Chapter 6
"The Periodic Table"

Section 6.1
Organizing the Elements

- OBJECTIVES:
  - Explain how elements are organized in a periodic table.

Section 6.1
Organizing the Elements

- OBJECTIVES:
  - Compare early and modern periodic tables.

Section 6.1
Organizing the Elements

- OBJECTIVES:
  - Identify three broad classes of elements.

Section 6.1
Organizing the Elements

- **A few elements, such as gold and copper, have been known for *thousands of years* - since ancient times**
- **Yet, only about 13 had been identified by the year 1700.**
- **As more were discovered, chemists realized they needed a way to <u>organize</u> the elements.**

Section 6.1
Organizing the Elements

- **Chemists used the** ***properties*** **of elements to sort them into groups.**
- **In 1829 J. W. Dobereiner arranged elements into <u>triads</u> – groups of three elements with similar properties**
  - One element in each triad had *properties* intermediate of the other two elements

Mendeleev's Periodic Table

- **By the mid-1800s, about 70 elements were known to exist**
- **Dmitri <u>Mendeleev</u> – a Russian chemist and teacher**
- **Arranged elements in order of increasing <u>*atomic mass*</u>**

Mendeleev

- *He left blanks* **for yet undiscovered elements**
  - When they were discovered, he had made good predictions
- **But, there were problems:**
  - Such as Co and Ni; Ar and K; Te and I

A better arrangement

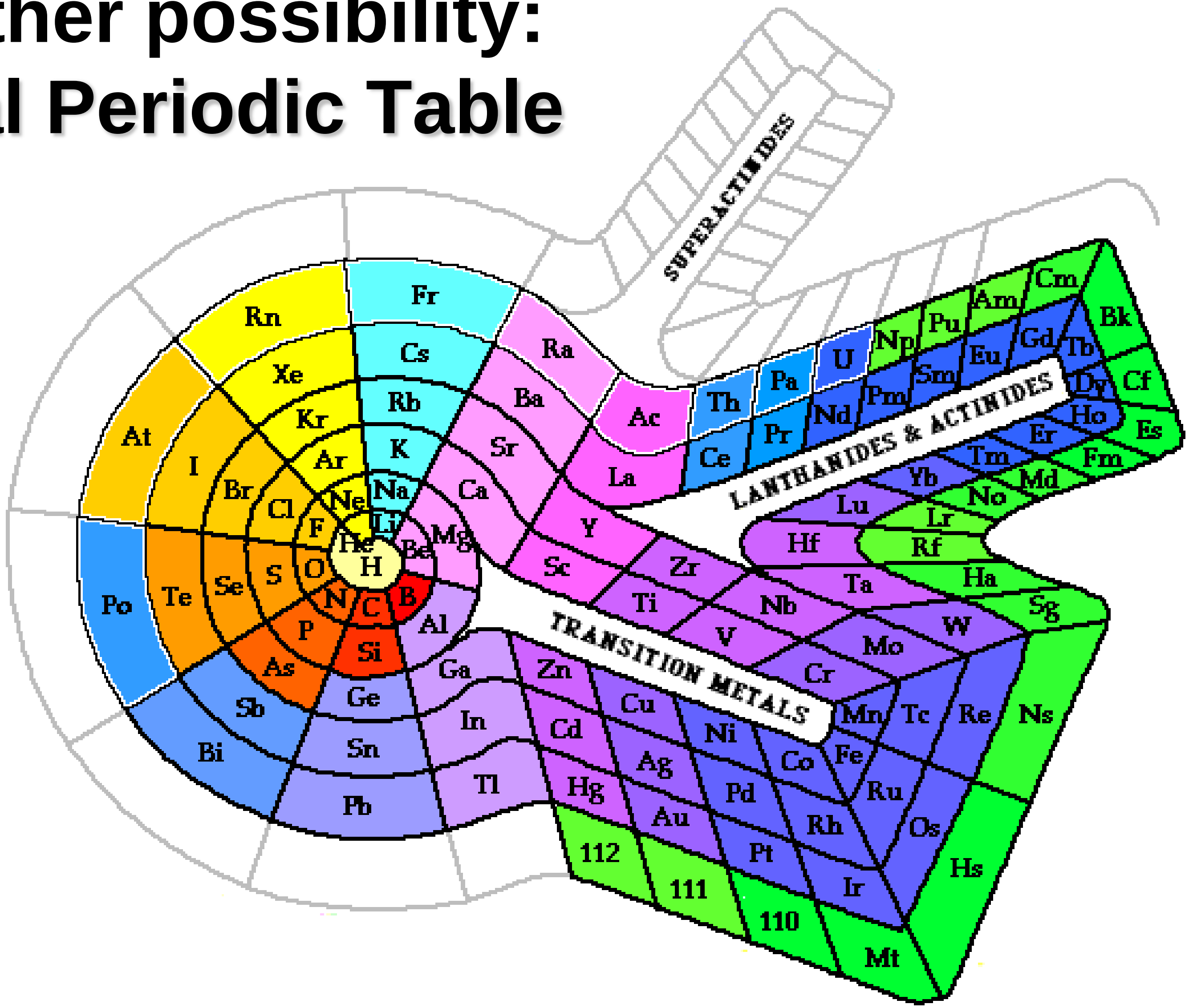
- **In 1913, Henry <u>Moseley</u> – British physicist, arranged elements according to increasing <u>*atomic number*</u>**
- **The arrangement used today**
- **The symbol, atomic number & mass are basic items included-textbook page 162 and 163**

Alkali metals (group 1); Alkaline earth metals (group 2); Transition metals (groups 3–12); Halogens (group 17); Noble gases (group 18)

1 1A	2 2A	3	4	5	6	7	8	9	10	11	12	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La*	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac†	104 Unq	105 Unp	106 Unh	107 Uns	108 Uno	109 Une	110 Uun	111 Uuu							

*Lanthanides	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
† Actinides	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Another possibility:
Spiral Periodic Table



The Periodic Law says:

- **When elements are arranged in order of increasing atomic number, there is a *periodic repetition* of their physical and chemical properties.**
- **Horizontal rows = <u>periods</u>**
  - There are 7 periods
- **Vertical column = <u>group</u> (or family)**
  - Similar physical & chemical prop.
  - Identified by number & letter (IA, IIA)

Areas of the periodic table

- **Three classes of elements are: 1) metals, 2) nonmetals, and 3) metalloids**

1) **<u>Metals</u>**: electrical conductors, have luster, ductile, malleable
2) **<u>Nonmetals</u>**: generally brittle and non-lustrous, poor conductors of heat and electricity

Areas of the periodic table

- **Some nonmetals are gases (O, N, Cl); some are brittle solids (S); one is a fuming dark red liquid (Br)**
- **Notice the heavy, stair-step line?**

3) Metalloids**: border the line-2 sides**

– Properties are *intermediate* between

Section 6.2
Classifying the Elements

- OBJECTIVES:
  - Describe the information in a periodic table.

Section 6.2
Classifying the Elements

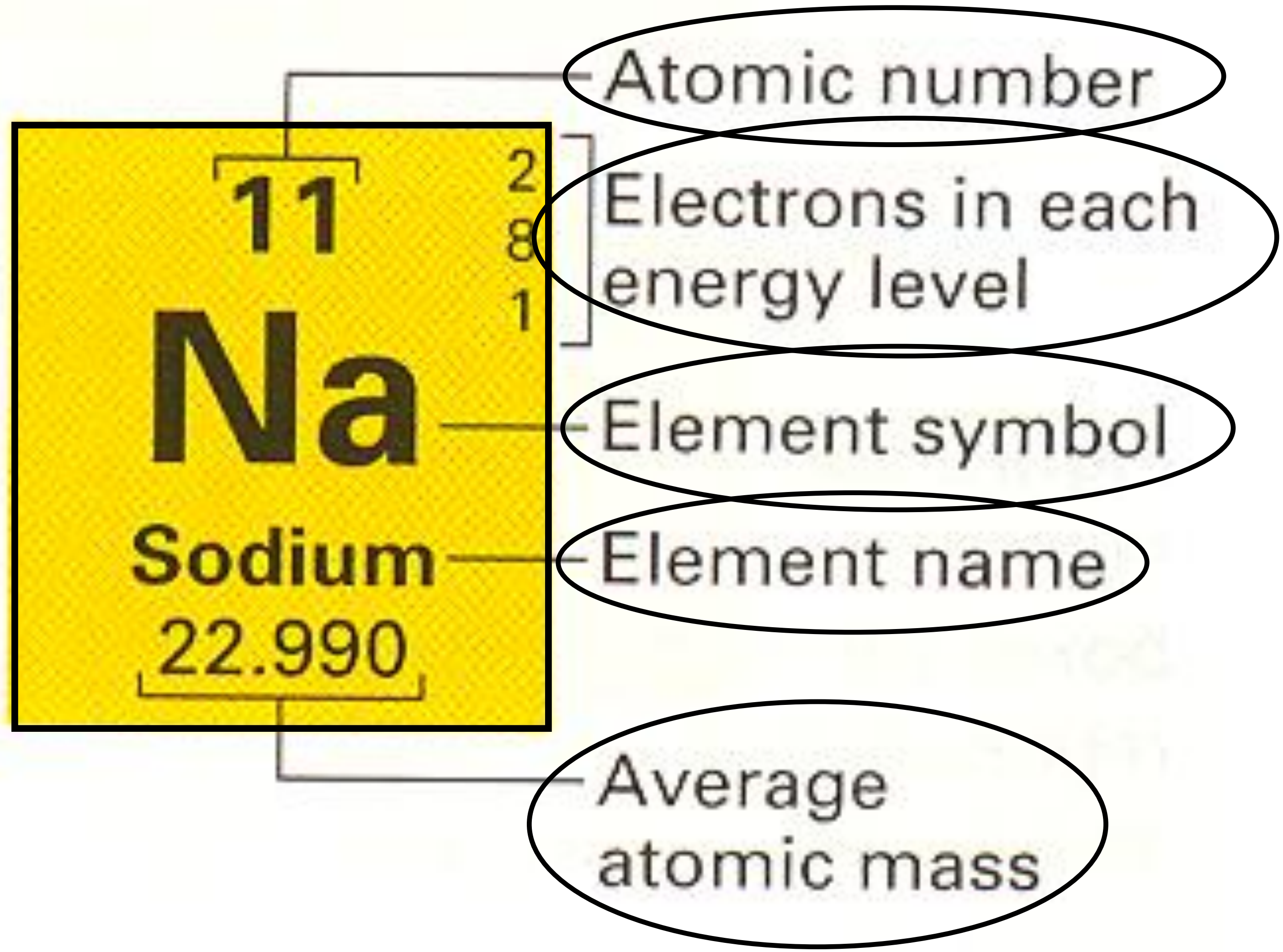
- OBJECTIVES:
  - <u>Classify</u> elements based on electron configuration.

Section 6.2
Classifying the Elements

- OBJECTIVES:
  - <u>Distinguish</u> representative elements and transition metals.

Squares in the Periodic Table

- **The periodic table displays the <u>symbols</u> and <u>names</u> of the elements, along with information about the structure of their atoms:**
  - Atomic number and atomic mass
  - **Black symbol = solid**; **red = gas**; **blue = liquid**
    (from the Periodic Table on our classroom wall)


11
2
8
1
Na
Sodium
22.990
Atomic number
Electrons in each energy level
Element symbol
Element name
Average atomic mass

Groups of elements **- family names**

- Group IA **– <u>alkali metals</u>**
  - Forms a “base” (or alkali) when <u>*reacting*</u> with water (not just dissolved!)
- Group 2A **– <u>alkaline earth metals</u>**
  - Also form bases with water; do not dissolve well, hence “earth metals”
- Group 7A **– <u>halogens</u>**
  - Means “salt-forming”

Electron Configurations in Groups

- **Elements can be sorted into 4 different groupings *__based on their electron configurations__*:**

1) Noble gases
2) Representative elements
3) Transition metals
4) Inner transition metals

Let's now take a closer look at these.

Electron Configurations in Groups

1) Noble gases **are the elements in Group 8A** **(also called Group18 or 0)**
   - Previously called "**inert gases**" because they rarely take part in a reaction; very stable = don't react
   - Noble gases have an electron configuration that has the outer s and p sublevels completely full

Electron Configurations in Groups

2) Representative Elements **are in Groups 1A through 7A**

- Display wide range of properties, thus a good “representative”
- Some are metals, or nonmetals, or metalloids; some are solid, others are gases or liquids
- Their outer s and p electron configurations are NOT filled

Electron Configurations in Groups

3) Transition metals **are in the "B" columns of the periodic table**

- Electron configuration has the outer s sublevel full, and is now filling the "d" sublevel
- A "transition" between the metal area and the nonmetal area
- Examples are gold, copper, silver

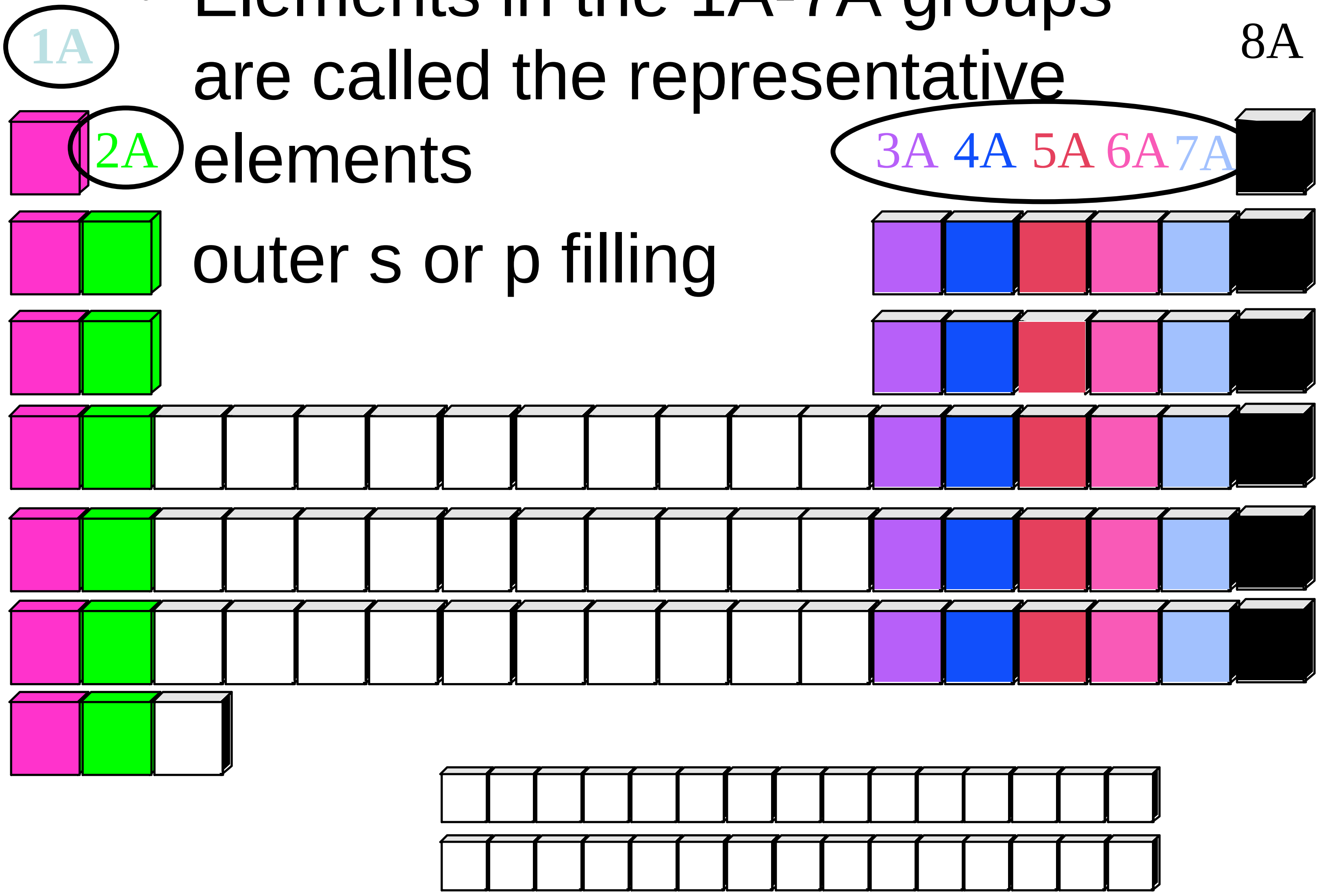
Electron Configurations in Groups

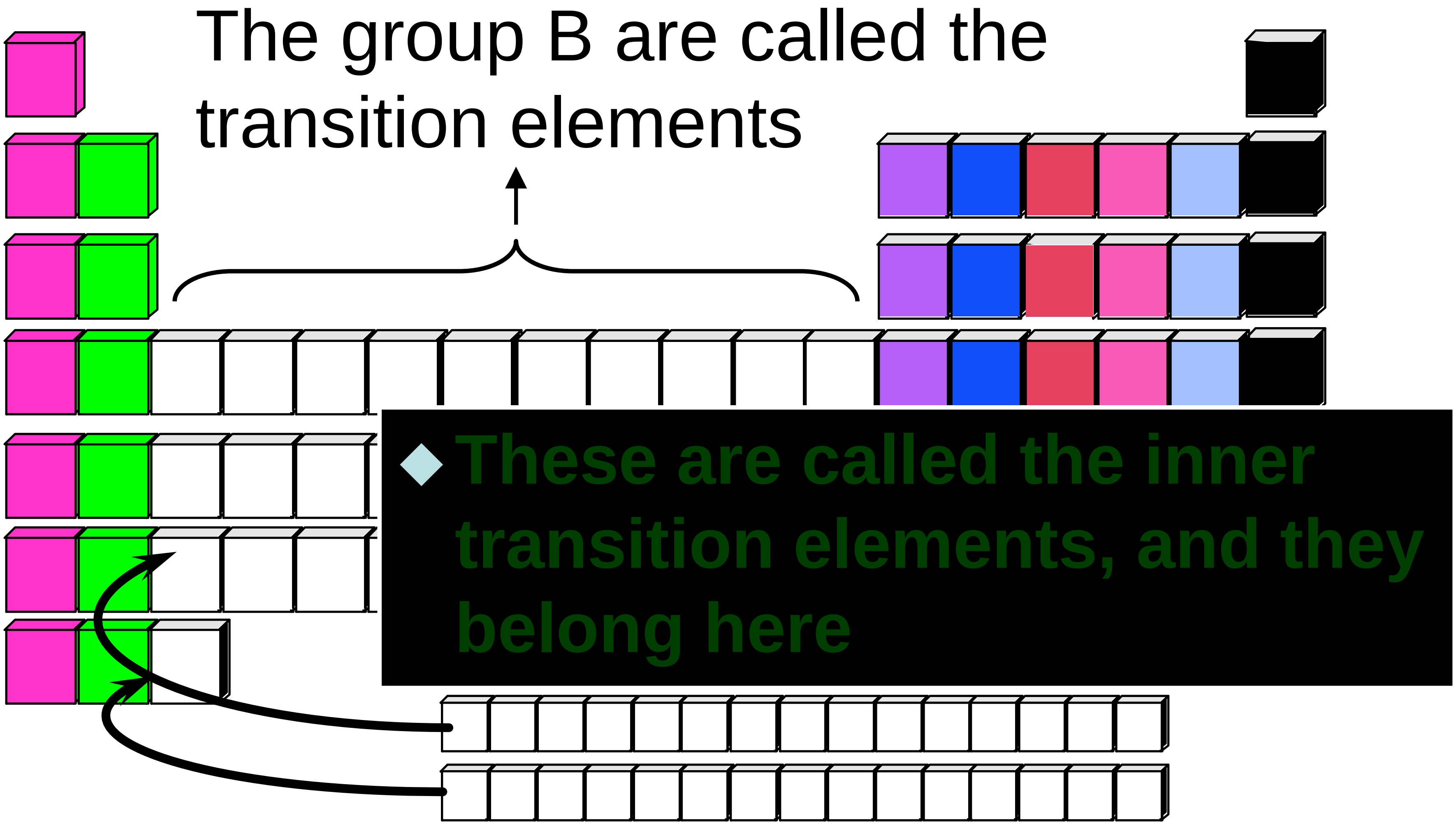
4) Inner Transition Metals **are located below the main body of the table, in two horizontal rows**

- Electron configuration has the outer s sublevel full, and is now filling the "f" sublevel
- Formerly called "rare-earth" elements, but this is not true because some are very abundant

- Elements in the 1A-7A groups are called the representative elements

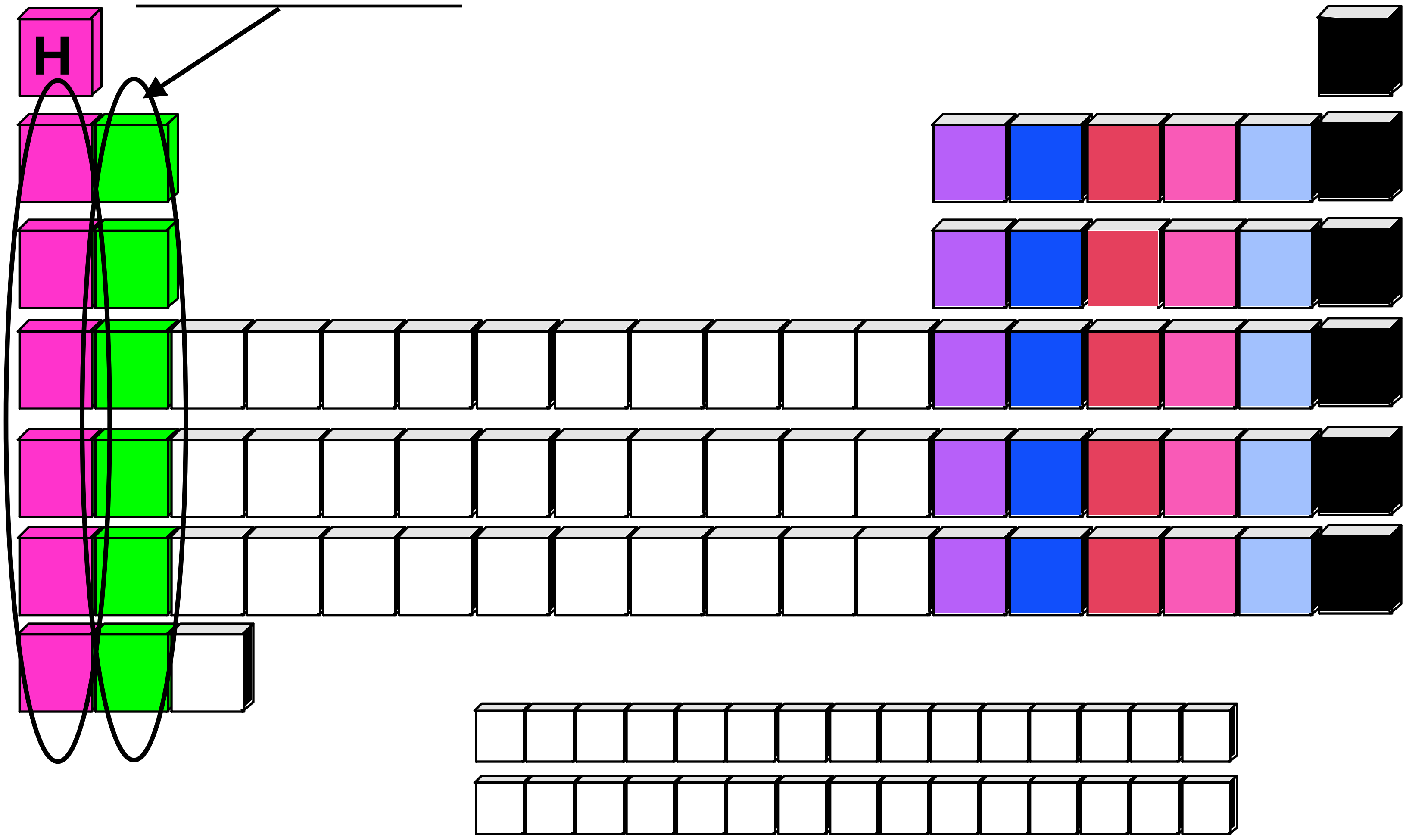
outer s or p filling



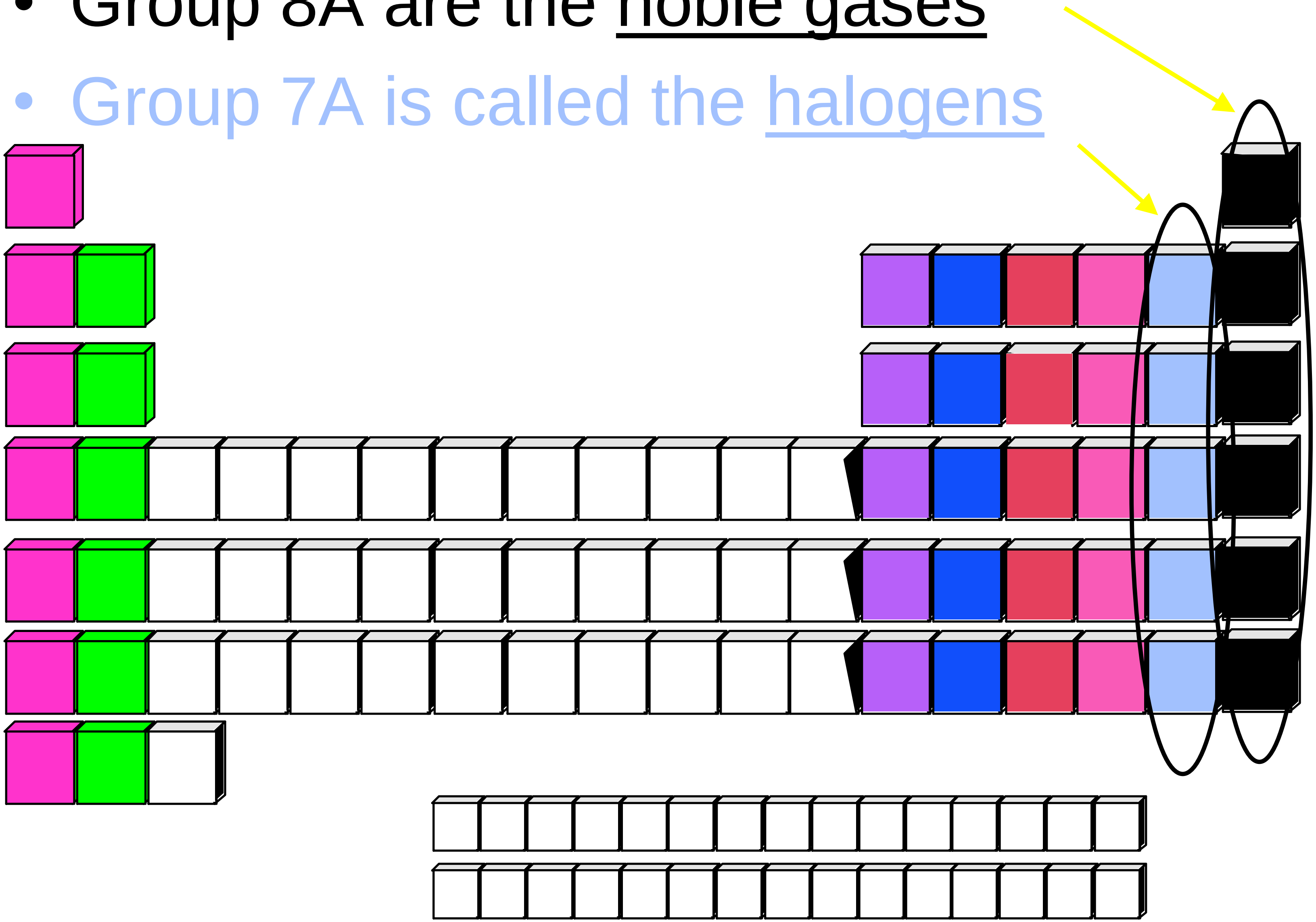

The group B are called the transition elements
These are called the inner transition elements, and they belong here

Group 1A are the alkali metals (but NOT H)

Group 2A are the alkaline earth metals



- Group 8A are the noble gases
- Group 7A is called the halogens



Do you notice any similarity in these configurations of the alkali metals?

Element	Configuration
H 1	$1s^1$
Li 3	$1s^2 \underline{2s^1}$
Na 11	$1s^2 2s^2 2p^6 \underline{3s^1}$
K 19	$1s^2 2s^2 2p^6 3s^2 3p^6 \underline{4s^1}$
Rb 37	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 \underline{5s^1}$
Cs 55	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 \underline{6s^1}$
Fr 87	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 \underline{7s^1}$

Do you notice any similarity in the configurations of the noble gases?

$\underline{1s^2}$ — He 2

$1s^2\underline{2s^22p^6}$ — Ne 10

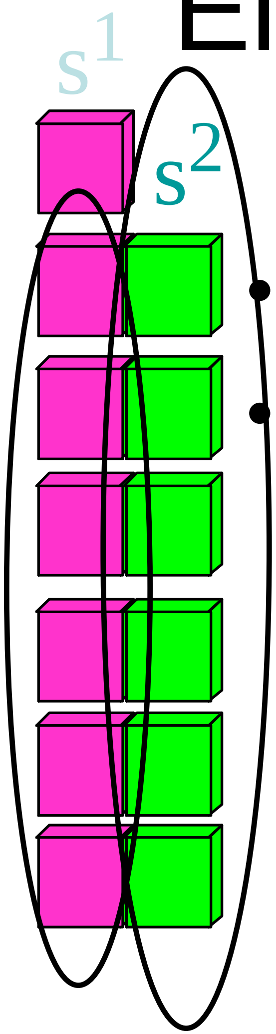
$1s^22s^22p^6\underline{3s^23p^6}$ — Ar 18

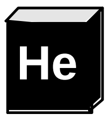
$1s^22s^22p^63s^23p^6\underline{4s^2}3d^{10}\underline{4p^6}$ — Kr 36

$1s^22s^22p^63s^23p^64s^23d^{10}4p^6\underline{5s^2}4d^{10}\underline{5p^6}$ — Xe 54

$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6\underline{6s^2}4f^{14}5d^{10}\underline{6p^6}$ — Rn 86

Elements in the s - blocks

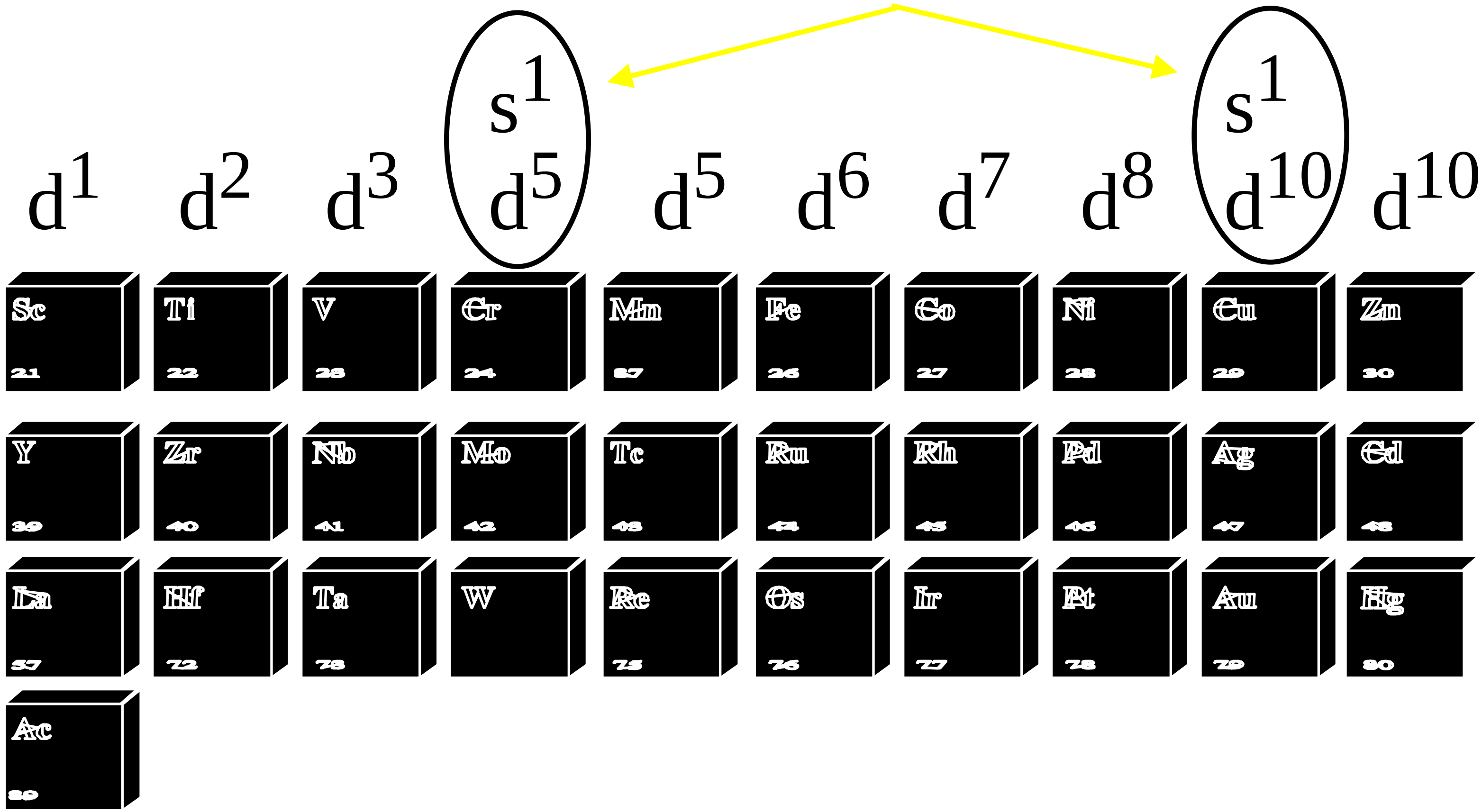




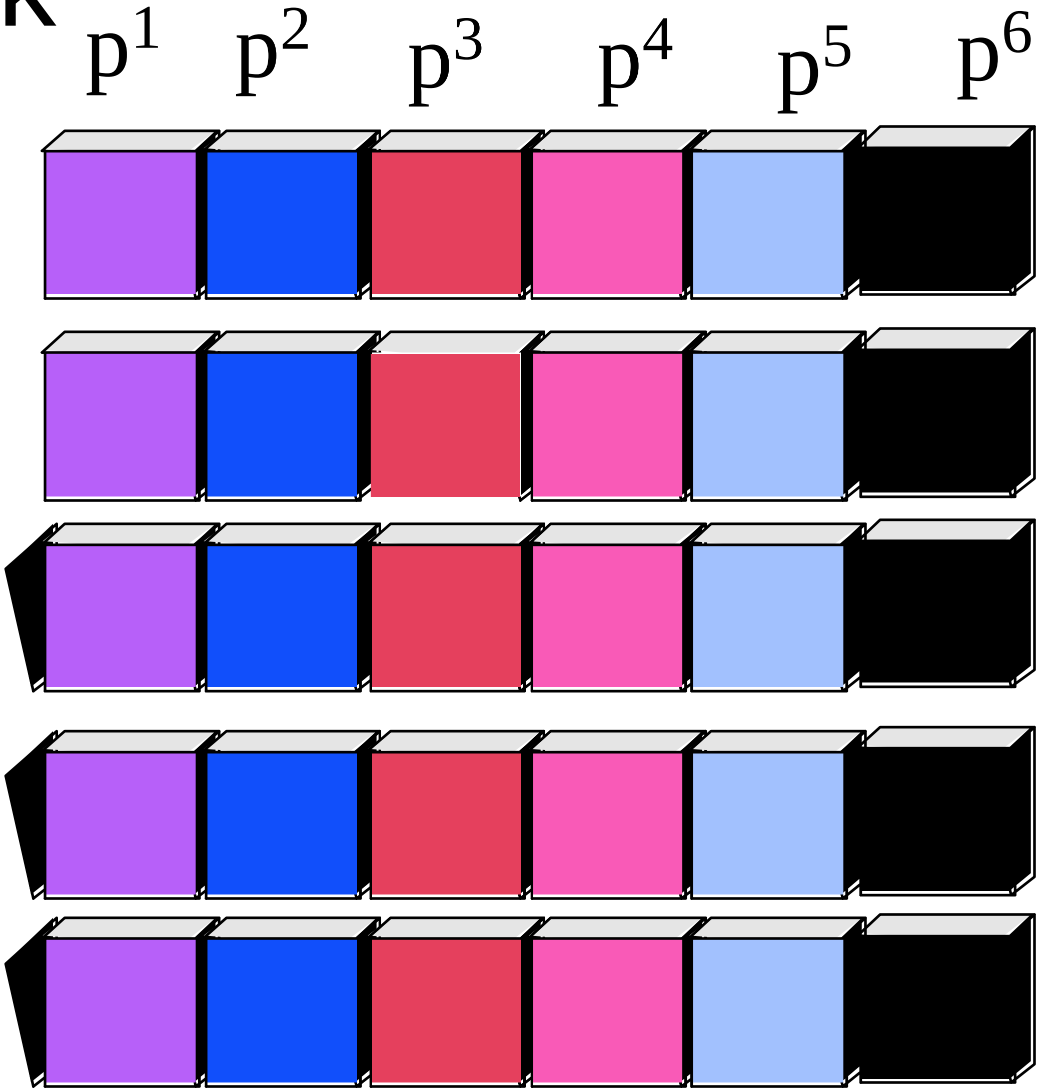
- Alkali metals all end in s^1
- Alkaline earth metals all end in s^2
  - really should include He, but it fits better in a different spot, since He has the properties of the noble *gases*, and has a full outer level of electrons.

Transition Metals - d block

Note the change in configuration.

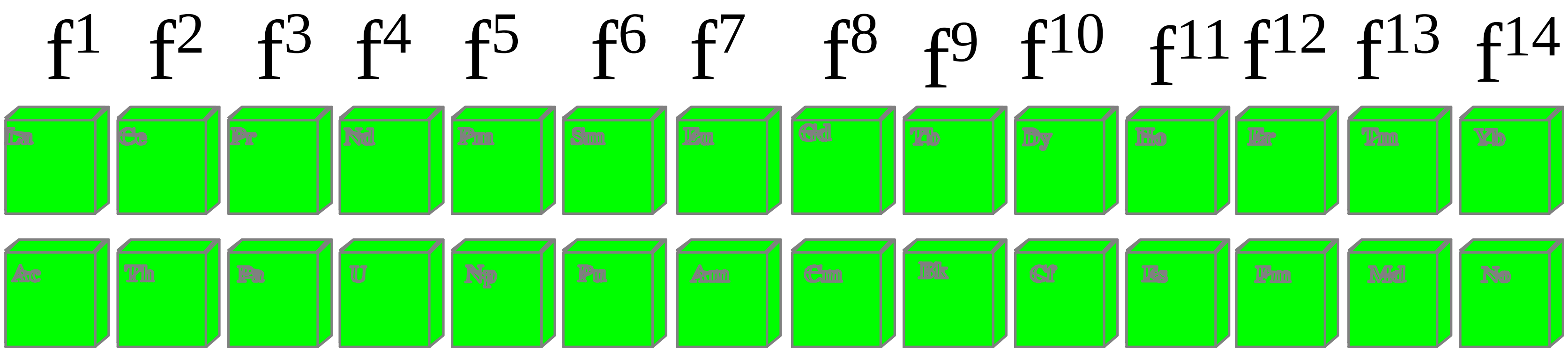


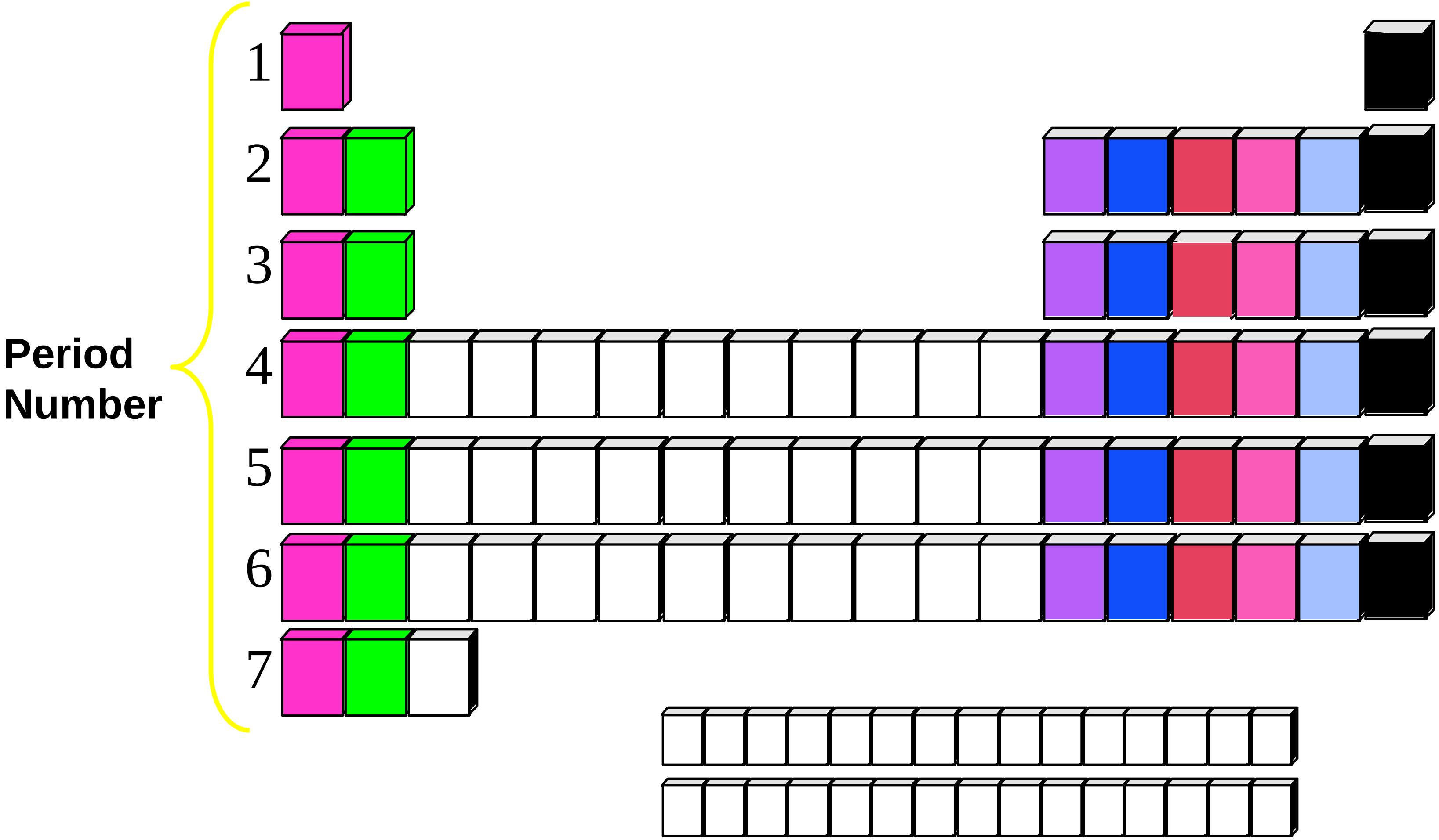
The P-block



F - block

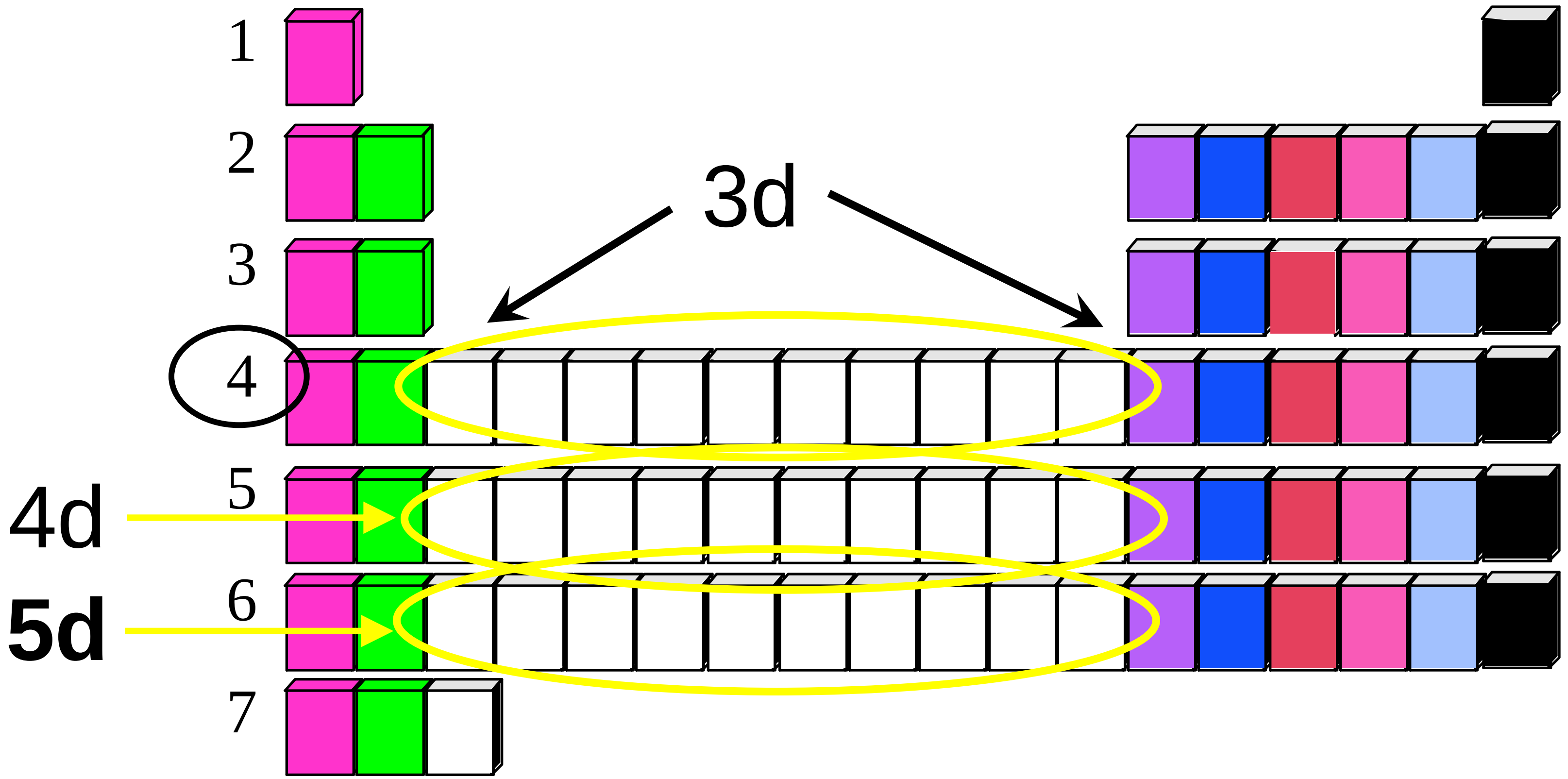
- Called the “inner transition elements”

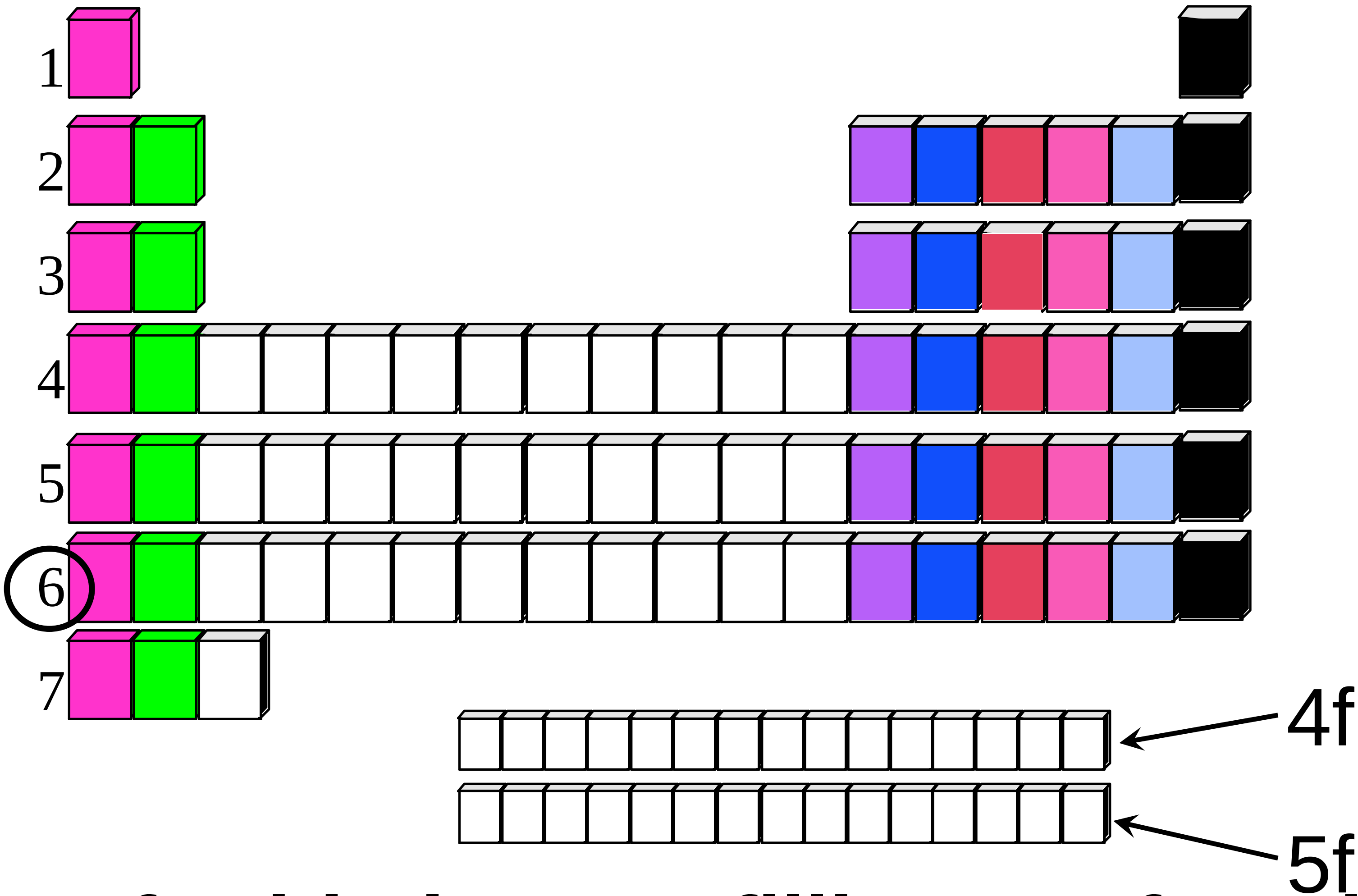




- Each row (or period) is the <u>energy level</u> for s and p orbitals.

- The “d” orbitals fill up in levels 1 less than the period number, so the first d is 3d even though it’s in row 4.





- **f orbitals start filling at 4f, and are <u>2 less</u> than the period number**

Section 6.3
Periodic Trends

- OBJECTIVES:
  - Describe *trends* among the elements for *atomic size*.

Section 6.3
Periodic Trends

- OBJECTIVES:
  - Explain how **ions** form.

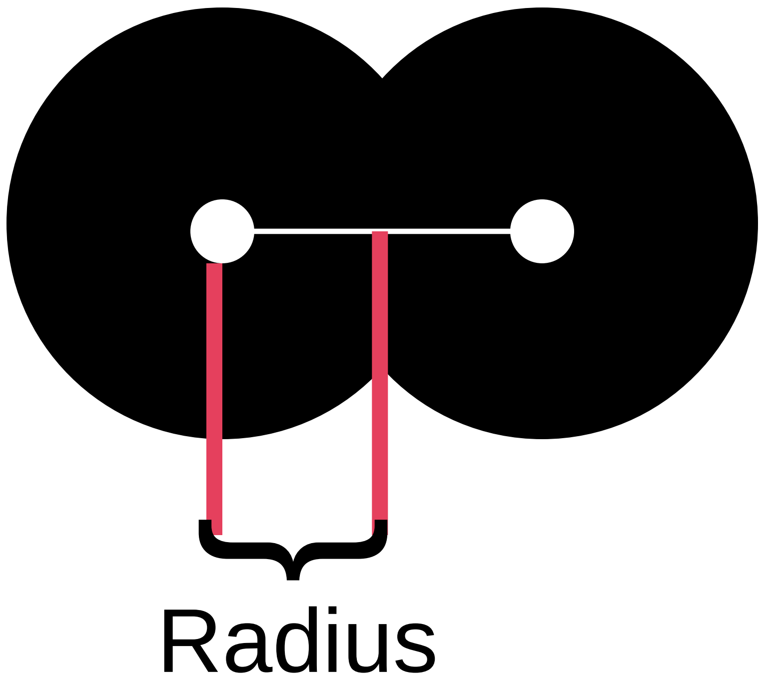
Section 6.3
Periodic Trends

- OBJECTIVES:
  - Describe periodic *trends* for first ionization energy, ionic size, and electronegativity.

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



- Measure the Atomic Radius - this is half the distance between the two nuclei of a diatomic molecule.

ALL Periodic Table Trends

- Influenced by three factors:

1. Energy Level
   - Higher energy levels are further away from the nucleus.
2. Charge on nucleus (# protons)
   - More charge pulls electrons in closer. (+ and – attract each other)

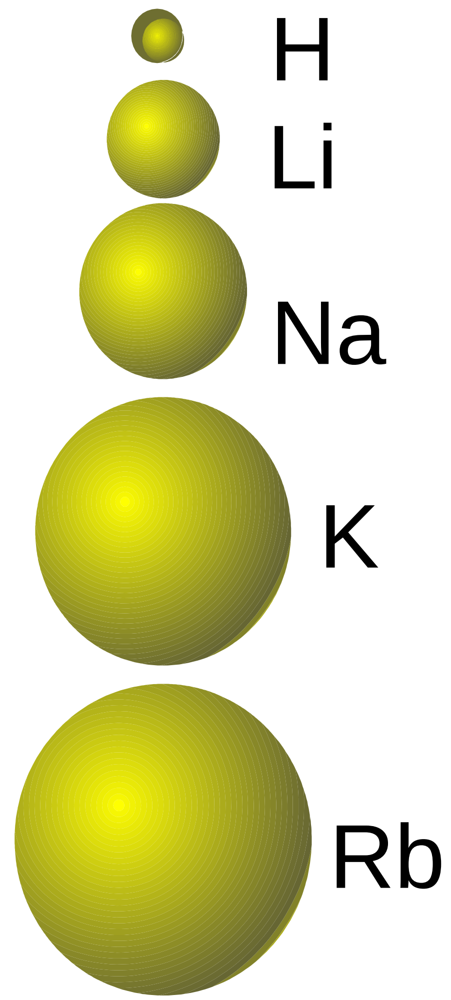
- 3. Shielding effect (blocking effect?)

What do they influence?

- ➢ Energy levels and Shielding have an effect on the *GROUP* (⮘)
- ➢ Nuclear charge has an effect on a *PERIOD* (⌦)

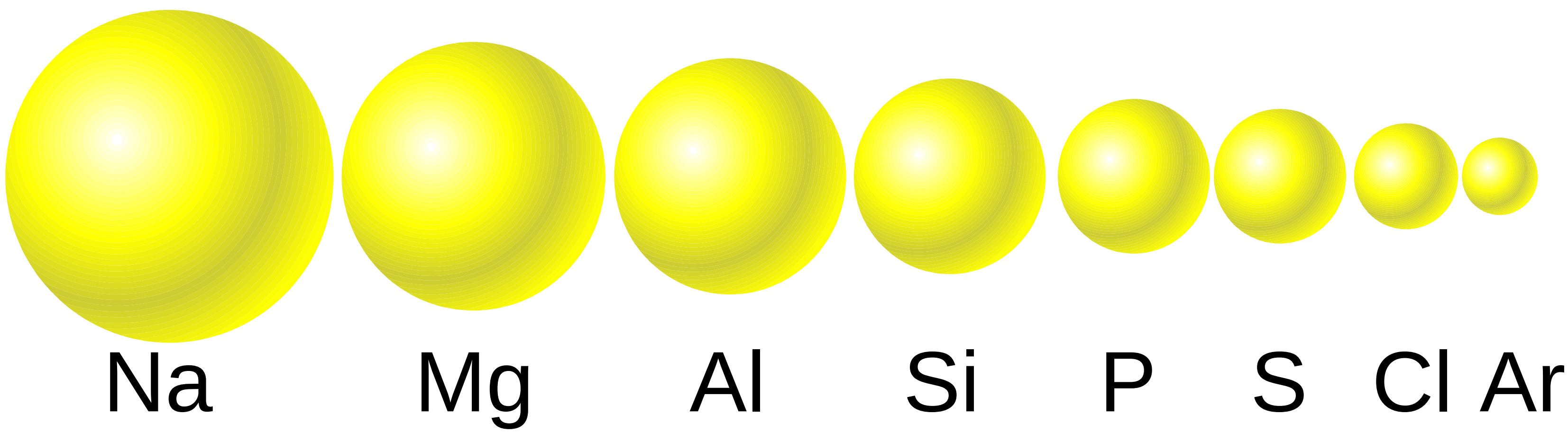
#1. Atomic Size - Group trends

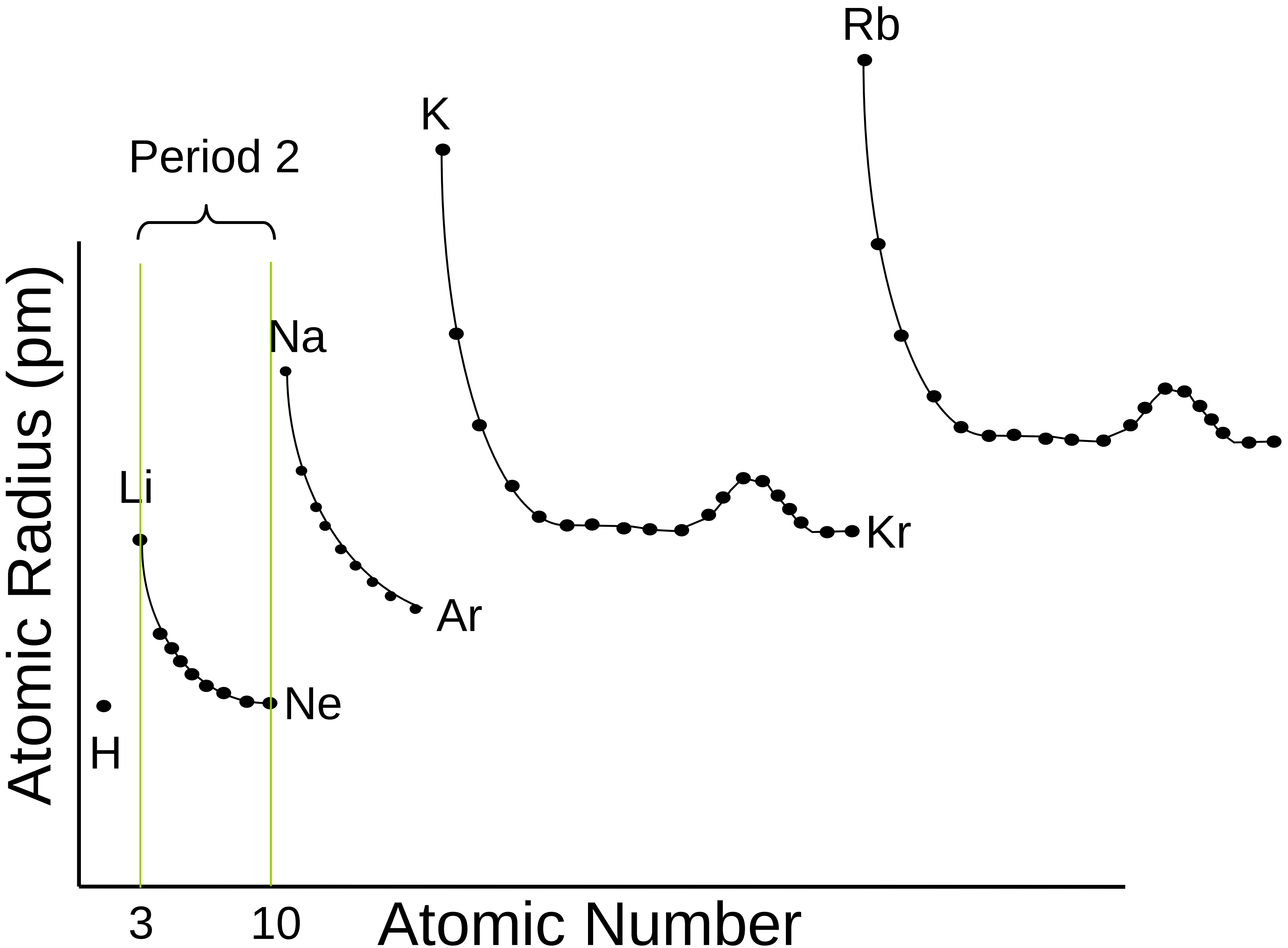
- As we increase the atomic number (or go down a group). . .
- each atom has another energy level,
- so the atoms get *bigger*.



#1. Atomic Size - Period Trends

- Going from left to right across a period, the size gets smaller.
- Electrons are in the same energy level.
- But, there is more nuclear charge.
- Outermost electrons are pulled closer.




Rb
K
Period 2
Na
Li
Ar
Kr
Ne
H
Atomic Radius (pm)
3
10
Atomic Number

Ions

- **Some compounds are composed of particles called “ions”**
  - An **ion** is an atom (or group of atoms) that has a *positive or negative charge*
- **Atoms** are neutral because the number of protons equals electrons
  - Positive and negative ions are formed when electrons are *transferred* (lost or gained) between atoms

Ions

- **<u>Metals tend to LOSE electrons</u>, from their outer energy level**
  - Sodium loses one: there are now more protons (11) than electrons (10), and thus a positively charged particle is formed = "**<u>cation</u>**"
  - The charge is written as a number followed by a plus sign: Na^{1+}
  - Now named a "**sodium ion**"

Ions

- **Nonmetals tend to GAIN one or more electrons**
  - Chlorine will gain one electron
  - Protons (17) no longer equals the electrons (18), so a charge of -1
  - Cl^{1-} is re-named a “chloride ion”
  - Negative ions are called “**anions**”

#2. Trends in Ionization Energy

- Ionization energy is 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 only the first electron is called the <u>first ionization energy.</u>

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.

Table 6.1, p. 173

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

Symbol	First	Second	Third
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O	1314	3391	5301
F	1681	3375	6045
Ne	2080	3963	6276

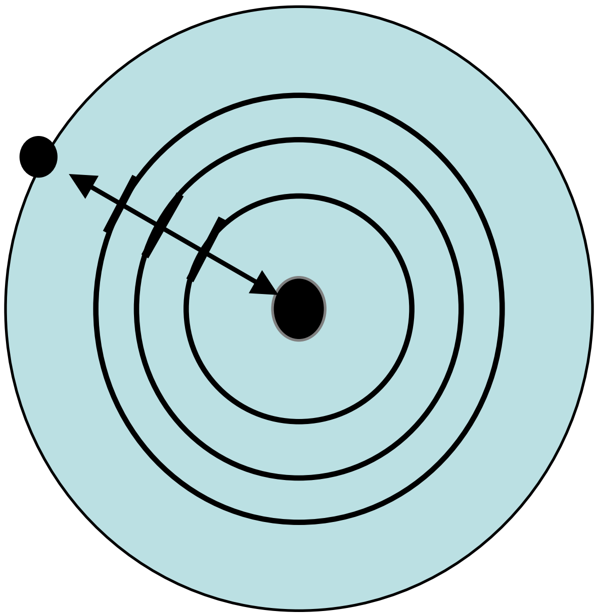
Why did these values increase **so much**?

What factors determine 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, lower 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 <u>same</u> shielding, if it is in the <u>same period</u>

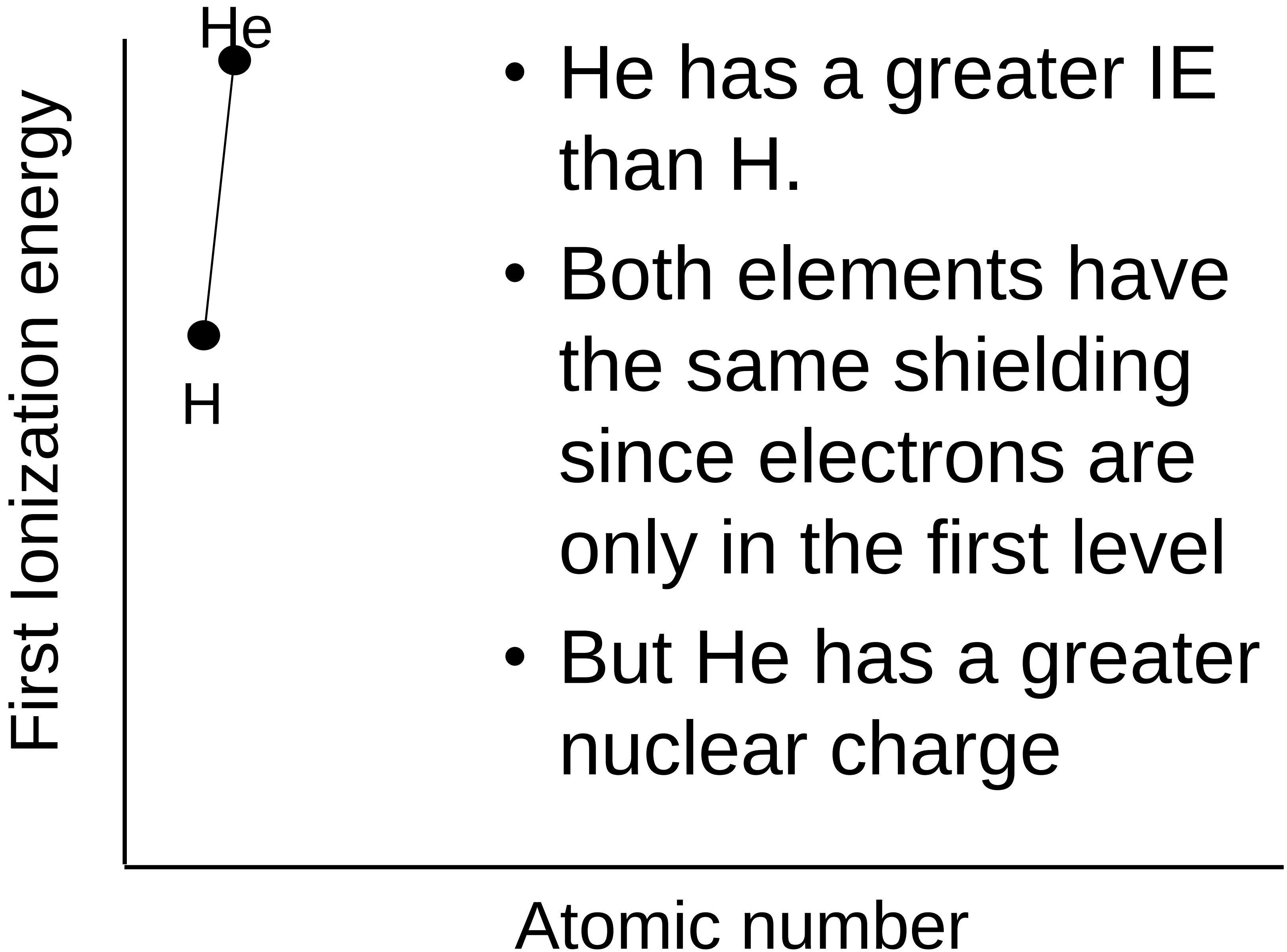


Ionization Energy - Group trends

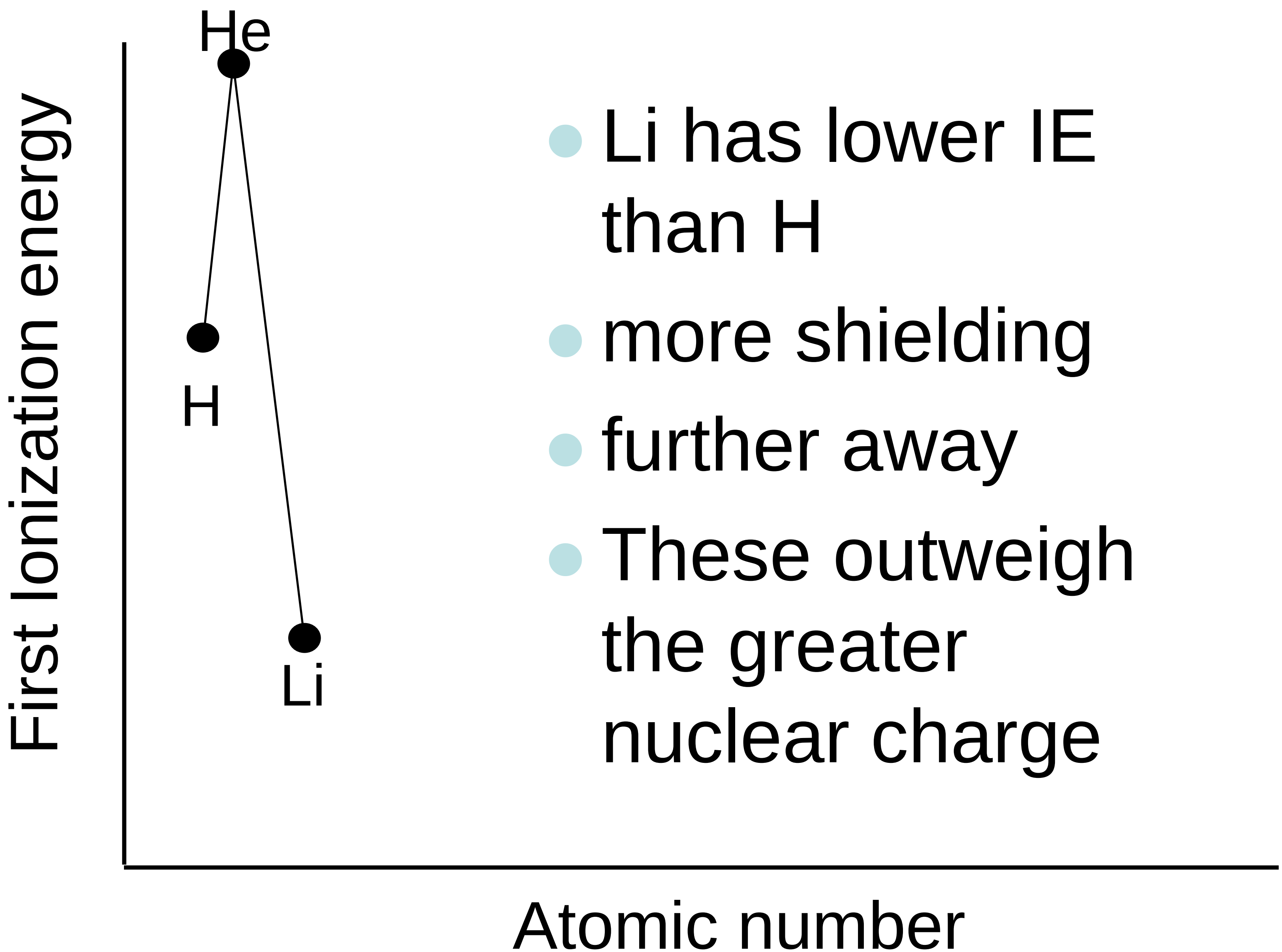
- As you go down a group, the first IE decreases because...
  - The electron is further away from the attraction of the nucleus, and
  - There is more shielding.

Ionization Energy - Period trends

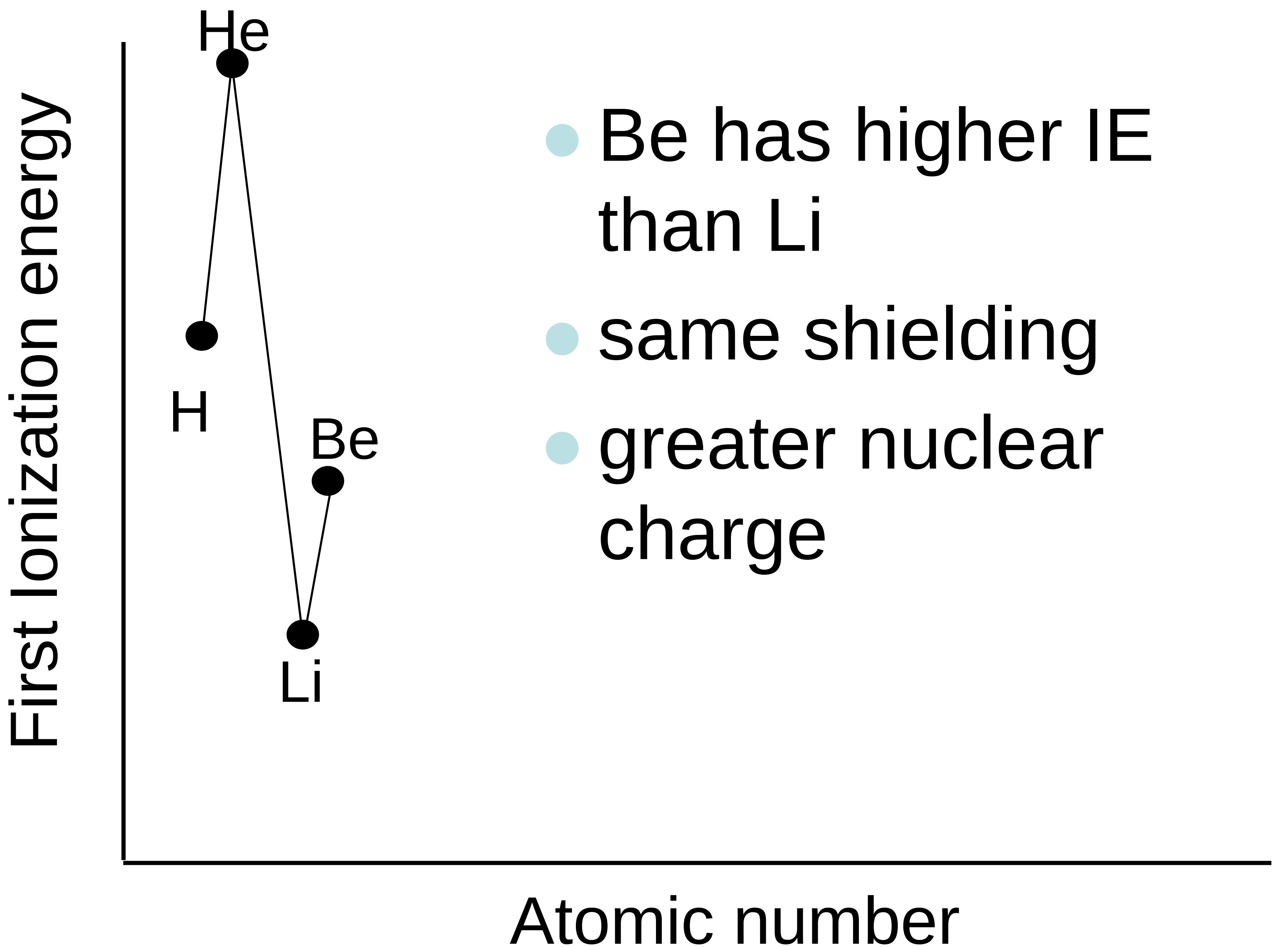
- All the atoms in the same period have the same energy level.
- Same shielding.
- But, increasing nuclear charge
- So IE generally increases from left to right.
- Exceptions at full and 1/2 full orbitals.



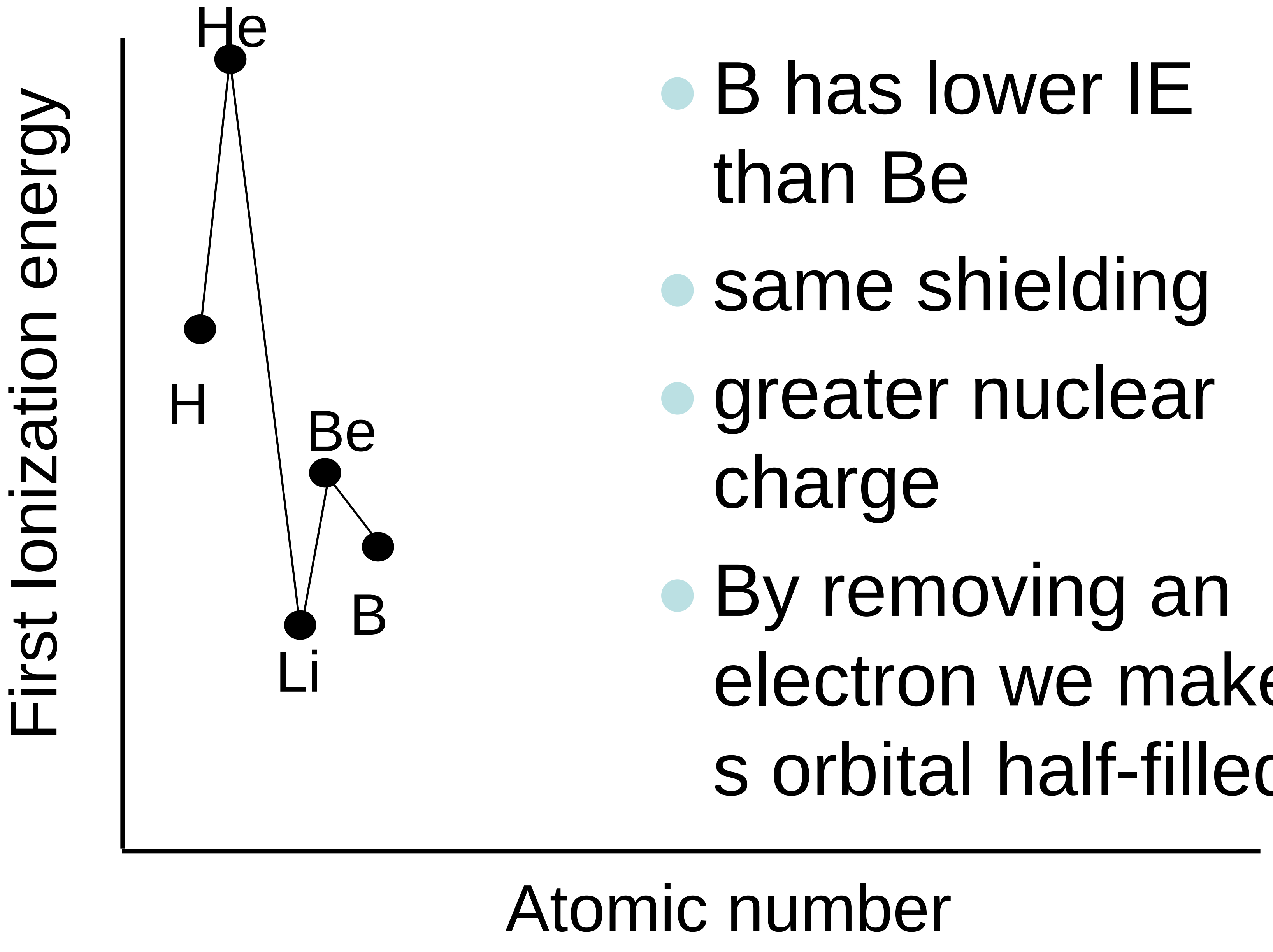
- He has a greater IE than H.
- Both elements have the same shielding since electrons are only in the first level
- But He has a greater nuclear charge



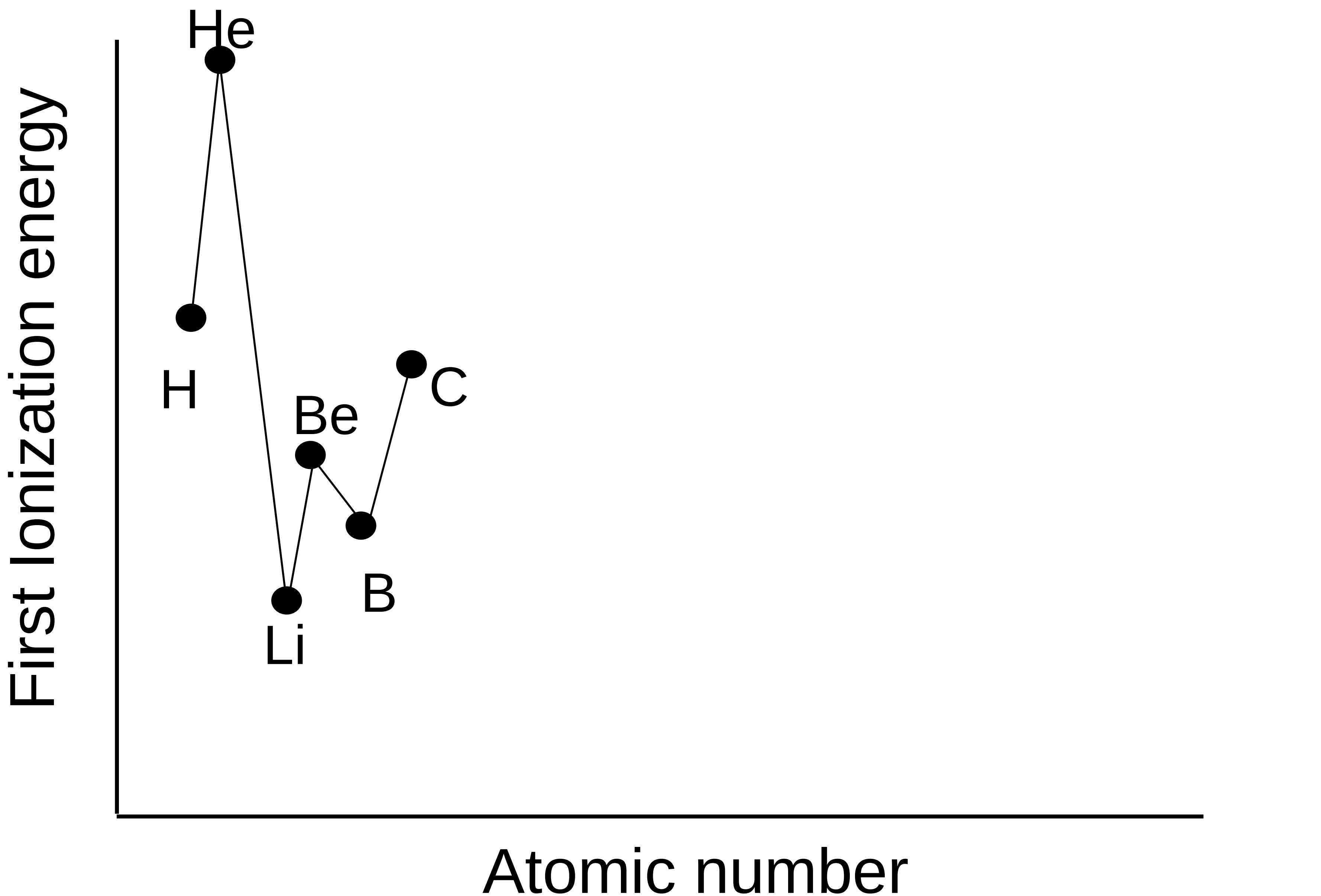
- Li has lower IE than H
- more shielding
- further away
- These outweigh the greater nuclear charge



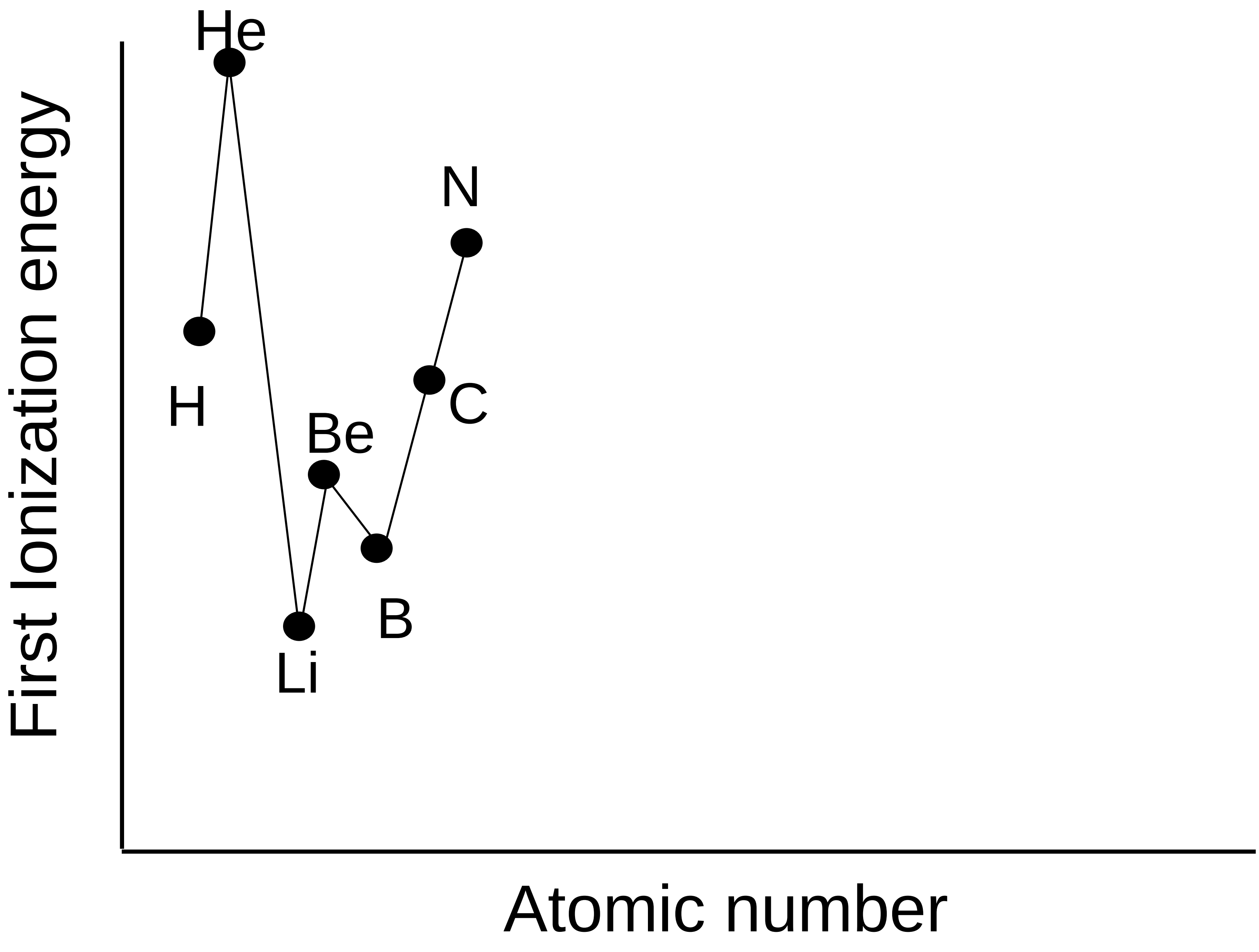
- Be has higher IE than Li
- same shielding
- greater nuclear charge

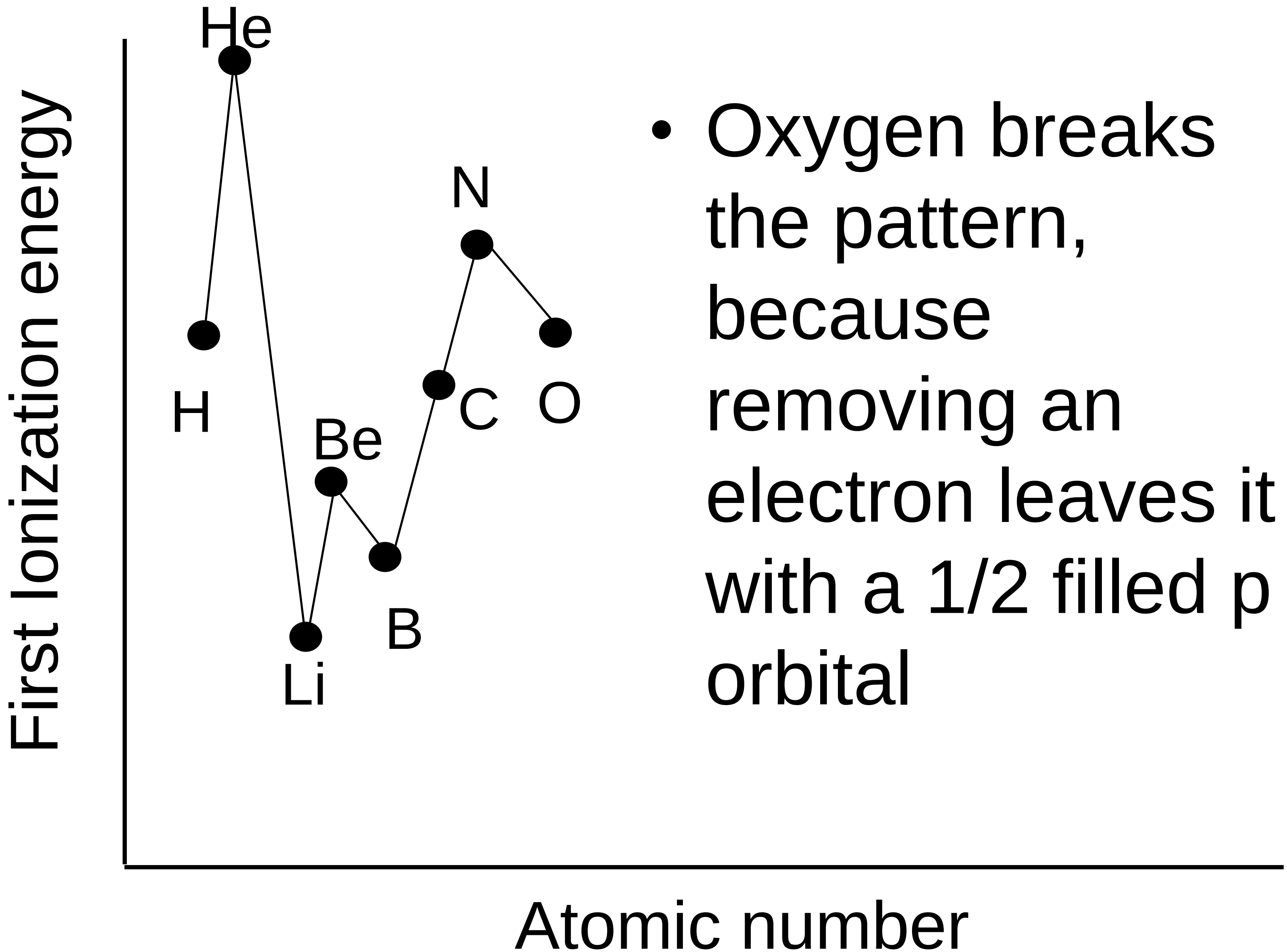


- B has lower IE than Be
- same shielding
- greater nuclear charge
- By removing an electron we make s orbital half-filled

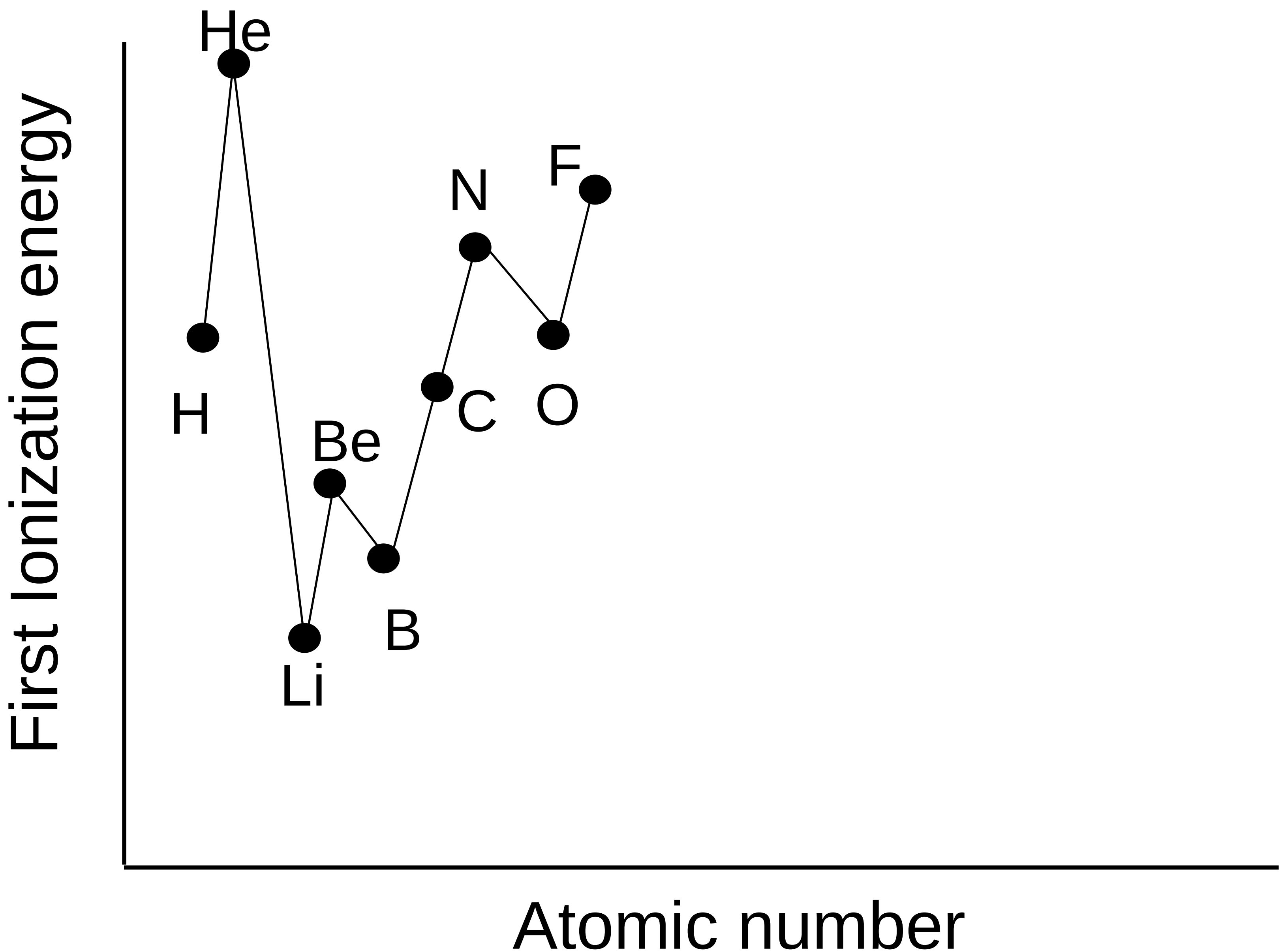


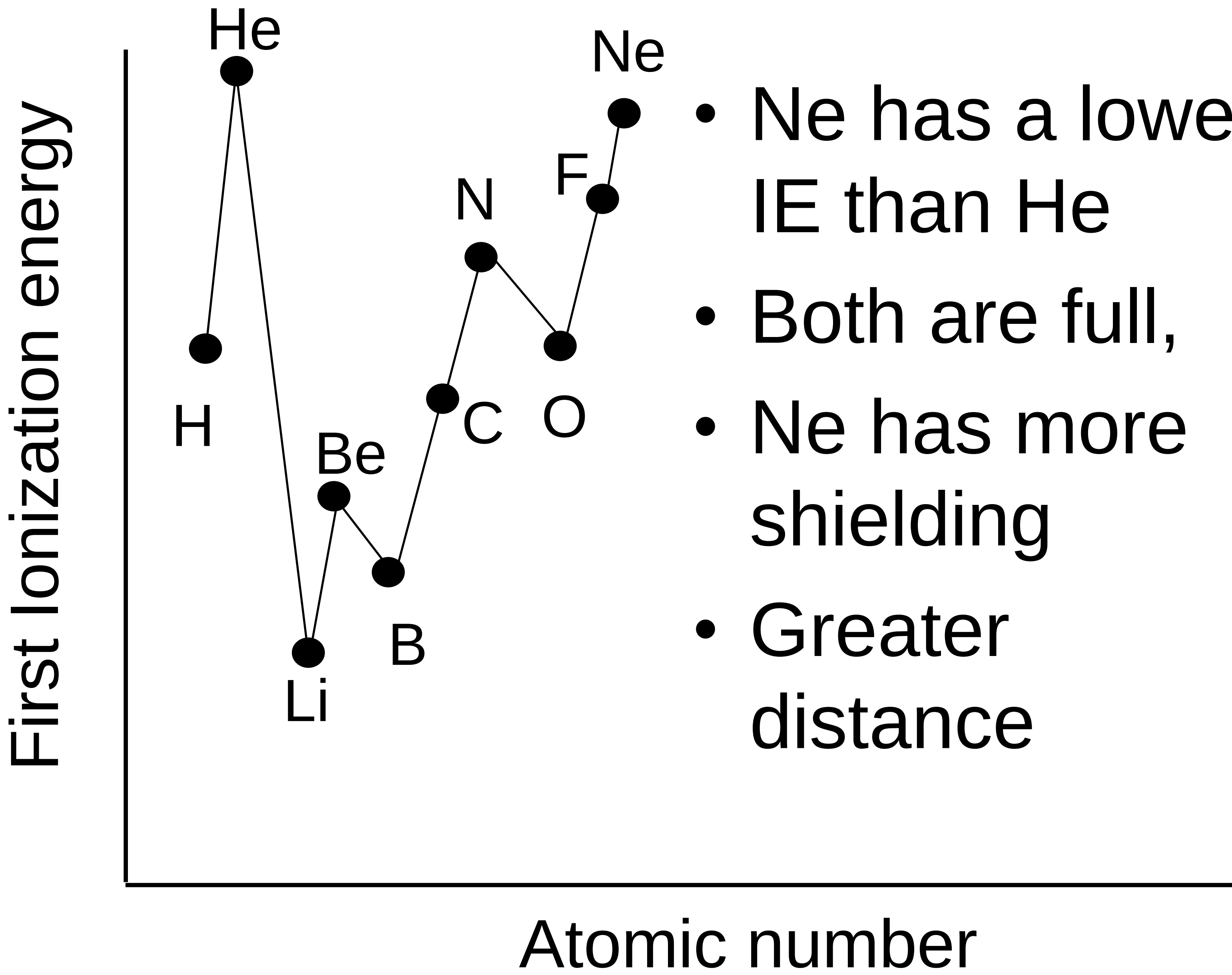
He
H
C
Be
B
Li
First Ionization energy
Atomic number


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

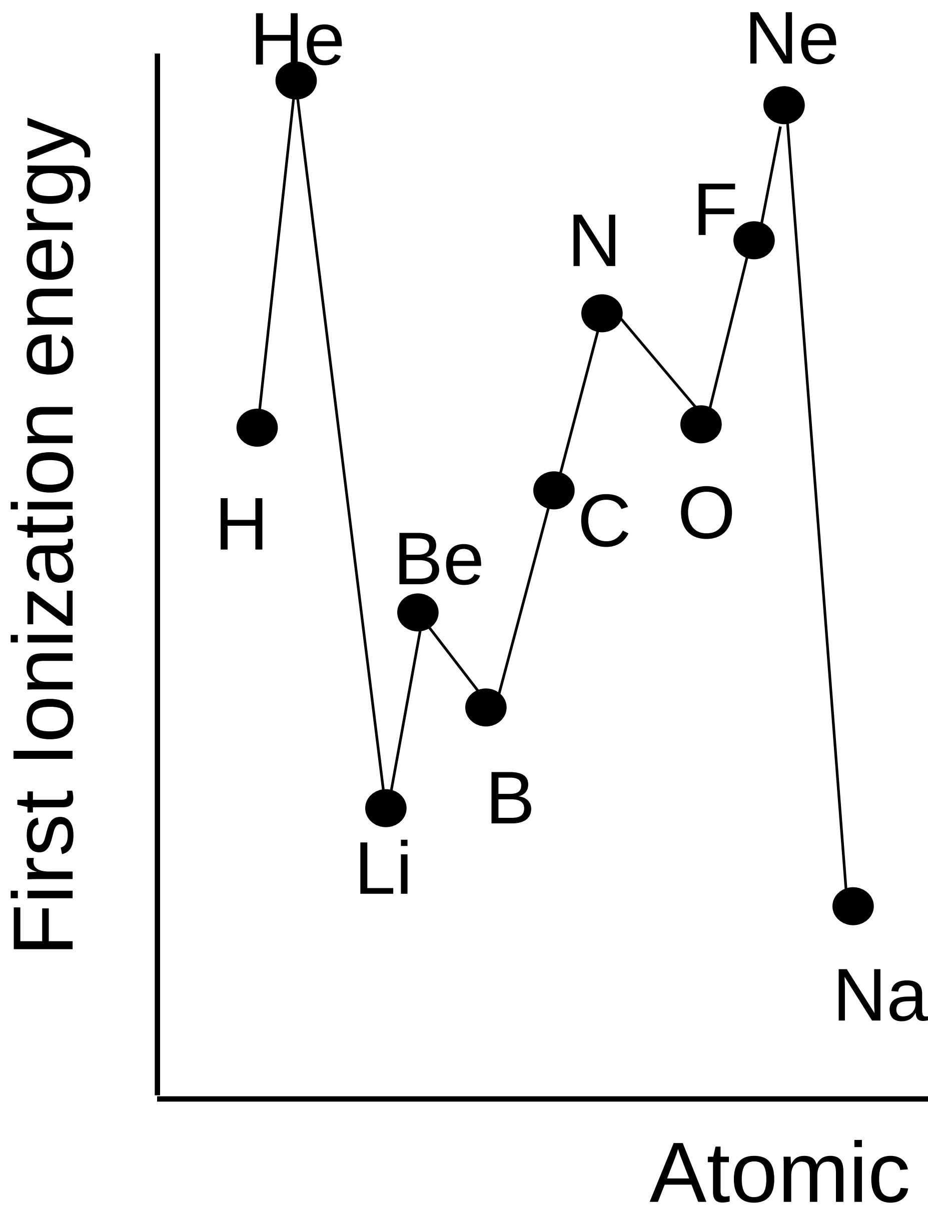


- Oxygen breaks the pattern, because removing an electron leaves it with a 1/2 filled p orbital

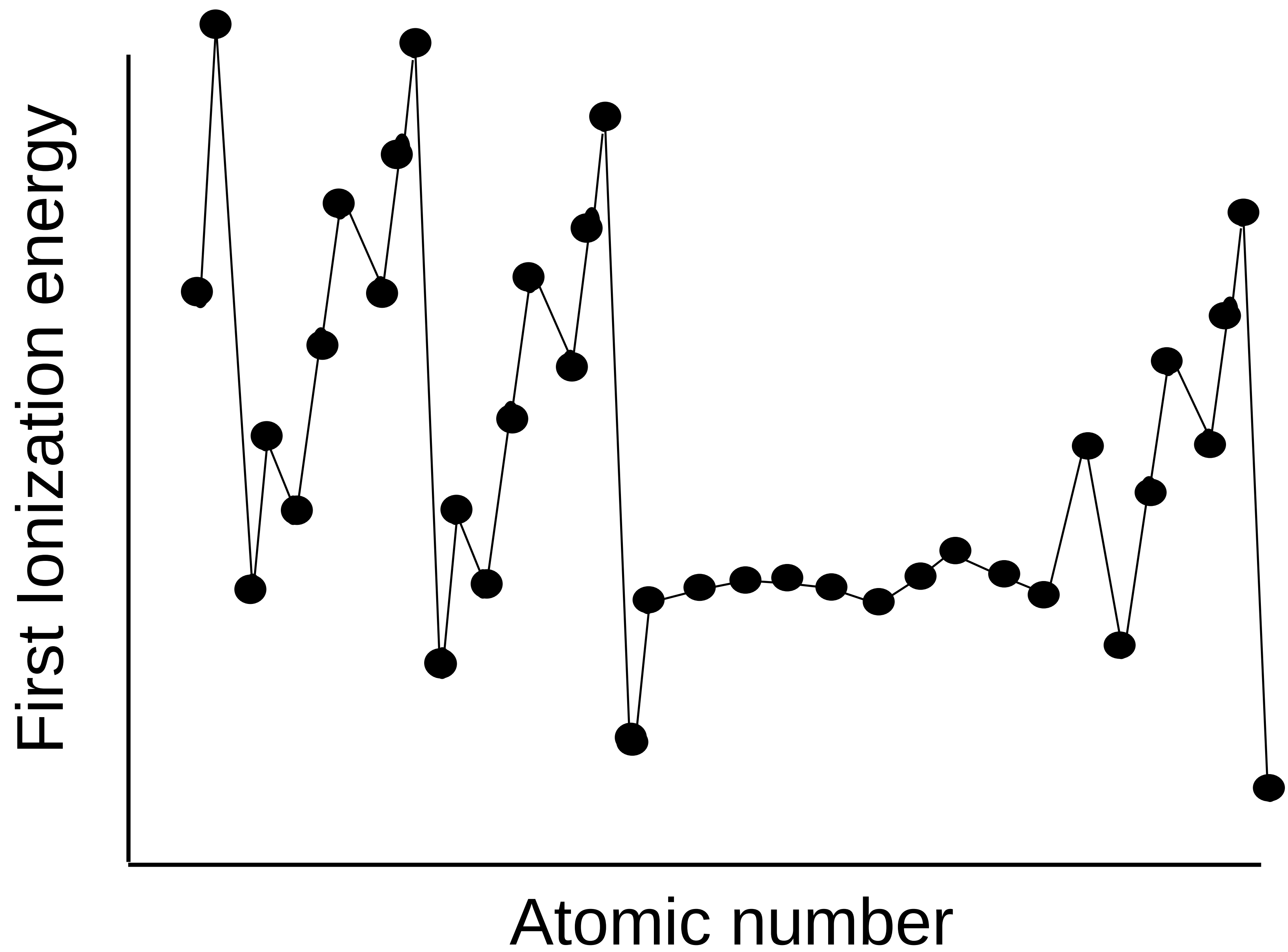

He
F
N
H
C
O
Be
B
Li
First Ionization energy
Atomic number



- 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 Forces

- *Full Energy Levels* require lots of energy to remove their electrons.
  - Noble Gases have full orbitals.
- Atoms behave in ways to try and achieve a 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 an low 3rd IE.
- Atoms in the aluminum family form 3+ ions.
- 2nd IE and 3rd IE are always higher than 1st IE!!!

Trends in Ionic Size: Cations

- Cations form by losing electrons.
- Cations are **smaller** than the atom they came from – not only do they lose electrons, they lose an *entire energy level*.
- Metals form cations.
- Cations of representative elements have the noble gas configuration *before* them.

Ionic size: Anions

- Anions form by <u>gaining</u> electrons.
- <u>Anions are **bigger** than the atom they came from</u> – have the same energy level, but a greater area the nuclear charge needs to cover
- Nonmetals form anions.
- Anions of representative elements have the noble gas configuration <u>*after*</u> them.

Configuration of Ions

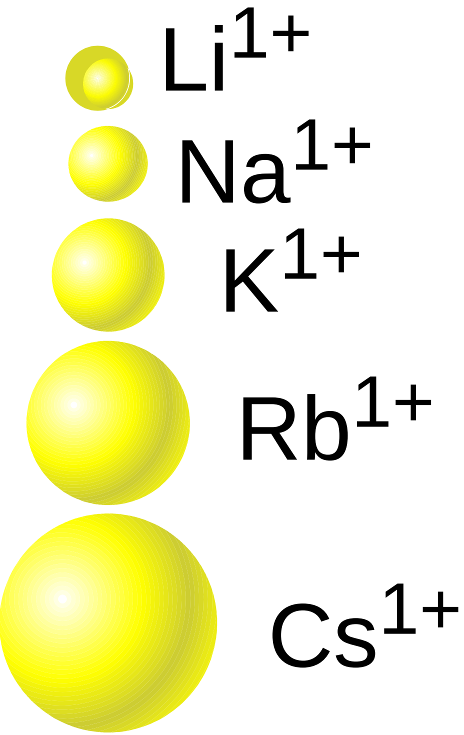
- Ions always have noble gas configurations (= a full outer level)
- Na atom is: $1s^2 2s^2 2p^6 3s^1$
- Forms a 1+ sodium ion: $1s^2 2s^2 2p^6$
- Same configuration as neon.
- Metals form ions with the configuration of the noble gas *before* them - they lose electrons.

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.

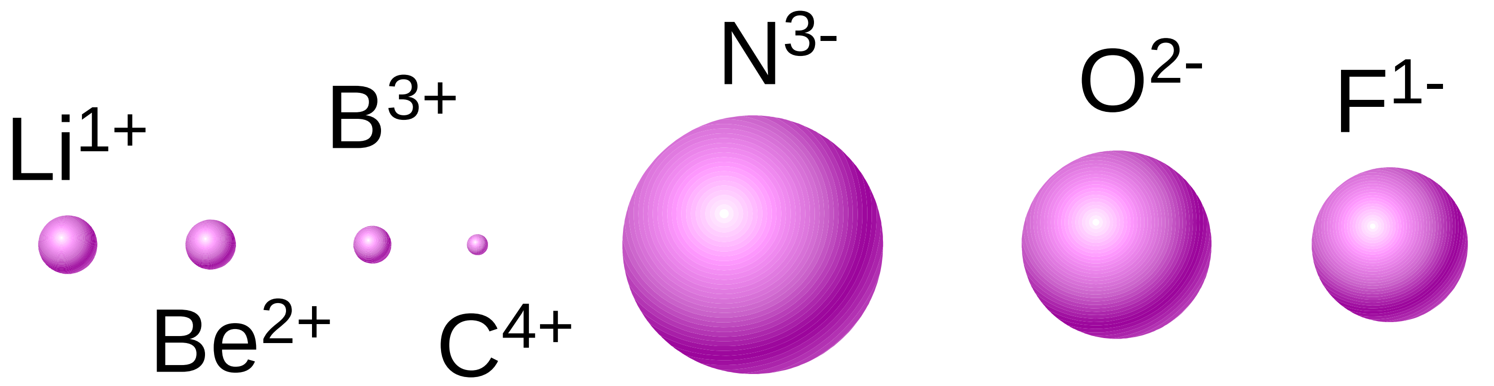
Ion Group trends

- Each step down a group is adding an energy level
- Ions therefore get <u>bigger</u> as you go down, because of the additional energy level.



Ion Period Trends

- Across the period from left to right, the nuclear charge increases - so they get smaller.
- Notice the *energy level changes* between anions and cations.

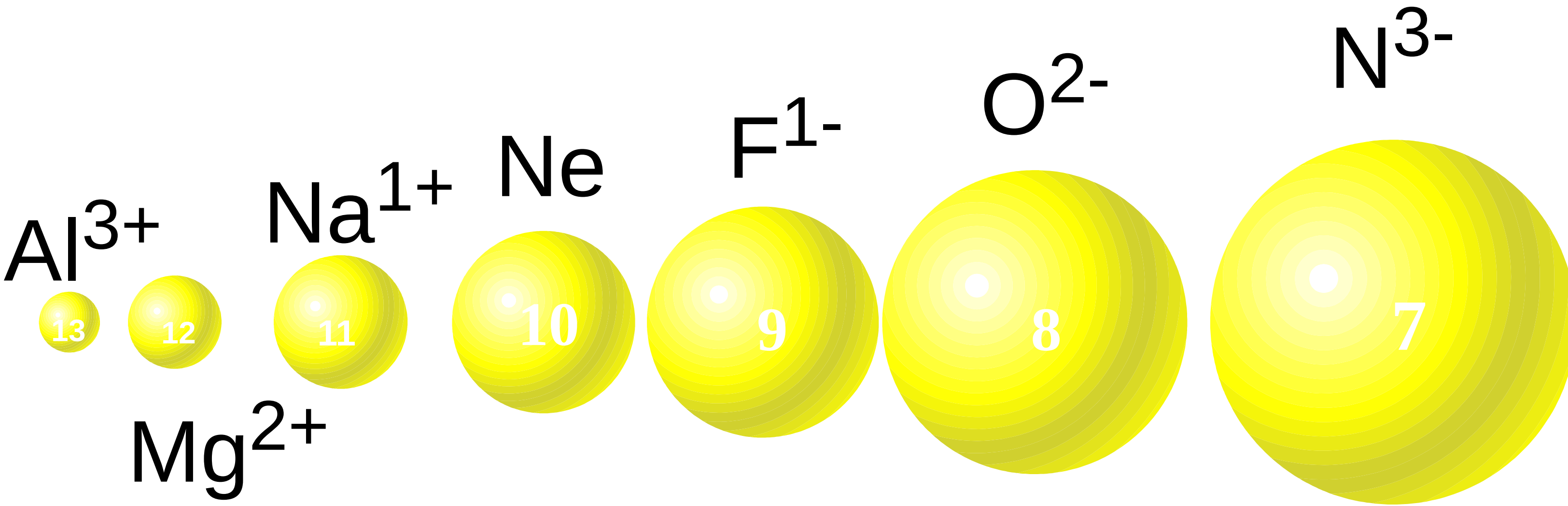


Size of Isoelectronic ions

- Iso- means “the same”
- Isoelectronic 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 same configuration: $1s^2 2s^2 2p^6$ **(which is the noble gas: neon)**

Size of Isoelectronic ions?

- Positive ions that have more protons would be <u>*smaller*</u> **(more protons would pull the same # of electrons in closer)**



#3. Trends in Electronegativity

- Electronegativity is the tendency for an atom to attract electrons to itself when it is chemically combined with another element.
- They share the electron, but how equally do they share it?
- An element with a big electronegativity means it pulls the electron towards itself strongly!

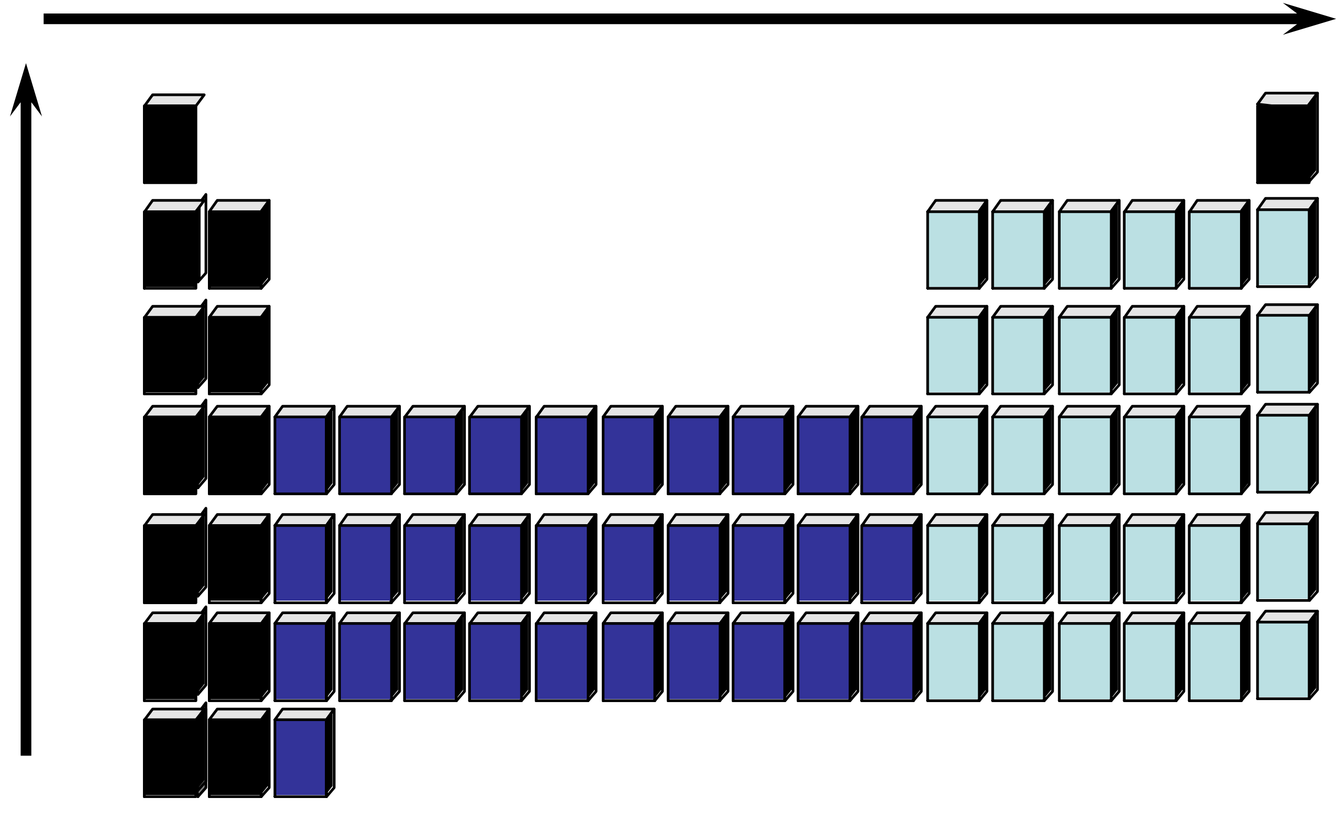
Electronegativity Group Trend

- The further down a group, the farther the electron is away from the nucleus, plus the more electrons an atom has.
- Thus, more willing to share.
- Low electronegativity.

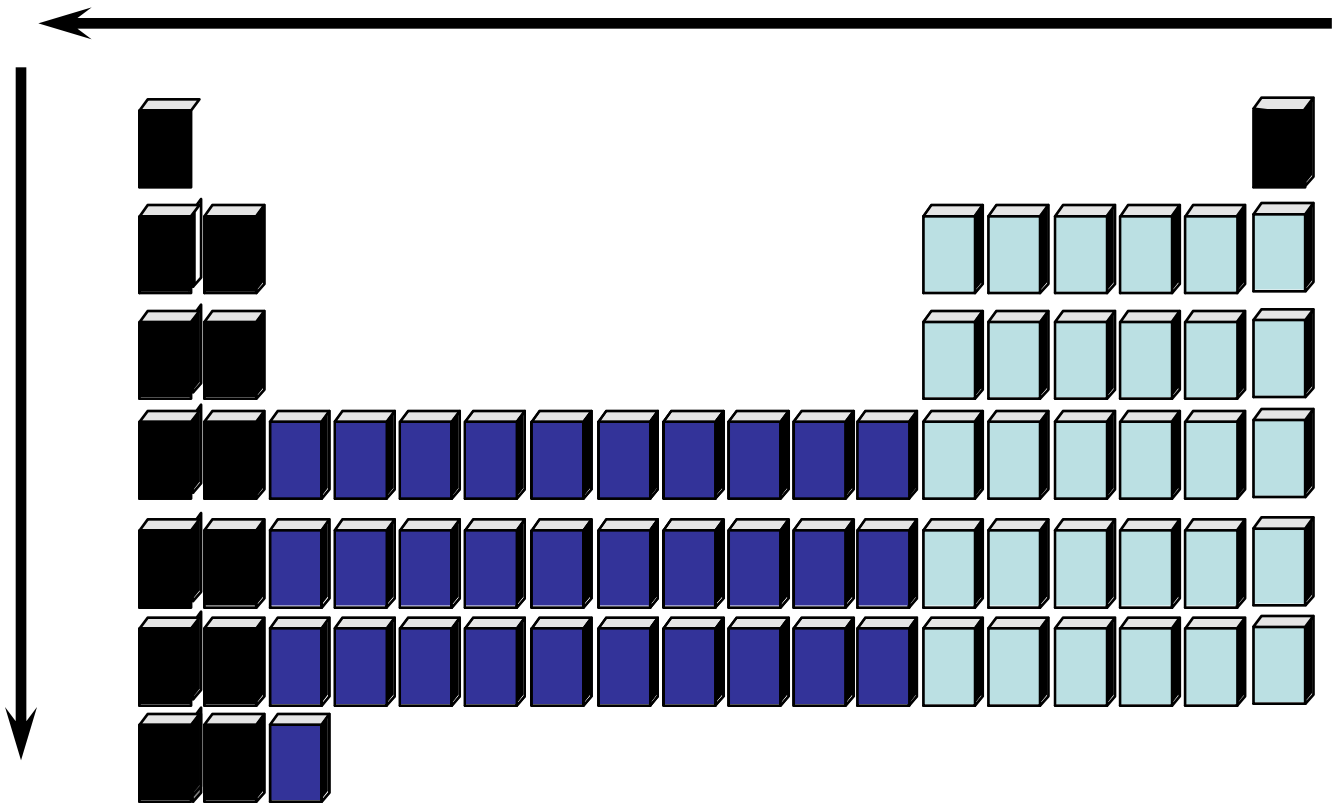
Electronegativity Period Trend

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

The arrows indicate the trend:
Ionization energy and Electronegativity
INCREASE in these directions



Atomic size and Ionic size increase in these directions:



Summary Chart of the trends:
Figure 6.22, p.178

End of Chapter 6