Chapter 6
The Periodic Table

The how and why

History

- **1829 German J. W. Dobereiner Grouped elements into triads**
  - Three elements with similar properties
  - Properties followed a pattern
  - The same element was in the middle of all trends
- **Not all elements had triads**

History

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

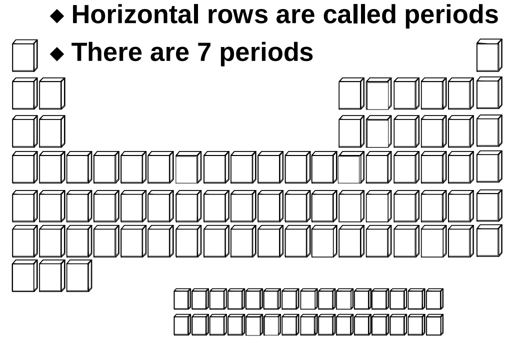
Mendeleev's Table

- **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**

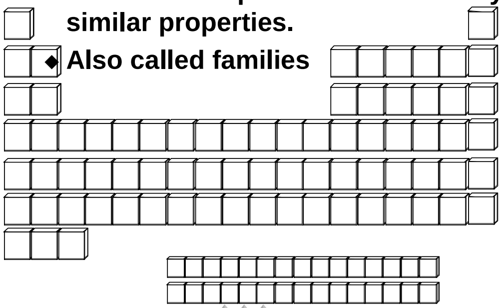
The Modern Table

- **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.**

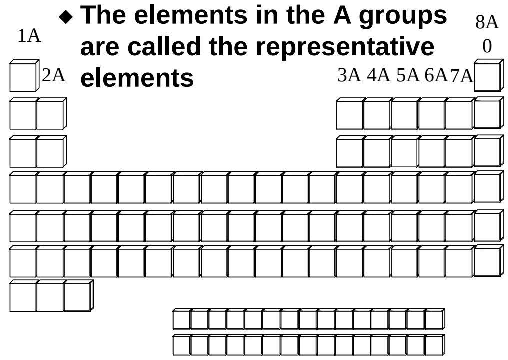
- **Horizontal rows are called periods**
- **There are 7 periods**



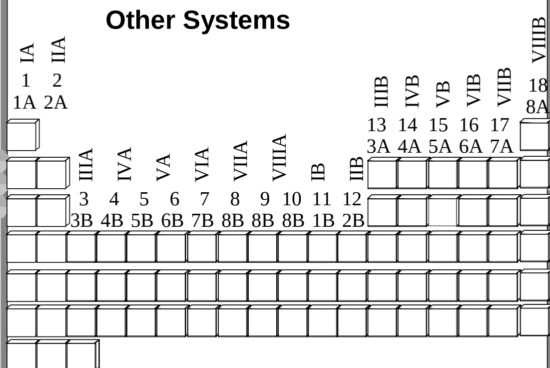
- Vertical columns are called groups.
- Elements are placed in columns by similar properties.
- Also called families



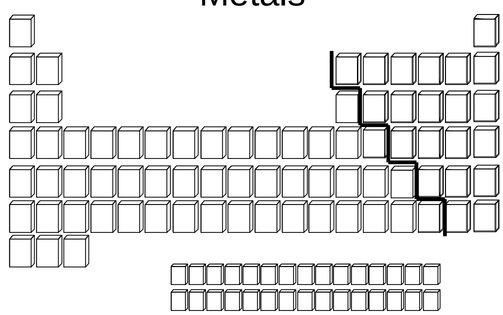
- The elements in the A groups are called the representative elements



Other Systems



Metals

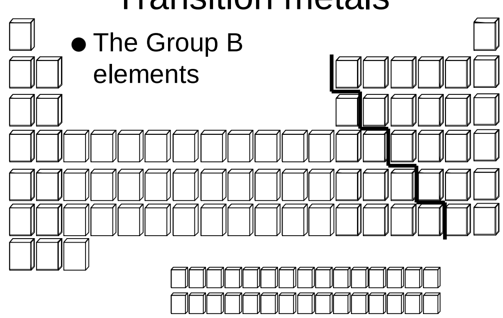


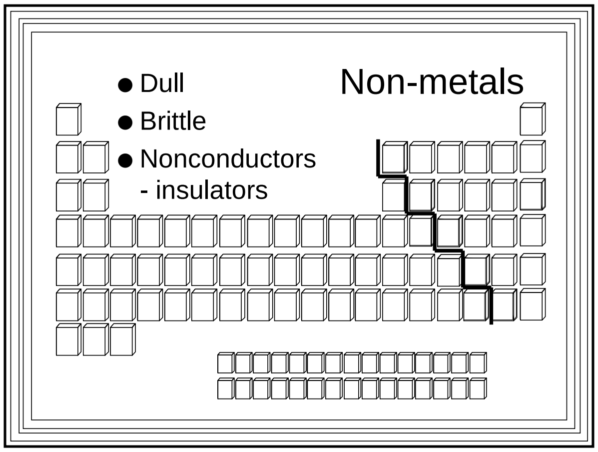
Metals

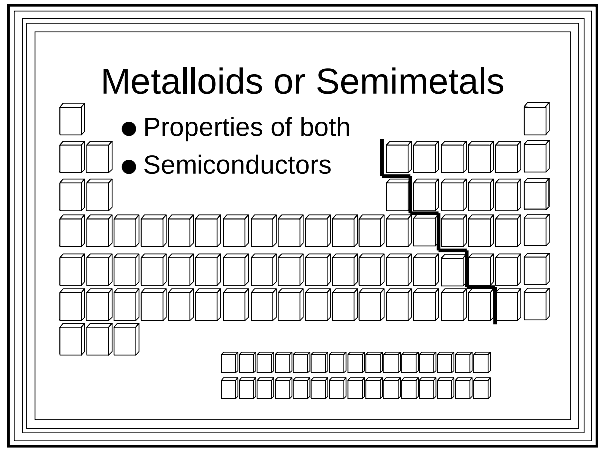
- Luster – shiny.
- Ductile – drawn into wires.
- Malleable – hammered into sheets.
- Conductors of heat and electricity.

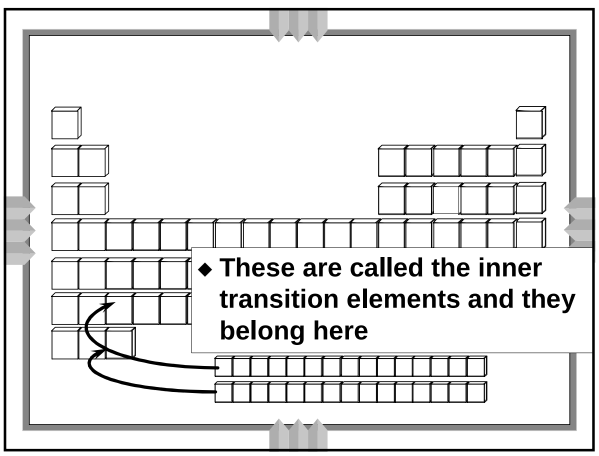
Transition metals

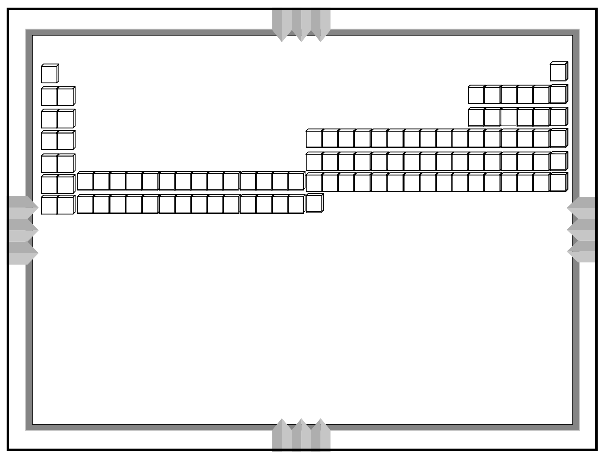
- The Group B elements

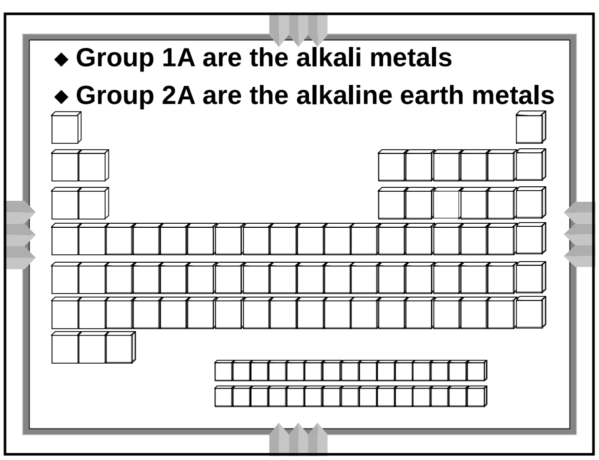


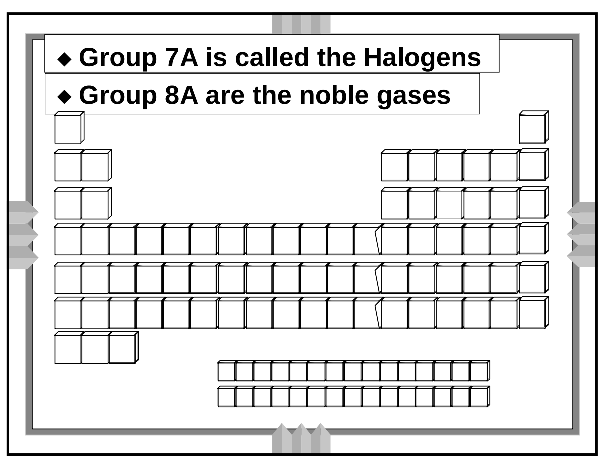

Non-metals
● Dull
● Brittle
● Nonconductors
- insulators


Metalloids or Semimetals
● Properties of both
● Semiconductors


◆ 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
- So.. 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 $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$

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

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

$1s^2$ He 2

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

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

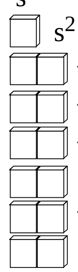
$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$ Kr 36

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

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

S- block

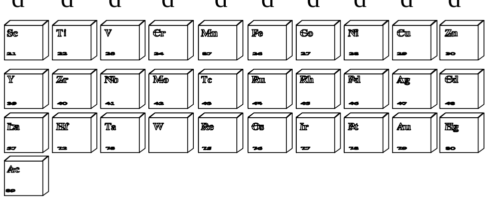
s^1 s^2



- 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

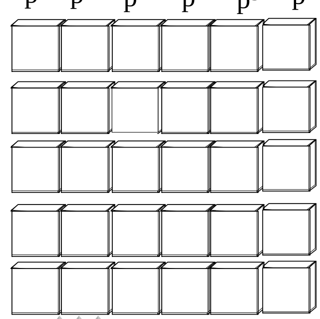
Transition Metals -d block

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



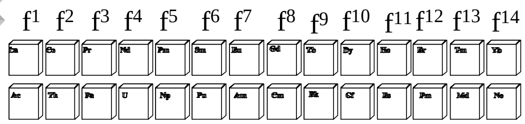
The P-block

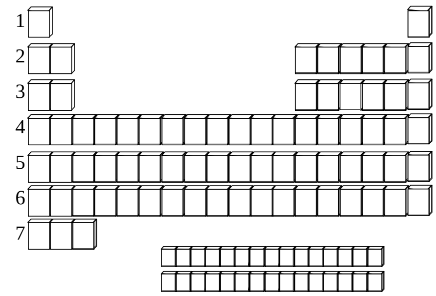
p^1 p^2 p^3 p^4 p^5 p^6



F - block

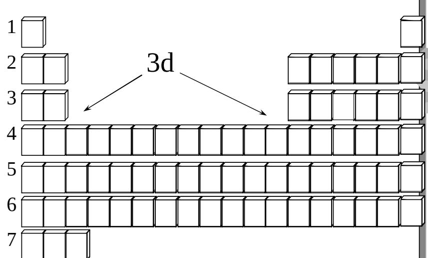
- inner transition elements

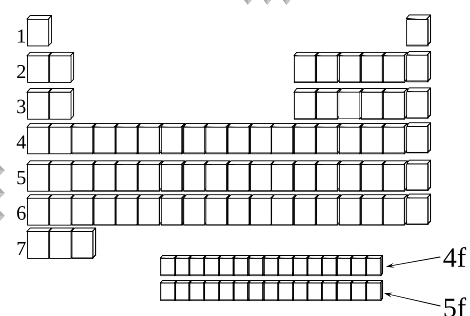




- Each row (or period) is the energy level for s and p orbitals

- 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

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 row the outermost energy level is full.
- This is the basis for our shorthand

The Shorthand

- Write the symbol of the noble gas before the element in brackets [ ]
- Then the rest of the electrons.
- Aluminum - full configuration
- $1s^2 2s^2 2p^6 3s^2 3p^1$
- Ne is $1s^2 2s^2 2p^6$
- so Al is [Ne] $3s^2 3p^1$

More examples

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

The Shorthand

Sn- 50 electrons

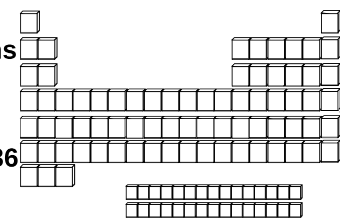
The noble gas before it is Kr

Takes care of 36

Next $5s^2$

Then $4d^{10}$

Finally $5p^2$



[Kr] $5s^2 4d^{10} 5p^2$

Practice

- Write the shorthand configuration for
- S
- Mn
- Mo
- W

Electron configurations and groups

- Representative elements have s and p orbitals as last filled
  - Group number = number of electrons in highest energy level
- Transition metals- d orbitals
- Inner transition- f orbitals
- Noble gases s and p orbitals full

Part 3
Periodic trends

Identifying the patterns

What we will investigate

- **Atomic size**
  - how big the atoms are
- **Ionization energy**
  - How much energy to remove an electron
- **Electronegativity**
  - The attraction for the electron in a compound
- **Ionic size**
  - How big ions are

What we will look for

- **Periodic trends-**
  - How those 4 things vary as you go across a period
- **Group trends**
  - How those 4 things vary as you go down a group
- **Why?**
  - Explain why they vary

The why first

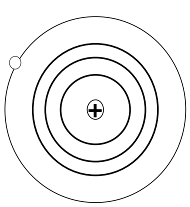
- **The positive nucleus pulls on electrons**
- **Periodic trends – as you go across a period**
  - The charge on the nucleus gets bigger
  - The outermost electrons are in the same energy level
  - So the outermost electrons are pulled stronger

The why first

- **The positive nucleus pulls on electrons**
- **Group Trends**
  - As you go down a group
    - You add energy levels
    - Outermost electrons not as attracted by the nucleus

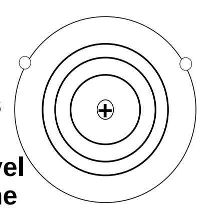
Shielding

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



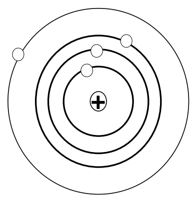
Shielding

- **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**
- **In the same energy level (period) shielding is the same**



Shielding

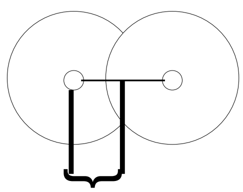
- As the energy levels changes the shielding changes
- Lower down the group
  - More energy levels
  - More shielding
  - Outer electron less attracted



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



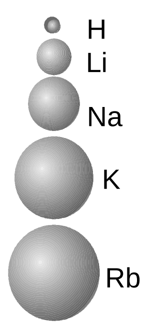
- **Atomic Radius = half the distance between two nuclei of molecule**

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

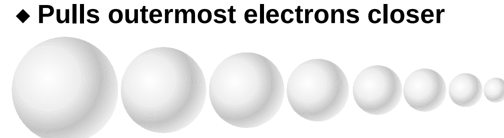
Group trends

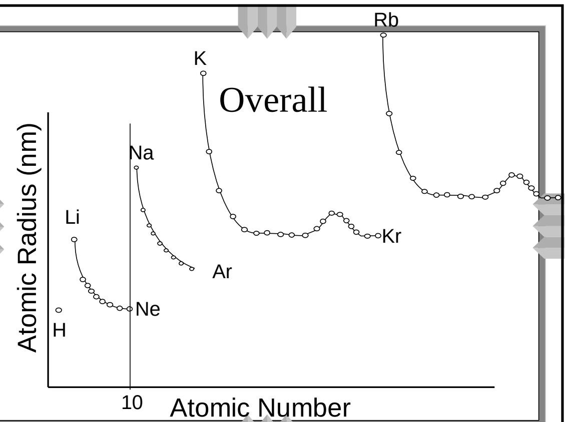
- As we go down a group
- Each atom has another energy level
- More shielding
- So the atoms get bigger

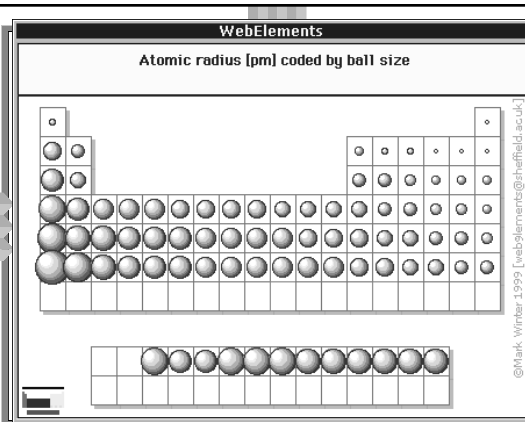


Periodic Trends

- As you go across a period the radius gets smaller.
- Same shielding and energy level
- More nuclear charge
- Pulls outermost electrons closer







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 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

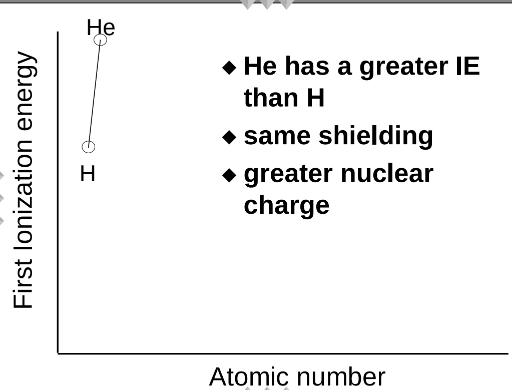
- The greater the nuclear charge the greater IE.
- Increased shielding decreases IE
- Filled and half filled orbitals have lower energy, so achieving them is easier, lower IE

Group trends

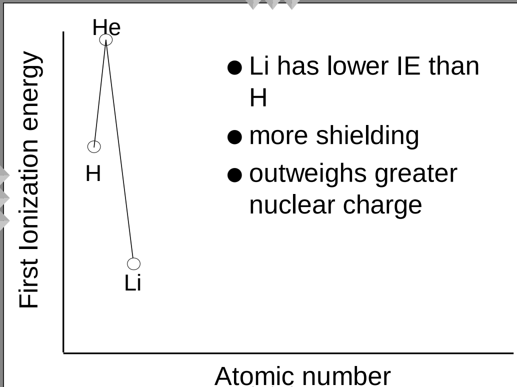
- **As you go down a group first IE decreases because of**
- **More shielding**
- **So outer electron less attracted**

Periodic trends

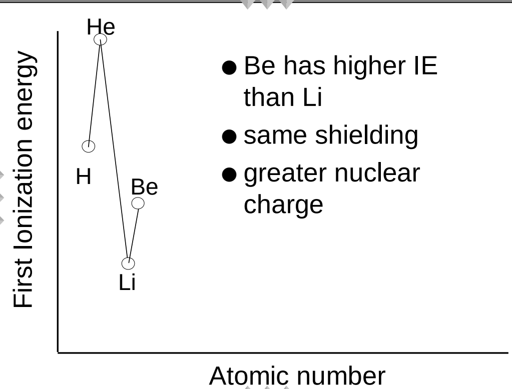
- **All the atoms in the same period**
  - Same shielding.
  - Increasing nuclear charge
- **So IE generally increases from left to right.**
- **Exceptions at full and 1/2 full orbitals**



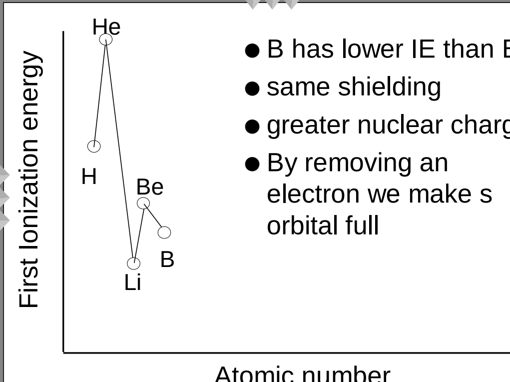
- **He has a greater IE than H**
- **same shielding**
- **greater nuclear charge**



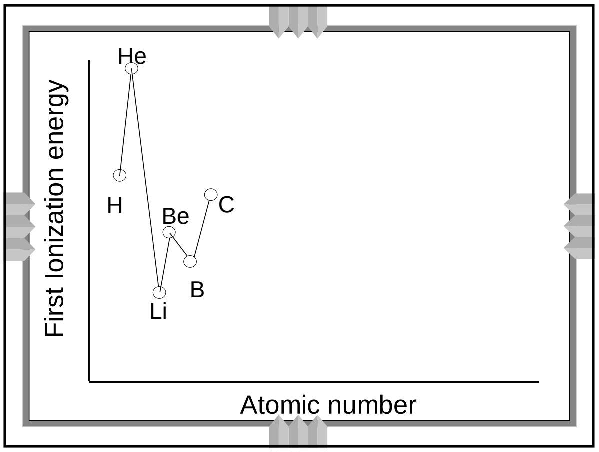
- Li has lower IE than H
- more shielding
- outweighs greater nuclear charge

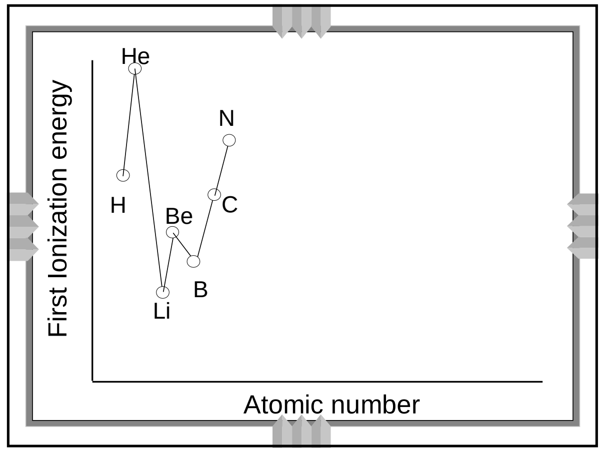


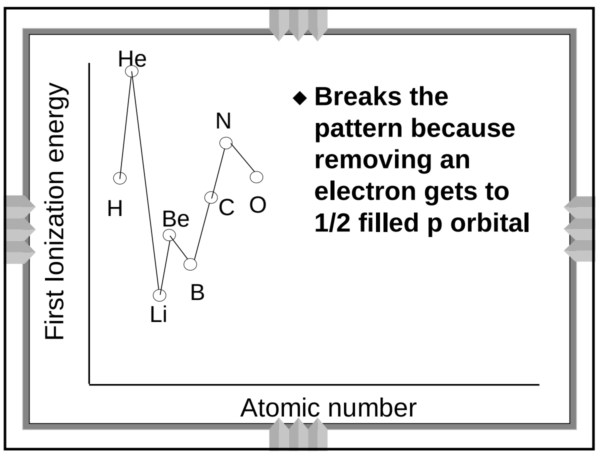
- Be has higher IE than Li
- same shielding
- greater nuclear charge

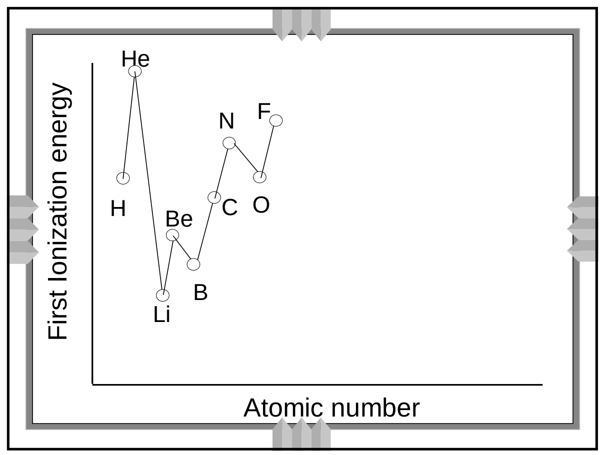


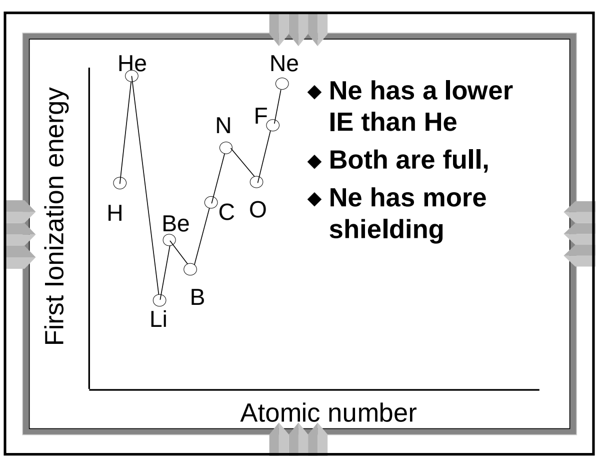
- B has lower IE than Be
- same shielding
- greater nuclear charge
- By removing an electron we make s orbital full

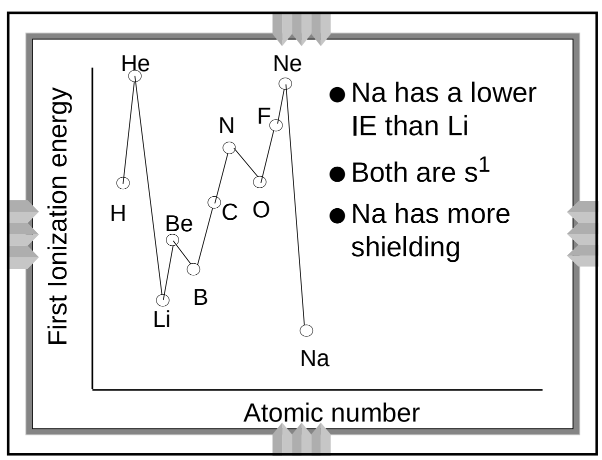

First Ionization energy
Atomic number
He
H
Be
C
B
Li


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

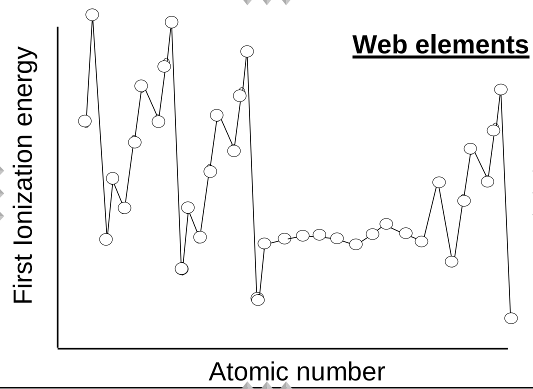

First Ionization energy
Atomic number
He
N
H
Be
C
O
B
Li
◆ Breaks the pattern because removing an electron gets to 1/2 filled p orbital


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


First Ionization energy
Atomic number
He
Ne
N
F
H
Be
C
O
B
Li
◆ Ne has a lower IE than He
◆ Both are full,
◆ Ne has more shielding


First Ionization energy
Atomic number
He
Ne
N
F
H
Be
C
O
B
Li
Na
● Na has a lower IE than Li
● Both are s1
● Na has more shielding

Web elements



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-full 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 boron family form 3+ ions
- 2nd IE and 3rd IE are always higher than 1st IE!!!

Web elements

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

Ionic Size

- Cations are positive ions
- 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 are negative ions
- 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

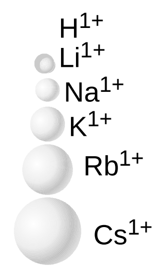
- Ions of representative elements have noble gas configuration
- Na is $1s^2 2s^2 2p^6 3s^1$
- Forms a 1+ ion - $1s^2 2s^2 2p^6$
- Same configuration as neon
- Metals form ions with the configuration of the noble gas before them - they lose electrons

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.

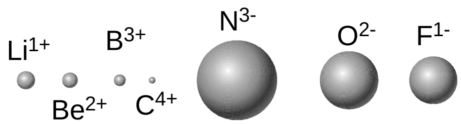
Group trends

- 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

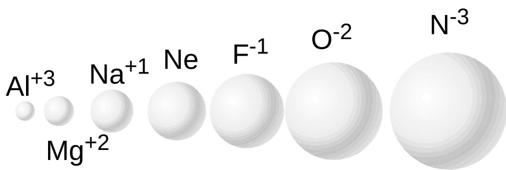


Size of Isoelectronic ions

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

Size of Isoelectronic ions

- **Positive 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 "greedy"**
- **Big electronegativity means it pulls the electron toward it.**

Group Trend

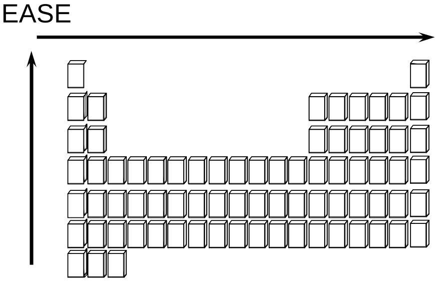
- **The further down a group**
  - More shielding
  - more electrons an atom has.
- **Less attraction for electrons**
- **Low electronegativity.**

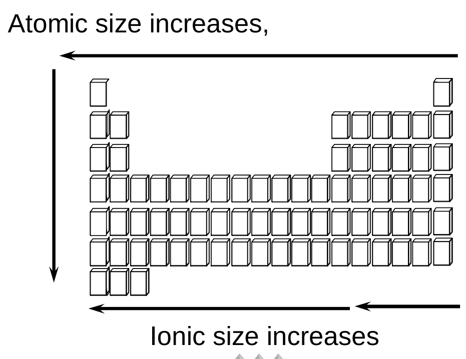
Periodic Trend

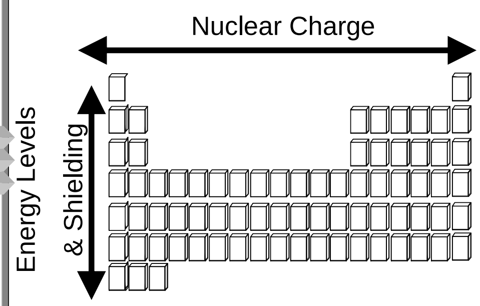
- **Metals - left end**
- **Low nuclear charge**
- **Low attraction**
- **Low electronegativity**
- **Right end - nonmetals**
- **High nuclear charge**
- **Large attraction**
- **High electronegativity**
- **Not noble gases- no compounds**

Ionization energy, electronegativity

INCREASE







How to answer why questions

- **Trend**
  - Periodic
  - Group
- **Reason**
  - Nuclear charge
  - Energy level and shielding
- **Result**
  - What happens to which electron

Trend

As you go across a period / down a group

Reason

the nuclear charge / energy level and shielding increases

Result

Making the
- **electron ______ to remove**
- **pull _____ on the other atom's electron**
- **outer electrons _______.**

so the ______ is _______.