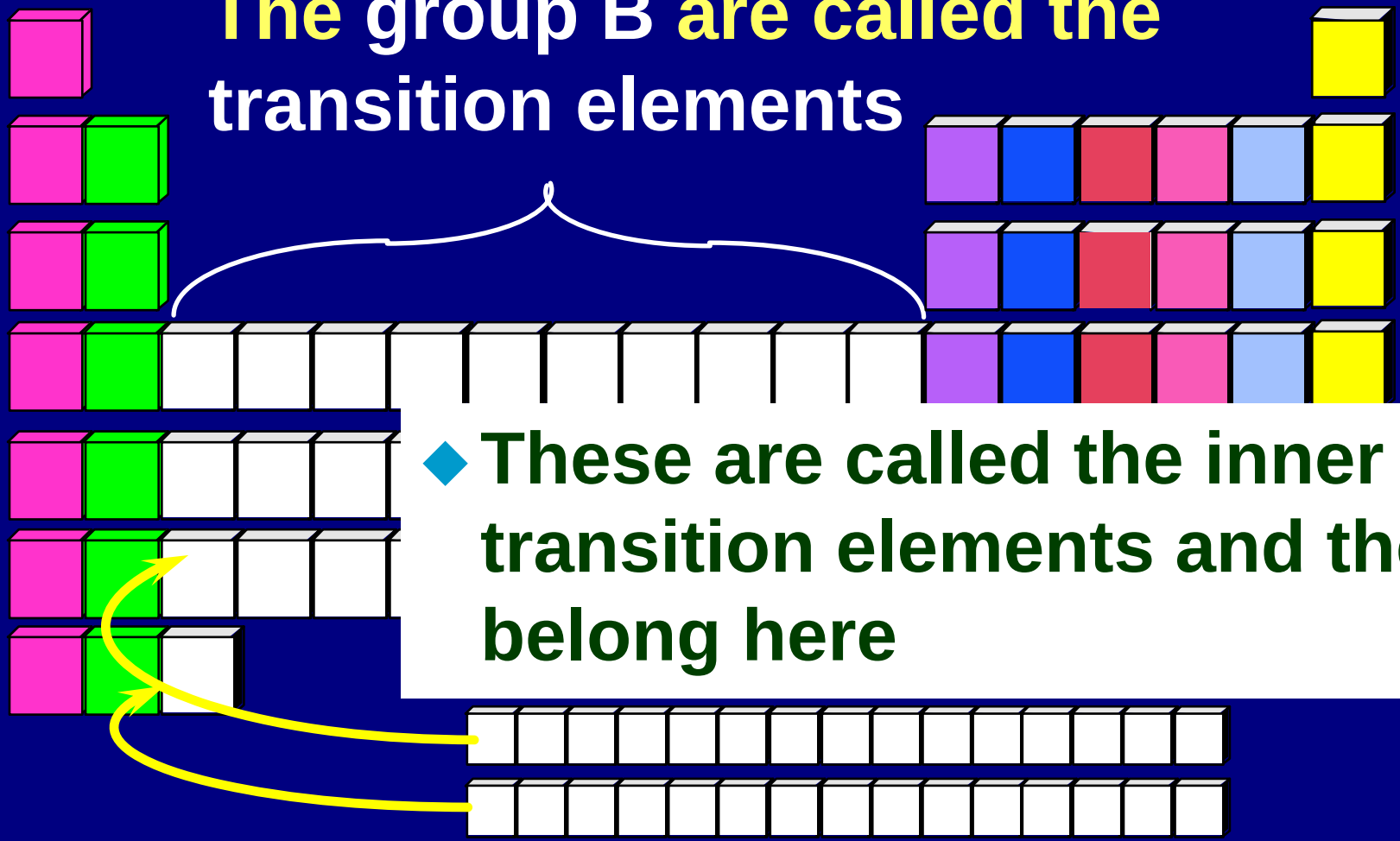
- Vertical columns are called groups.
- Elements are placed in columns by similar properties.
- Also called families



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

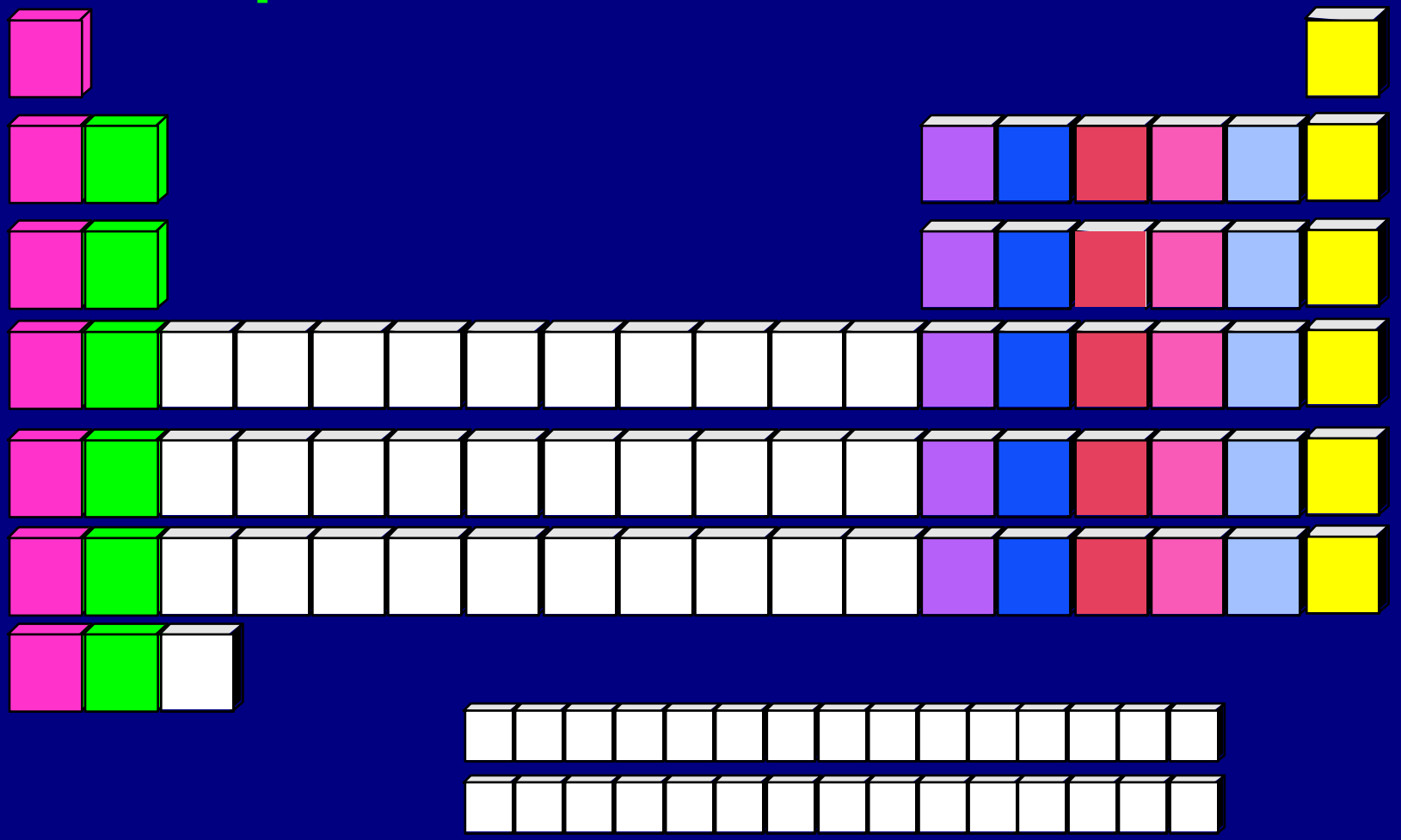


The group B are called the transition elements



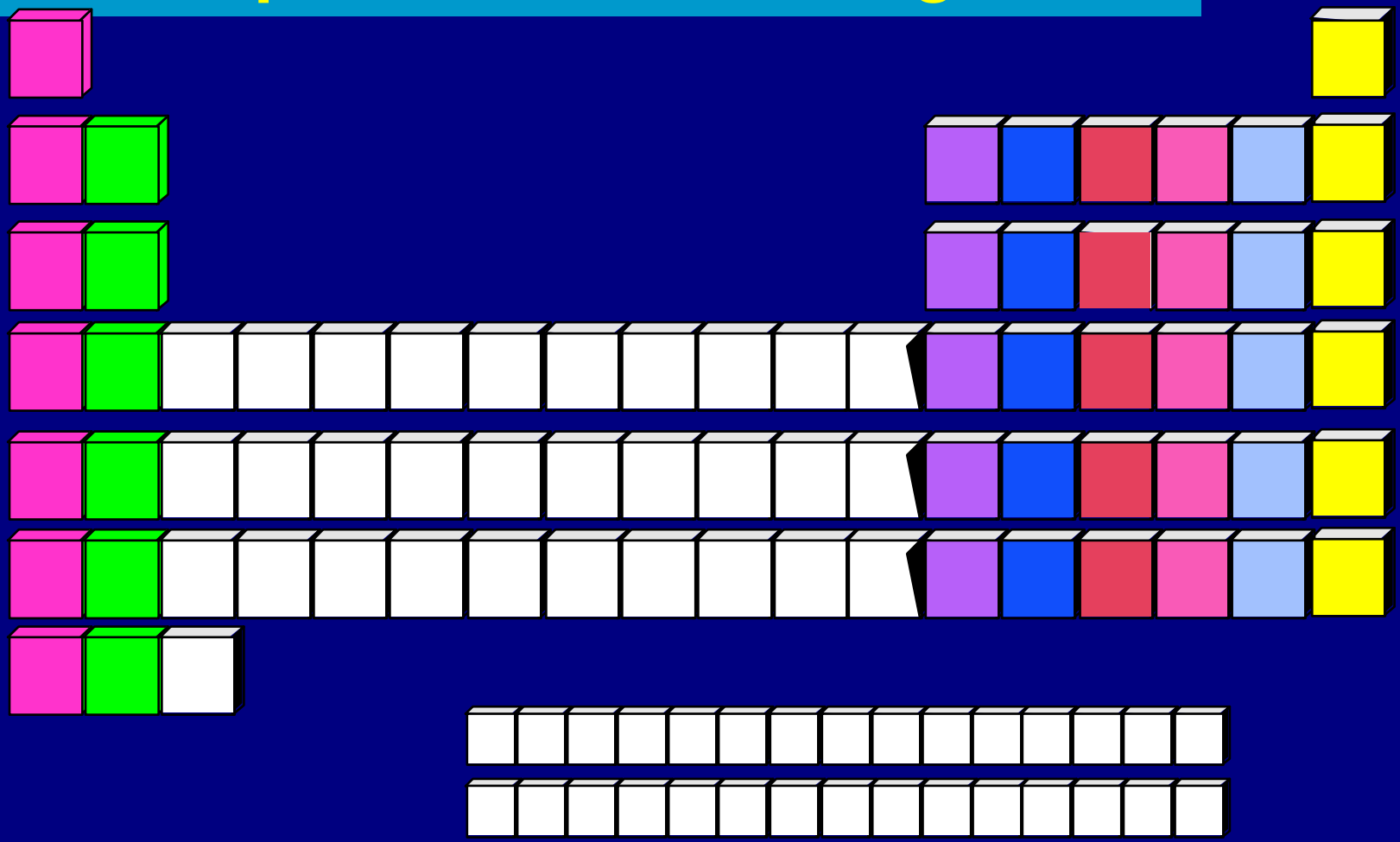
- These are called the inner transition elements and they belong here

- Group 1A are the alkali metals
- Group 2A are the alkaline earth metals



- Group 7A is called the Halogens

Group 8A are the noble gases



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

$1s^2$ He 2

$1s^2 2s^2 2p^6$ Ne 10

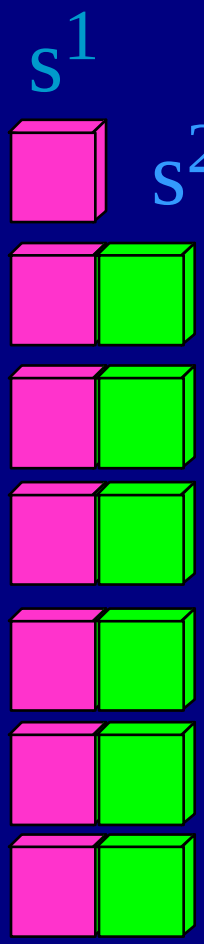
$1s^2 2s^2 2p^6 3s^2 3p^6$ Ar 18

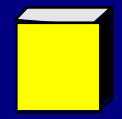
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$ Kr 36

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$ Xe 54

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6$ Rn 86

S- block



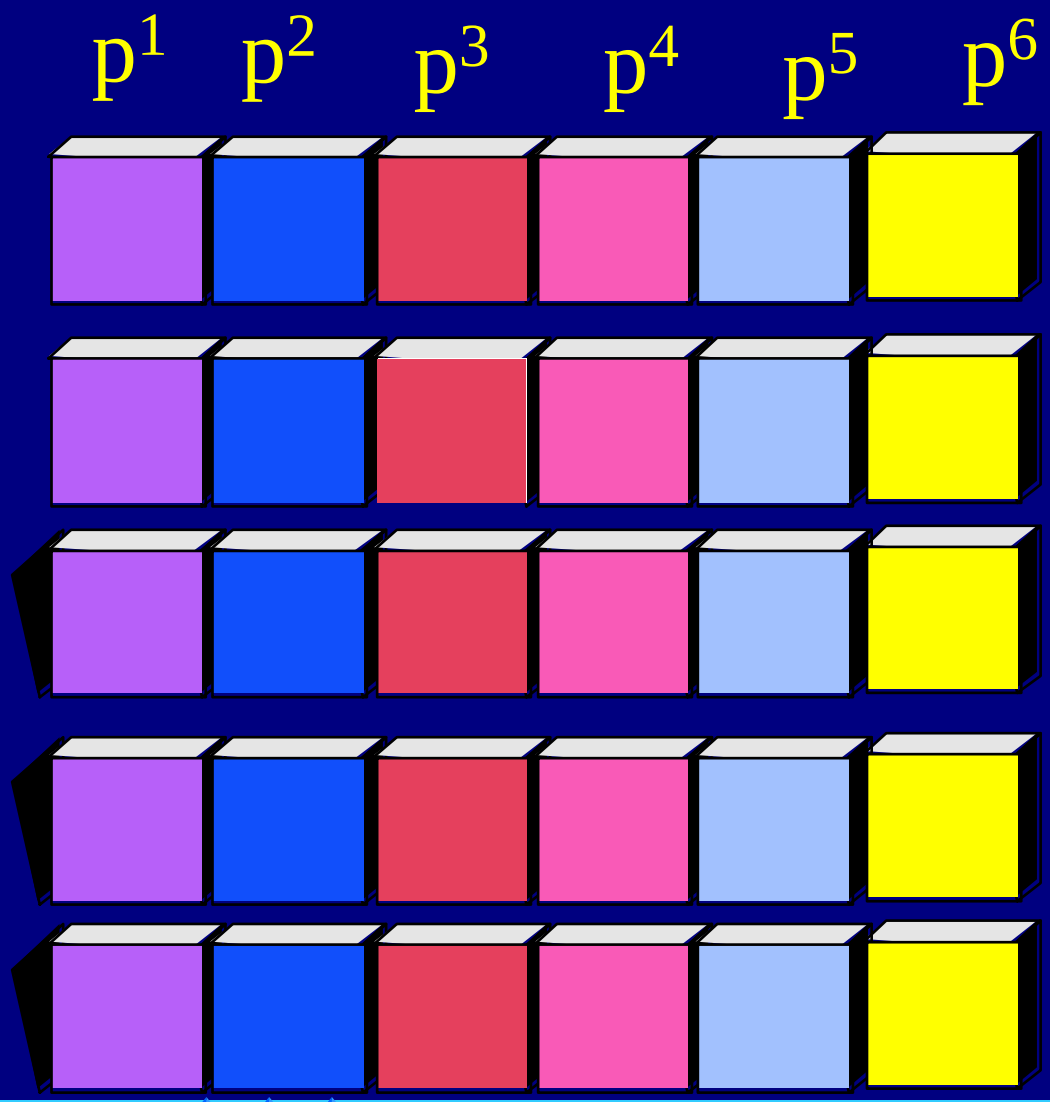


- **Alkali metals all end in s^1**
- **Alkaline earth metals all end in s^2**
- **really have to include He but it fits better later.**
- **He has the properties of the noble gases.**

Transition Metals -d block

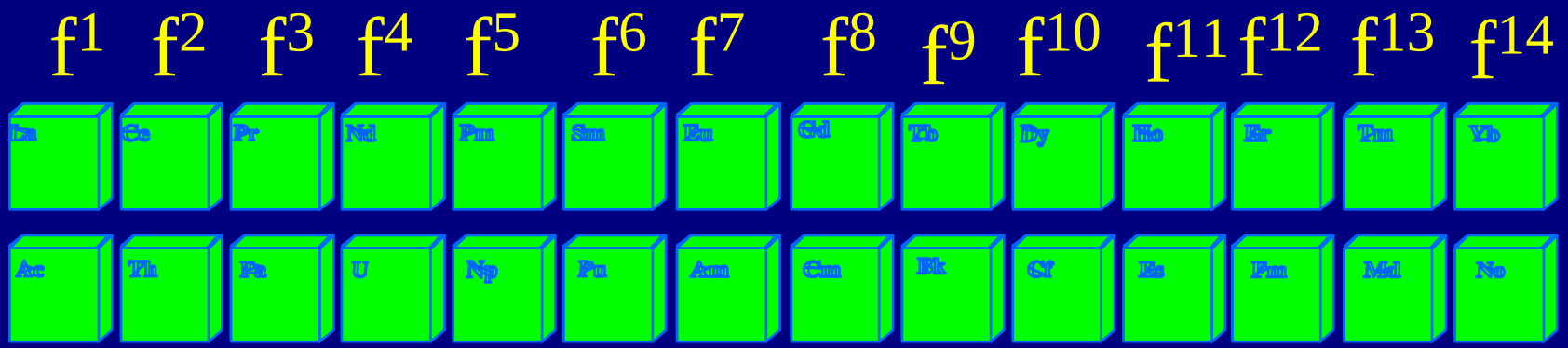
			s^1					s^1	
d^1	d^2	d^3	d^5	d^5	d^6	d^7	d^8	d^{10}	d^{10}
Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30
Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48
La 57	Hf 72	Ta 73	W	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80
Ac 89									

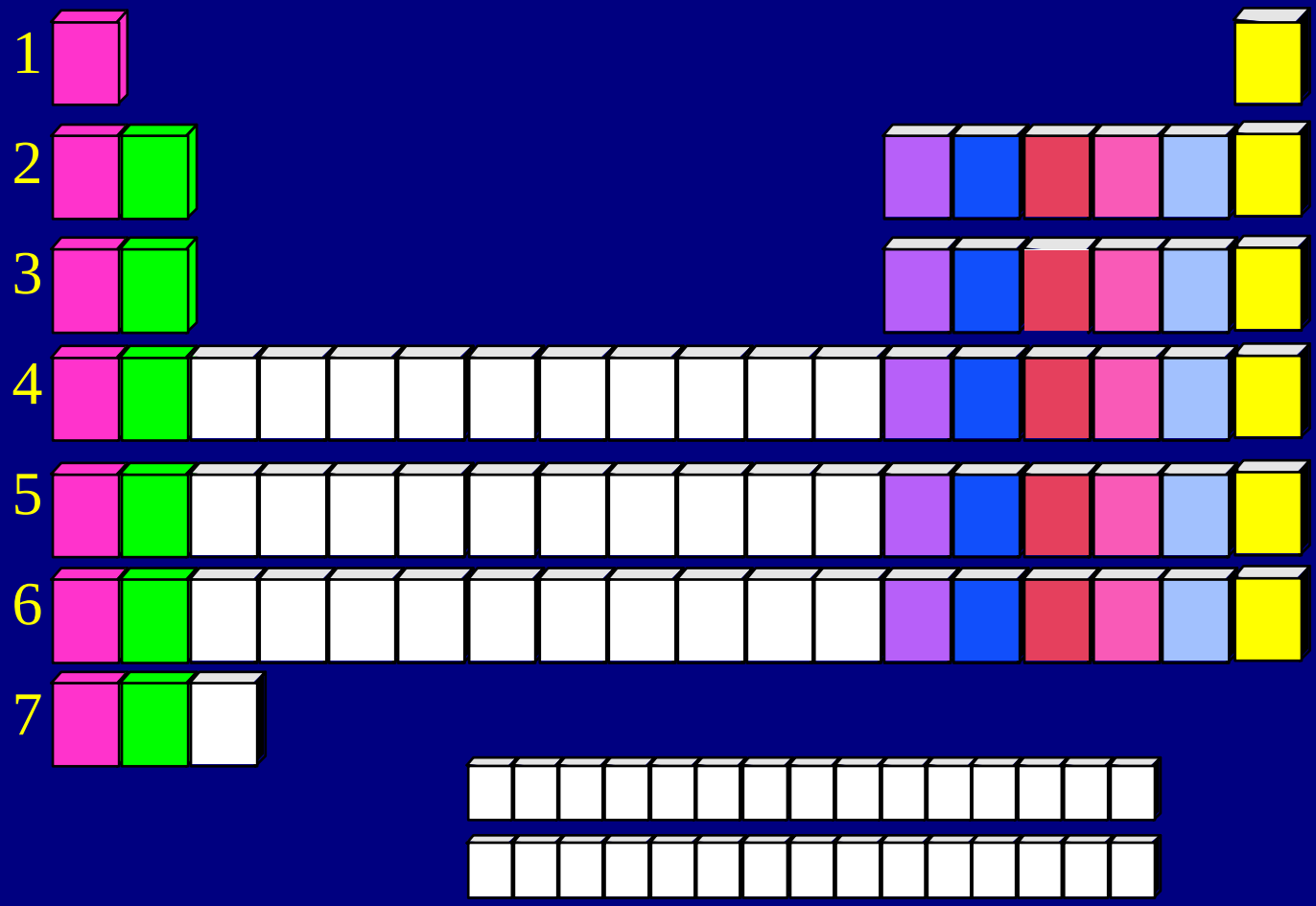
The P-block



F - block

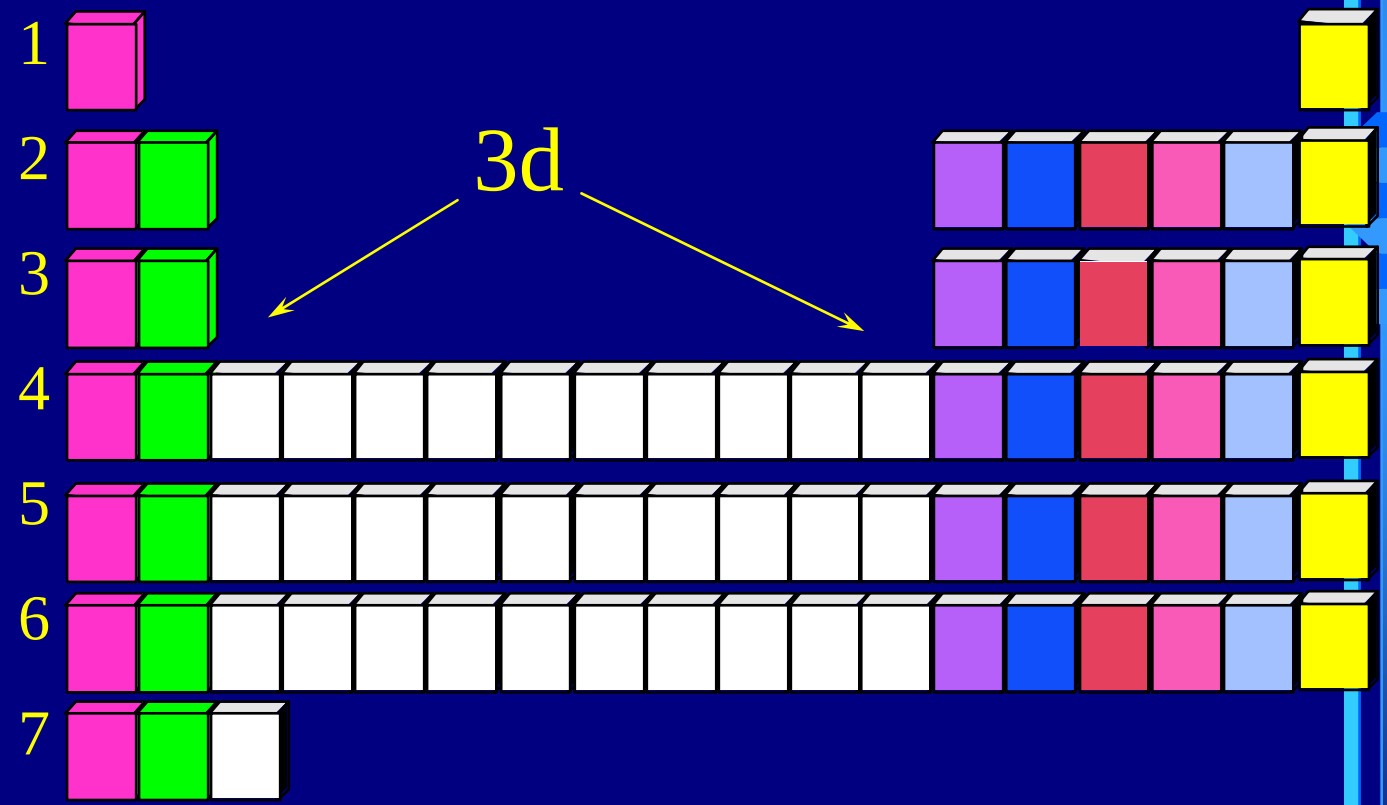
- **inner transition elements**

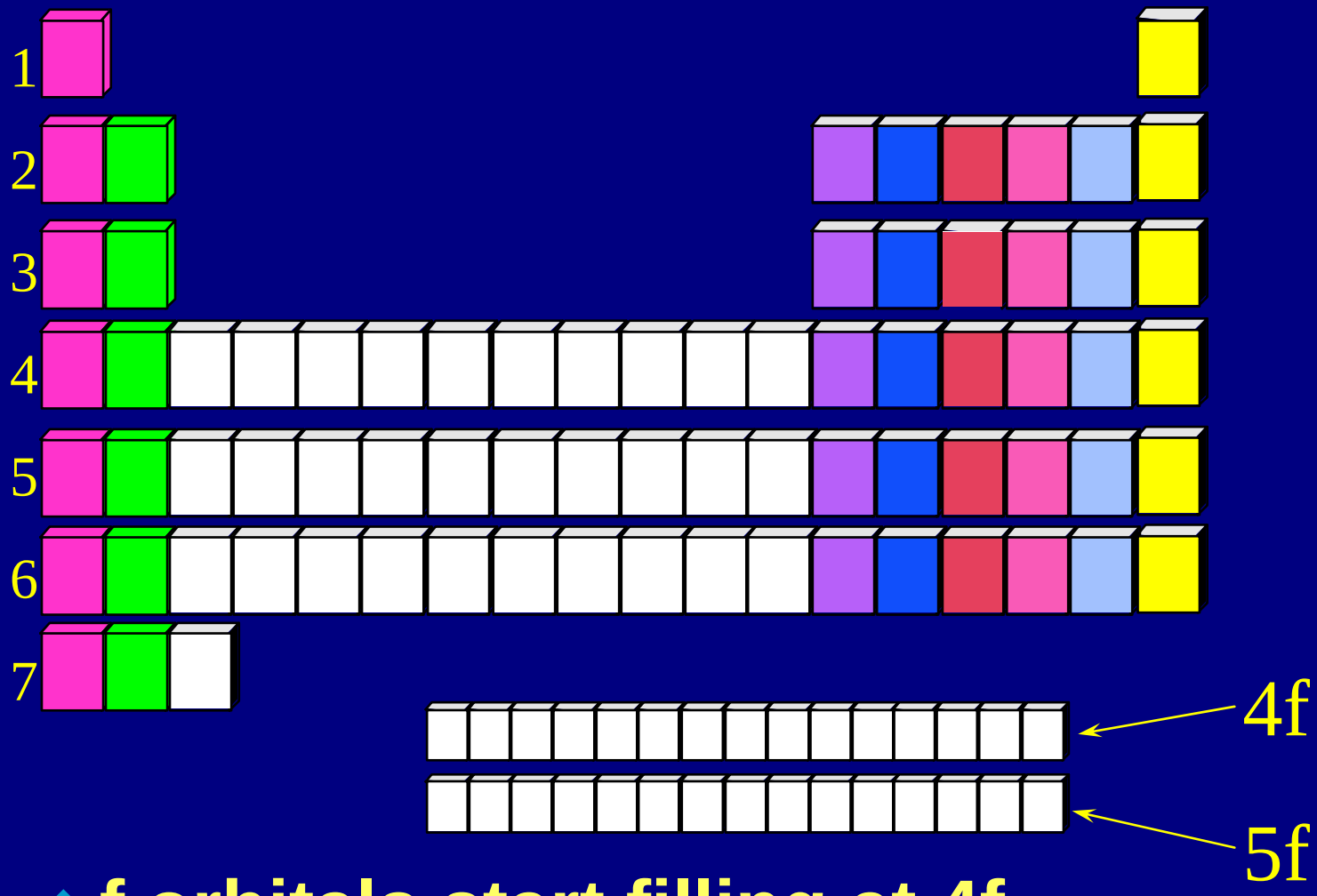




- **Each row (or period) is the energy level for s and p orbitals.**

- **D orbitals fill up after previous energy level so first d is 3d even though it's in row 4.**



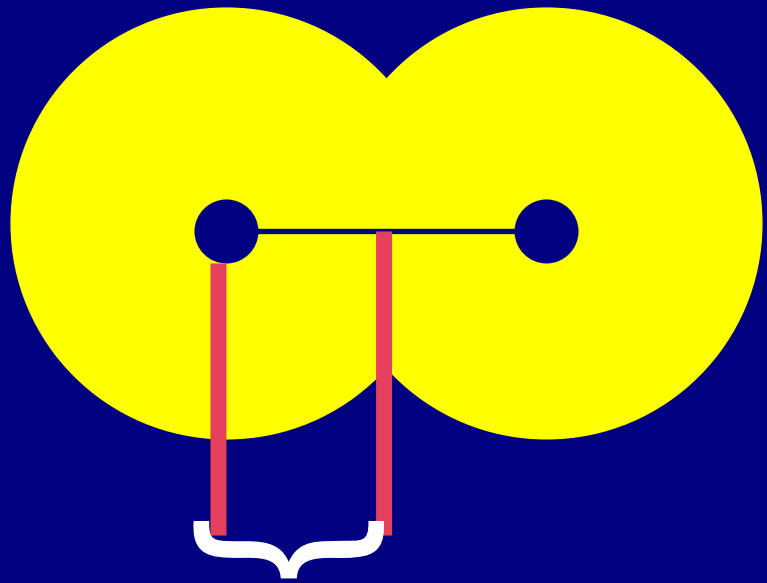


- **f orbitals start filling at 4f**

Atomic Size

- **First problem where do you start measuring.**
- **The electron cloud doesn’t have a definite edge.**
- **They get around this by measuring more than 1 atom at a time.**

Atomic Size



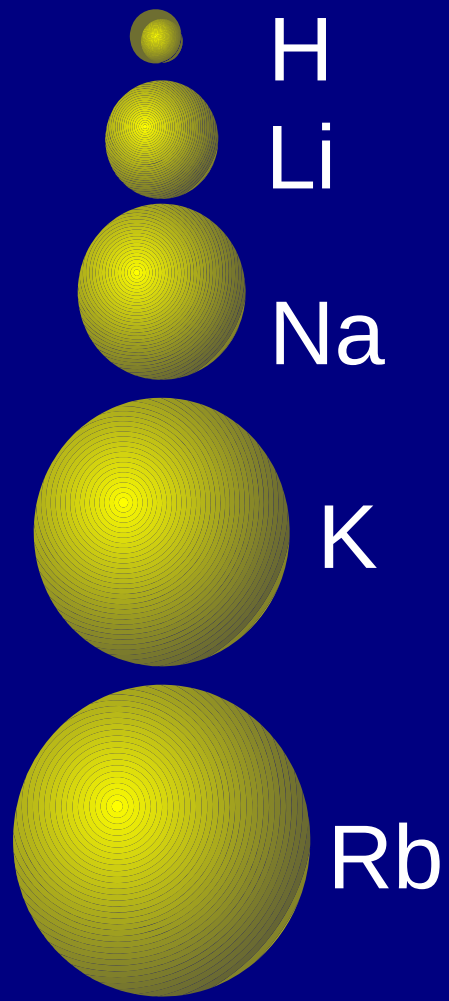
- **Atomic Radius = half the distance between two nuclei of a diatomic molecule.**

Trends in Atomic Size

- **Influenced by two factors.**
- **Energy Level**
- **Higher energy level is further away.**
- **Charge on nucleus**
- **More charge pulls electrons in closer.**

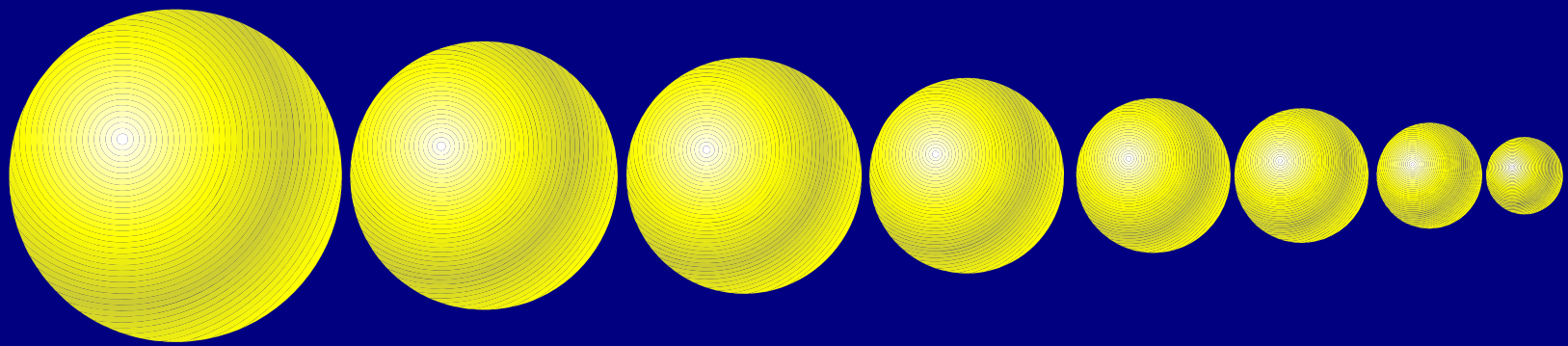
Group trends

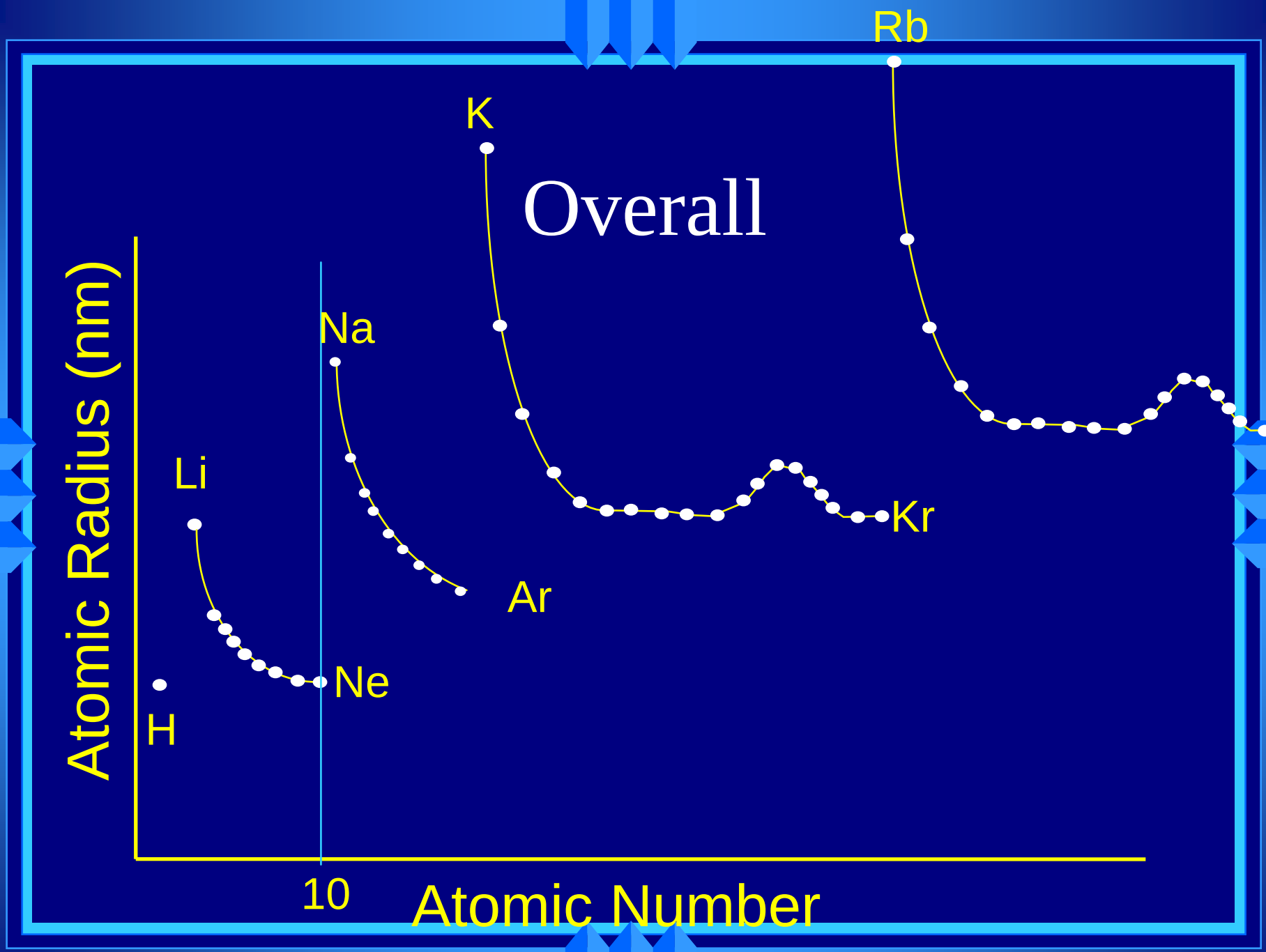
- **As we go down a group**
- **Each atom has another energy level,**
- **So the atoms get bigger.**



Periodic Trends

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




Overall
Atomic Radius (nm)
Atomic Number
H
Li
Ne
Na
Ar
K
Kr
Rb
10

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 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 of 2nd IE.**

Symbol	First	Second	Third
H	1312		
He	2731	5247	11
Li	520	7297	810
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

Symbol	First	Second	Third
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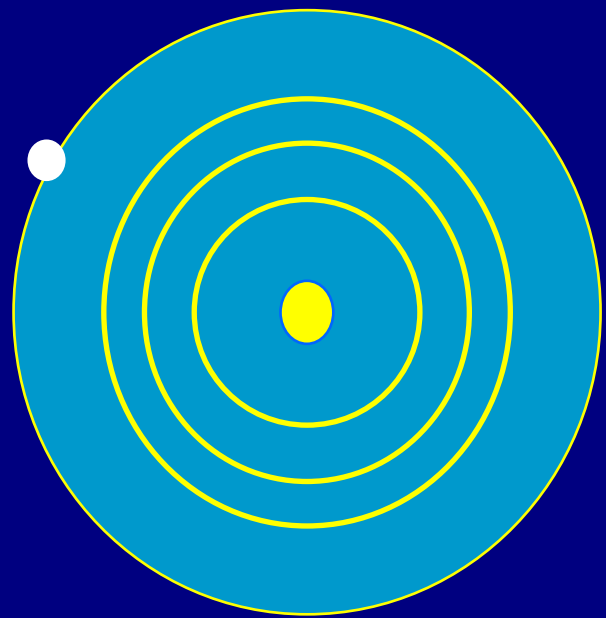
What determines IE

- **The greater the nuclear charge the greater IE.**
- **Distance form nucleus decreases IE**
- **Shielding**

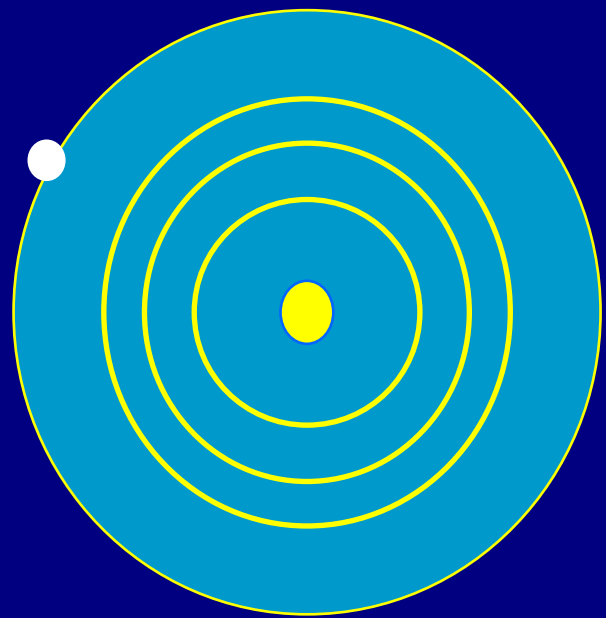
Shielding

- **The electron on the outside energy level has to look through all the other energy levels to see the nucleus**



Shielding

- **The electron on the outside energy level has to look through all the other energy levels to see the nucleus.**
- **A second electron has the same shielding.**

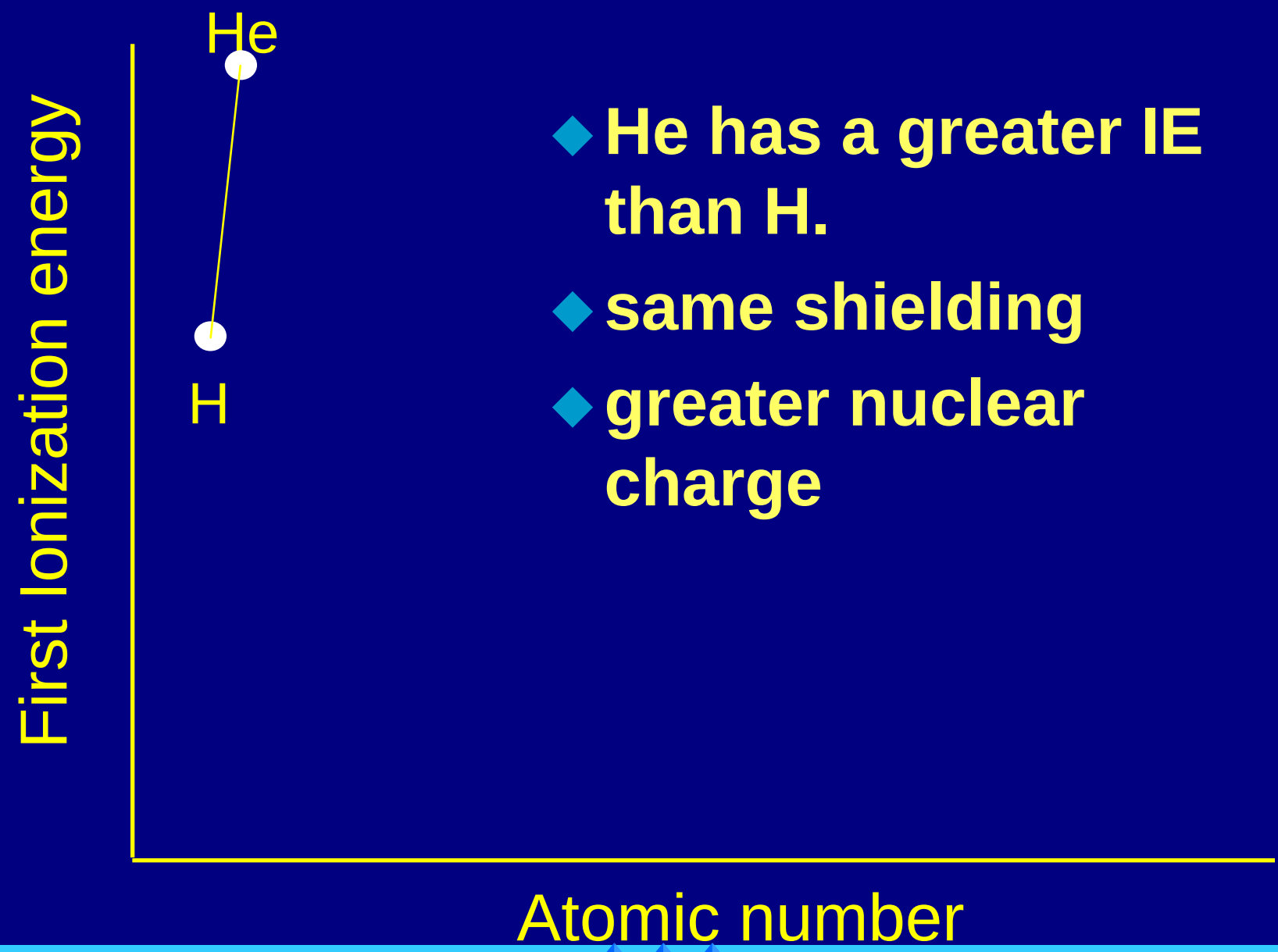


Group trends

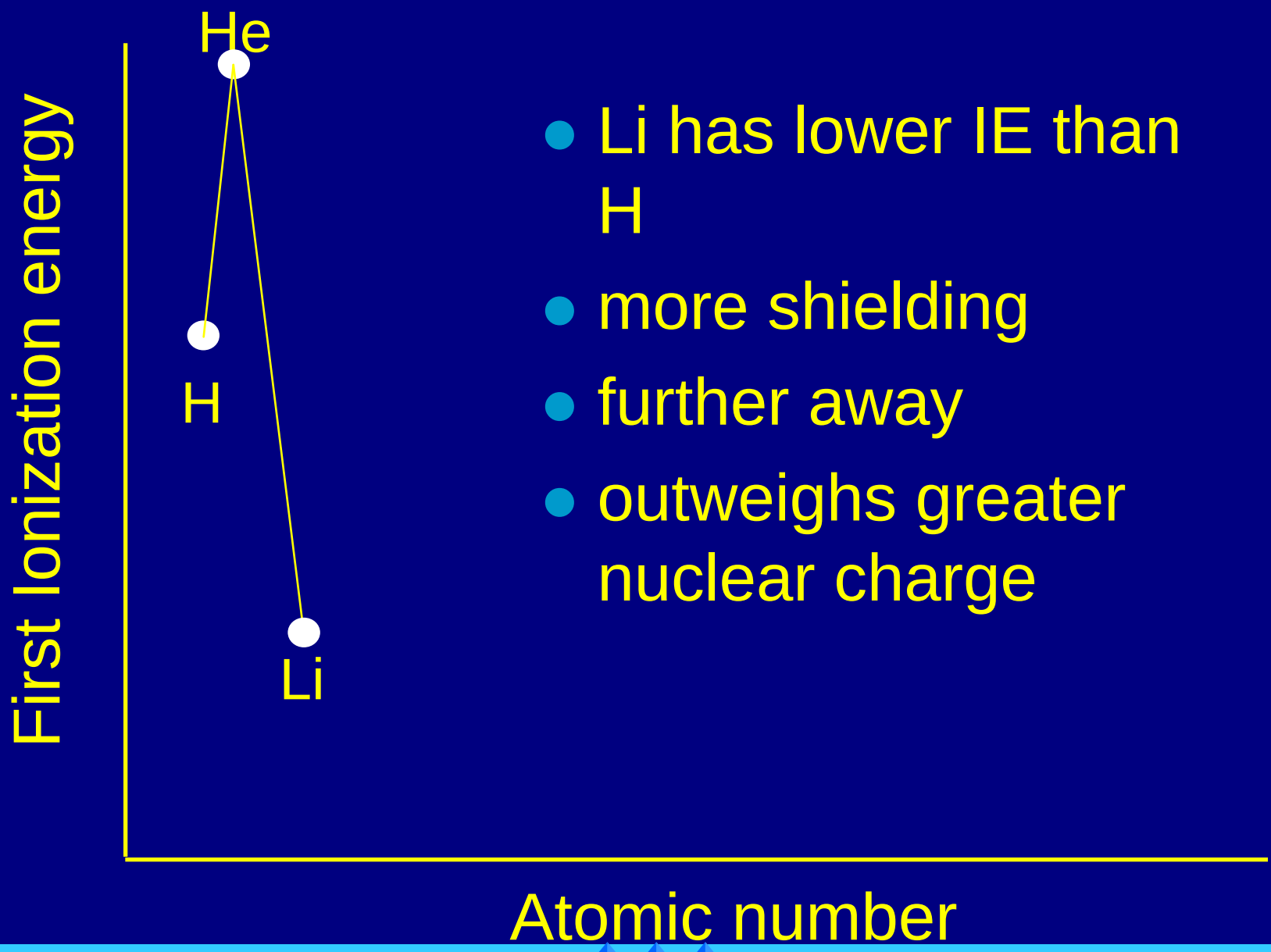
- **As you go down a group first IE decreases because**
- **The electron is further away.**
- **More shielding.**

Periodic trends

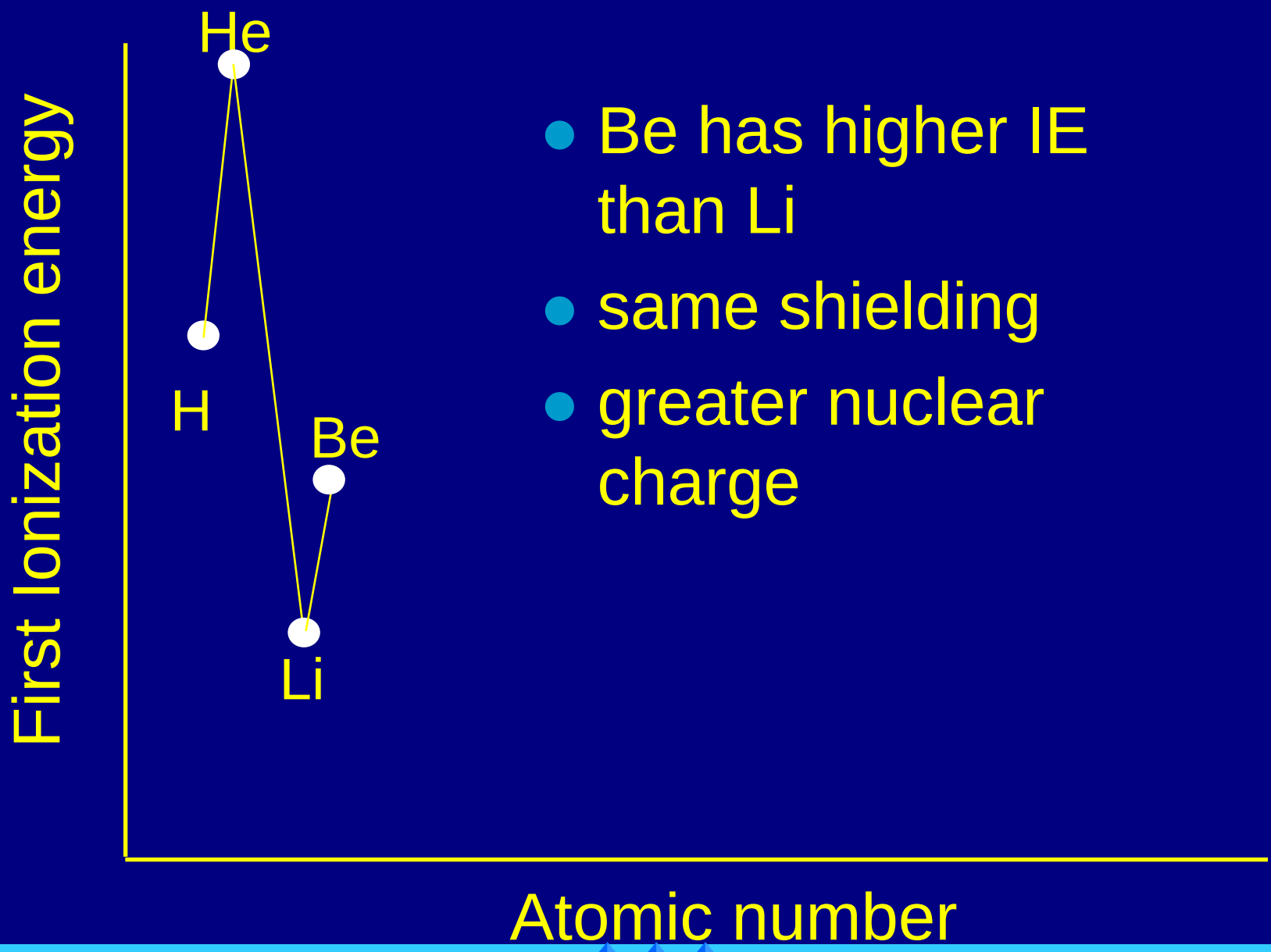
- **All the atoms in the same period have the same energy level.**
- **Same shielding.**
- **Increasing nuclear charge**
- **So IE generally increases from left to right.**
- **Exceptions at full and 1/2 fill orbitals.**



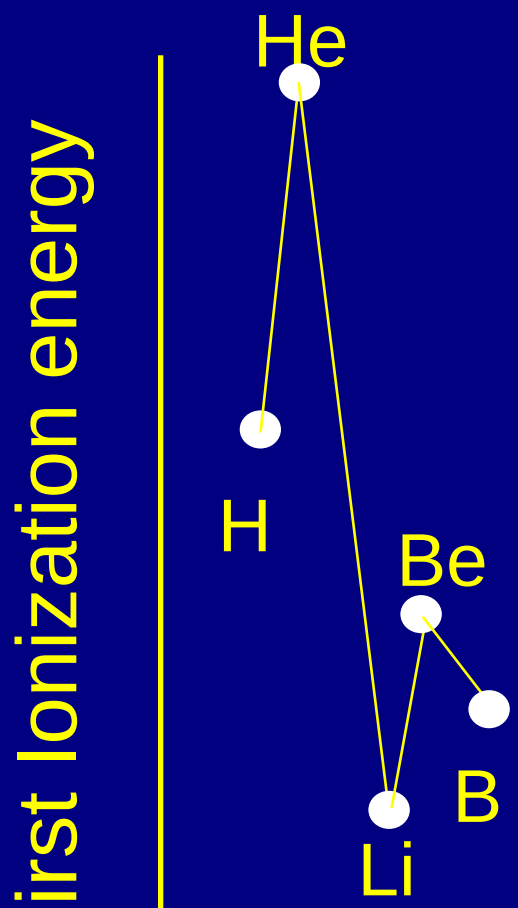
- **He has a greater IE than H.**
- **same shielding**
- **greater nuclear charge**



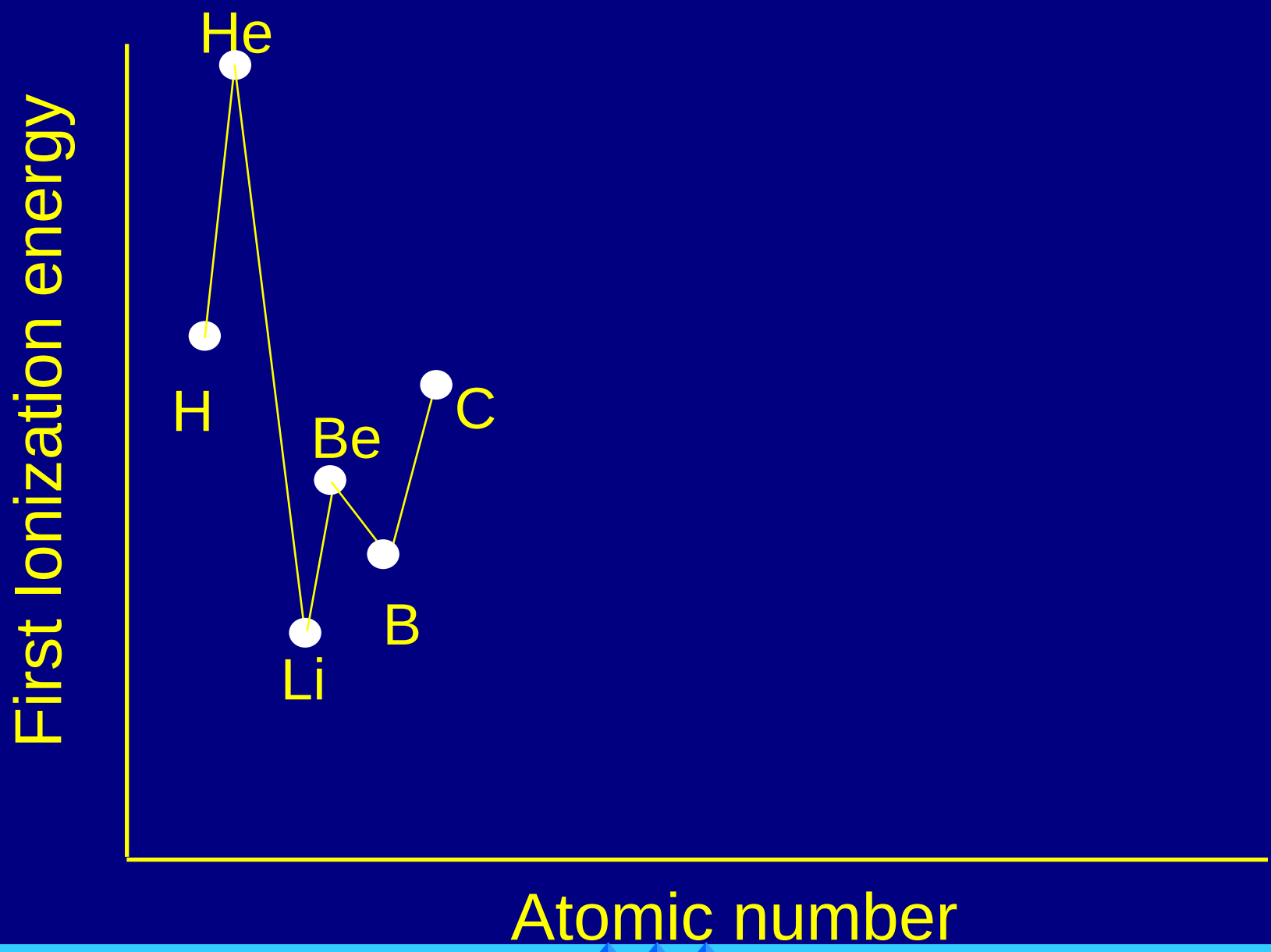
- Li has lower IE than H
- more shielding
- further away
- outweighs greater nuclear charge

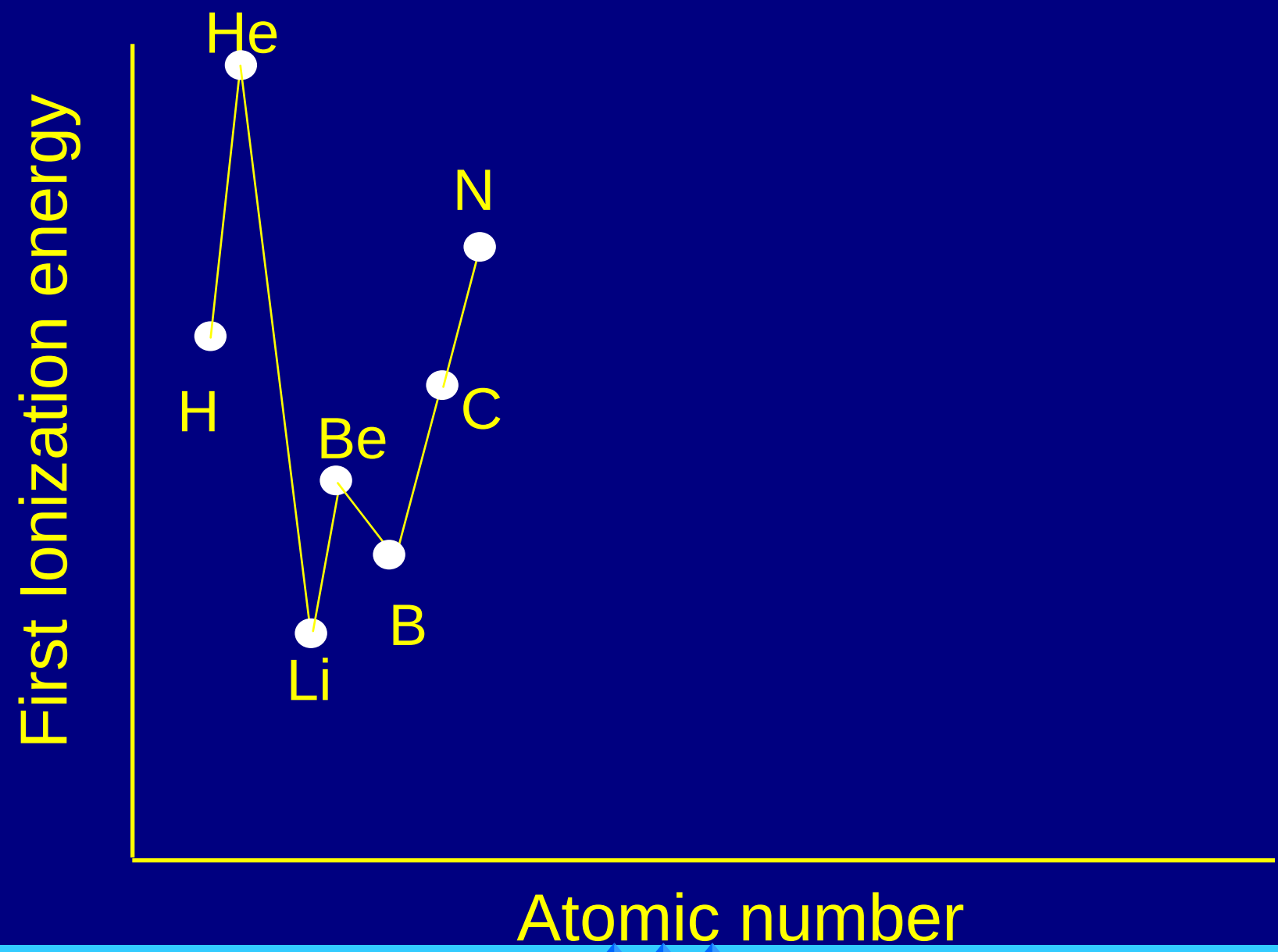


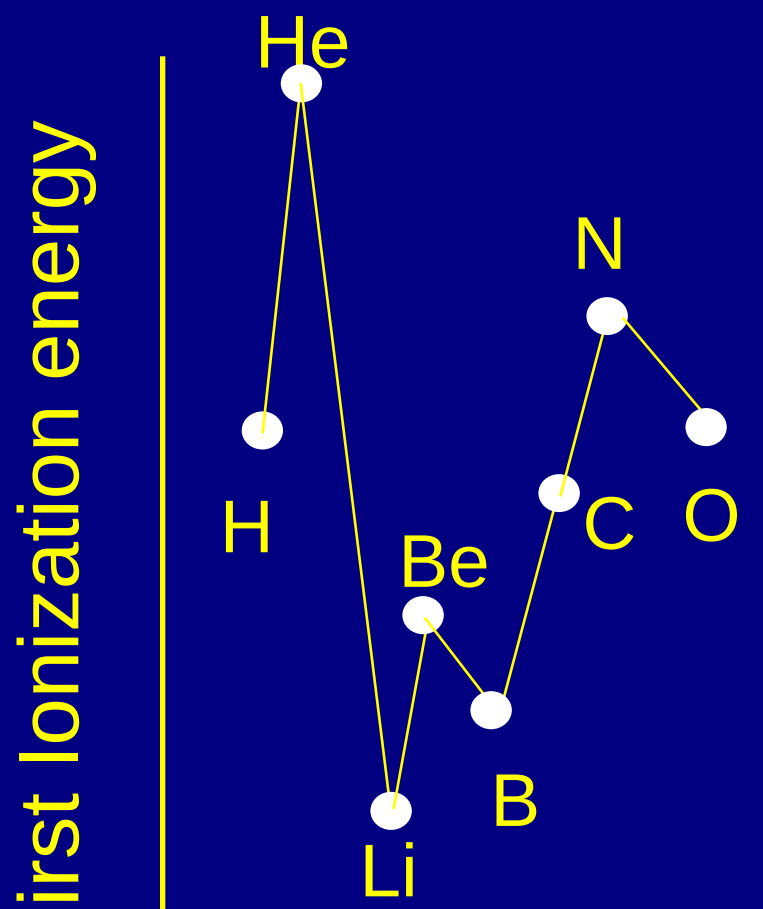
- Be has higher IE than Li
- same shielding
- greater nuclear charge



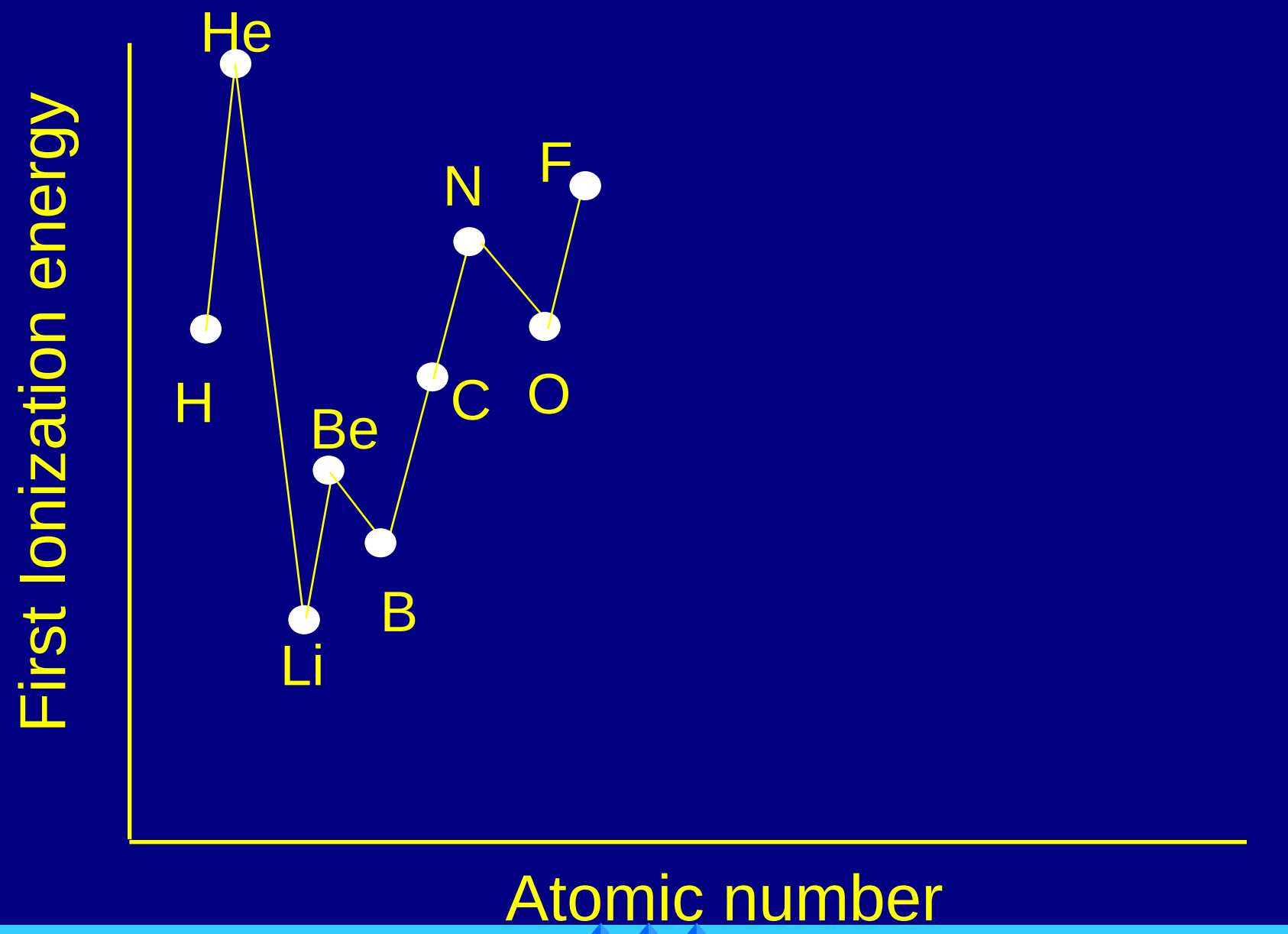
- B has lower IE than Be
- same shielding
- greater nuclear charge
- By removing an electron we make s orbital half filled

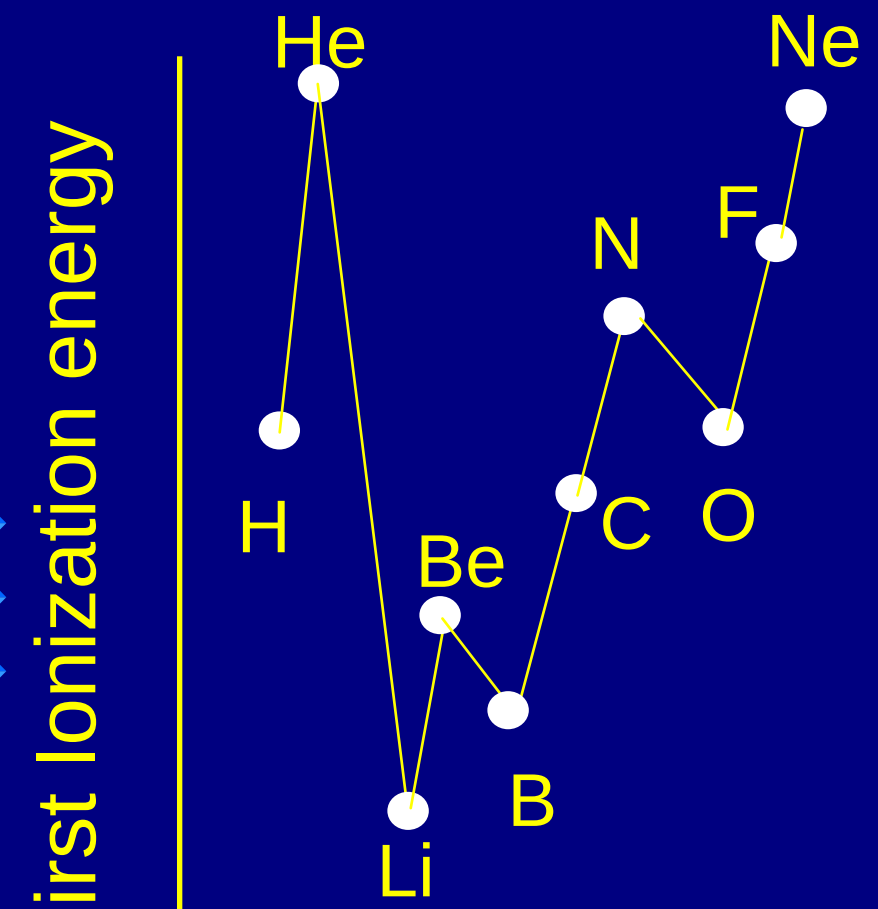

First Ionization energy
He
H
Be
C
B
Li
Atomic number


First Ionization energy
Atomic number
He
H
Li
Be
B
C
N

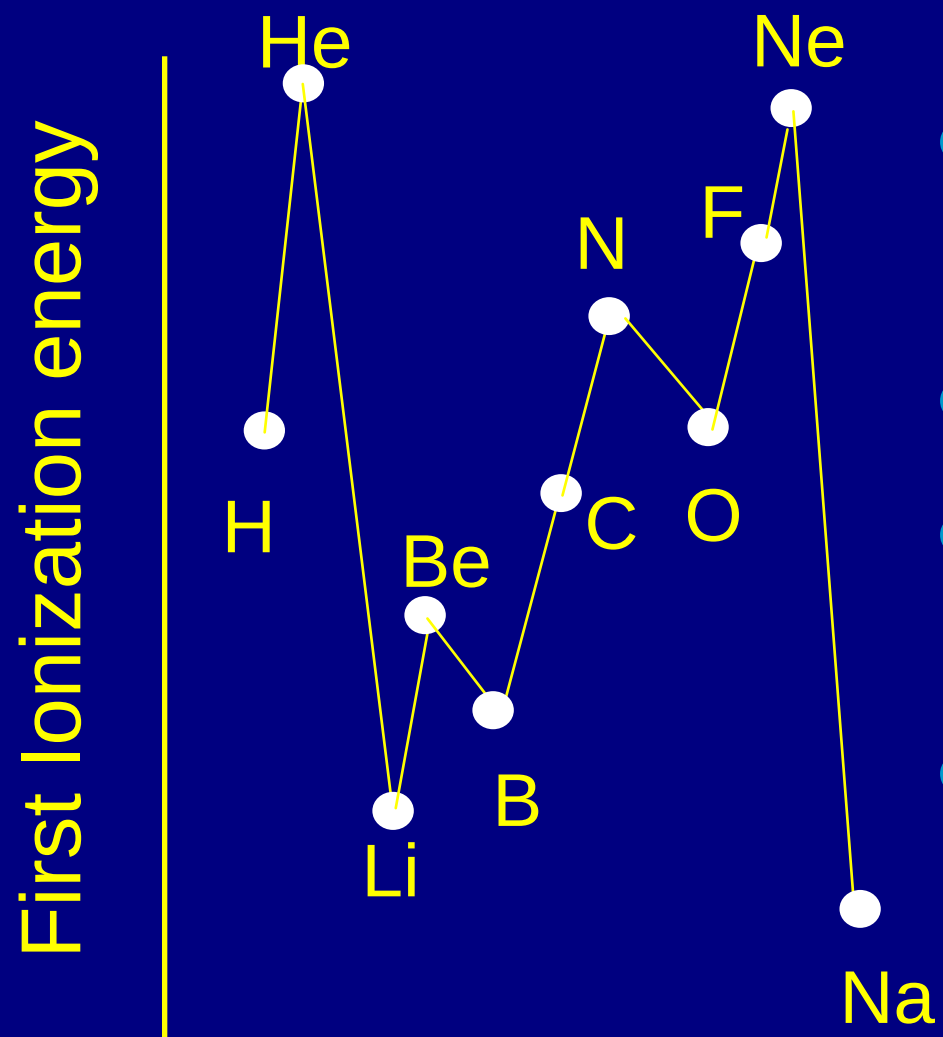


- **Breaks the pattern because removing an electron gets to 1/2 filled p orbital**

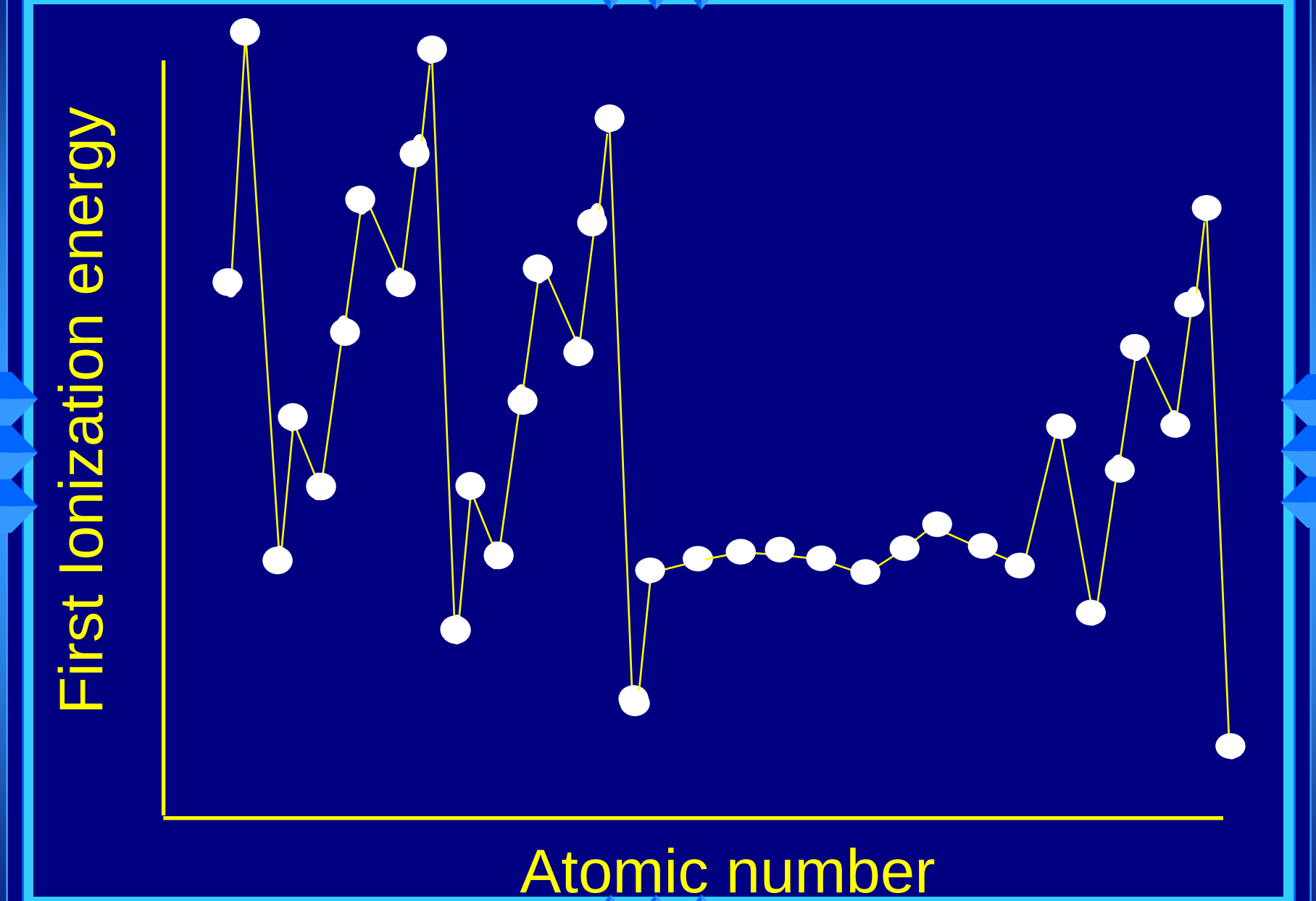

First Ionization energy
Atomic number
H
He
Li
Be
B
C
N
O
F



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



- Na has a lower IE than Li
- Both are s^1
- Na has more shielding
- Greater distance


First Ionization energy
Atomic number

Driving Force

- **Full Energy Levels are very low energy.**
- **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**
- **Alkali earth metals form +2 ions.**

3rd IE

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

Electron Affinity

- **The energy released when 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.**

Ionic Size

- **Cations form by losing electrons.**
- **Cations are smaller that 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 that the atom they come from.**
- **Nonmetals form anions.**
- **Anions of representative elements have noble gas configuration.**

Configuration of Ions

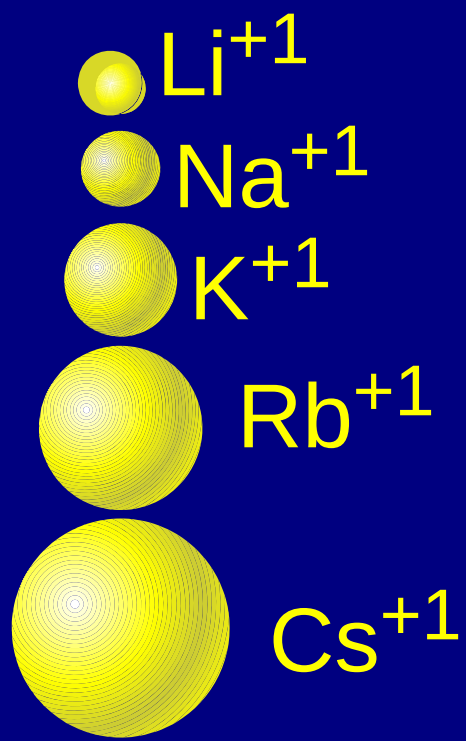
- **Ions always have noble gas configuration.**
- **Na is $1s^2 2s^2 2p^6 3s^1$**
- **Forms a +1 ion - $1s^2 2s^2 2p^6$**
- **Same configuration as neon.**
- **Metals form ions with the configuration of the noble gas before them - they lose electrons.**

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.**

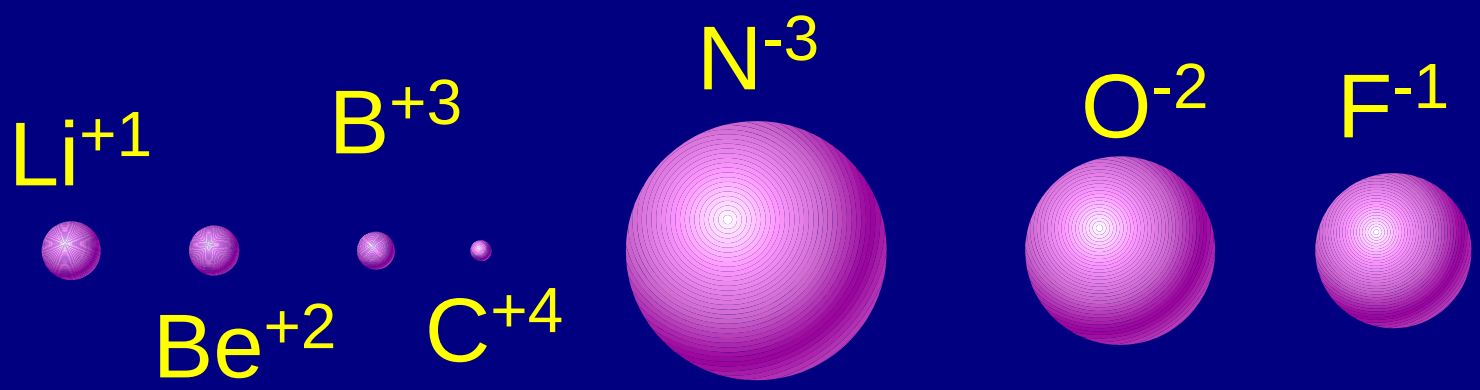
Group trends

- **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.**

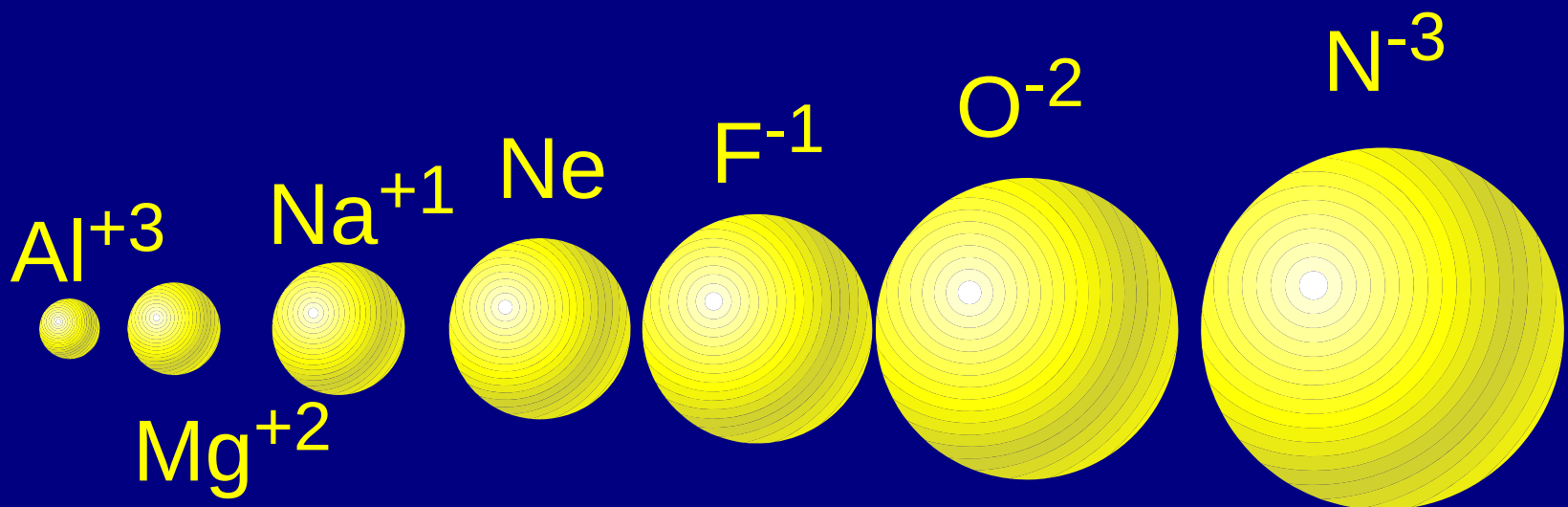


Size of Isoelectronic ions

- **Iso - 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 $1s^1 2s^2 2p^6$**

Size of Isoelectronic ions

- **Positvie ions have more protons so they are smaller.**



Electronegativity

Electronegativity

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

Group Trend

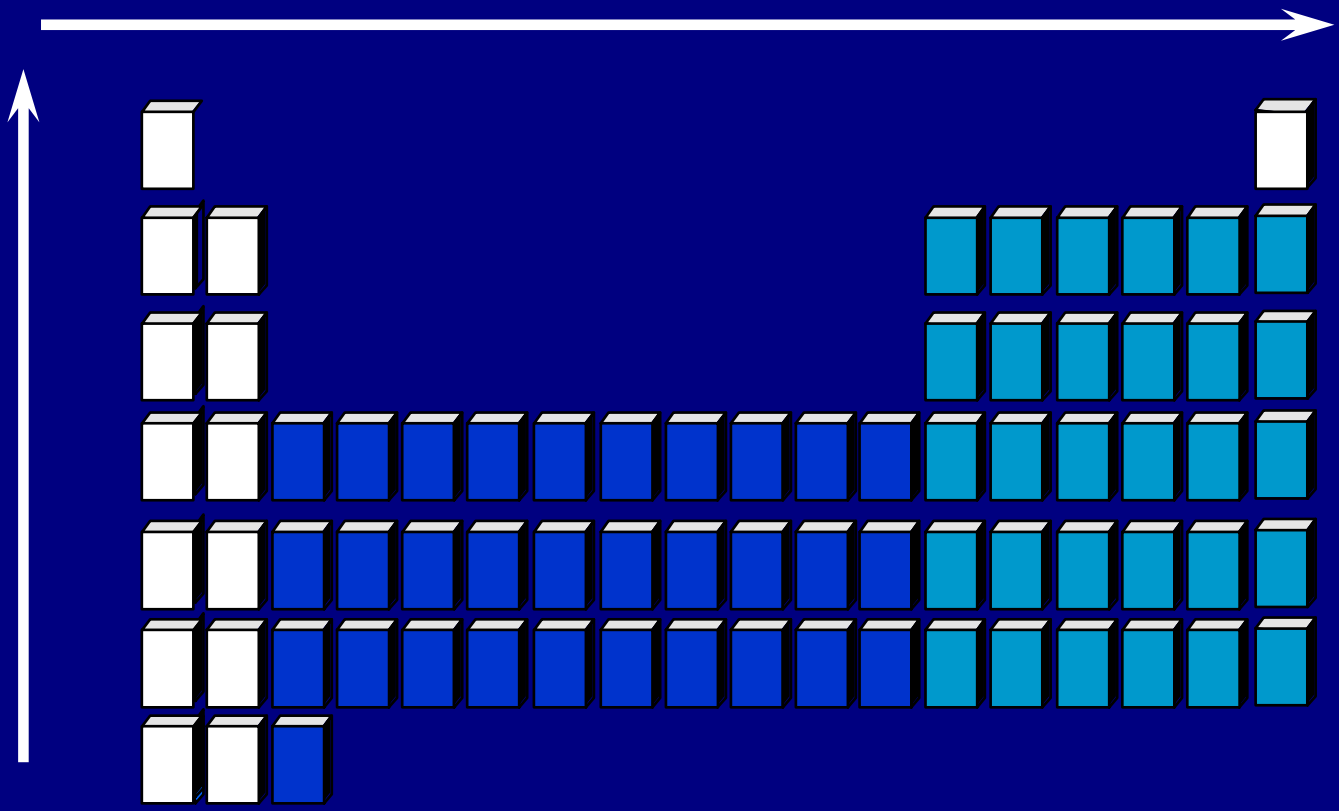
- **The further down a group the farther the electron is away and the more electrons an atom has.**
- **More willing to share.**
- **Low electronegativity.**

Periodic Trend

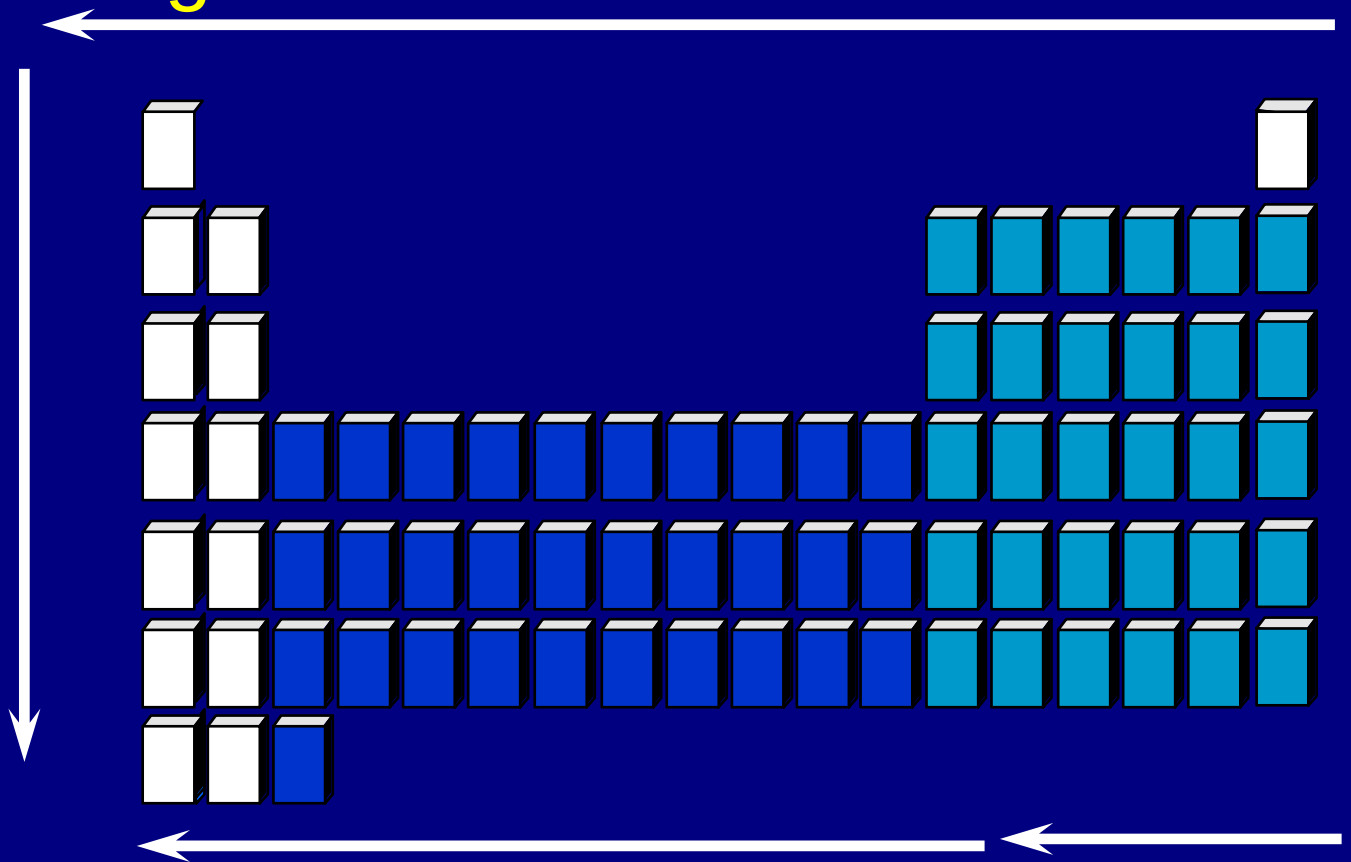
- Metals are at the left end.
- They let their electrons go easily
- Low electronegativity
- At the right end are the nonmetals.
- They want more electrons.
- Try to take them away.
- High electronegativity.

Ionization energy, electronegativity

Electron affinity INCREASE



Atomic size increases,
shielding constant



Ionic size increases