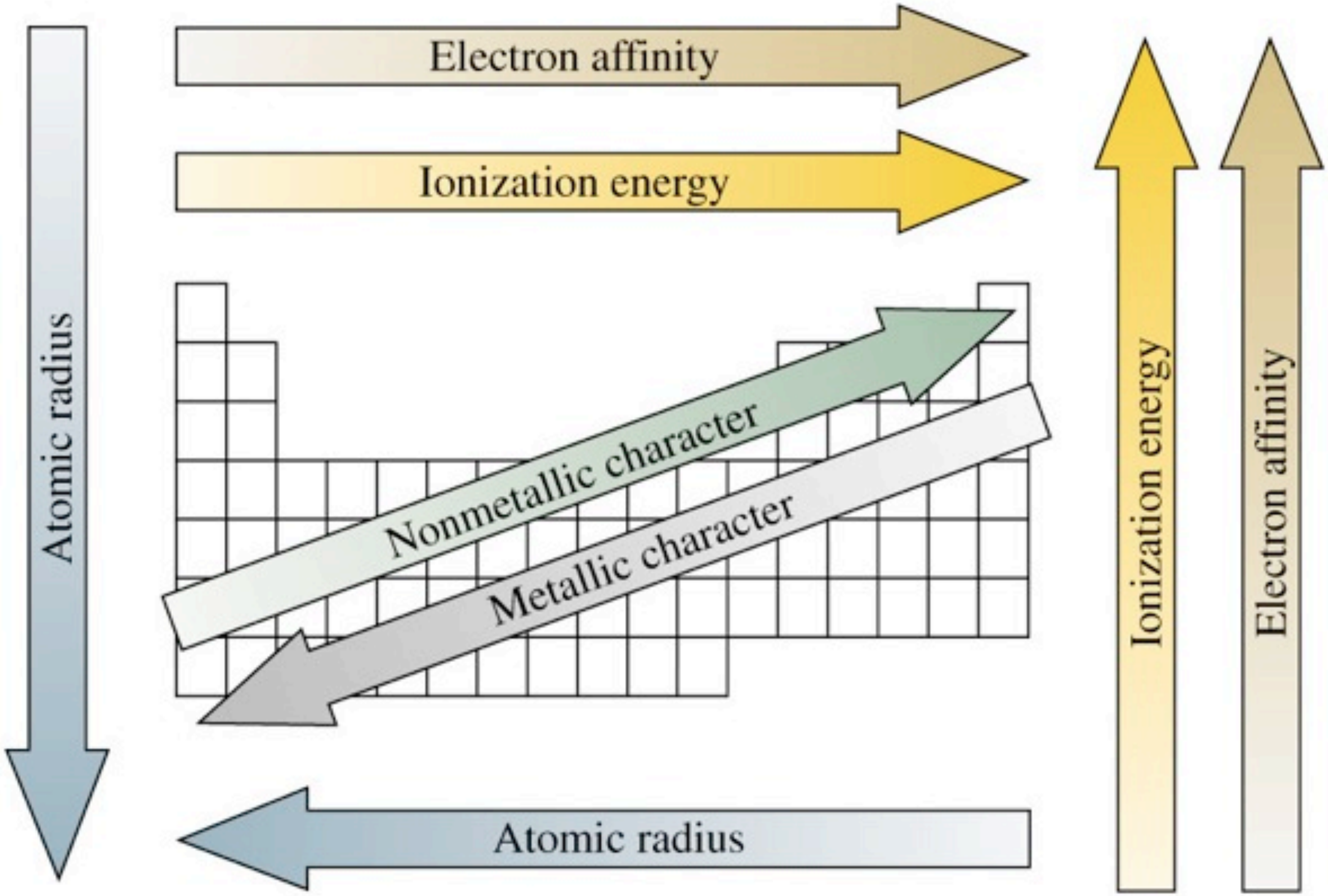
Summation of Periodic Trends



Factors Affecting Atomic Orbital Energies

The Effect of Nuclear Charge ($Z_{effective}$)

Higher nuclear charge lowers orbital energy (stabilizes the system) by increasing nucleus-electron attractions.

The Effect of Electron Repulsions (Shielding)

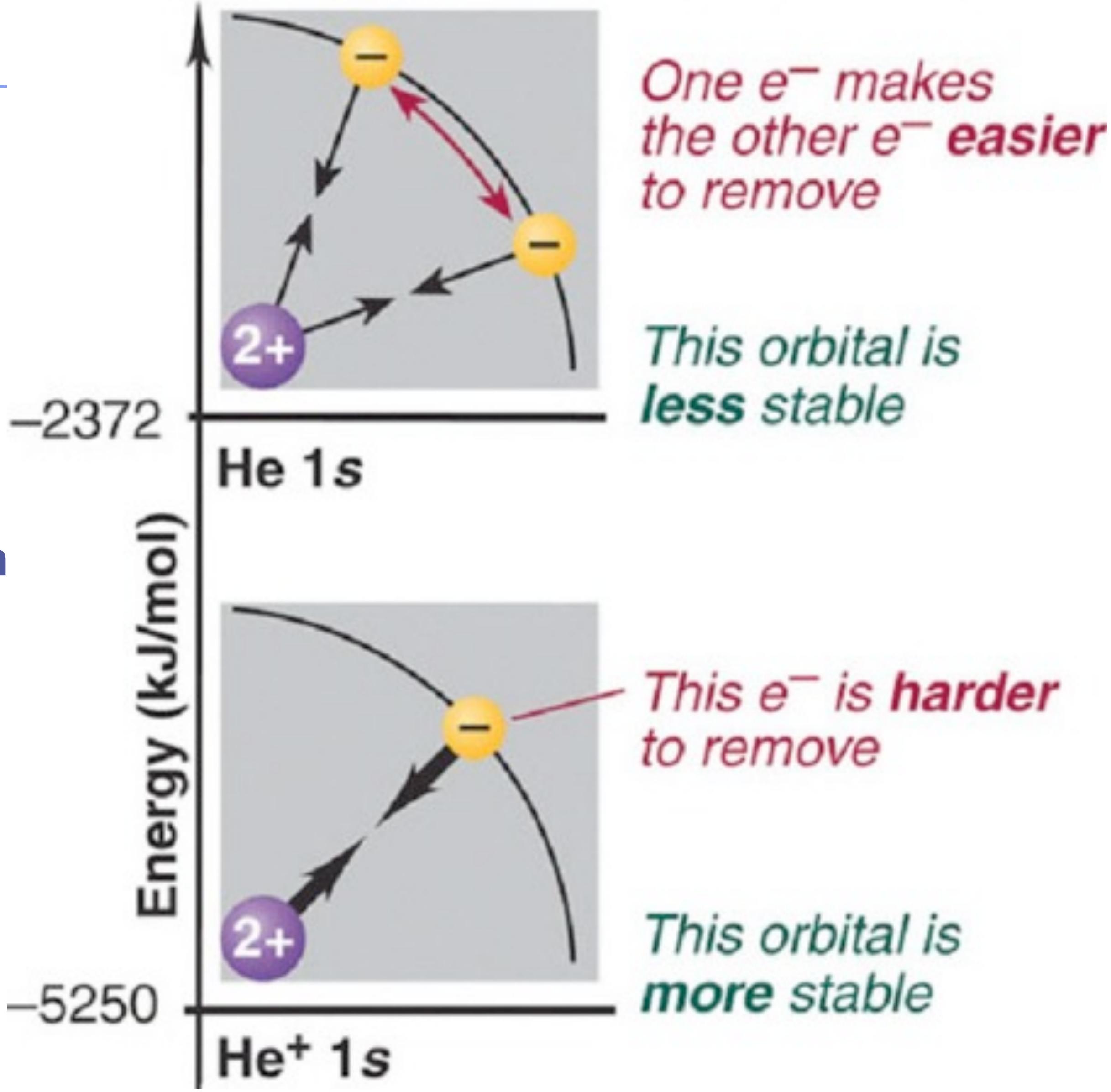
Additional electron in the same orbital (makes less stable)

An additional electron raises the orbital energy through electron-electron repulsions.

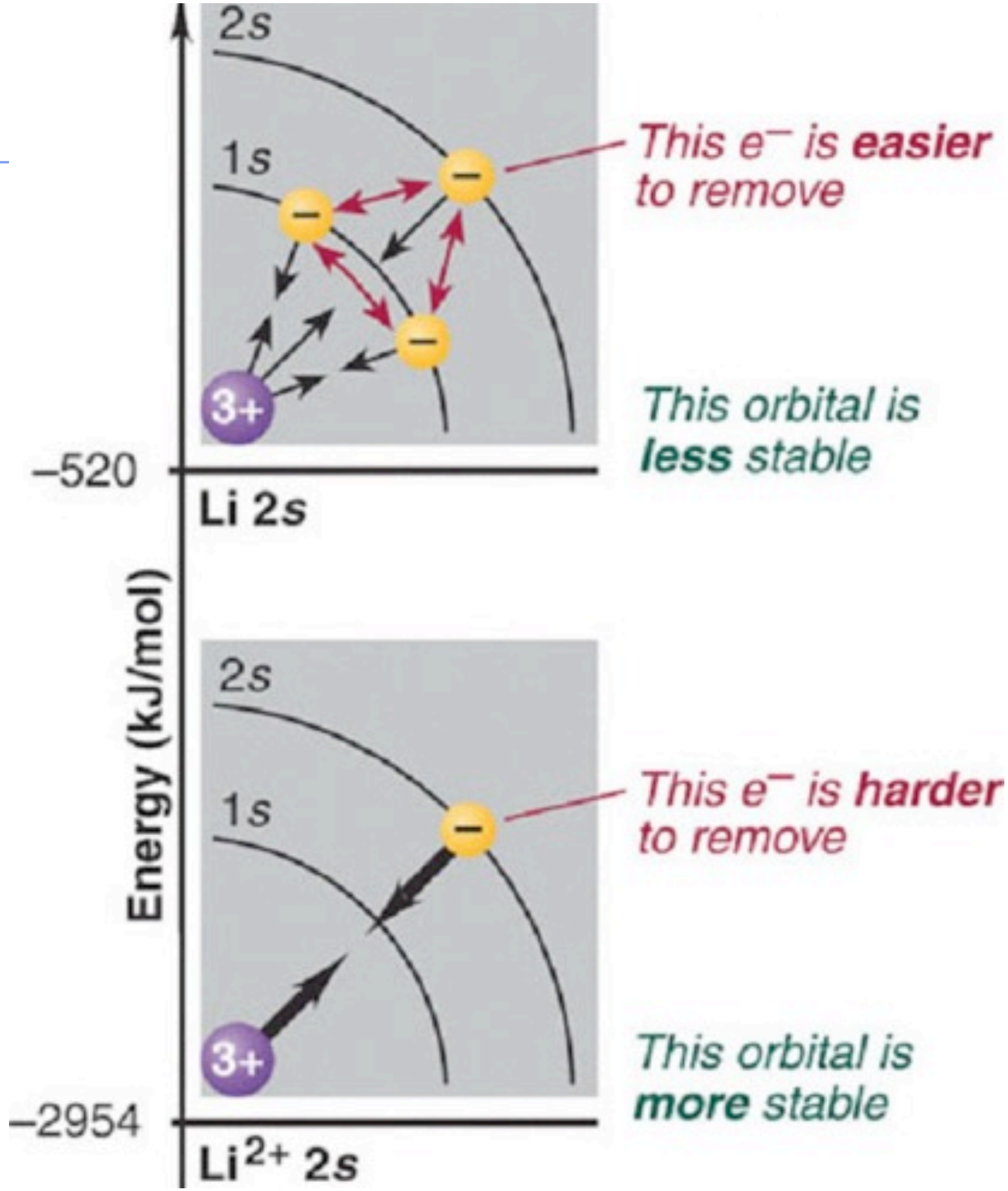
Additional electrons in inner orbitals (makes outer orbital less stable)

Inner electrons shield outer electrons more effectively than do electrons in the same sublevel.

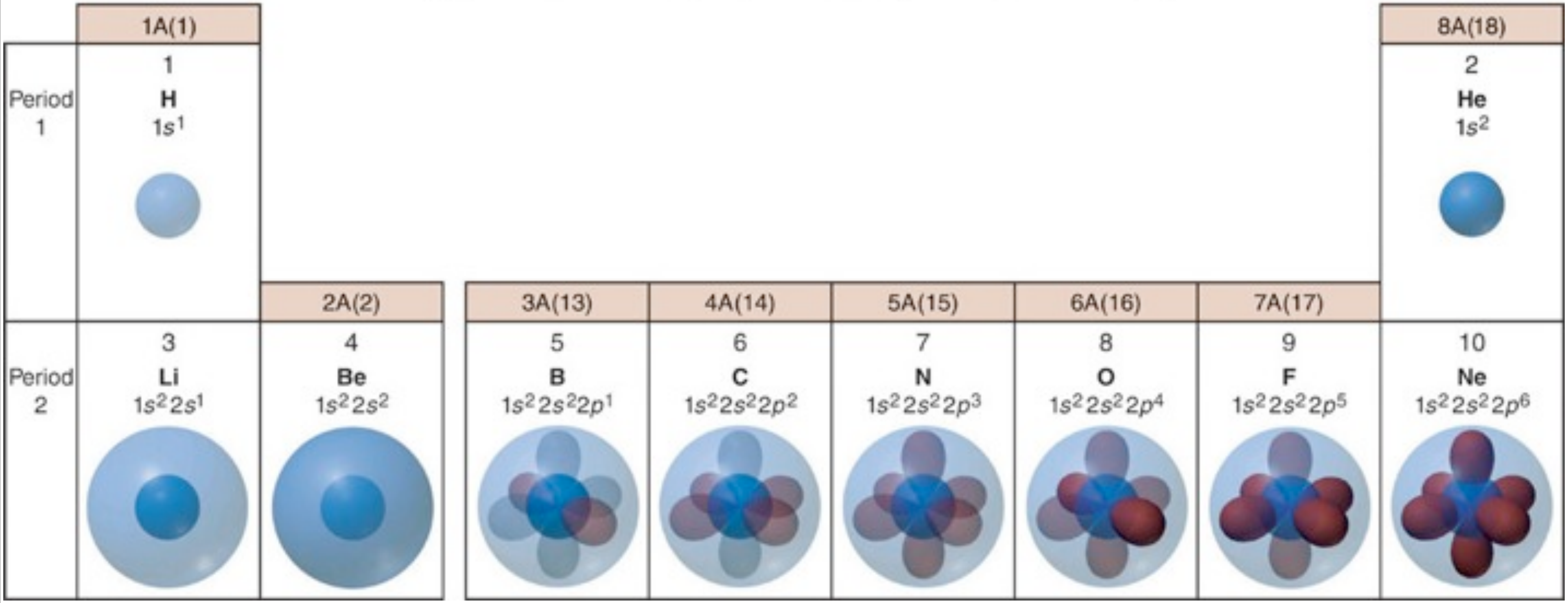
The effect of another electron in the same orbital



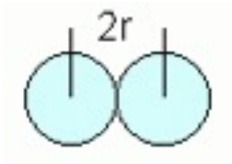
The effect of other electrons in inner orbitals



Orbital occupancy for the first 10 elements, H through Ne.



Atomic Radius



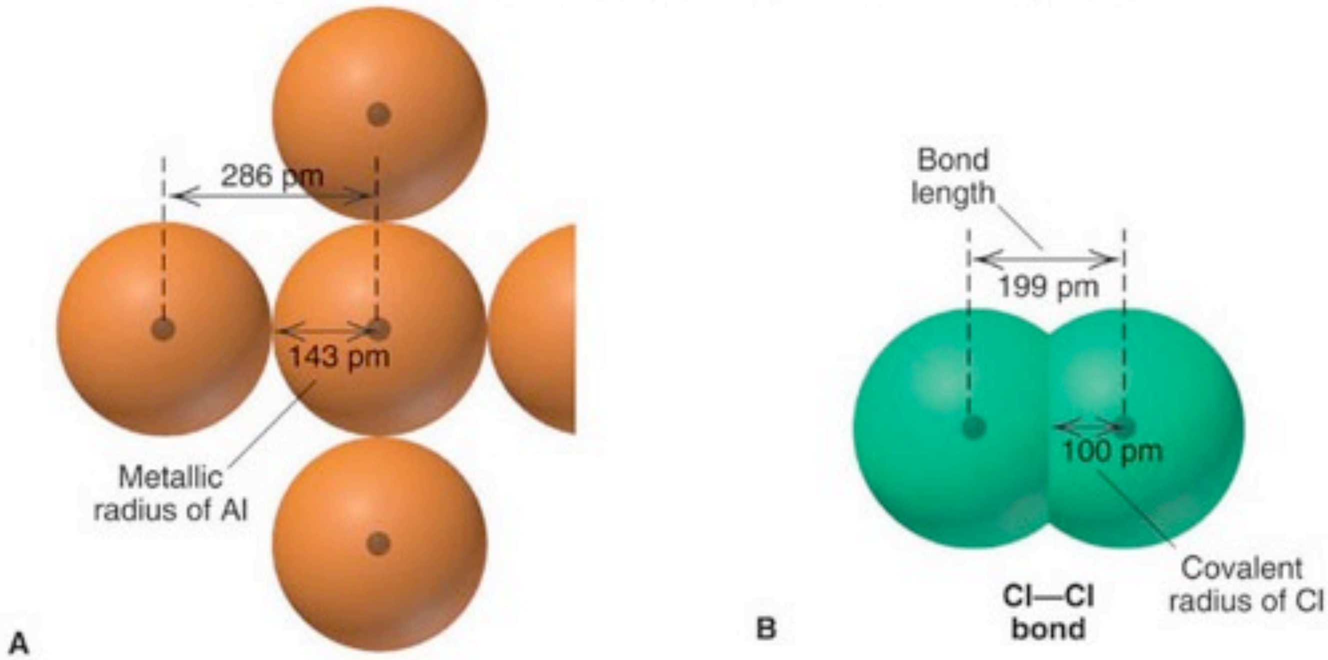
Half of the distance between nuclei in covalently bonded diatomic molecule

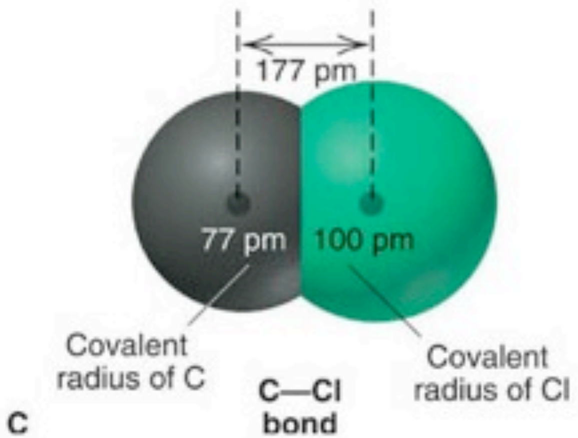
"covalent atomic radii"

Periodic Trends in Atomic Radius

- Radius decreases across a period
  Increased effective nuclear charge due to decreased shielding
- Radius increases down a group
  Addition of principal quantum levels

Defining metallic and covalent radii

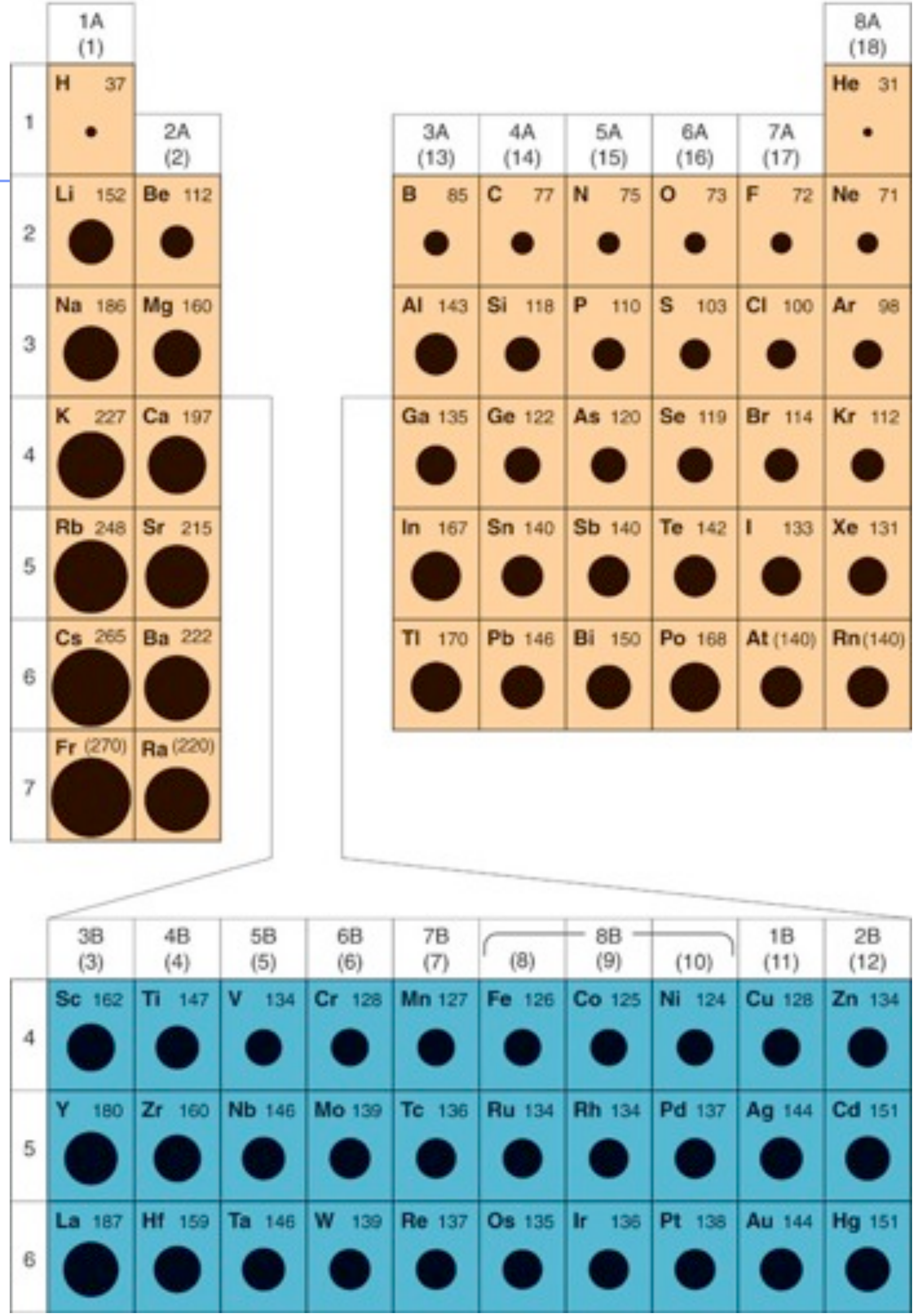




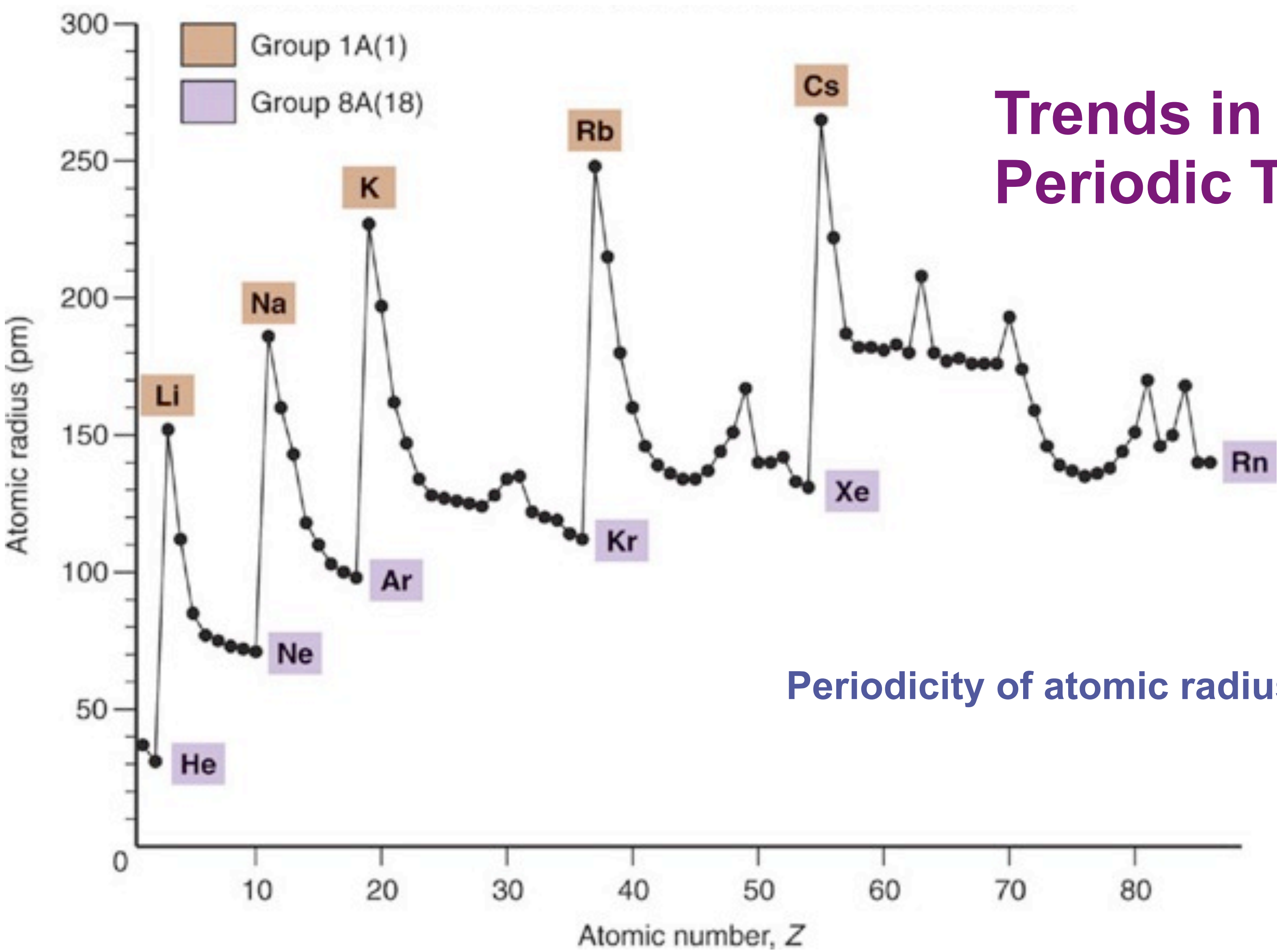
Knowing the Cl radius and the C-Cl bond length, the C radius can be determined.

Trends in the Periodic Table

Atomic radii of the main-group and transition elements.



Trends in the Periodic Table



Periodicity of atomic radius

Ranking Elements by Atomic Size

PROBLEM: Using only the periodic table rank each set of main group elements in order of *decreasing* atomic size:

(a) Ca, Mg, Sr **(b)** K, Ga, Ca **(c)** Br, Rb, Kr **(d)** Sr, Ca, Rb

PLAN: Elements in the same group decrease in size as you go up; elements decrease in size as you go across a period.

SOLUTION:

(a) Sr > Ca > Mg These elements are in Group 2A(2).

(b) K > Ca > Ga These elements are in Period 4.

(c) Rb > Br > Kr Rb has a higher energy level and is far to the left.
Br is to the left of Kr.

(d) Rb > Sr > Ca Ca is one energy level smaller than Rb and Sr.
Rb is to the left of Sr.

Ionization Energy

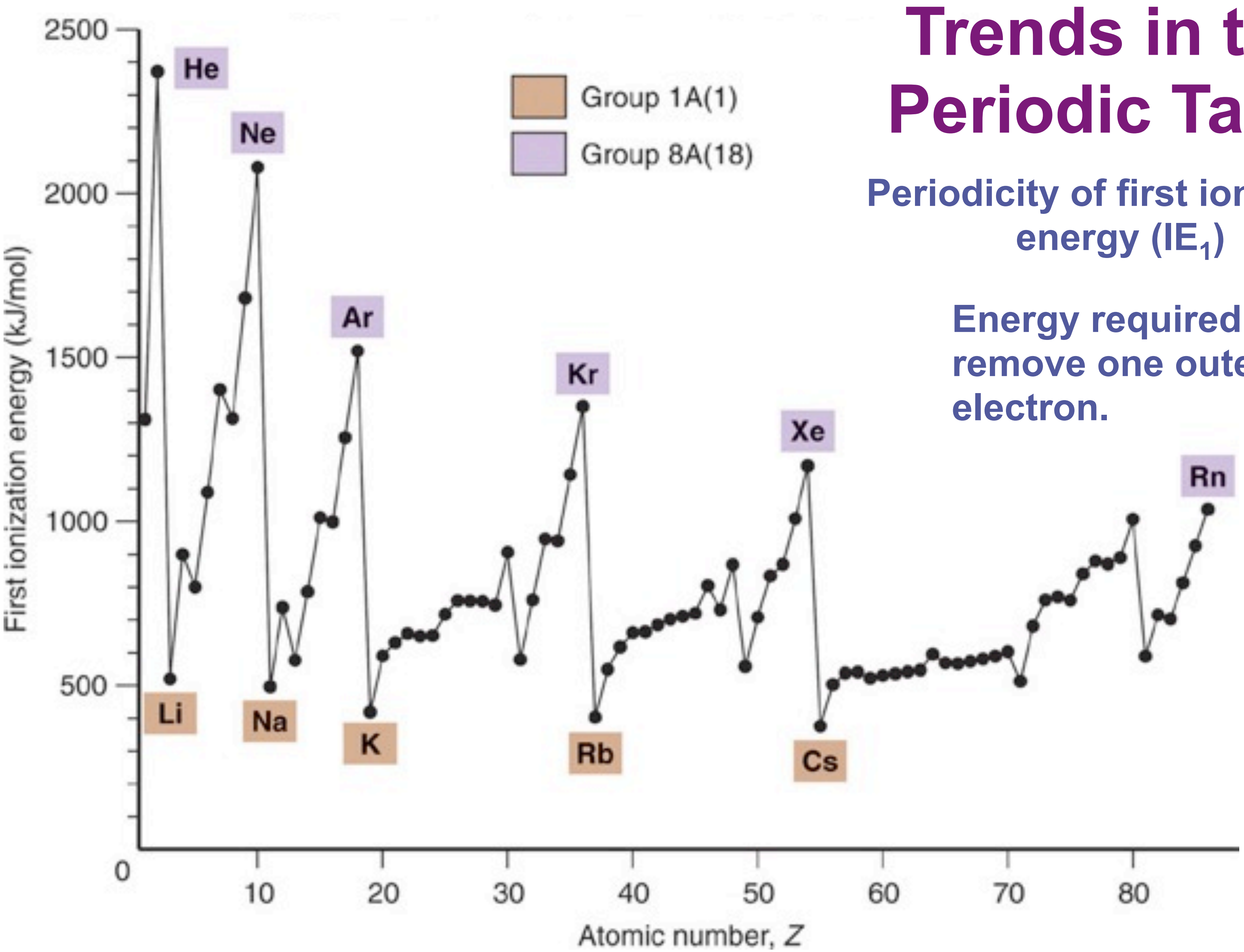
the energy required to remove an electron from an atom

- Increases for successive electrons taken from the same atom
- Tends to increase across a period
  - Electrons in the same quantum level do not shield as effectively as electrons in inner levels
  - Irregularities at half filled and filled sublevels due to extra repulsion of electrons paired in orbitals, making them easier to remove
- Tends to decrease down a group
  - Outer electrons are farther from the nucleus

Trends in the Periodic Table

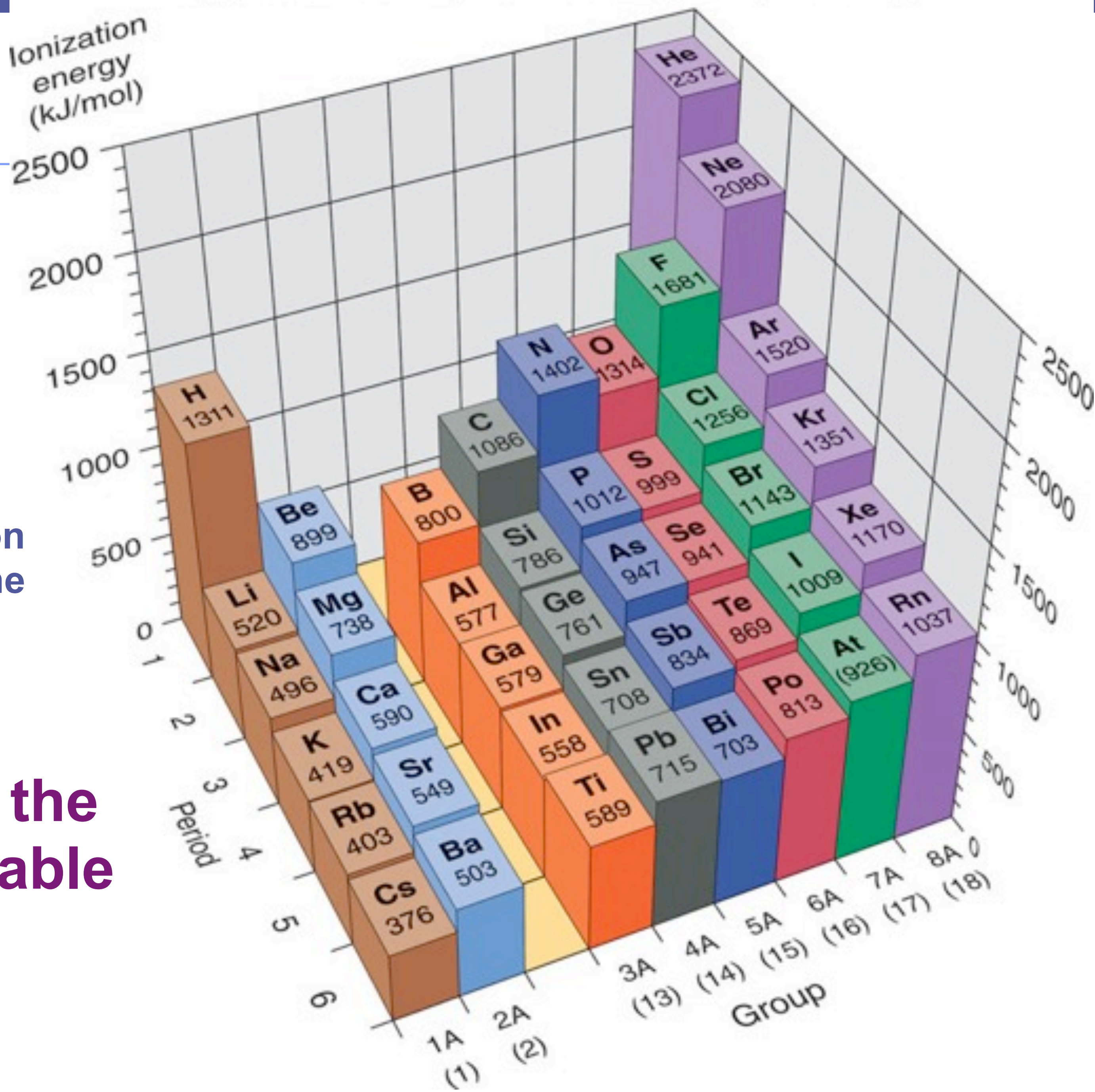
Periodicity of first ionization energy (IE_1)

Energy required to remove one outermost electron.



Trends in the Periodic Table

First ionization energies of the main-group elements



Ranking Elements by First Ionization Energy

PROBLEM: Using the periodic table only, rank the elements in each of the following sets in order of *decreasing* IE_1:

(a) Kr, He, Ar **(b)** Sb, Te, Sn **(c)** K, Ca, Rb **(d)** I, Xe, Cs

PLAN: IE increases as you proceed up in a group; IE increases as you go across a period.

SOLUTION:

(a) He > Ar > Kr — Group 8A(18) - IE decreases down a group.

(b) Te > Sb > Sn — Period 5 elements - IE increases across a period.

(c) Ca > K > Rb — Ca is to the right of K; Rb is below K.

(d) Xe > I > Cs — I is to the left of Xe; Cs is further to the left and down one period.

Trends in the Periodic Table

The first three ionization energies of beryllium (in MJ/mol)

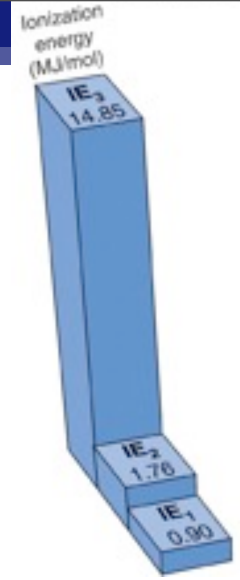


Table 8.5 Successive Ionization Energies of the Elements Lithium Through Sodium

			Ionization Energy (MJ/mol)*									
Z	Element	Number of Valence Electrons	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8	IE_9	IE_{10}
3	Li	1	0.52	7.30	11.81							
4	Be	2	0.90	1.76	14.85	21.01						
5	B	3	0.80	2.43	3.66	25.02	32.82					
6	C	4	1.09	2.35	4.62	6.22	37.83	47.28				
7	N	5	1.40	2.86	4.58	7.48	9.44	53.27	64.36			
8	O	6	1.31	3.39	5.30	7.47	10.98	13.33	71.33	84.08		
9	F	7	1.68	3.37	6.05	8.41	11.02	15.16	17.87	92.04	106.43	
10	Ne	8	2.08	3.95	6.12	9.37	12.18	15.24	20.00	23.07	115.38	131.43
11	Na	1	0.50	4.56	6.91	9.54	13.35	16.61	20.11	25.49	28.93	141.37

Core electrons

*MJ/mol, or megajoules per mole = 10^3 kJ/mol.

Identifying an Element from Successive Ionization Energies

PROBLEM: Name the Period 3 element with the following ionization energies (in kJ/mol) and write its electron configuration:

IE_1	IE_2	IE_3	IE_4	IE_5	IE_6
1012	1903	2910	4956	6278	22,230

PLAN: Look for a large increase in energy which indicates that all of the valence electrons have been removed.

SOLUTION:

The largest increase occurs after IE_5, that is, after the 5th valence electron has been removed. Five electrons would mean that the valence configuration is $3s^2 3p^3$ and the element must be phosphorous, P (Z = 15).

The complete electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^3$.

Electronegativity

A measure of the ability of an atom in a chemical compound to attract electrons

- Electronegativities tend to increase across a period
- Electronegativities tend to decrease down a group or remain the same

Periodic Table of Electronegativities

- below 1.0
- 1.0–1.4
- 1.5–1.9
- 2.0–2.4
- 2.5–2.9
- 3.0–4.0

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
H 2.1																
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.8	Ba 0.9	La* 1.1	Hf 1.3	Ta 1.5	W 2.4	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac† 1.1														

*Lanthanides: 1.1–1.3
†Actinides: 1.3–1.5

Trends in the Periodic Table

Electron affinities of the main-group elements

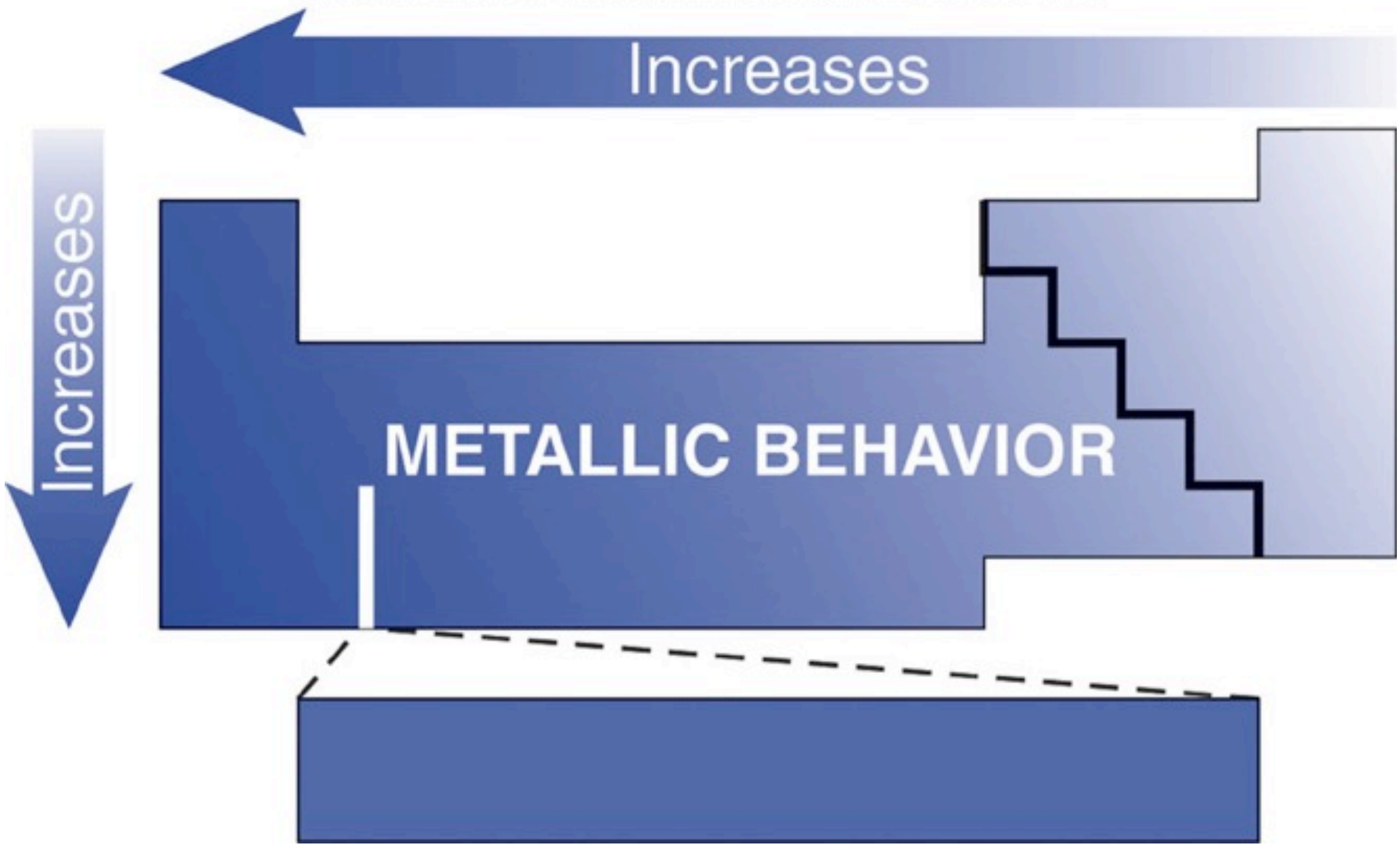
1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
H −72.8							He (0.0)
Li −59.6	Be (+18)	B −26.7	C −122	N +7	O −141	F −328	Ne (+29)
Na −52.9	Mg (+21)	Al −42.5	Si −134	P −72.0	S −200	Cl −349	Ar (+35)
K −48.4	Ca (+186)	Ga −28.9	Ge −119	As −78.2	Se −195	Br −325	Kr (+39)
Rb −46.9	Sr (+146)	In −28.9	Sn −107	Sb −103	Te −190	I −295	Xe (+41)
Cs −45.5	Ba (+46)	Tl −19.3	Pb −35.1	Bi −91.3	Po −183	At −270	Rn (+41)

Electron Affinity:

Energy change to add one electron.

In most cases, EA negative (energy released because electron attracted to nucleus

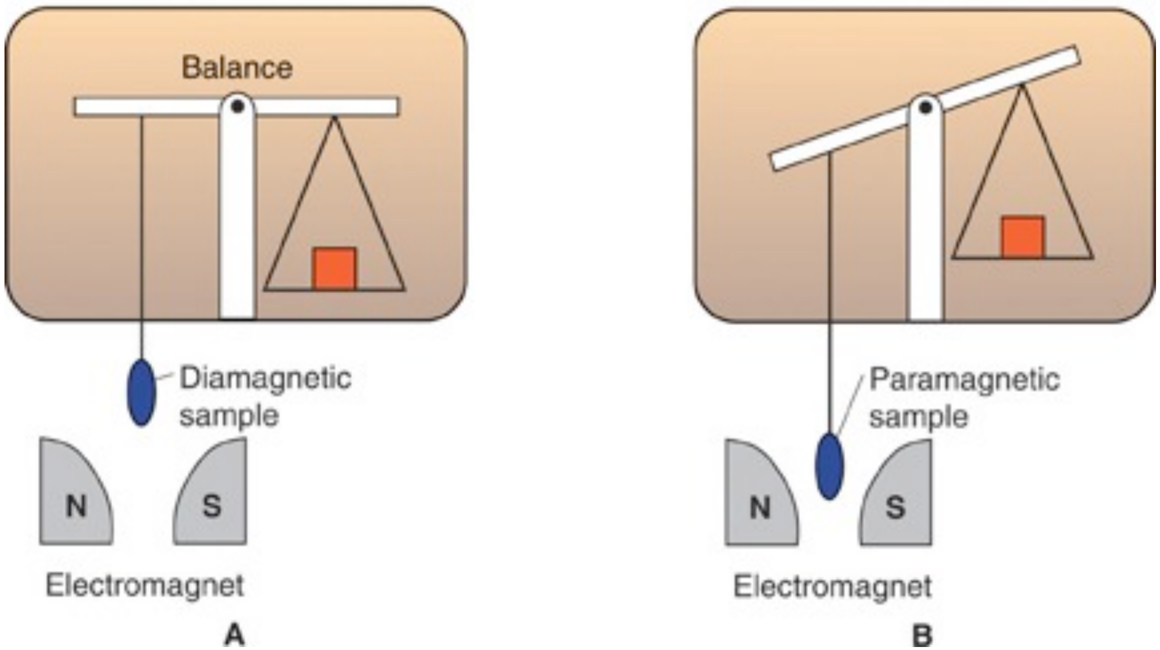
Trends in metallic behavior



Magnetic Properties of Transition Metal Ions

A species with unpaired electrons exhibits **paramagnetism**. It is attracted by an external magnetic field.

Species with all paired e's, not attracted........**diamagnetic**



Writing Electron Configurations and Predicting Magnetic Behavior of Transition Metal Ions

PROBLEM: Use condensed electron configurations to write the reaction for the formation of each transition metal ion, and predict whether the ion is paramagnetic.

(a) Mn^{2+}(Z = 25) **(b)** Cr^{3+}(Z = 24) **(c)** Hg^{2+}(Z = 80)

PLAN: Write the electron configuration and remove electrons starting with ns to match the charge on the ion. If the remaining configuration has unpaired electrons, it is paramagnetic.

SOLUTION:

(a) Mn^{2+}(Z = 25) $Mn([Ar]4s^23d^5) \longrightarrow Mn^{2+}([Ar]\,3d^5) + 2e^-$ paramagnetic

(b) Cr^{3+}(Z = 24) $Cr([Ar]4s^23d^4) \longrightarrow Cr^{3+}([Ar]\,3d^3) + 3e^-$ paramagnetic

(c) Hg^{2+}(Z = 80) $Hg([Xe]6s^24f^{14}5d^{10}) \longrightarrow Hg^{2+}([Xe]\,4f^{14}5d^{10}) + 2e^-$

not paramagnetic (is diamagnetic)

Ionic Radii

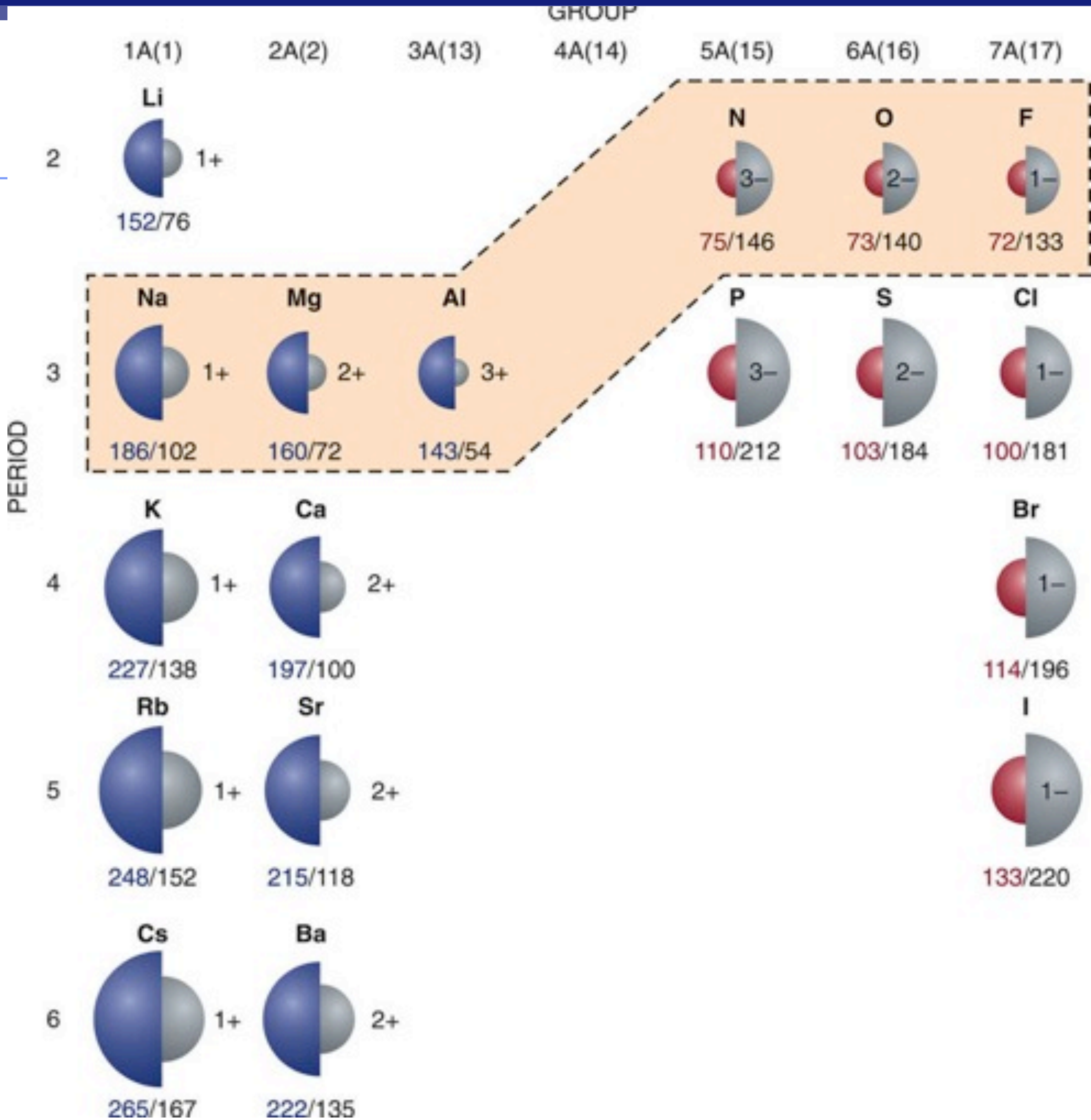
Cations

- Positively charged ions formed when an atom of a metal loses one or more electrons
- Smaller than the corresponding atom

Anions

- Negatively charged ions formed when nonmetallic atoms gain one or more electrons
- Larger than the corresponding atom

Ionic vs. atomic radius



Ranking Ions by Size

PROBLEM: Rank each set of ions in order of *decreasing* size, and explain your ranking:

(a) Ca^{2+}, Sr^{2+}, Mg^{2+} **(b)** K^{+}, S^{2-}, Cl^{-} **(c)** Au^{+}, Au^{3+}

PLAN: Compare positions in the periodic table, formation of positive and negative ions and changes in size due to gain or loss of electrons.

SOLUTION:

(a) $Sr^{2+} > Ca^{2+} > Mg^{2+}$ These are members of the same Group (2A/2) and therefore decrease in size going up the group.

(b) $S^{2-} > Cl^{-} > K^{+}$ The ions are isoelectronic; S^{2-} has the smallest Z_{eff} and therefore is the largest while K^{+} is a cation with a large Z_{eff} and is the smallest.

(c) $Au^{+} > Au^{3+}$ The higher the + charge, the smaller the ion.