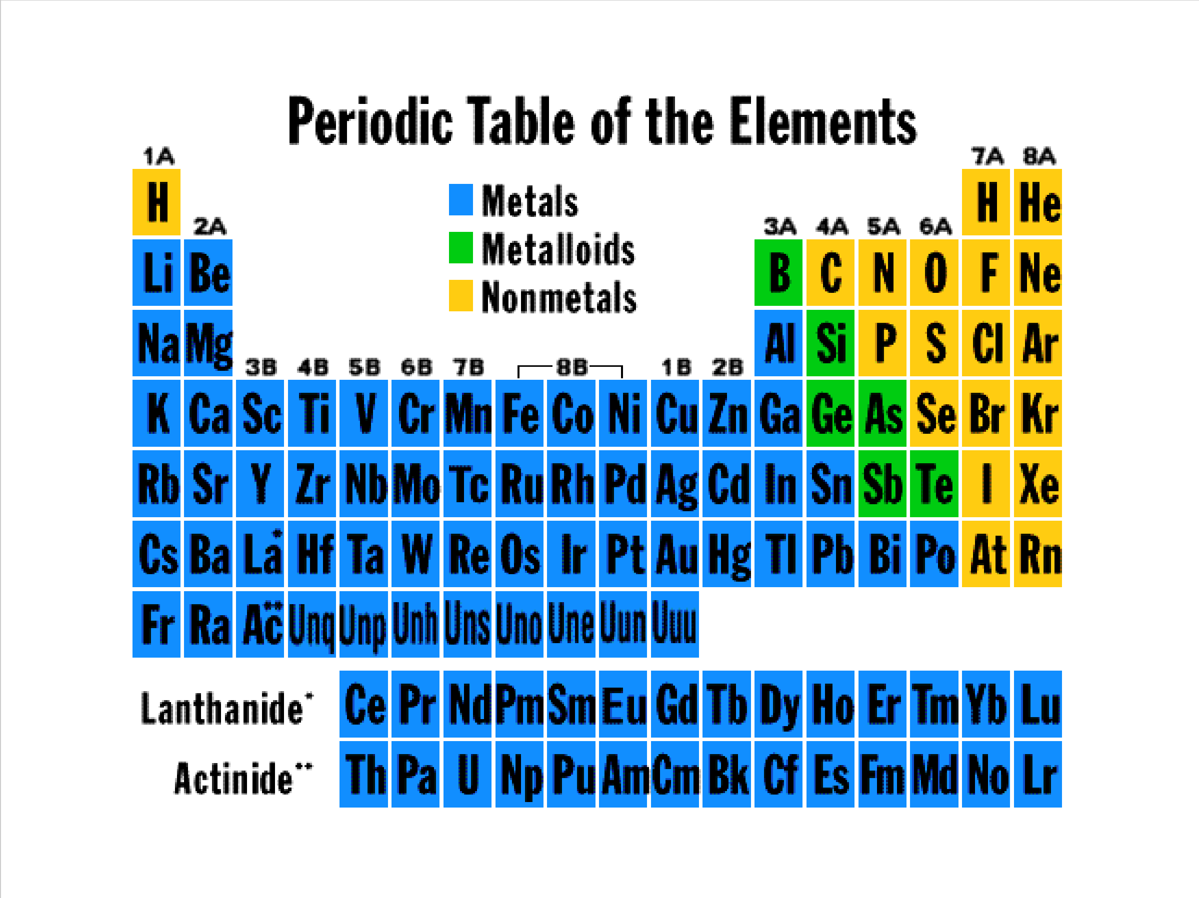
CHAPTER 6

Chemical Periodicity



Chapter Goals

1. More About the Periodic Table

Periodic Properties of the Elements

2. Atomic Radii
3. Ionization Energy (IE)
4. Electron Affinity (EA)
5. Ionic Radii
6. Electronegativity

Chemical Reactions and Periodicity

7. Hydrogen & the Hydrides
8. Oxygen & the Oxides

More About the Periodic Table

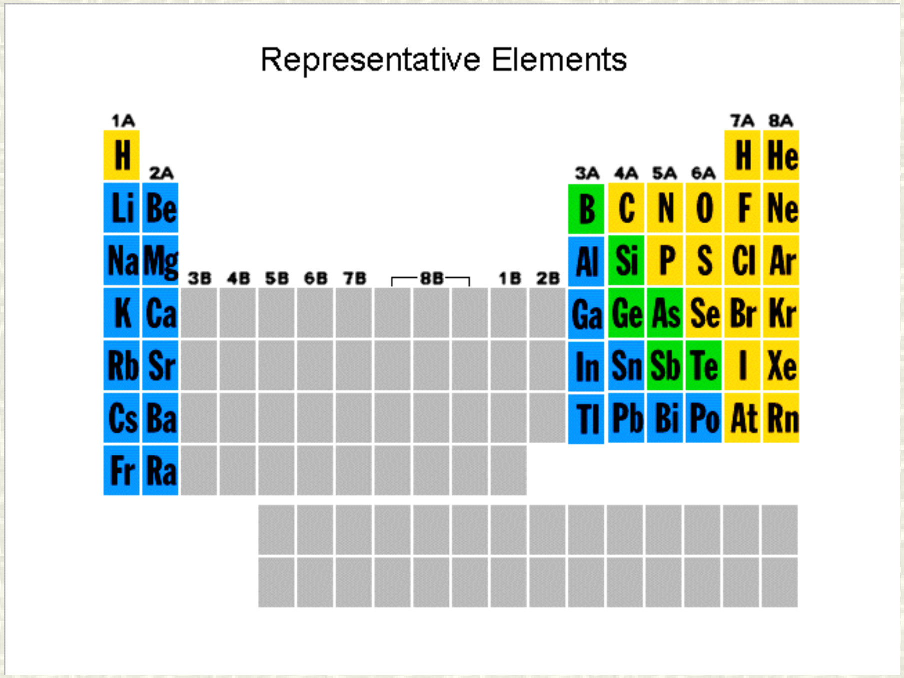
Noble Gases

- All of them have completely filled electron shells.

- Since they have similar electronic structures, **full *s* and *p* orbitals**, their chemical reactions are similar.
  - He $1s^2$
  - Ne [He] $2s^2$ $2p^6$
  - Ar [Ne] $3s^2$ $3p^6$
  - Kr [Ar] $4s^2$ $4p^6$
  - Xe [Kr] $5s^2$ $5p^6$
  - Rn [Xe] $6s^2$ $6p^6$

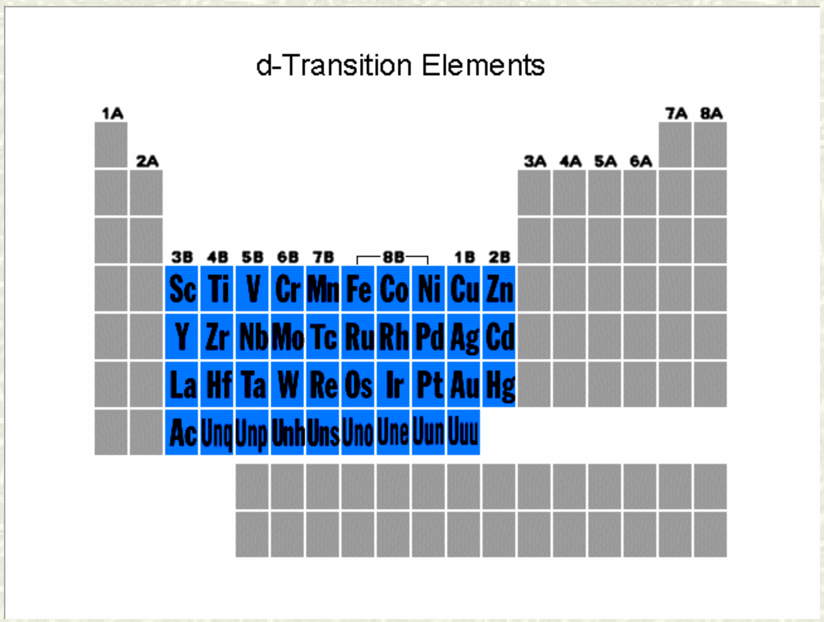
More About the Periodic Table

- **Representative Elements**
  - Are the elements in A groups on periodic chart.
- These elements will have their "last" electron in an outer *s* or *p* orbital.



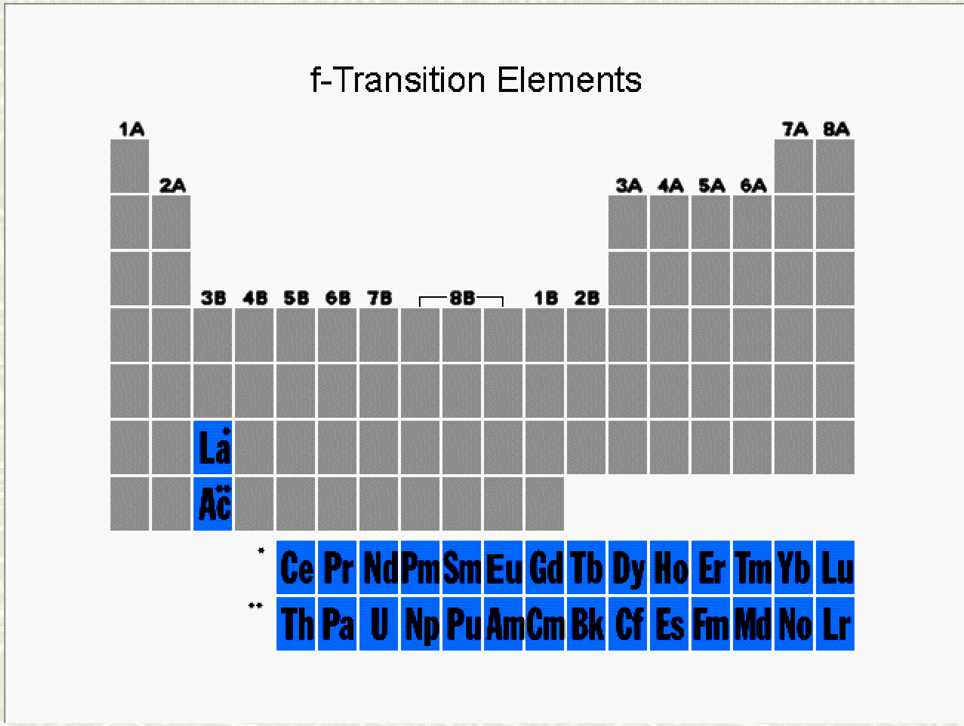
More About the Periodic Table

- **<u>*d*-Transition Elements</u>**
- Each metal has *d* electrons.
  - n*s* (n-1)*d* configurations
- These elements make the transition from metals to nonmetals.



More About the Periodic Table

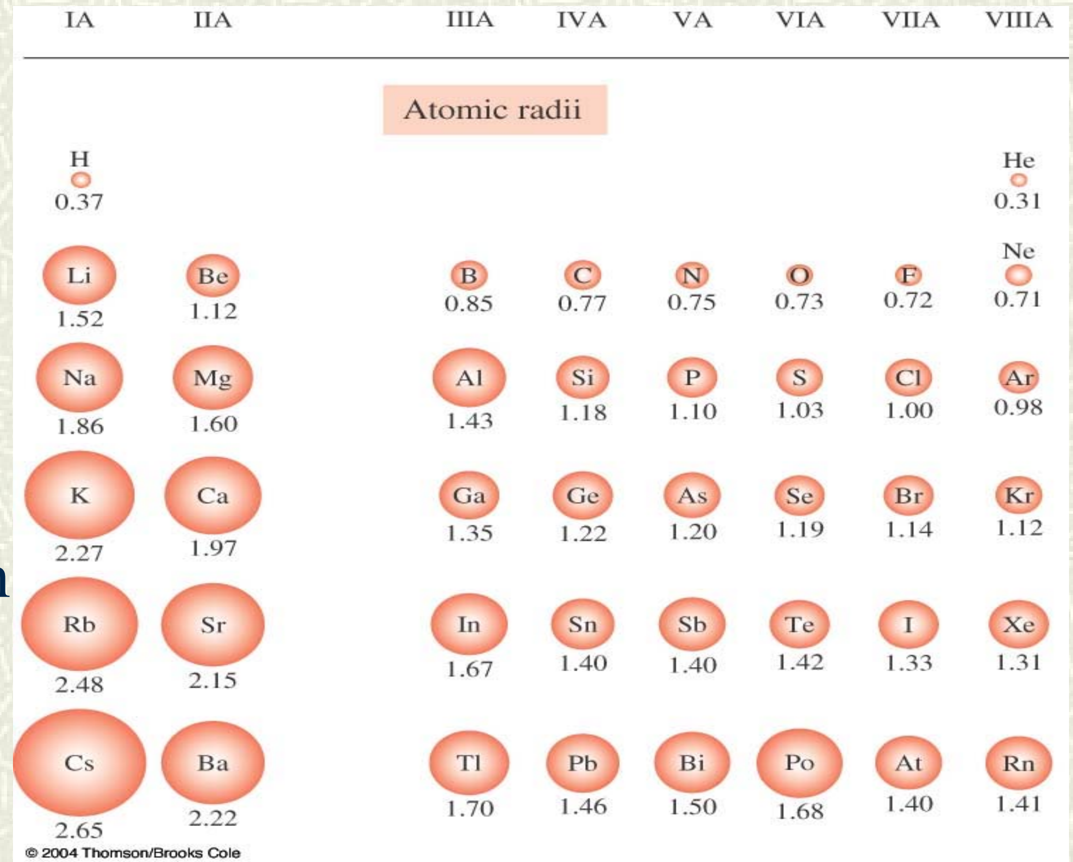
- ***f* - transition metals**
  - Sometimes called inner transition metals.
- Electrons are being added to *f* orbitals.
- Very slight variations of properties from one element to another.

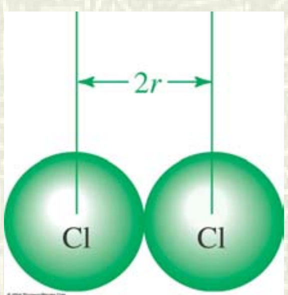


Periodic Properties of the Elements

Atomic Radii

- Atomic radii **increase** within a column going from the **top to the bottom** of the periodic table.
- Atomic radii **decrease** within a row going from **left to right** on the periodic table.





Atomic Radii

- The reason the atomic radii decrease across a period is due to ***shielding*** **or** ***screening*** **effect**.
  - Effective nuclear charge, Z_{eff}, experienced by an electron is less than the actual nuclear charge, Z.
  - The inner electrons block the nuclear charge's effect on the outer electrons.
  - Consequently, the outer electrons feel a stronger effective nuclear charge.
  - **For Li, Z_{eff} ~ +1**
  - **For Be, Z_{eff} ~ +2**

Atomic Radii

- **Example: Arrange these elements based on their atomic radii.**
  - **Se, S, O, Te**

O < S < Se < Te

- **Example: Arrange these elements based on their atomic radii.**
  - **P, Cl, S, Si**

Cl < S < P < Si

- **Example: Arrange these elements based on their atomic radii.**
  - **Ga, F, S, As**

F < S < As < Ga

Ionization Energy

- **First ionization energy (IE_1)**
  - The minimum amount of energy required to remove the most loosely bound electron from an isolated gaseous atom to form a 1+ ion.
- Symbolically:

$$Atom_{(g)} + energy \rightarrow ion^{+}_{(g)} + e^{-}$$

$$Mg_{(g)} + 738kJ/mol \rightarrow Mg^{+} + e^{-}$$

Ionization Energy

- **Second ionization energy (IE_2)**
  - The amount of energy required to remove the second electron from a gaseous 1+ ion.
- Symbolically:
  - **$ion^+ + energy \rightarrow ion^{2+} + e^-$**

$Mg^+ + 1451\ kJ/mol \rightarrow Mg^{2+} + e^-$

•Atoms can have 3rd (IE_3), 4th (IE_4), etc. ionization energies.

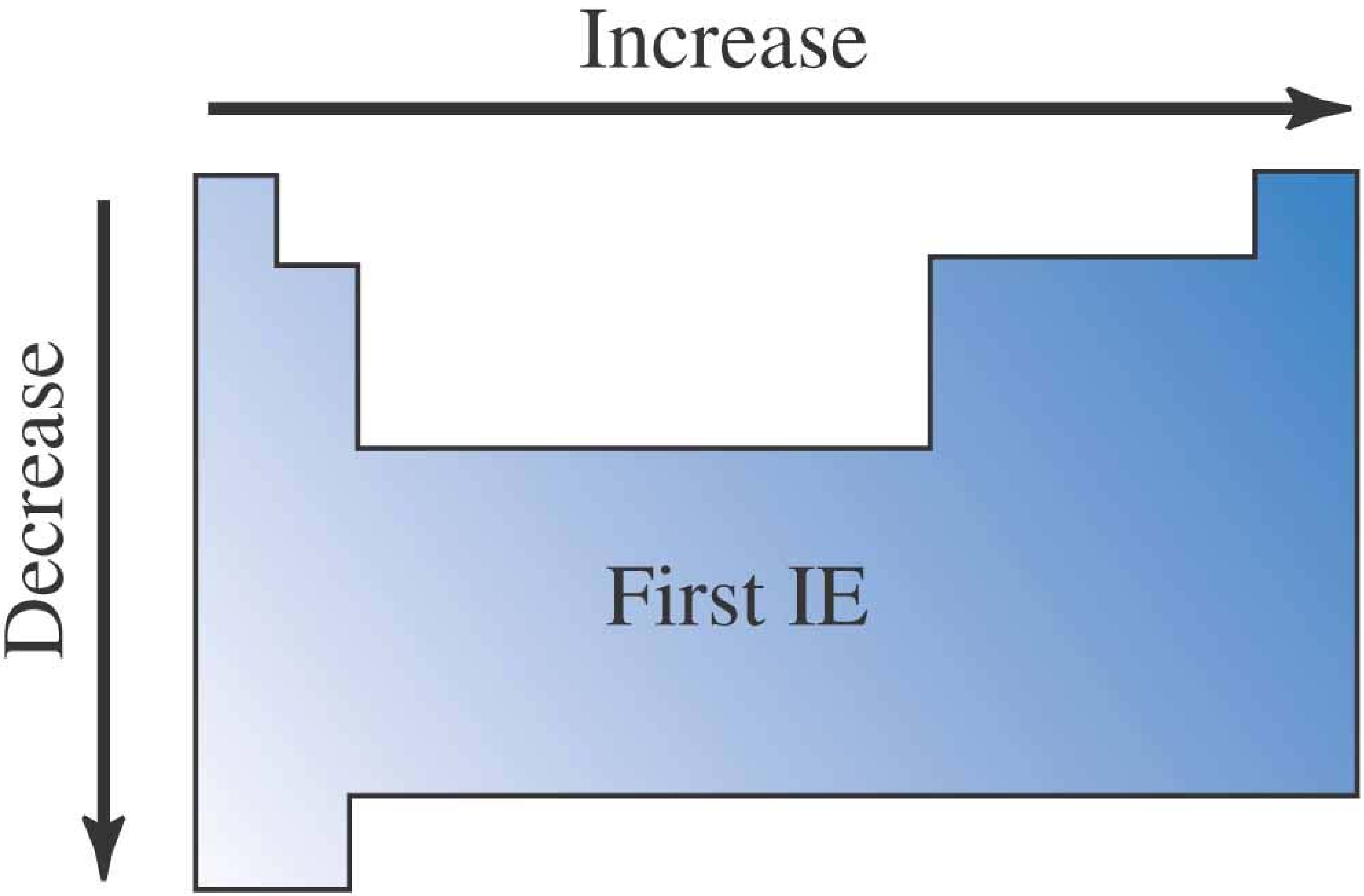
Ionization Energy

- $IE_2 > IE_1$ It always takes more energy to remove a second electron from an ion than from a neutral atom.
- IE_1 generally increases moving from IA elements to VIIIA elements.

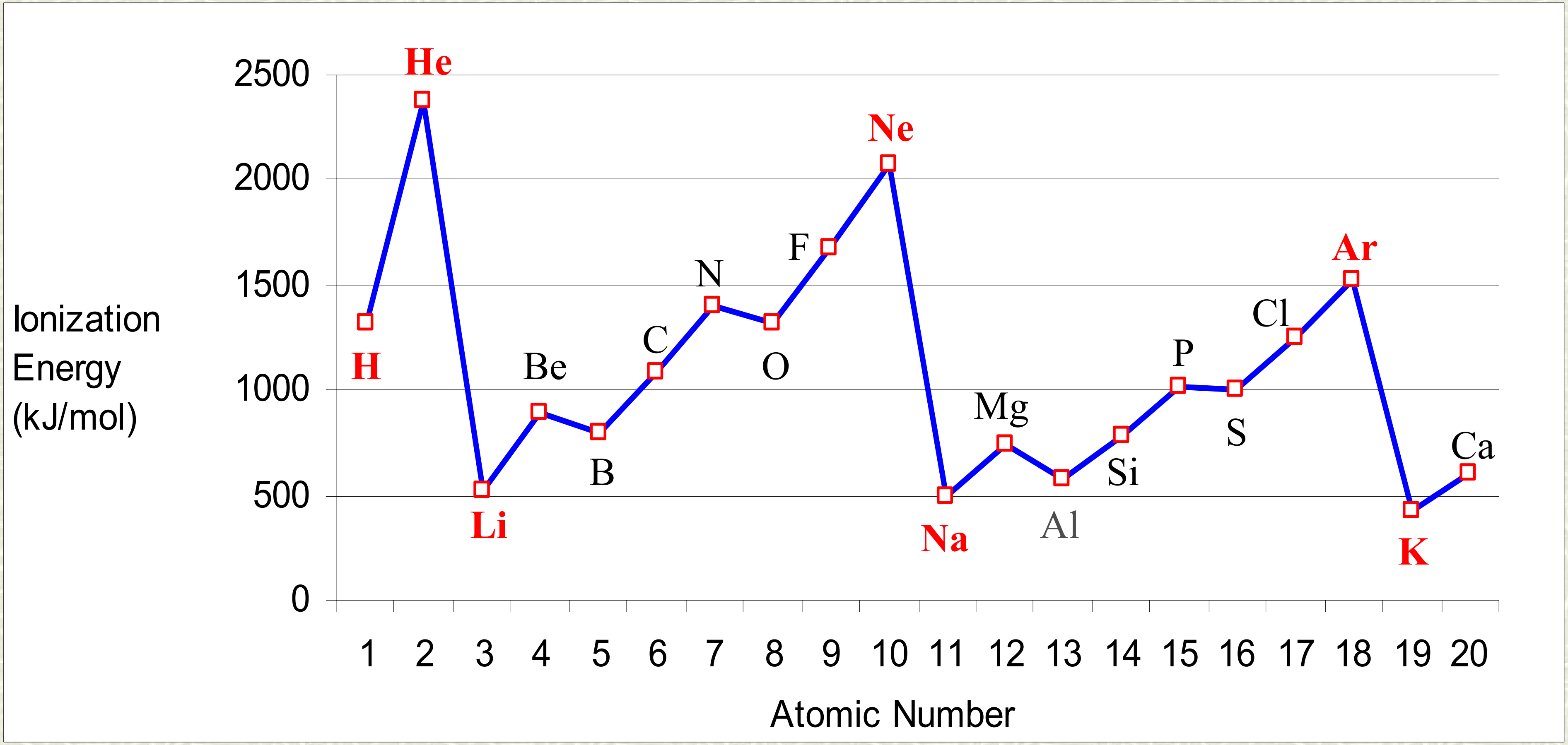
  Important exceptions at Be & Mg, N & P, etc. due to filled and half-filled subshells.

  - IE_1 generally decreases moving down a family.

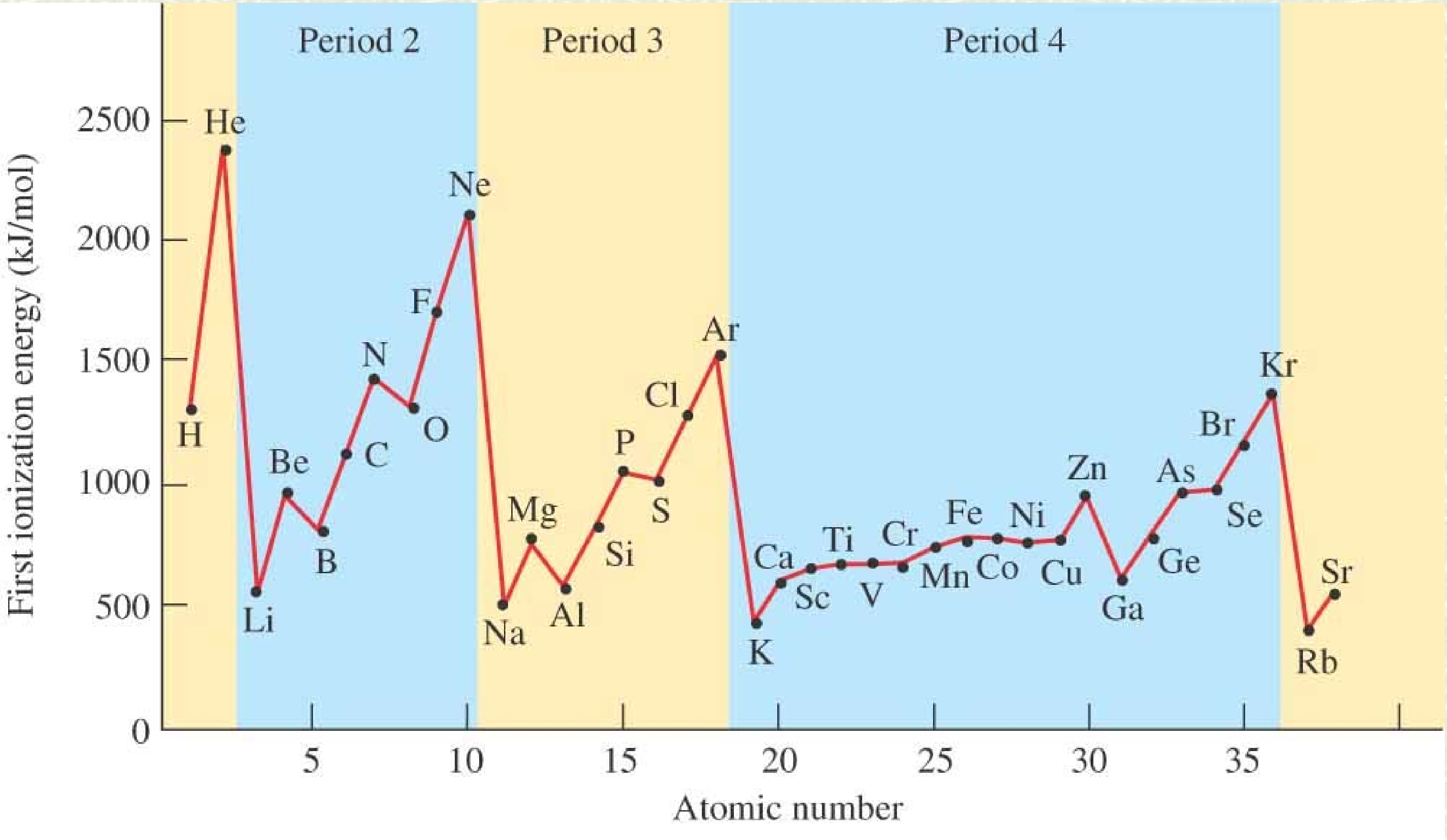
  IE_1 for Li > IE_1 for Na, etc.


Increase
Decrease
First IE

First Ionization Energies of Some Elements



First Ionization Energies of Some Elements



Ionization Energy

- **Example: Arrange these elements based on their first ionization energies.**
  - **Sr, Be, Ca, Mg**

Sr < Ca < Mg < Be

- **Example: Arrange these elements based on their first ionization energies.**
  - **Al, Cl, Na, P**

Na < Al < P < Cl

- **Example: Arrange these elements based on their first ionization energies.**
  - **B, O, Be, N**

B < Be < O < N

Ionization Energy

- First, second, third, etc. ionization energies exhibit periodicity as well.

IE_1 (kJ/mol)	1680
IE_2 (kJ/mol)	3370
IE_3 (kJ/mol)	6050
IE_4 (kJ/mol)	8410
IE_5 (kJ/mol)	11020
IE_6 (kJ/mol)	15160
IE_7 (kJ/mol)	17870
IE_8 (kJ/mol)	92040

Electron Affinity

- Electron affinity is the amount of energy ***absorbed*** when an electron is added to an isolated gaseous atom to form an ion with a 1- charge.
- Electron affinity is a measure of an atom's ability to form negative ions.
- Symbolically:

$$\text{atom}(g) + e^- + \text{EA} \rightarrow \text{ion}^-(g)$$

Electron Affinity

<u>Sign conventions for electron affinity</u>

If electron affinity > 0 energy is absorbed.

If electron affinity < 0 energy is released.

$Mg_{(g)} + e^- + 231\ kJ/mol \rightarrow Mg^-_{(g)}$

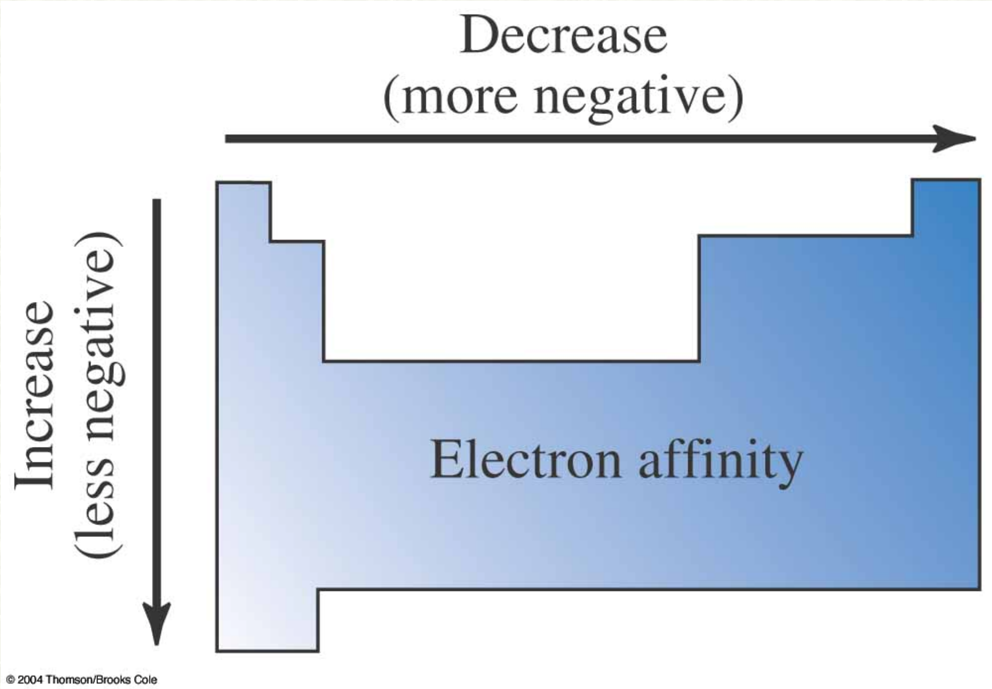
EA = +231 kJ/mol

$Br_{(g)} + e^- \rightarrow Br^-_{(g)} + 323\ kJ/mol$

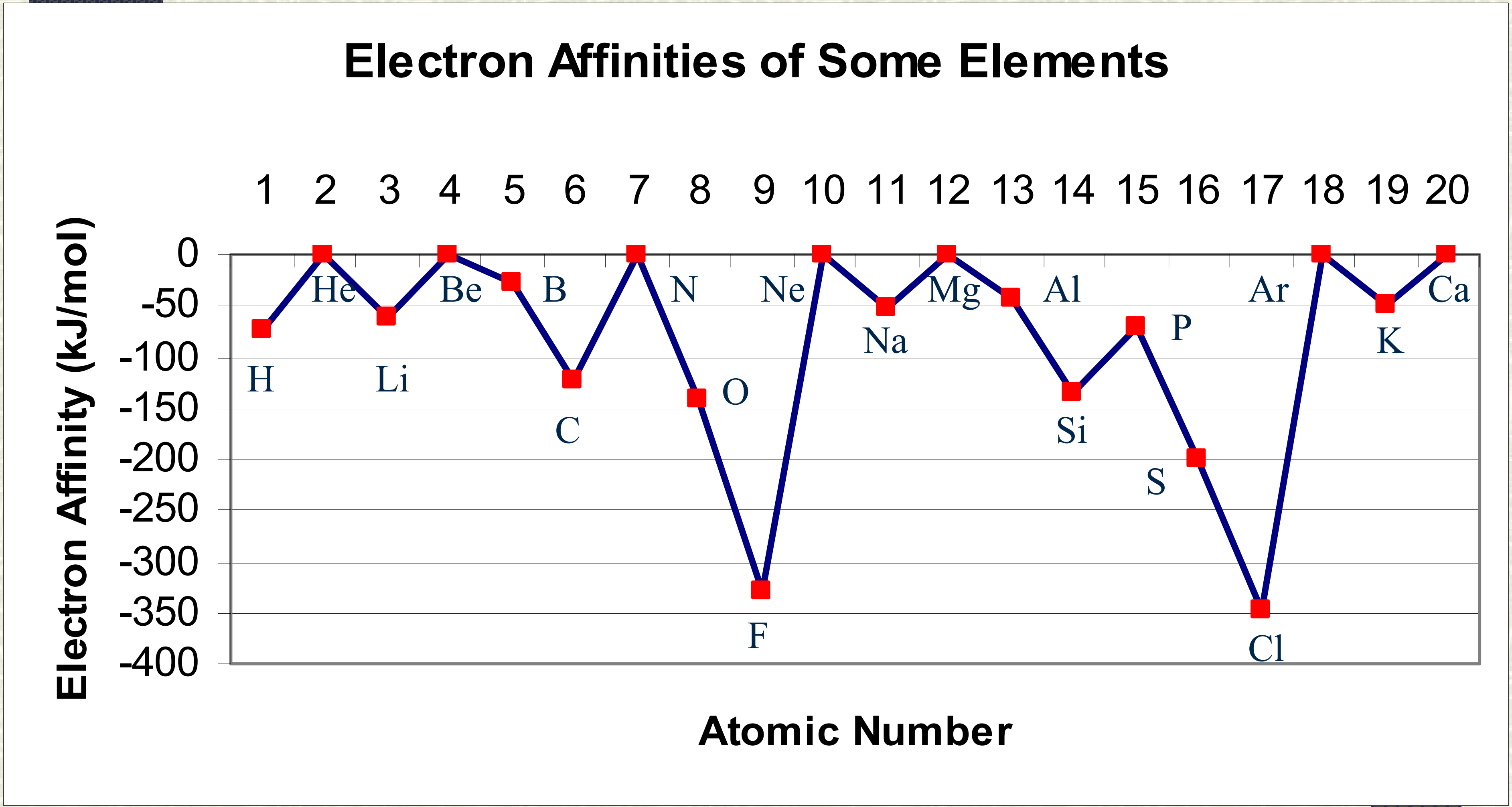
EA = -323 kJ/mol

Electron Affinity

- the values become more negative from left to right across a period on the periodic chart.
- the values become more negative from bottom to top up a row on the periodic chart.



Electron Affinity



Electron Affinity

1	H −73								He 0
2	Li −60	Be (~0)		B −29	C −122	N 0	O −141	F −328	Ne 0
3	Na −53	Mg (~0)		Al −43	Si −134	P −72	S −200	Cl −349	Ar 0
4	K −48	Ca (~0)	Cu −118	Ga −29	Ge −119	As −78	Se −195	Br −324	Kr 0
5	Rb −47	Sr (~0)	Ag −125	In −29	Sn −107	Sb −101	Te −190	I −295	Xe 0
6	Cs −45	Ba (~0)	Au −282	Tl −19	Pb −35	Bi −91			

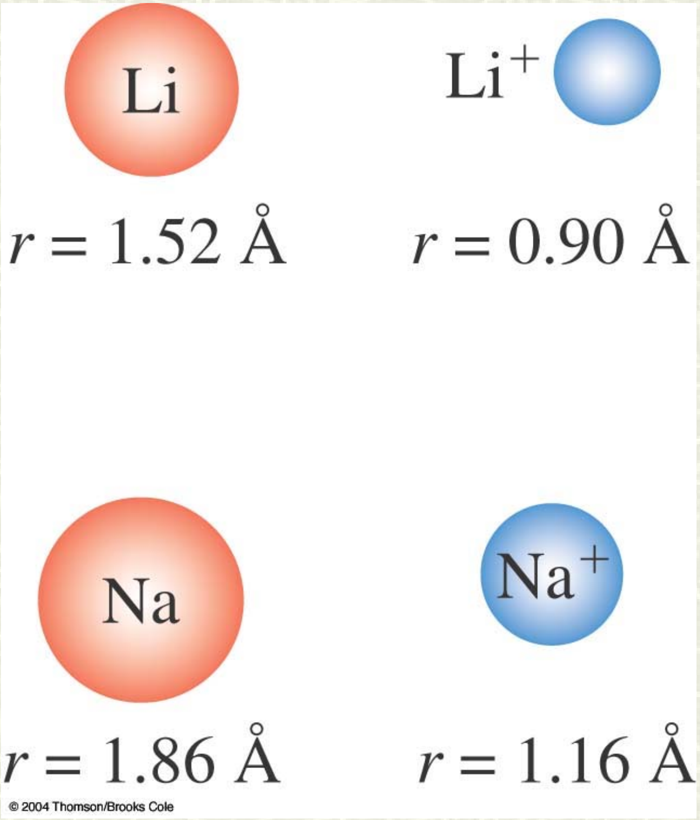
Electron Affinity

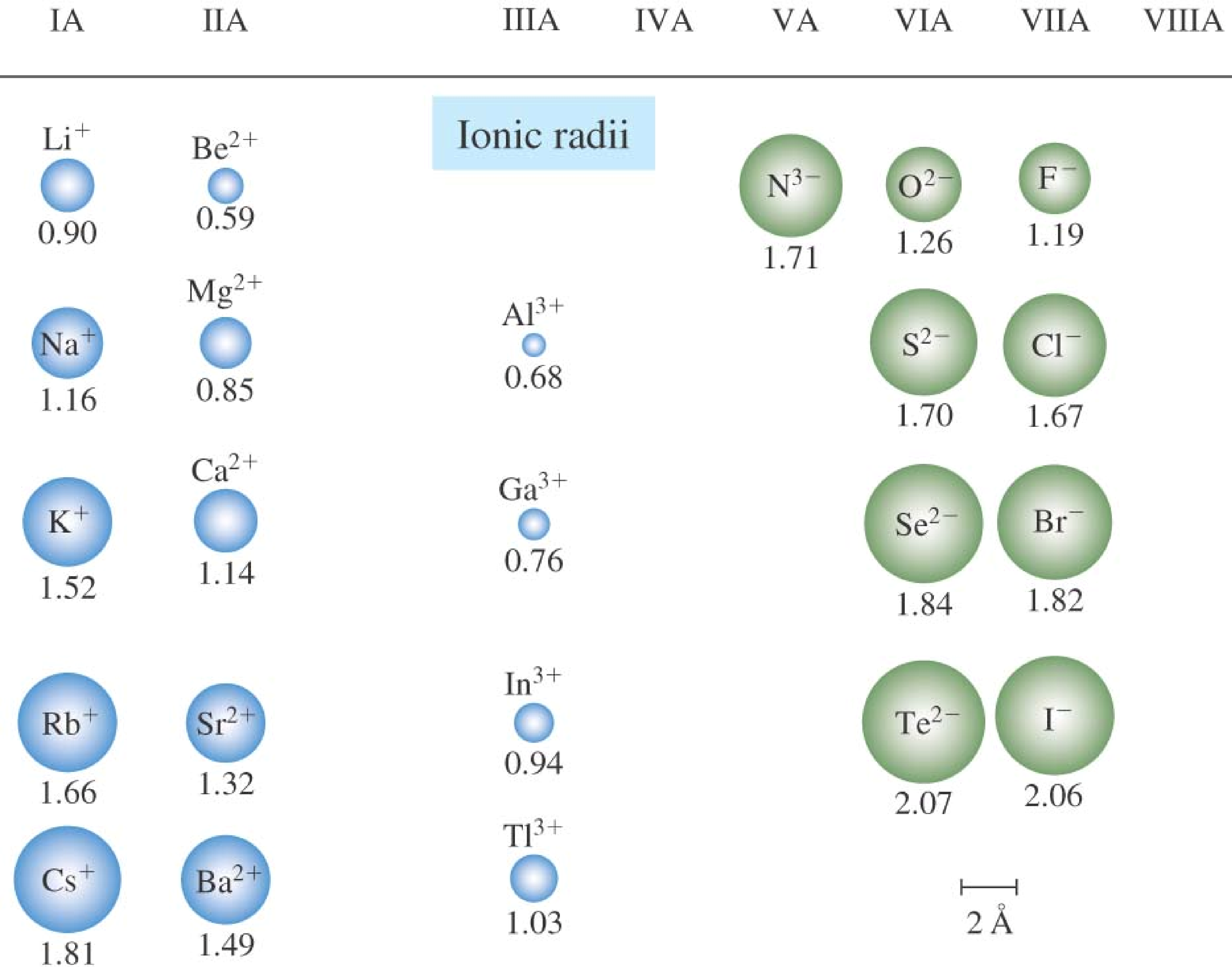
- Example: Arrange these elements based on their electron affinities.
  - Al, Mg, Si, Na

Si < Al < Na < Mg

Ionic Radii

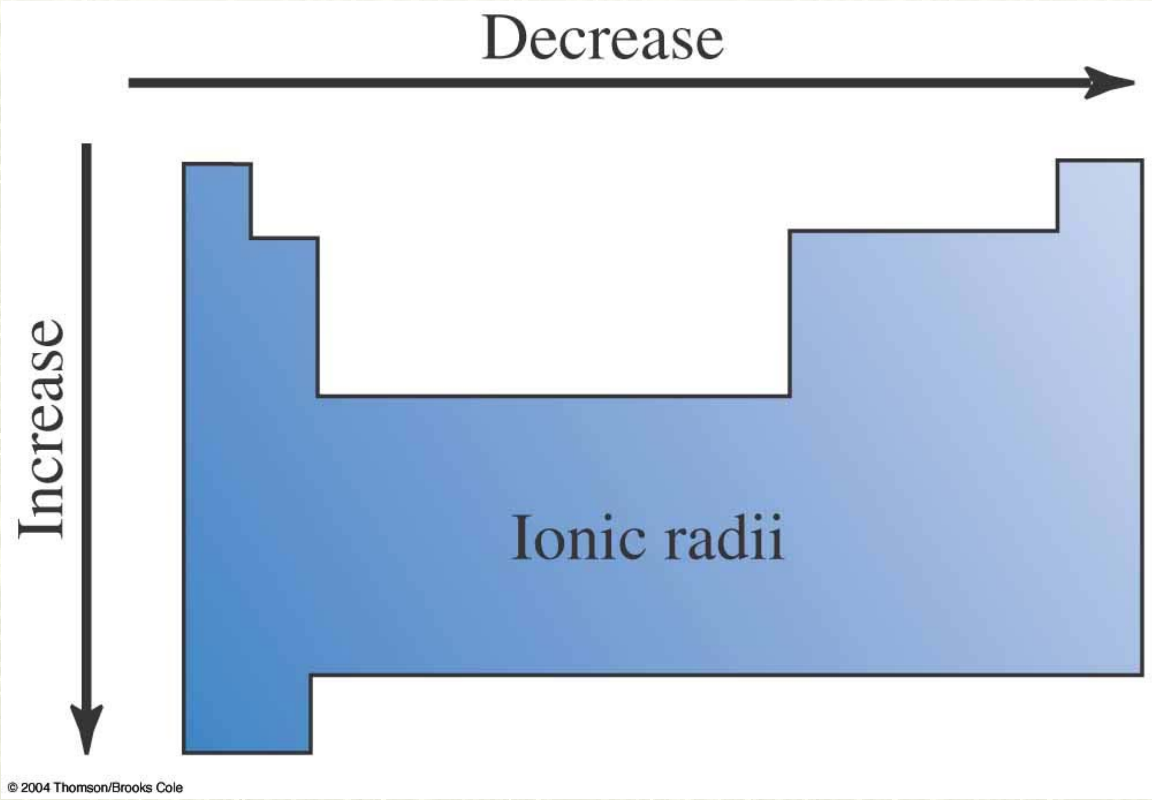
- Cations are always ***smaller*** than their respective neutral atoms.




IA
IIA
IIIA
IVA
VA
VIA
VIIA
VIIIA
Ionic radii
Li+ 0.90
Be2+ 0.59
N3− 1.71
O2− 1.26
F− 1.19
Na+ 1.16
Mg2+ 0.85
Al3+ 0.68
S2− 1.70
Cl− 1.67
K+ 1.52
Ca2+ 1.14
Ga3+ 0.76
Se2− 1.84
Br− 1.82
Rb+ 1.66
Sr2+ 1.32
In3+ 0.94
Te2− 2.07
I− 2.06
Cs+ 1.81
Ba2+ 1.49
Tl3+ 1.03
2 Å

Ionic Radii

- Anions are always ***larger*** than their neutral atoms.



Ionic Radii

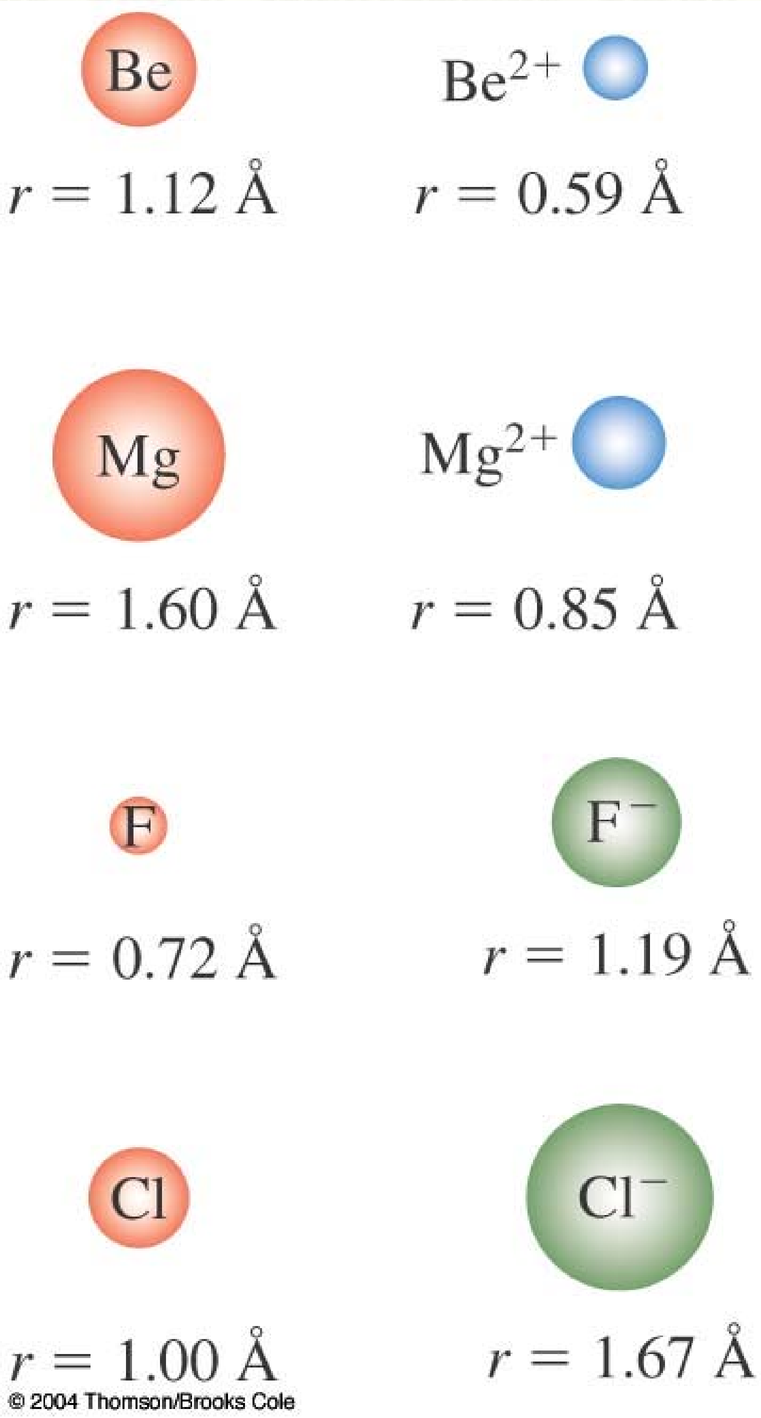
- Cations radii decrease from left to right across a period.
  - Increasing nuclear charge attracts the electrons and decreases the radius.

Ion	Rb^{+}	Sr^{2+}	In^{3+}
Ionic Radii(Å)	1.66	1.32	0.94

Ionic Radii

- Anions radii decrease from left to right across a period.
  - Increasing electron numbers in highly charged ions cause the electrons to repel and increase the ionic radius.

Ion	N^{3-}	O^{2-}	F^{1-}
Ionic Radii(Å)	1.71	1.26	1.19


Be
Be^{2+}
$r = 1.12$ Å
$r = 0.59$ Å
Mg
Mg^{2+}
$r = 1.60$ Å
$r = 0.85$ Å
F
F^-
$r = 0.72$ Å
$r = 1.19$ Å
Cl
Cl^-
$r = 1.00$ Å
$r = 1.67$ Å
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Ionic Radii

- Example: Arrange these elements based on their ionic radii.
  - Ga, K, Ca

$$K^{1+} < Ca^{2+} < Ga^{3+}$$

- Example: Arrange these elements based on their ionic radii.
  - Cl, Se, Br, S

$$Cl^{1-} < S^{2-} < Br^{1-} < Se^{2-}$$

Isoelectronic ions

An isoelectronic series of ions

	N^{3-}	O^{2-}	F^-	Na^+	Mg^{2+}	Al^{3+}
Ionic radius (Å)	1.71	1.26	1.19	1.16	0.85	0.68
No. of electrons	10	10	10	10	10	10
Nuclear charge	+7	+8	+9	+11	+12	+13

Electronegativity

- **Electronegativity** is a measure of the relative tendency of an atom to attract electrons to itself when ***chemically combined with another element***.
  - Electronegativity is measured on the **Pauling** scale.
  - Fluorine is the most electronegative element.
  - Cesium and francium are the least electronegative elements.
- For the representative elements, electronegativities usually increase from left to right across periods and decrease from top to bottom within groups.

Electronegativity

Metals

Nonmetals

Metalloids

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB	VIIIB	VIIIB	IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	1 H 2.1																	2 He
2	3 Li 1.0	4 Be 1.5											5 B 2.0	6 C 2.5	7 N 3.0	8 O 3.5	9 F 4.0	10 Ne
3	11 Na 1.0	12 Mg 1.2											13 Al 1.5	14 Si 1.8	15 P 2.1	16 S 2.5	17 Cl 3.0	18 Ar
4	19 K 0.9	20 Ca 1.0	21 Sc 1.3	22 Ti 1.4	23 V 1.5	24 Cr 1.6	25 Mn 1.6	26 Fe 1.7	27 Co 1.7	28 Ni 1.8	29 Cu 1.8	30 Zn 1.6	31 Ga 1.7	32 Ge 1.9	33 As 2.1	34 Se 2.4	35 Br 2.8	36 Kr

Electronegativity

- Example: Arrange these elements based on their electronegativity.
  - Se, Ge, Br, As

Ge < As < Se < Br

- Example: Arrange these elements based on their electronegativity.
  - Be, Mg, Ca, Ba

Chemistry is fun!

Ba < Ca < Mg < Be