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 organize 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 triads – groups of three elements with similar properties
  - One element in each triad had *properties* intermediate of the other two elements

Mendeleev's Periodic Table

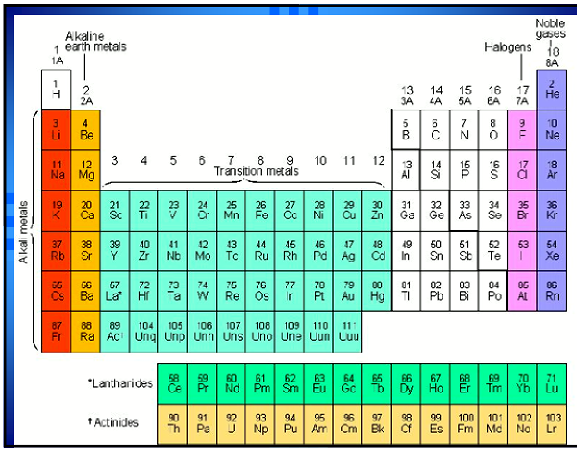
- By the mid-1800s, about 70 elements were known to exist
- Dmitri Mendeleev – a Russian chemist and teacher
- Arranged elements in order of increasing *atomic mass*
- Thus, the first "Periodic Table"

Mendeleev

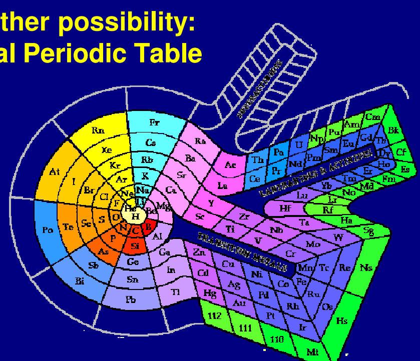
- ***He left blanks*** for yet undiscovered elements
  - When they were discovered, he had made good predictions
- But, there were problems:
  - Such as Co and Ni; Ar and K; Te and I

A better arrangement

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



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 = periods
  - There are 7 periods
- Vertical column = group (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) **Metals**: electrical conductors, have luster, ductile, malleable
2) **Nonmetals**: 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 metals and nonmetals

Section 6.2 Classifying the Elements

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

Section 6.2 Classifying the Elements

- **OBJECTIVES:**
  - Classify elements based on electron configuration.

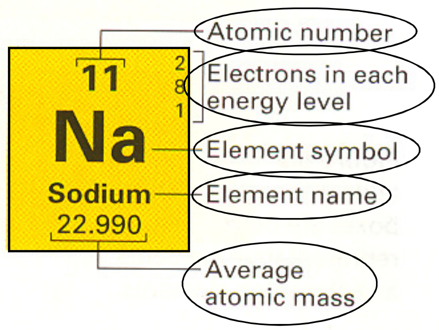
Section 6.2 Classifying the Elements

- **OBJECTIVES:**
  - Distinguish representative elements and transition metals.

Squares in the Periodic Table

- **The periodic table displays the symbols and names 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)



Groups of elements - **family names**

- **Group IA** – alkali metals
  - Forms a "base" (or alkali) when *reacting* with water (not just dissolved!)
- **Group 2A** – alkaline earth metals
  - Also form bases with water; do not dissolve well, hence "earth metals"
- **Group 7A** – halogens
  - 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

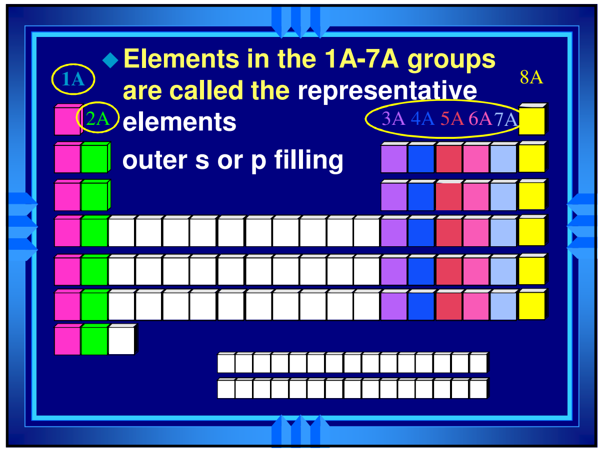
Electron Configurations in Groups

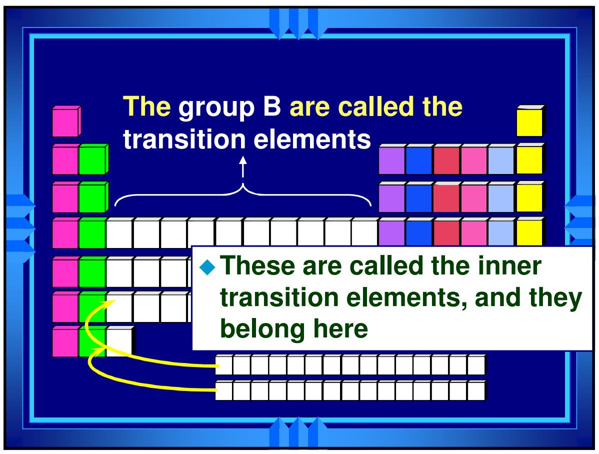
3) **Transition metals** are in the "B" columns of the periodic table
   - Electron configuration has the outer s sublevel full, and is now filling the "d" sublevel
   - A "transition" between the metal area and the nonmetal area
   - Examples are gold, copper, silver

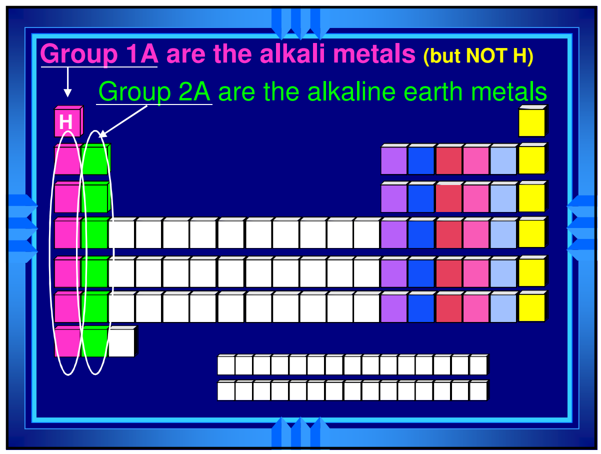
Electron Configurations in Groups

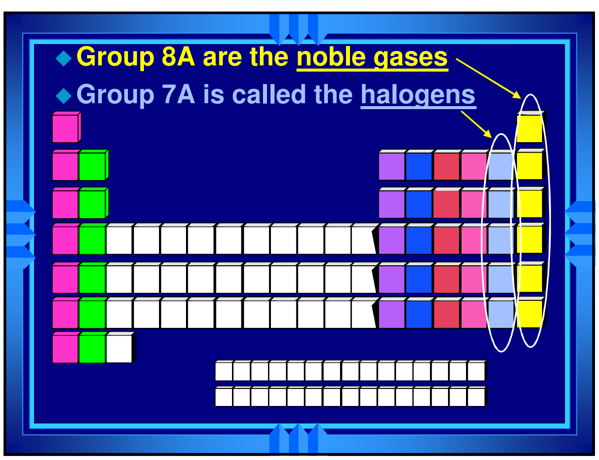
4) **Inner Transition Metals** are located below the main body of the table, in two horizontal rows

- 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



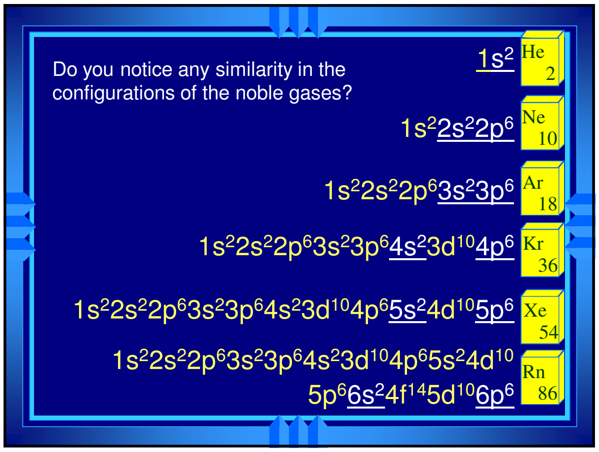


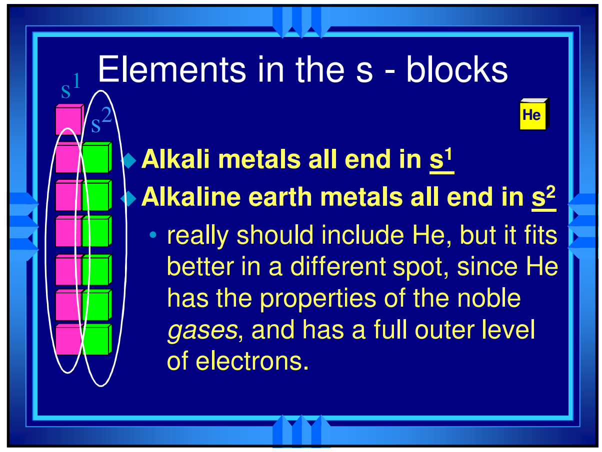


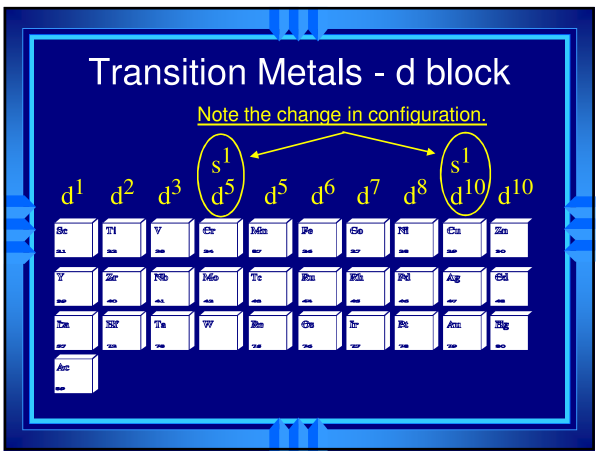


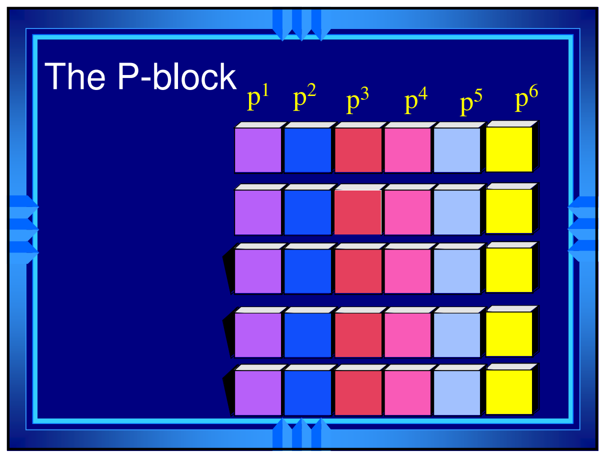
Do you notice any similarity in these configurations of the alkali metals?

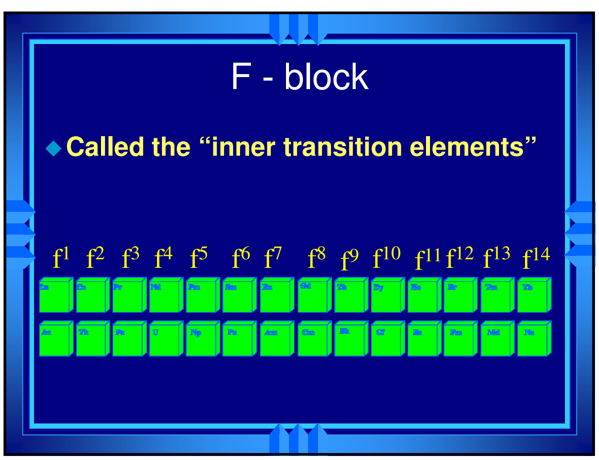
H 1	$1s^1$
Li 3	$1s^2 2s^1$
Na 11	$1s^2 2s^2 2p^6 3s^1$
K 19	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
Rb 37	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$
Cs 55	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$
Fr 87	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14}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

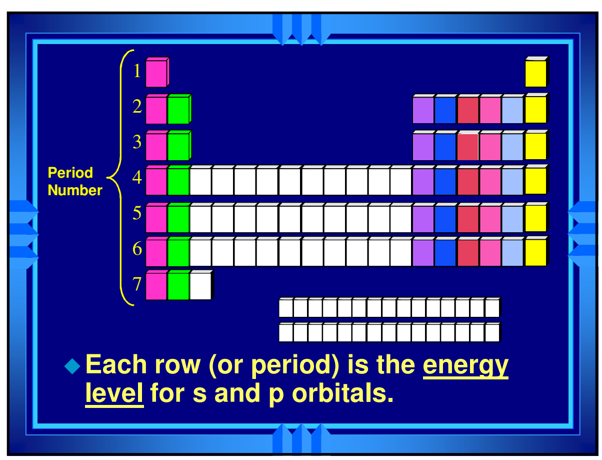

Do you notice any similarity in the configurations of the noble gases?
1s2 He 2
1s22s22p6 Ne 10
1s22s22p63s23p6 Ar 18
1s22s22p63s23p64s23d104p6 Kr 36
1s22s22p63s23p64s23d104p65s24d105p6 Xe 54
1s22s22p63s23p64s23d104p65s24d10 5p66s24f145d106p6 Rn 86


Elements in the s - blocks
s1
s2
He
Alkali metals all end in s1
Alkaline earth metals all end in s2
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.

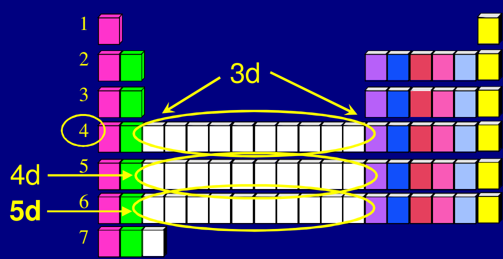

Transition Metals - d block
Note the change in configuration.
s1 d5
s1 d10
d1 d2 d3 d5 d6 d7 d8 d10

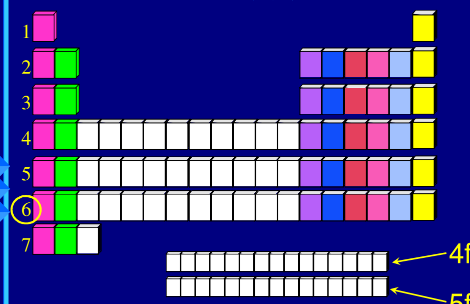

The P-block
p1 p2 p3 p4 p5 p6


F - block
Called the "inner transition elements"
f1 f2 f3 f4 f5 f6 f7 f8 f9 f10 f11 f12 f13 f14


Period Number
1
2
3
4
5
6
7
Each row (or period) is the energy level for s and p orbitals.

- **The “d” orbitals fill up in levels <u>1 less</u> 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:**
  - <u>Describe</u> *trends* among the elements for *atomic size*.

Section 6.3
Periodic Trends

- **OBJECTIVES:**
  - <u>Explain</u> how **ions** form.

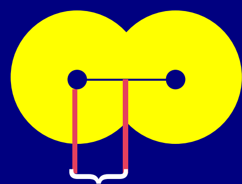
Section 6.3
Periodic Trends

- **OBJECTIVES:**
  - <u>Describe</u> 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

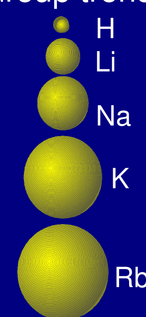
- **Influenced by three factors:**
  - **1. Energy Level**
    - Higher energy levels are further away from the nucleus.
  - **2. Charge on nucleus (# protons)**
    - More charge pulls electrons in closer. (+ and – attract each other)
- **3. Shielding effect** (blocking effect?)

What do they influence?

- **Energy levels and Shielding have an effect on the *GROUP* (∢)**
- **Nuclear charge has an effect on a *PERIOD* (⌧)**

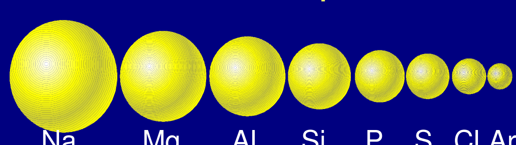
#1. Atomic Size - Group trends

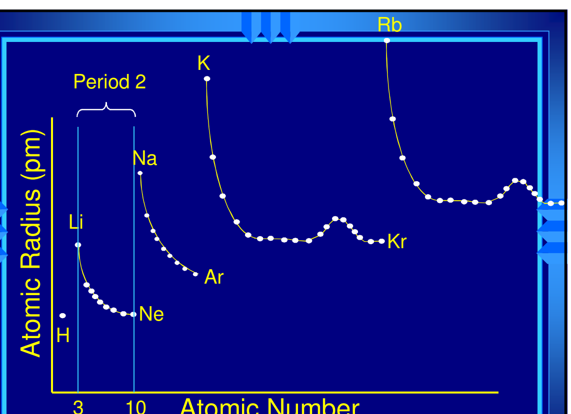
- **As we increase the atomic number (or go down a group). . .**
- **each atom has another energy level,**
- **so the atoms get *bigger*.**



#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.**





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

- Metals tend to LOSE electrons, 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 = "**cation**"
  - 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 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.

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

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Ne	2080	3963	6276

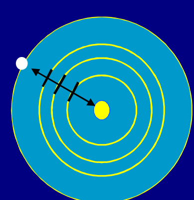
Why did these values increase **so much**?

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>**

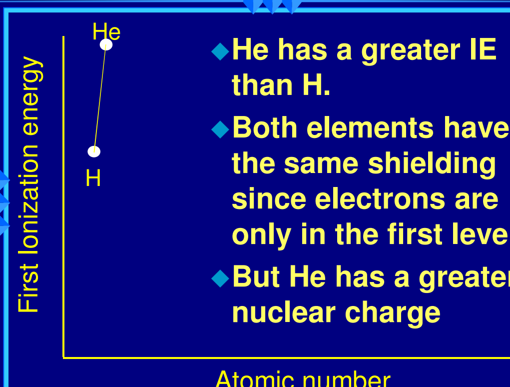


Ionization Energy - Group trends

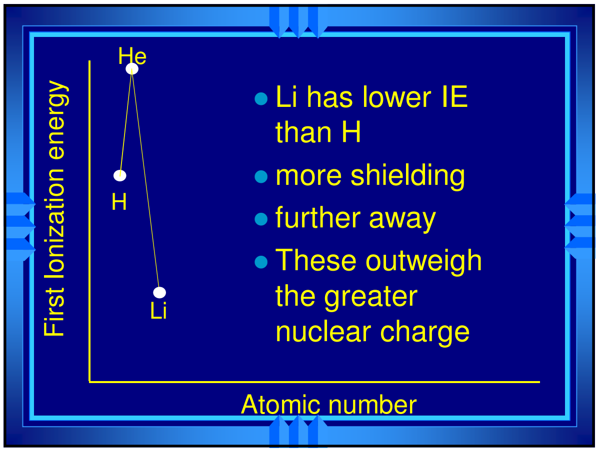
- **As you go down a group, the first IE decreases because...**
  - The electron is further away from the attraction of the nucleus, and
  - There is more shielding.

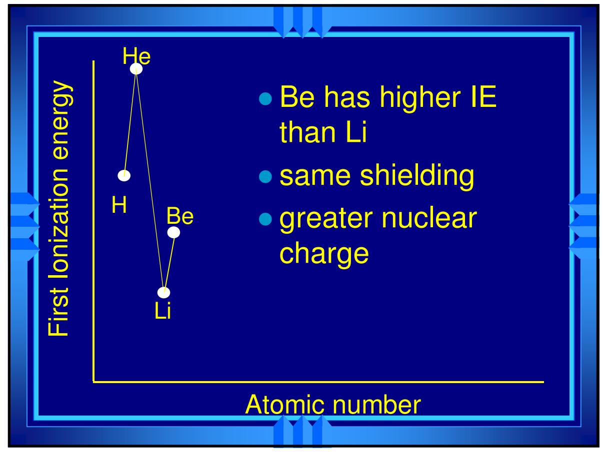
Ionization Energy - Period trends

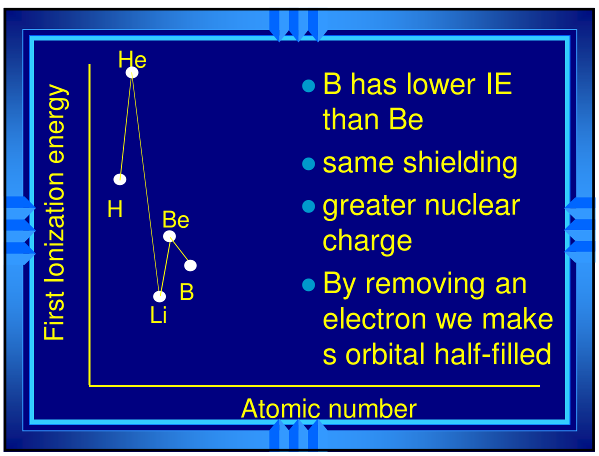
- **All the atoms in the same period have the same energy level.**
- **Same shielding.**
- **But, increasing nuclear charge**
- **So IE generally <u>increases</u> from left to right.**
- **Exceptions at full and 1/2 full orbitals.**

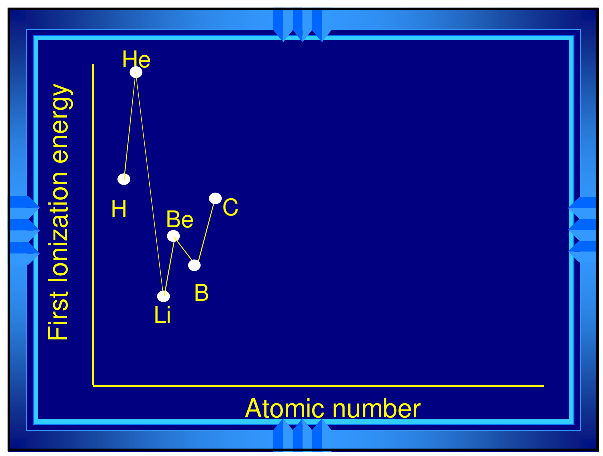


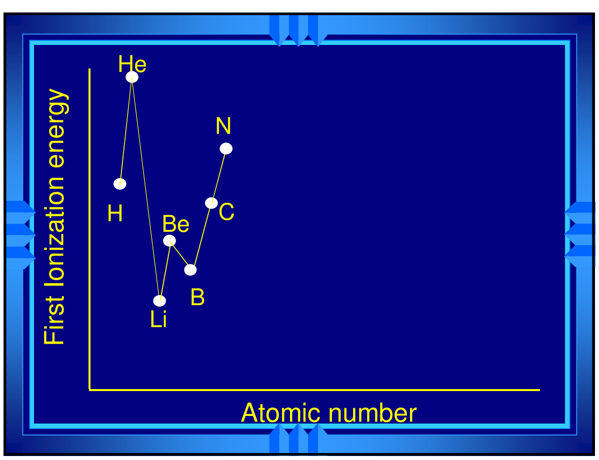
- **He has a greater IE than H.**
- **Both elements have the same shielding since electrons are only in the first level**
- **But He has a greater nuclear charge**

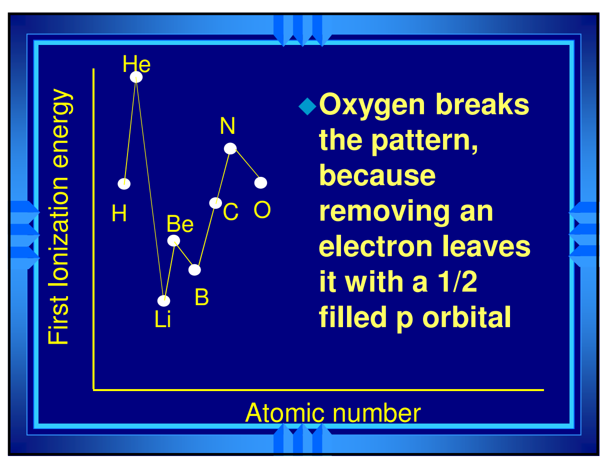

He
H
Li
First Ionization energy
Atomic number
Li has lower IE than H
more shielding
further away
These outweigh the greater nuclear charge

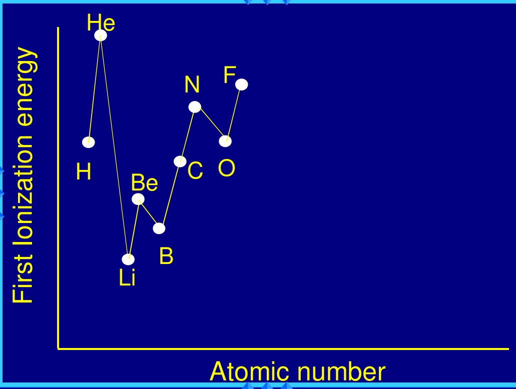

He
H
Be
Li
First Ionization energy
Atomic number
Be has higher IE than Li
same shielding
greater nuclear charge

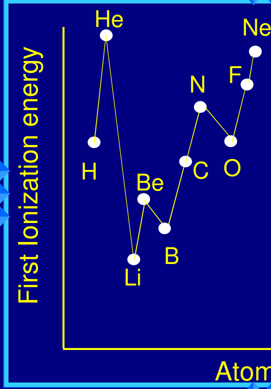

He
H
Be
B
Li
First Ionization energy
Atomic number
B has lower IE than Be
same shielding
greater nuclear charge
By removing an electron we make s orbital half-filled


He
H
Be
C
B
Li
First Ionization energy
Atomic number

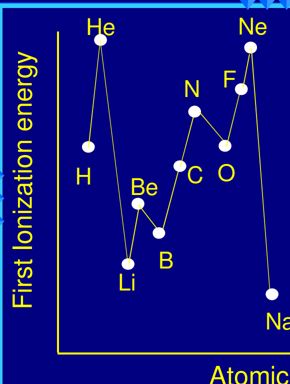

He
N
H
Be
C
B
Li
First Ionization energy
Atomic number


He
N
H
Be
C
O
B
Li
First Ionization energy
Atomic number
Oxygen breaks the pattern, because removing an electron leaves it with a 1/2 filled p orbital

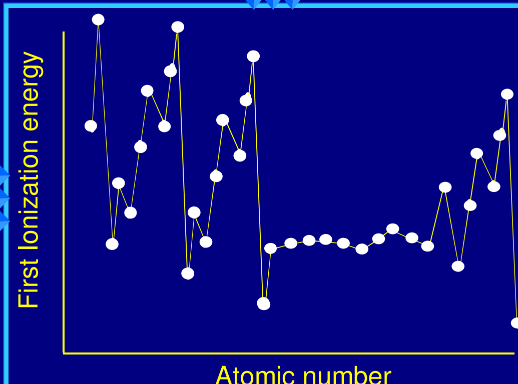




- **Ne has a lower IE than He**
- **Both are full,**
- **Ne has more shielding**
- **Greater distance**



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



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 gaining electrons.
- Anions are **bigger** than the atom they came from – 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 *after* them.

Configuration of Ions

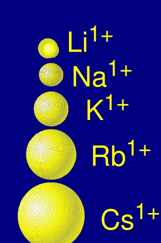
- Ions always have noble gas configurations (= a full outer level)
- Na atom is: $1s^22s^22p^63s^1$
- Forms a 1+ sodium ion: $1s^22s^22p^6$
- Same configuration as neon.
- Metals form ions with the configuration of the noble gas *before* them - they lose electrons.

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.

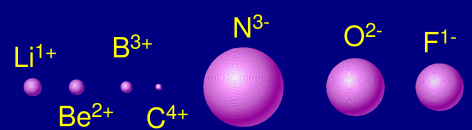
Ion Group trends

- Each step down a group is adding an energy level
- Ions therefore get bigger 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 <u>smaller.</u>**
- **Notice the *energy level changes* between anions and cations.**

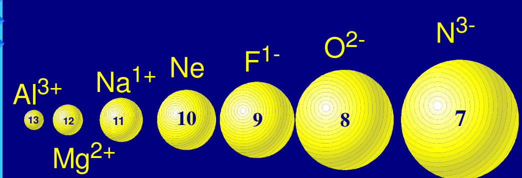


Size of Isoelectronic ions

- **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

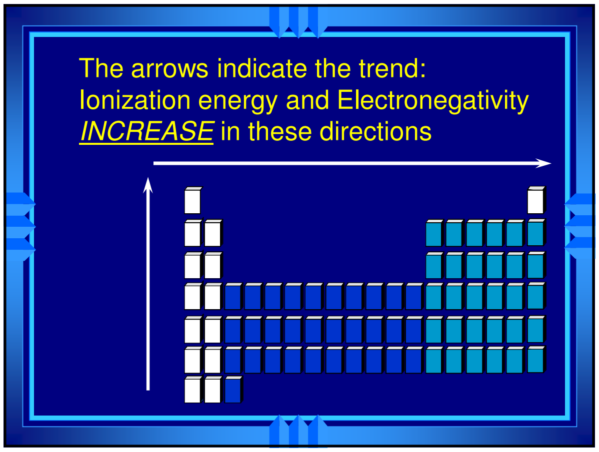
- **Electronegativity is the tendency for an atom to <u>attract</u> electrons to itself when it is <u>chemically combined</u> with another element.**
- **They share the electron, but how equally do they share it?**
- **An element with a big electronegativity means it pulls the electron towards itself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

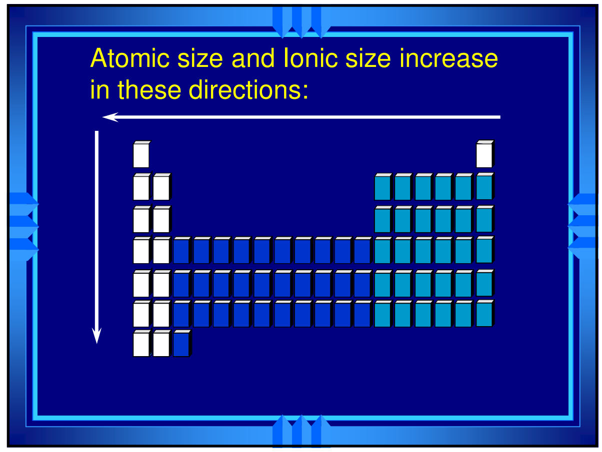
Electronegativity Group Trend

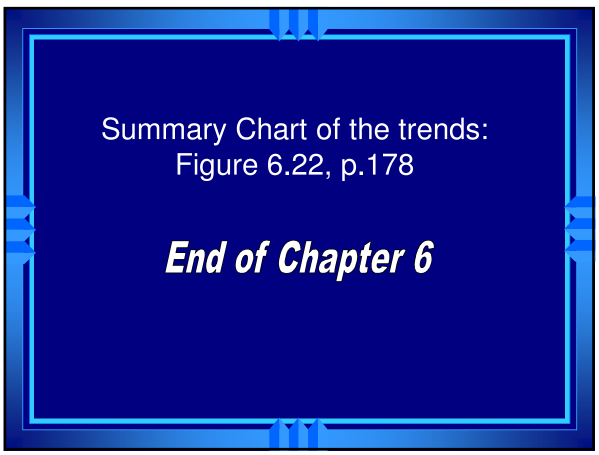
- **The further down a group, the farther the electron is away from the nucleus, plus the more electrons an atom has.**
- **Thus, more willing to share.**
- **Low electronegativity.**

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 <u>more</u> 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