Chapter 7

Electron Configuration and the Periodic Table

Topics

- Development of the periodic table
- The modern periodic table
- Effective nuclear charge
- Periodic trends in properties of elements
- Electron configuration of ions
- Ionic radius
- Periodic trends in chemical properties of the main group elements (Self study)

7.1 Development of the Periodic Table

- 1864 - John Newlands - Law of Octaves-
  - every 8th element had similar properties when arranged by atomic masses (not true past Ca)
- 1869 - Dmitri Mendeleev & Lothar Meyer - independently proposed idea of periodicity (repetition of properties)
  - They didn't know much about atom.
  - Elements were arranged in columns by similar properties.
  - Properties of missing elements were predicted.

Mendeleev

- Grouped elements (66) according to properties
- Predicted properties for elements not yet discovered
- Though a good model, Mendeleev could not explain inconsistencies, for instance, all elements were not in order according to atomic mass

Mendeleev's Early Periodic Table, Published in 1872

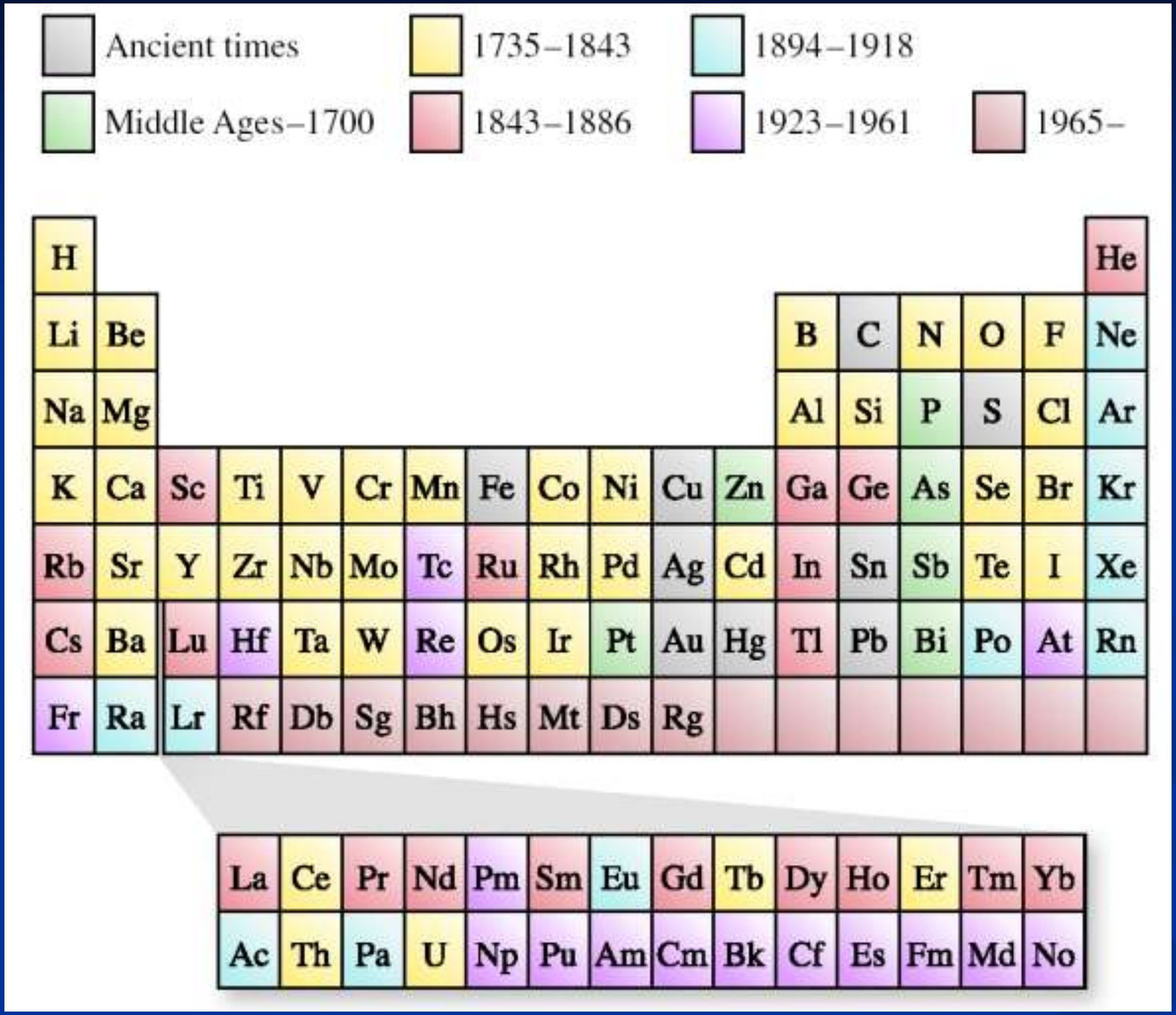
TABELLE II

REIHEN	GRUPPE I. — R^2O	GRUPPE II. — RO	GRUPPE III. — R^2O^3	GRUPPE IV. RH^4 RO^2	GRUPPE V. RH^3 R^2O^5	GRUPPE VI. RH^2 RO^3	GRUPPE VII. RH R^2O^7	GRUPPE VIII. — RO^4
1	H = 1							
2	Li = 7	Be = 9,4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27,3	Si = 28	P = 31	S = 32	Cl = 35,5	
4	K = 39	Ca = 40	— = 44	Ti = 48	V = 51	Cr = 52	Mn = 55	Fe = 56, Co = 59, Ni = 59, Cu = 63.
5	(Cu = 63)	Zn = 65	— = 68	— = 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	?Yt = 88	Zr = 90	Nb = 94	Mo = 96	— = 100	Ru = 104, Rh = 104, Pd = 106, Ag = 108.
7	(Ag = 108)	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	J = 127	
8	Cs = 133	Ba = 137	?Di = 138	?Ce = 140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er = 178	?La = 180	Ta = 182	W = 184	—	Os = 195, Ir = 197, Pt = 198, Au = 199.
11	(Au = 199)	Hg = 200	Tl = 204	Pb = 207	Bi = 208	—	—	
12	—	—	—	Th = 231	—	U = 240	—	— — — —

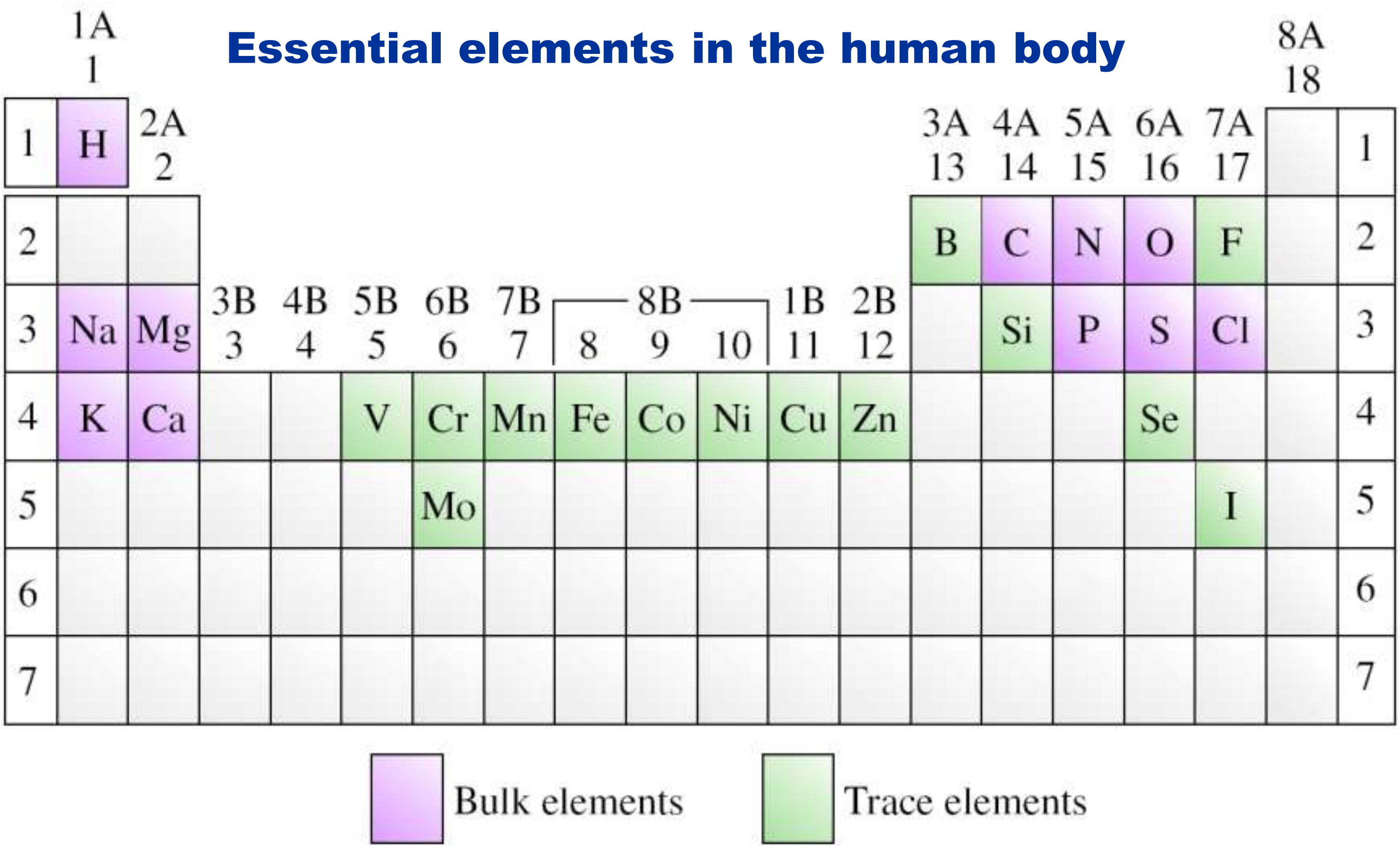
1913 - Henry Moseley explained the discrepancy

- He noticed that the frequency of X-rays generated by bombarding an element with high energy electrons increases with increasing atomic mass with exceptions like Ar and K
- He discovered a correlation between atomic number and frequency of X rays
- Then he realized that elements should be arranged in order of increasing atomic number (# of protons = # electrons)rather than atomic mass.

Periodic Table by Dates of Discovery



Essential Elements in the Human Body



The Modern Periodic Table
& Valence electron configuration

1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
1 **H** $1s^1$																	2 **He** $1s^2$
3 **Li** $2s^1$	4 **Be** $2s^2$											5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$
11 **Na** $3s^1$	12 **Mg** $3s^2$											13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$
19 **K** $4s^1$	20 **Ca** $4s^2$	21 **Sc** $3d^14s^2$	22 **Ti** $3d^24s^2$	23 **V** $3d^34s^2$	24 **Cr** $3d^54s^1$	25 **Mn** $3d^54s^2$	26 **Fe** $3d^64s^2$	27 **Co** $3d^74s^2$	28 **Ni** $3d^84s^2$	29 **Cu** $3d^{10}4s^1$	30 **Zn** $3d^{10}4s^2$	31 **Ga** $4s^24p^1$	32 **Ge** $4s^24p^2$	33 **As** $4s^24p^3$	34 **Se** $4s^24p^4$	35 **Br** $4s^24p^5$	36 **Kr** $4s^24p^6$
37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $4d^15s^2$	40 **Zr** $4d^25s^2$	41 **Nb** $4d^45s^1$	42 **Mo** $4d^55s^1$	43 **Tc** $4d^55s^2$	44 **Ru** $4d^75s^1$	45 **Rh** $4d^85s^1$	46 **Pd** $4d^{10}$	47 **Ag** $4d^{10}5s^1$	48 **Cd** $4d^{10}5s^2$	49 **In** $5s^25p^1$	50 **Sn** $5s^25p^2$	51 **Sb** $5s^25p^3$	52 **Te** $5s^25p^4$	53 **I** $5s^25p^5$	54 **Xe** $5s^25p^6$
55 **Cs** $6s^1$	56 **Ba** $6s^2$	57 ***La** $5d^16s^2$	72 **Hf** $5d^26s^2$	73 **Ta** $5d^36s^2$	74 **W** $5d^46s^2$	75 **Re** $5d^56s^2$	76 **Os** $5d^66s^2$	77 **Ir** $5d^76s^2$	78 **Pt** $5d^96s^1$	79 **Au** $5d^{10}6s^1$	80 **Hg** $5d^{10}6s^2$	81 **Tl** $6s^26p^1$	82 **Pb** $6s^26p^2$	83 **Bi** $6s^26p^3$	84 **Po** $6s^26p^4$	85 **At** $6s^26p^5$	86 **Rn** $6s^26p^6$
87 **Fr** $7s^1$	88 **Ra** $7s^2$	89 **†Ac** $6d^17s^2$	104 **Rf** $6d^27s^2$	105 **Db** $6d^37s^2$	106 **Sg** $6d^47s^2$	107 **Bh**	108 **Hs**	109 **Mt**	110 **Ds**	111	112	Unknown	114	Unknown	116		

*	58 **Ce** $4f^26s^2$	59 **Pr** $4f^36s^2$	60 **Nd** $4f^46s^2$	61 **Pm** $4f^56s^2$	62 **Sm** $4f^66s^2$	63 **Eu** $4f^76s^2$	64 **Gd** $4f^75d^16s^2$	65 **Tb** $4f^96s^2$	66 **Dy** $4f^{10}6s^2$	67 **Ho** $4f^{11}6s^2$	68 **Er** $4f^{12}6s^2$	69 **Tm** $4f^{13}6s^2$	70 **Yb** $4f^{14}6s^2$	71 **Lu** $4f^{14}5d^16s^2$
†	90 **Th** $6d^27s^2$	91 **Pa** $5f^26d^17s^2$	92 **U** $5f^36d^17s^2$	93 **Np** $5f^46d^17s^2$	94 **Pu** $5f^67s^2$	95 **Am** $5f^77s^2$	96 **Cm** $5f^76d^17s^2$	97 **Bk** $5f^97s^2$	98 **Cf** $5f^{10}7s^2$	99 **Es** $5f^{11}7s^2$	100 **Fm** $5f^{12}7s^2$	101 **Md** $5f^{13}7s^2$	102 **No** $5f^{14}7s^2$	103 **Lr** $5f^{14}6d^17s^2$

7.2 The Modern Periodic Table

Classification of Elements

- Main group elements - "representative elements" Group 1A-7A
- Noble gases - Group 8A all have ns^2np^6 configuration(exception-He)
- Transition elements - 1B, 3B - 8B "*d*-block"
- Lanthanides/actinides - "*f*-block"

Classification of Elements

Elements in which the orbitals being filled in the aufbau process are either *s* or *p* orbitals of the outermost shell are called main group (Representative) elements

"A" group designation on the periodic table

The first 20 elements are all main group elements

Transition Elements

In transition elements, the sublevel (shell) being filled in the aufbau process is in an inner principal shell (d or f)

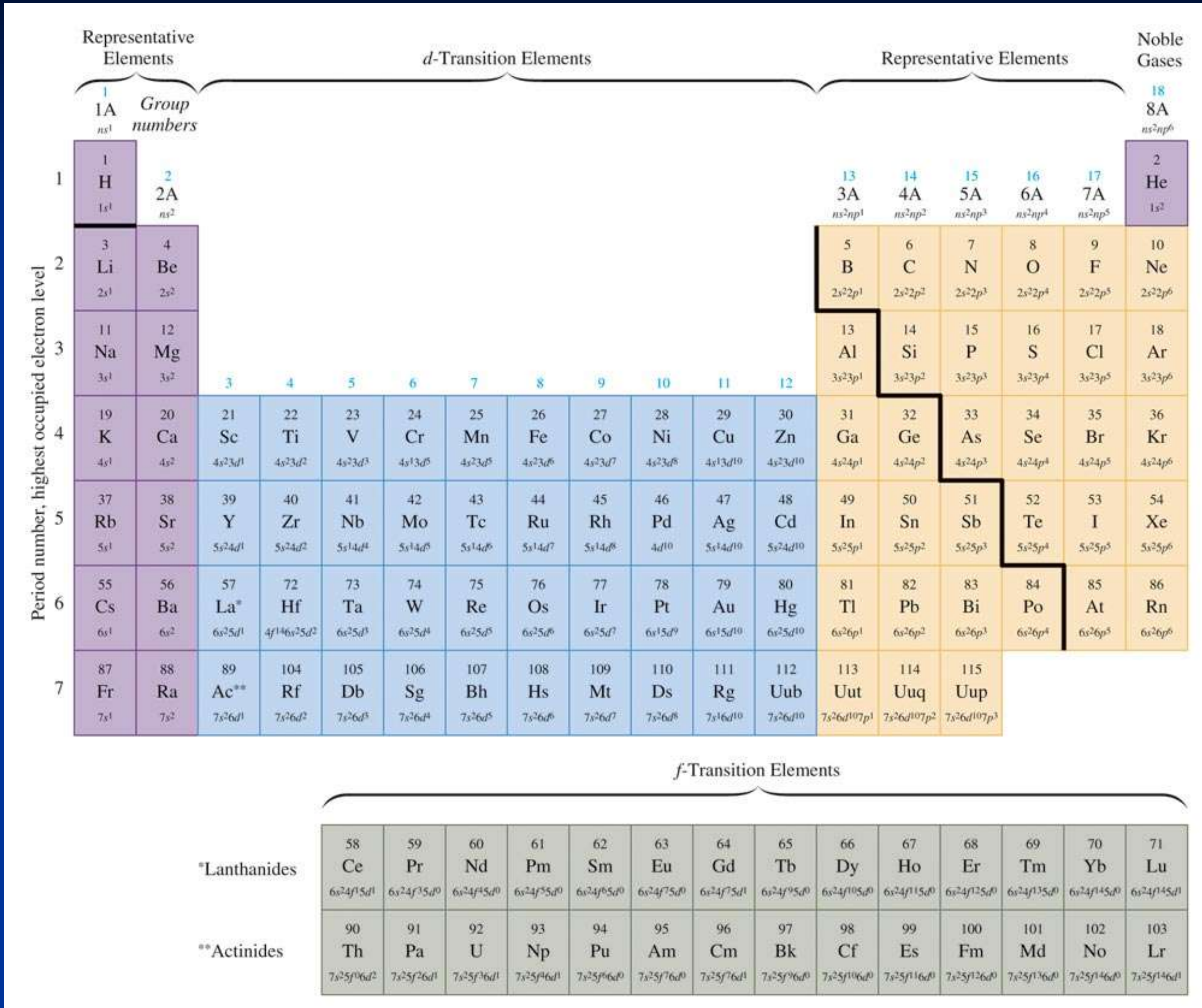
d-Block **transition elements: Electrons enter the d-sublevels (incomplete d-subshells**

f-Block **transition elements: *d* sublevels are completely filled. Electrons enter f-sublevels**

Lanthanides: electrons fill 4f subleves

Actinides: electrons fill 5f sublevels

The Periodic Table



Periodic Table Colored Coded By Main Classifications

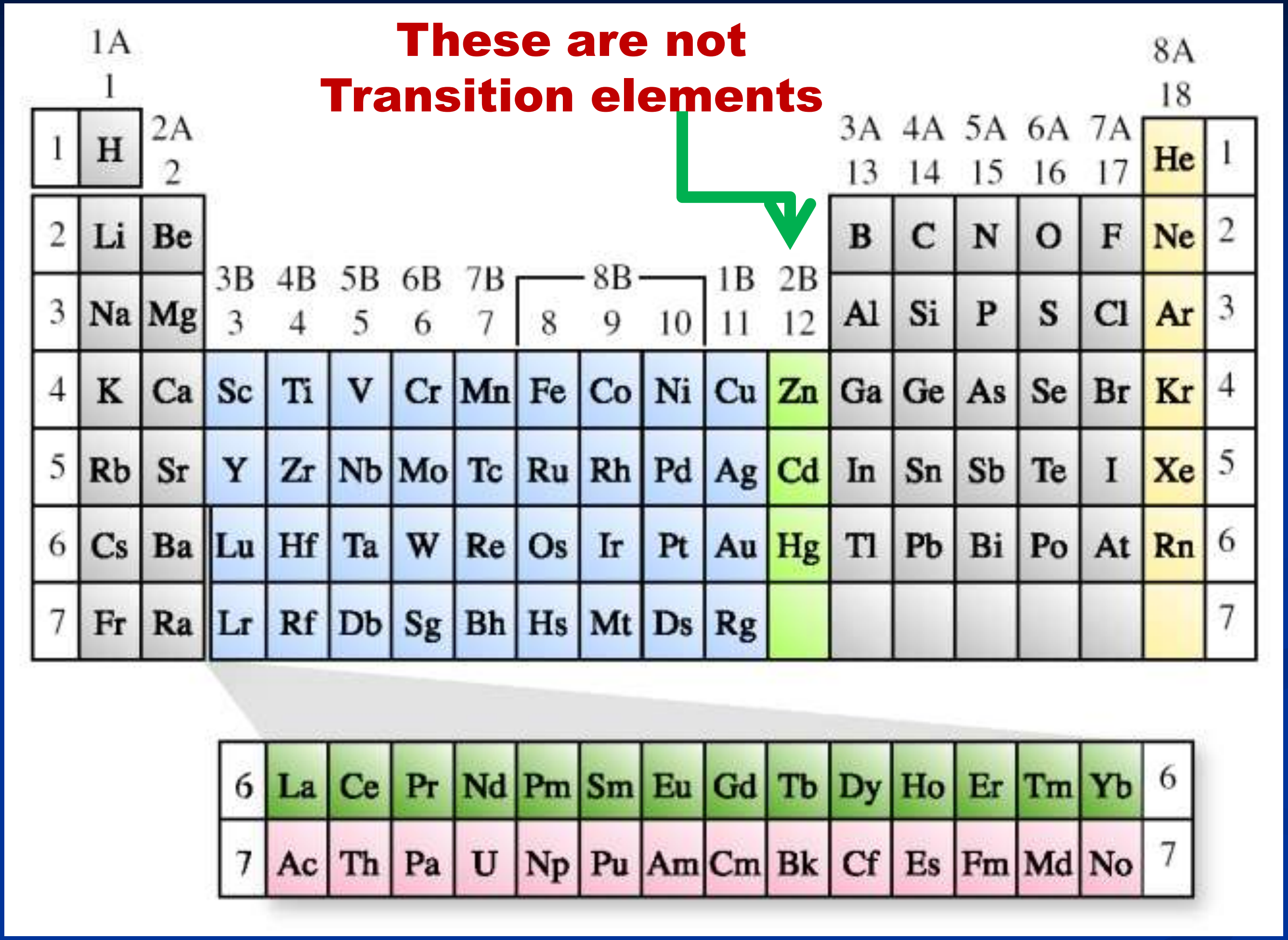


TABLE 7.1 Electron Configurations of Group 1A and Group 2A Elements

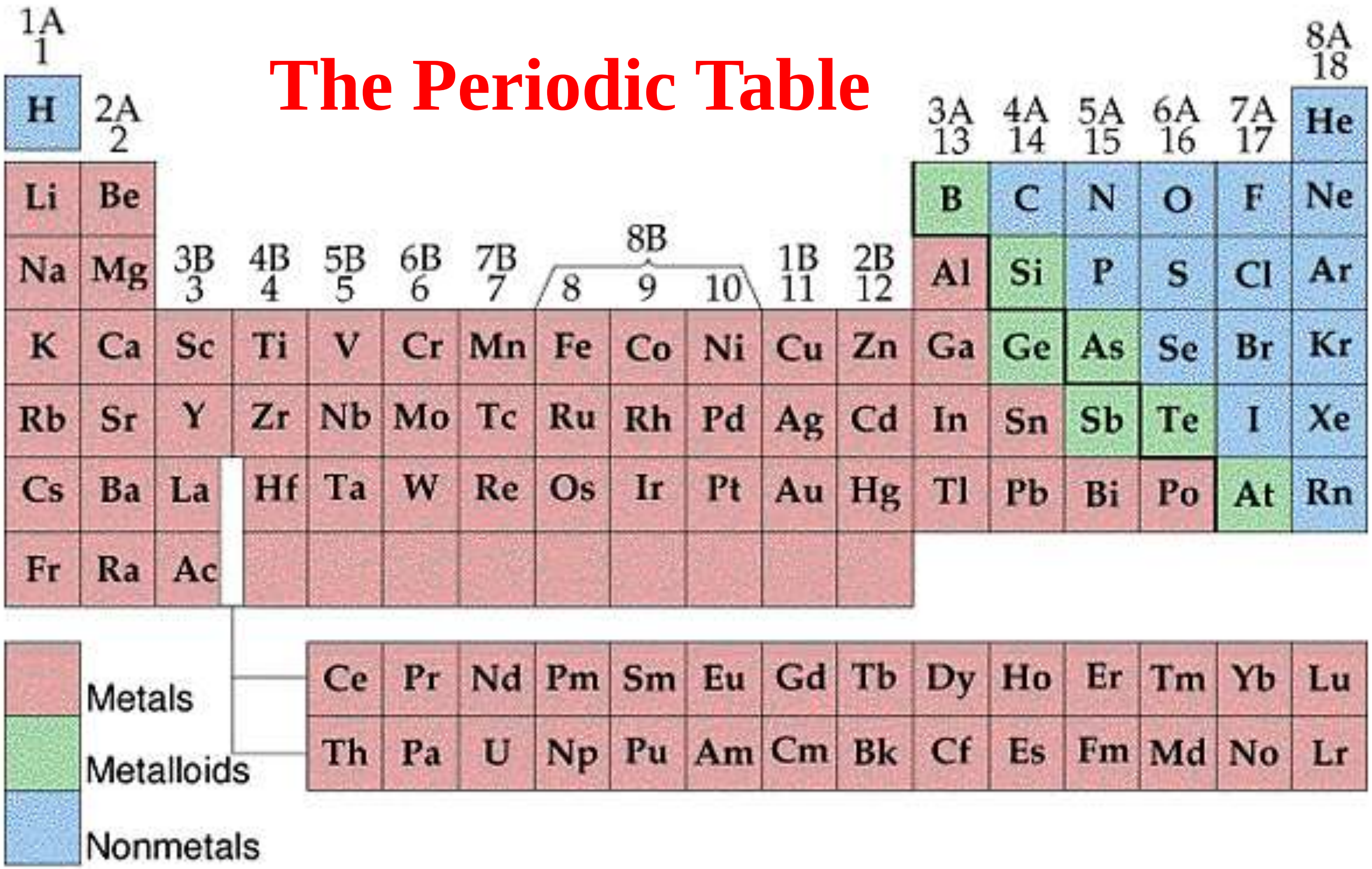
Group 1A		Group 2A	
Li	$[He]2s^1$	Be	$[He]2s^2$
Na	$[Ne]3s^1$	Mg	$[Ne]3s^2$
K	$[Ar]4s^1$	Ca	$[Ar]4s^2$
Rb	$[Kr]5s^1$	Sr	$[Kr]5s^2$
Cs	$[Xe]6s^1$	Ba	$[Xe]6s^2$
Fr	$[Rn]7s^1$	Ra	$[Rn]7s^2$

Predicting properties

- **Valence electrons** are the outermost electrons and are involved **in bonding**
- **Similarity of valence electron configurations** help predict chemical properties
- Group 1A, 2A and 8A all have similar properties to other members of their respective group

- Groups 3A - 7A show considerable variation among properties from metallic to nonmetallic
- Transition metals do not always exhibit regular patterns in their electron configurations but have some similarities as a whole such as colored compounds and multiple oxidation states.

The Periodic Table



Representing Free Elements in Chemical Equations

- Metals are always represented by their empirical formulas (same as symbol for element)
- Nonmetals may be written as empirical formula (C) or as polyatomic molecules (H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2, and P_4).
- Sulfur usually S instead of S_8

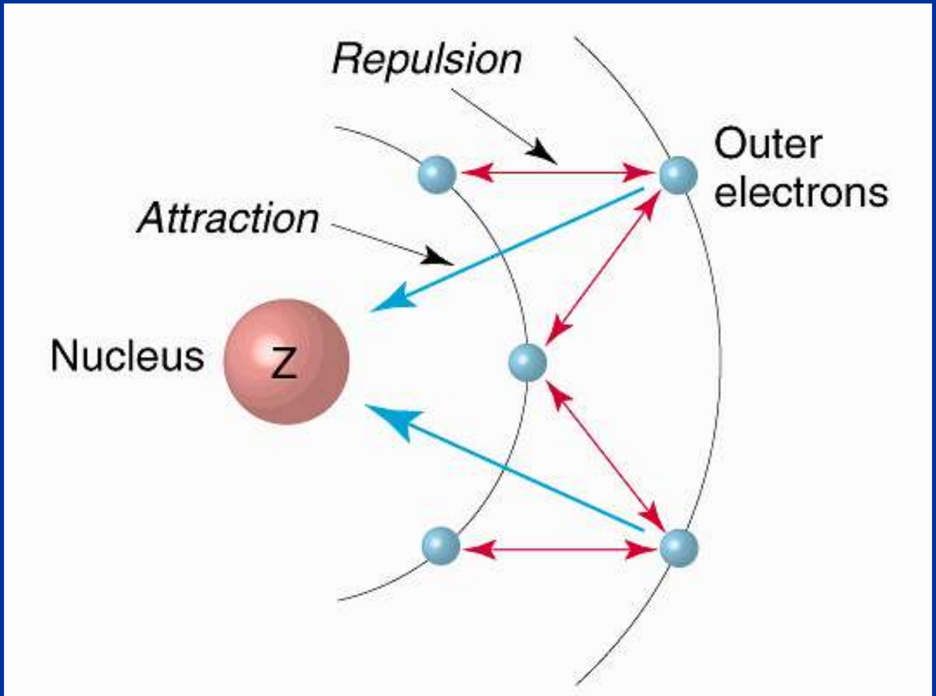
- **Noble Gases** all exist as isolated atoms, so use symbols (Xe, He, etc.)
- **Metalloids** are represented with empirical formulas (B, Si, Ge, etc.)

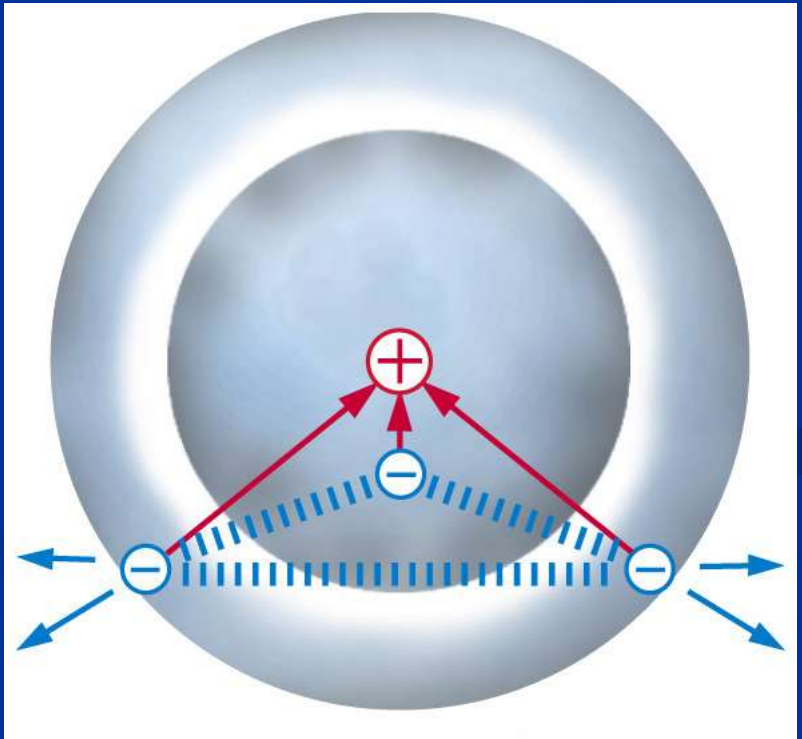
7.3 Effective Nuclear Charge

- *Z* (*nuclear charge*) = the number of protons in the nucleus of an atom
- Z_{eff} (*effective nuclear charge*) = the magnitude of positive charge "experienced" by an electron in the atom
- Z_{eff} increases from left to right across a period; changes very little down a column

Electron Shielding

- occurs when an electron (outer) is not attracted to the nucleus
- because of electrons (core) in lower energy levels repelling it.





- Shielding occurs when an electron in a many-electron atom is partially shielded from the positive charge of the nucleus by other electrons in the atom.
- However, core electrons (inner electrons) shield the most and are constant across a period.

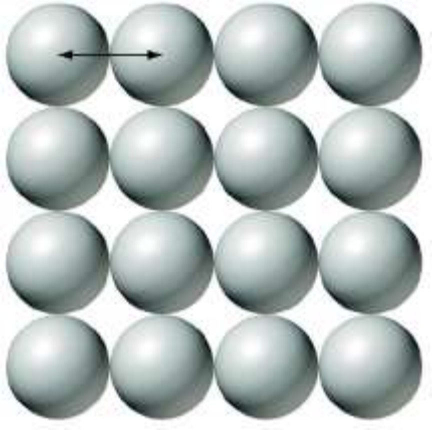
- $Z_{eff} = Z - \sigma$
  - σ represents the shielding constant (greater than 0 but less than Z)
  - Example:

	Li	Be	B	C	N
Z	3	4	5	6	7
Z_{eff}	1.28	1.91	2.42	3.14	3.83

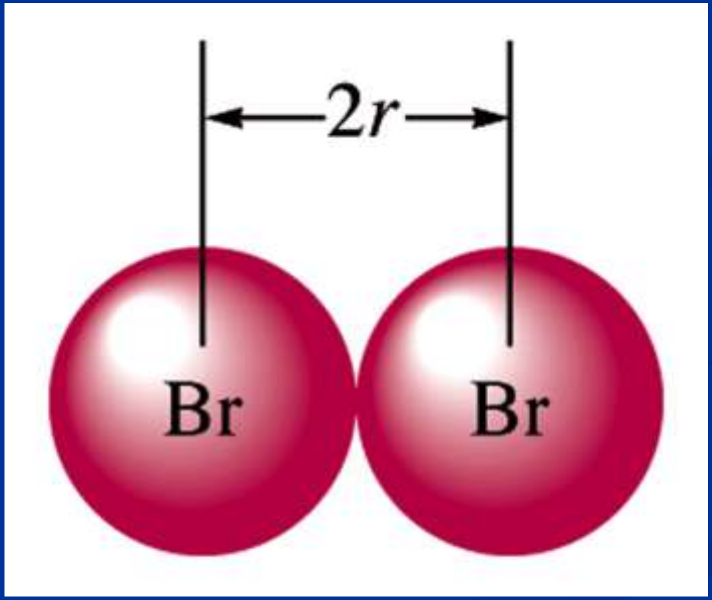
7.4 Periodic Trends in Properties of Elements

Atomic radius

- *Atomic radius*: distance between nucleus of an atom and its valence shell
- Size of shells can not be specified exactly, neither can the size of atom
- *Metallic radius*: half the distance between nuclei of two adjacent, identical metal atoms



- **Covalent radius**: half the distance between adjacent, identical nuclei in a molecule



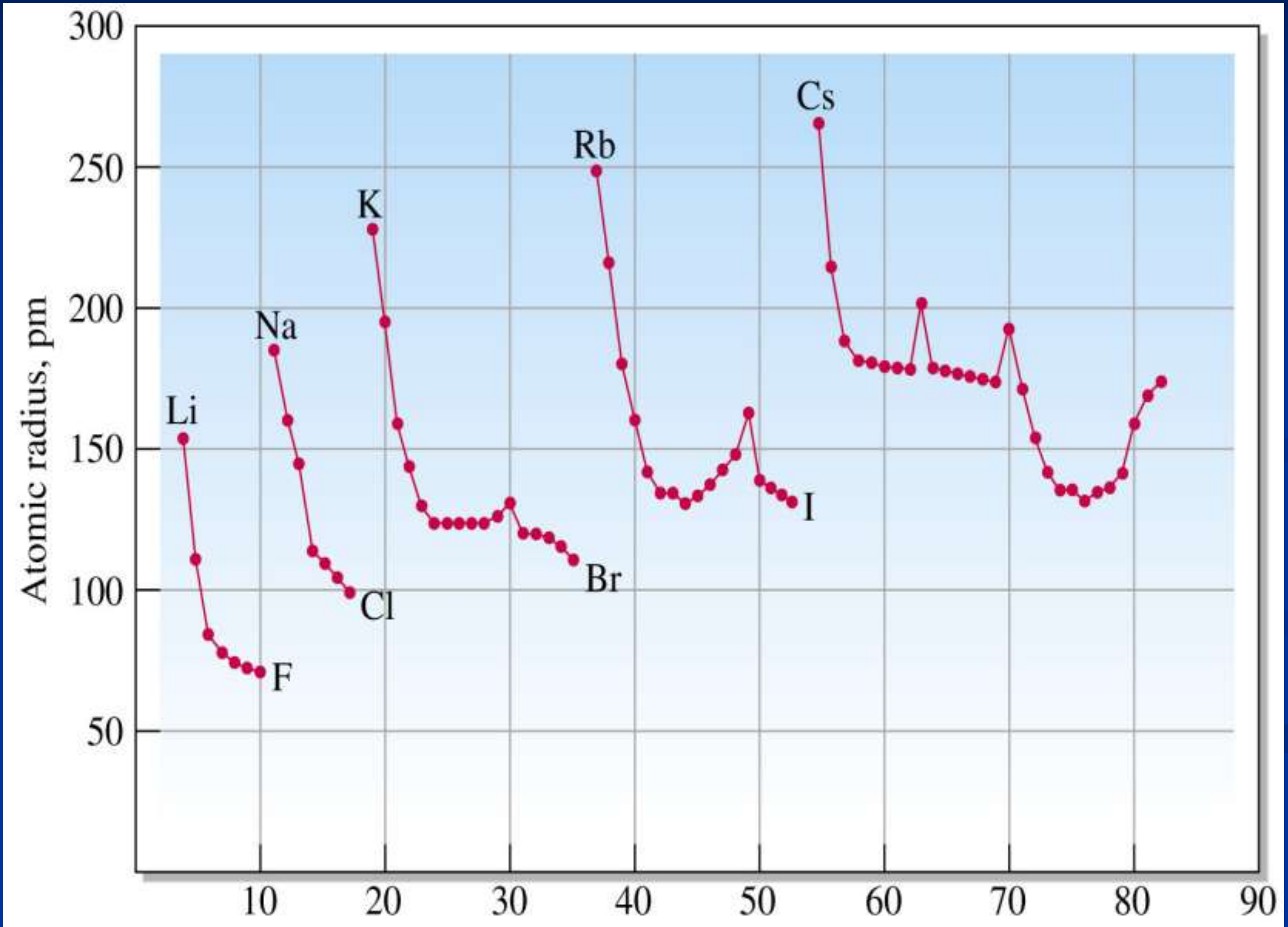
Periodic trends in atomic radii

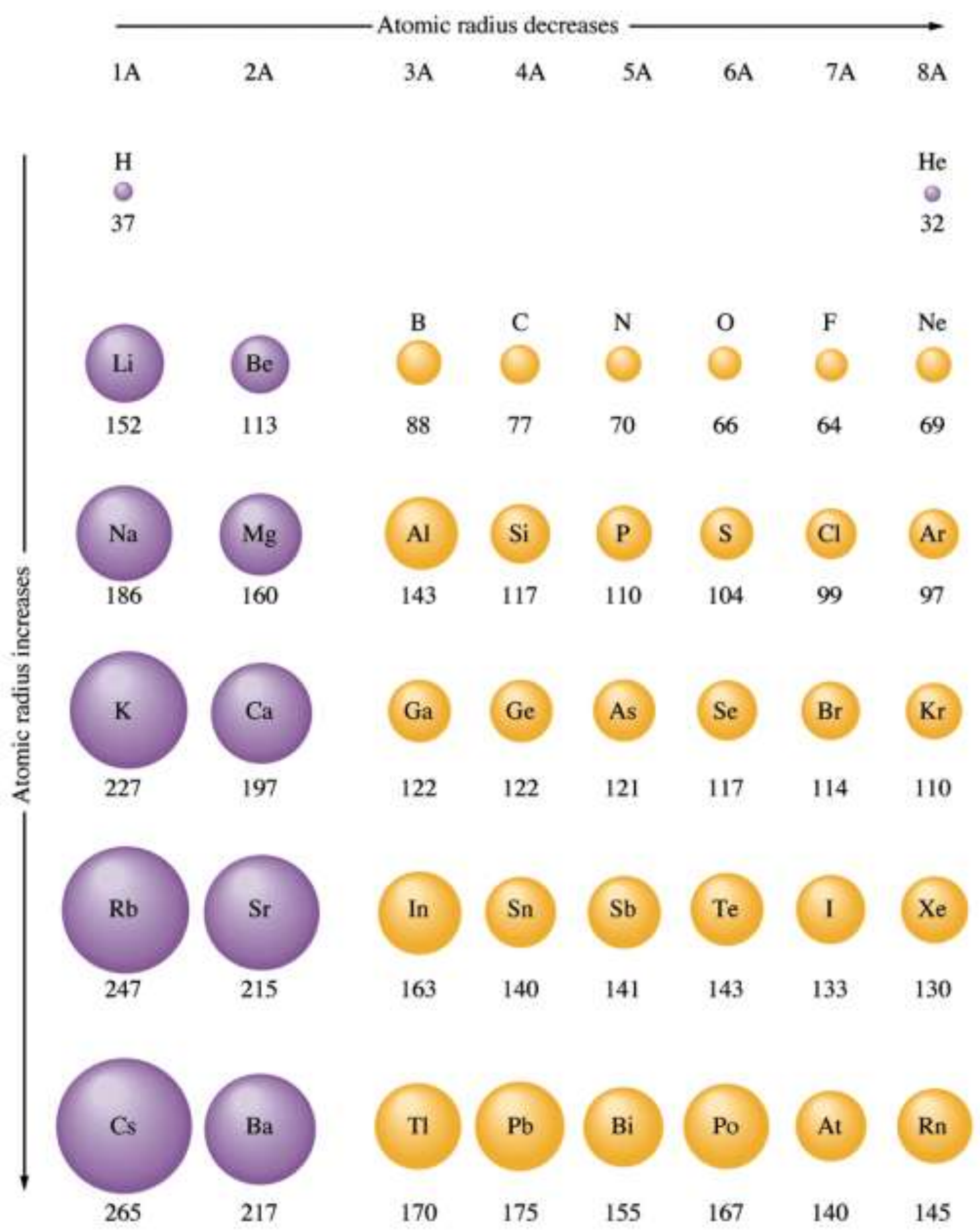
- **Across Period:**
  - atoms get smaller

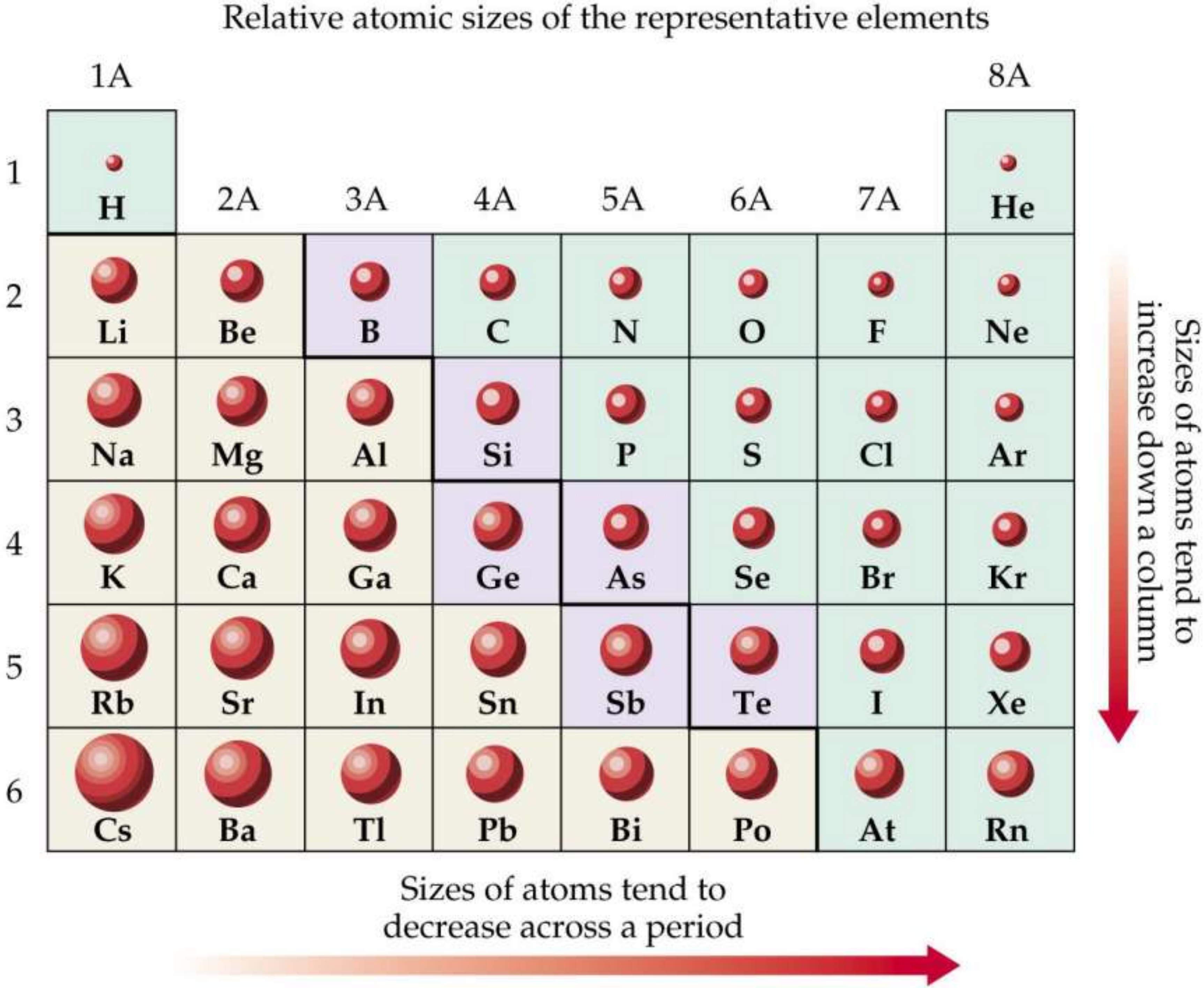
    because of the increased number of protons attracting the electrons
  - the electrons added in the same energy level do not shield electrons from nuclear charge

- **Down Group:**
  - atoms get larger
  - increases
  - because the energy levels being added to the atom

Periodic trends in atomic radii




Atomic radius decreases
Atomic radius increases
1A
2A
3A
4A
5A
6A
7A
8A
H
37
He
32
Li
152
Be
113
B
88
C
77
N
70
O
66
F
64
Ne
69
Na
186
Mg
160
Al
143
Si
117
P
110
S
104
Cl
99
Ar
97
K
227
Ca
197
Ga
122
Ge
122
As
121
Se
117
Br
114
Kr
110
Rb
247
Sr
215
In
163
Sn
140
Sb
141
Te
143
I
133
Xe
130
Cs
265
Ba
217
Tl
170
Pb
175
Bi
155
Po
167
At
140
Rn
145



Relative atomic sizes of the representative elements
1A
8A
1
H
2A
3A
4A
5A
6A
7A
He
2
Li
Be
B
C
N
O
F
Ne
3
Na
Mg
Al
Si
P
S
Cl
Ar
4
K
Ca
Ga
Ge
As
Se
Br
Kr
5
Rb
Sr
In
Sn
Sb
Te
I
Xe
6
Cs
Ba
Tl
Pb
Bi
Po
At
Rn
Sizes of atoms tend to increase down a column
Sizes of atoms tend to decrease across a period

Explain

- What do you notice about the atomic radius across a period? Why? (hint: Z_{eff})

- What do you notice about the atomic radius down a column? Why? (hint: *n*)

- What do you notice about the atomic radius across a period? Why? (hint: Z_{eff})

Atomic radius decreases from left to right across a period due to increasing Z_{eff}.

- What do you notice about the atomic radius down a column? Why? (hint: n)

Atomic radius increases down a column of the periodic table because the distance of the electron from the nucleus increases as n increases.

Ionization energy (*IE*)

minimum energy needed to remove an electron from an atom in the gas phase

- Representation:

$$Na_{(g)} \rightarrow Na^{+}_{(g)} + e^{-}$$

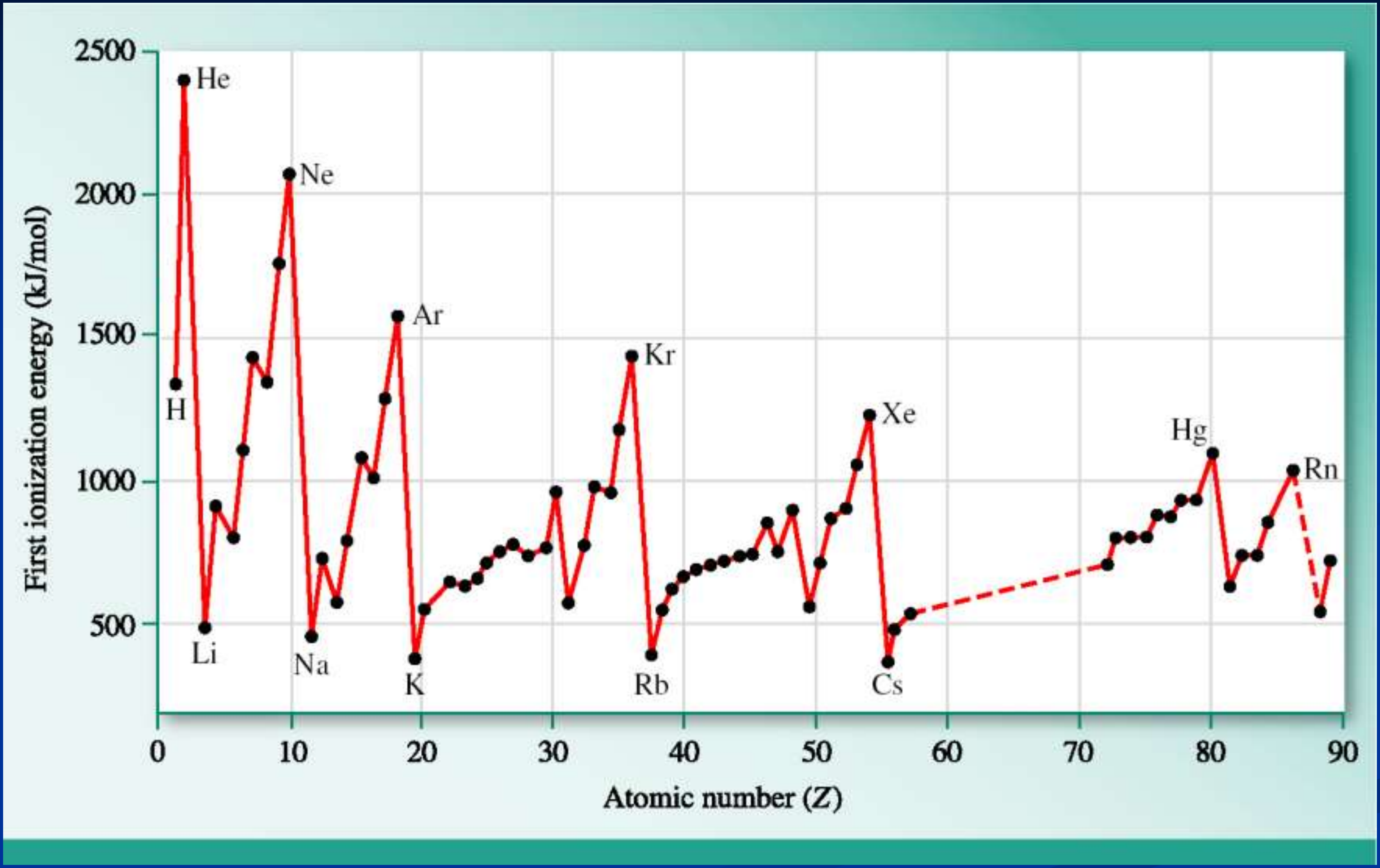
- *IE* for this 1st ionization = 495.8 kJ/mol

- In general, ionization energy increases as Z_{eff} increases
  - Exceptions occur due to the stability of specific electron configurations

IE_1 (kJ/mol) Values for Main Group Elements

1A 1	2A 2	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
H 1312							**He** 2372
Li 520	**Be** 899	**B** 800	**C** 1086	**N** 1402	**O** 1314	**F** 1681	**Ne** 2080
Na 496	**Mg** 738	**Al** 577	**Si** 786	**P** 1012	**S** 999	**Cl** 1256	**Ar** 1520
K 419	**Ca** 590	**Ga** 579	**Ge** 761	**As** 947	**Se** 941	**Br** 1143	**Kr** 1351
Rb 403	**Sr** 549	**In** 558	**Sn** 708	**Sb** 834	**Te** 869	**I** 1009	**Xe** 1170
Cs 376	**Ba** 503	**Tl** 589	**Pb** 715	**Bi** 703	**Po** 813	**At** (926)	**Rn** 1037

Periodic Trends in IE_1



Explain

- What do you notice about the 1st *IE* across a period? Why? (hint: Z_{eff})

- What do you notice about the 1st *IE* down a column? Why? (hint: *n*)

- What do you notice about the 1st *IE* across a period? Why? (hint: Z_{eff})

IE_1 increases from left to right across a period due to increasing Z_{eff}.

- What do you notice about the 1st *IE* down a column? Why? (hint: *n*)

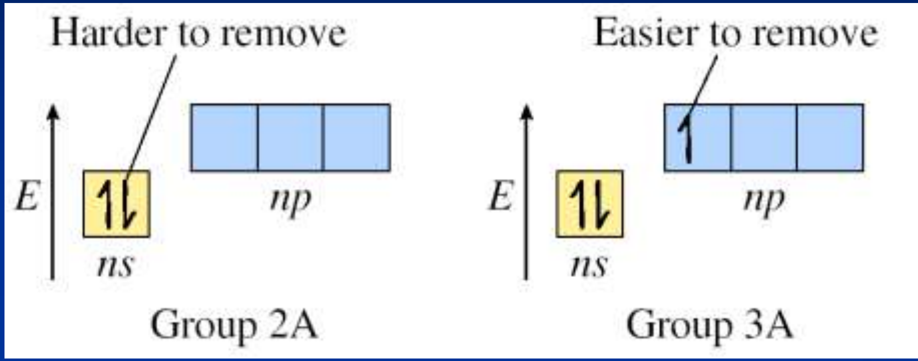
IE_1 decreases down a column of the periodic table because the distance of the electron from the nucleus increases as *n* increases.

Explain

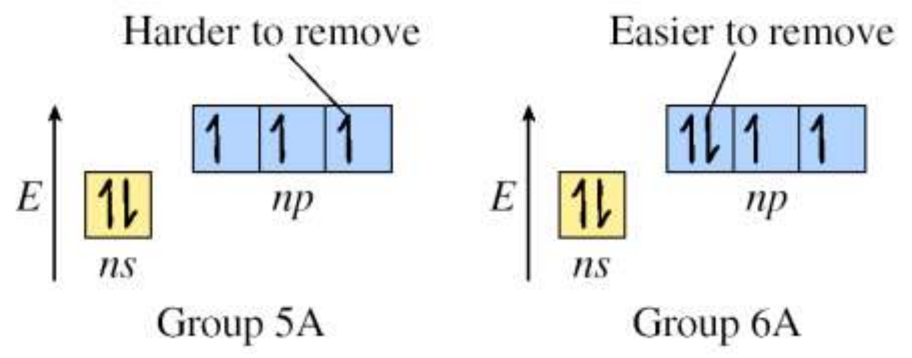
- What do you notice about the 1st *IE* between 2A and 3A? Why? (hint: draw the electron configuration)

- What do you notice about the 1st *IE* between 5A and 6A? Why? (hint: draw the electron configuration)

- What do you notice about the 1st *IE* between 2A and 3A? Why? (hint: draw the electron configuration)



- What do you notice about the 1st *IE* between 5A and 6A? Why? (hint: draw the electron configuration)



Multiple Ionizations

- it takes **more energy to remove the 2nd, 3rd, 4th, etc.** electron and much more energy to remove core electrons
- Why?
  - Core electrons are closer to nucleus
  - Core electrons experience greater Z_{eff}

Successive ionization Energies

- for Mg
  - I_1 = 735 kJ/mole
  - I_2 = 1445 kJ/mole
  - I_3 = 7730 kJ/mole
- The effective nuclear charge increases as electrons are removed
- It takes much more energy to remove a core electron than a valence electron because there is less shielding.

Successive ionization Energies

Continual removal of electrons increases ionization energy greatly

$B \rightarrow B^{+} + e^{-}$ $I = 801\ \text{kJ mol}^{-1}$

$B^{+} \rightarrow B^{+2} + e^{-}$ $I = 2427\ \text{kJ mol}^{-1}$

$B^{+2} \rightarrow B^{+3} + e^{-}$ $I = 3660\ \text{kJ mol}^{-1}$

Core electron

$B^{+3} \rightarrow B^{+4} + e^{-}$ $I = 25{,}025\ \text{kJ mol}^{-1}$

$B^{+4} \rightarrow B^{+5} + e^{-}$ $I = 32{,}822\ \text{kJ mol}^{-1}$

TABLE 7.3	Ionization Energies (in kJ/mol) for Elements 3 through 11*

	Z	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8	IE_9	IE_{10}
Li	3	520	7,298	11,815							
Be	4	899	1,757	14,848	21,007	21,007					
B	5	800	2,427	3,660	25,026	32,827					
C	6	1,086	2,353	4,621	6,223	37,831	47,277				
N	7	1,402	2,856	4,578	7,475	9,445	53,267	64,360			
O	8	1,314	3,388	5,301	7,469	10,990	13,327	71,330	84,078		
F	9	1,681	3,374	6,050	8,408	11,023	15,164	17,868	92,038	106,434	
Ne	10	2,080	3,952	6,122	9,371	12,177	15,238	19,999	23,069	115,380	131,432
Na	11	496	4,562	6,910	9,543	13,354	16,613	20,117	25,496	28,932	141,362

*Cells shaded with blue represent the removal of core electrons.

Electron Affinity (*EA*)

Energy released when an atom in the gas phase accepts an electron

- **Representation:**

$Cl_{(g)} + e^{-} \rightarrow Cl^{-}_{(g)}$ (ΔH = - 349.0 kJ/mol)

- ***EA* for this equation +349.0 kJ/mol energy released (ΔH = negative)**

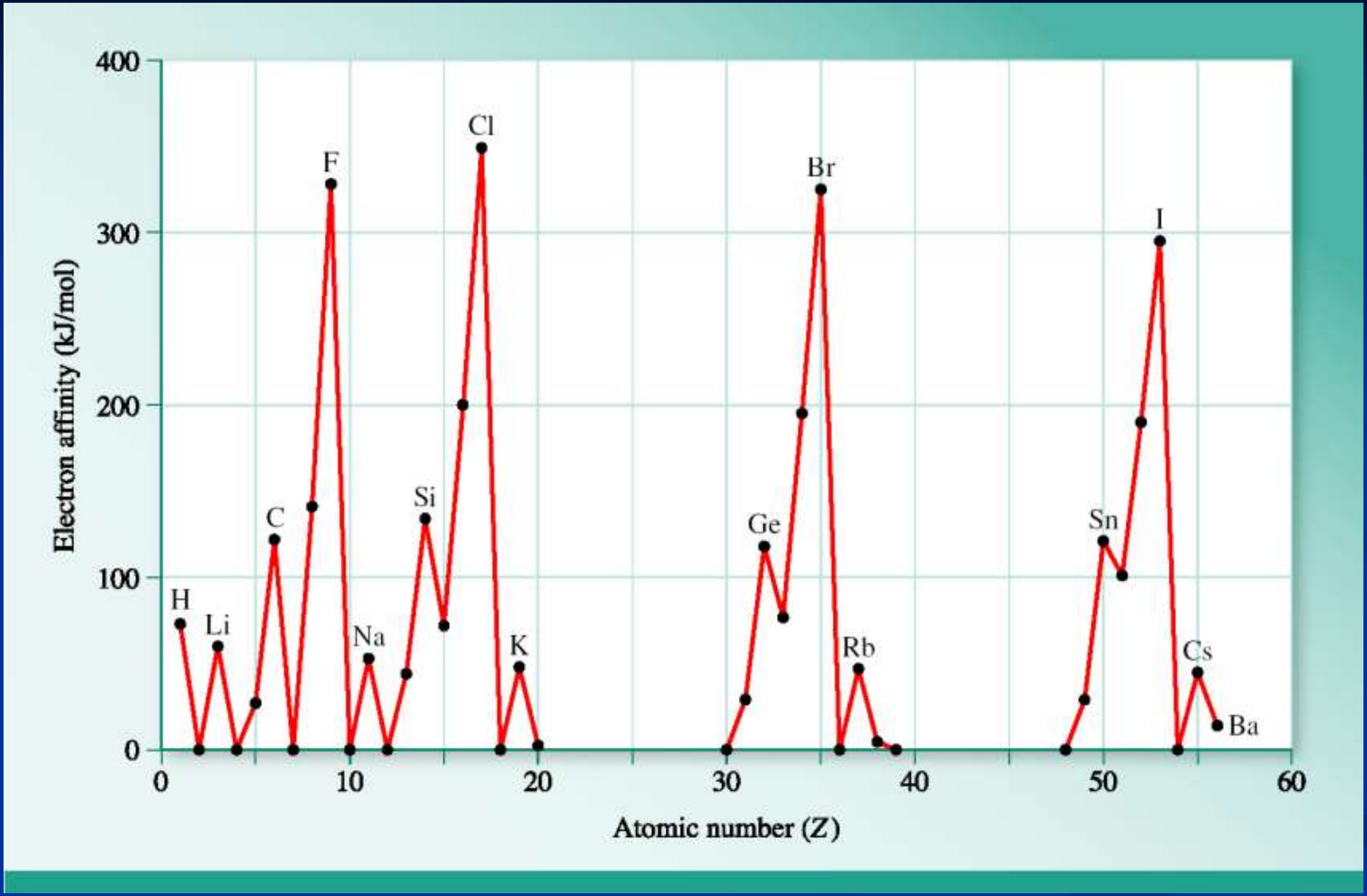
Electron Affinity

- **Across Period:**
  - releases more energy so number increases (gets more negative)
  - because electrons added in the same energy level do not shield electrons from nuclear charge
- **Down Group:**
  - releases less energy so number decreases (gets less negative)
  - because the electrons being added are farther away from the attracting protons

EA (kJ/mol) Values for Main Group Elements

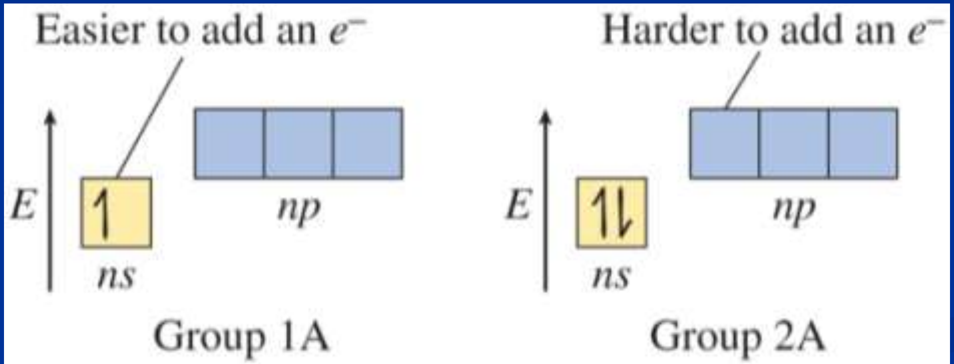
1A 1	2A 2	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
H +72.8							**He** (0.0)
Li +59.6	**Be** ≤0	**B** +26.7	**C** +122	**N** −7	**O** +141	**F** +328	**Ne** (−29)
Na +52.9	**Mg** ≤0	**Al** +42.5	**Si** +134	**P** +72.0	**S** +200	**Cl** +349	**Ar** (−35)
K +48.4	**Ca** +2.37	**Ga** +28.9	**Ge** +119	**As** +78.2	**Se** +195	**Br** +325	**Kr** (−39)
Rb +46.9	**Sr** +5.03	**In** +28.9	**Sn** +107	**Sb** +103	**Te** +190	**I** +295	**Xe** (−41)
Cs +45.5	**Ba** +13.95	**Tl** +19.3	**Pb** +35.1	**Bi** +91.3	**Po** +183	**At** +270	**Rn** (−41)

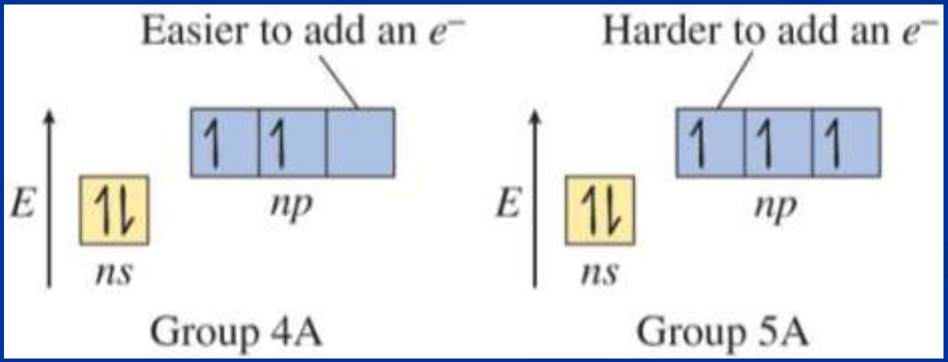
Periodic Trends in *EA*



Periodic Interruptions in *EA*

- Explained in much the same way as *IE* except not the same elements!





Metallic Character

- Metals
  - Shiny, lustrous, malleable
  - Good conductors
  - Low *IE* (form cations)
  - Form ionic compounds with chlorine
  - Form basic, ionic compounds with oxygen
  - Metallic character increases top to bottom in group and decreases left to right across a period

Nonmetals

- Vary in color, not shiny
- Brittle
- Poor conductors
- Form acidic, molecular compounds with oxygen
- High *EA* (form anions)

- Metalloids
  - Properties between the metals and nonmetals

7.5 Electron Configuration of Ions

- Follow Hund's rule and Pauli exclusion principle as for atoms
- Writing electron configurations helps explain charges memorized earlier

Ions of main group elements

- Noble gases (8A) almost completely unreactive due to electron configuration
  - ns^2np^6 (except He $1s^2$)
- Main group elements tend to gain or lose electrons to become isoelectronic (same valence electron configuration as nearest noble gas)

Na: $1s^2 2s^2 2p^6 3s^1 \rightarrow$ **Na^+** $1s^2 2s^2 2p^6$

Na: $[Ne]3s^1 \rightarrow$ **Na^+** $[Ne]$

(Na^+ 10 electrons - isoelectronic with Ne)

Cl: $1s^2 2s^2 2p^6 3s^2 3p^5 \rightarrow$ **Cl^-** $1s^2 2s^2 2p^6 3s^2 3p^6$

Cl: $[Ne]3s^2 3p^5 \rightarrow Cl^-$ $[Ar]$

(Cl^- 18 electrons - isoelectronic with Ar)

Ions of *d*-Block Elements

- Recall that the 4*s* orbital fills before the 3*d* orbital in the first row of transition metals
- Electrons are always lost from the highest "*n*" value (then from *d*)

Fe: $[Ar]4s^2 3d^6 \rightarrow Fe^{2+}: [Ar]3d^6$

Fe: $[Ar]4s^2 3d^6 \rightarrow Fe^{3+}: [Ar]3d^5$

7.6 Ionic Radius

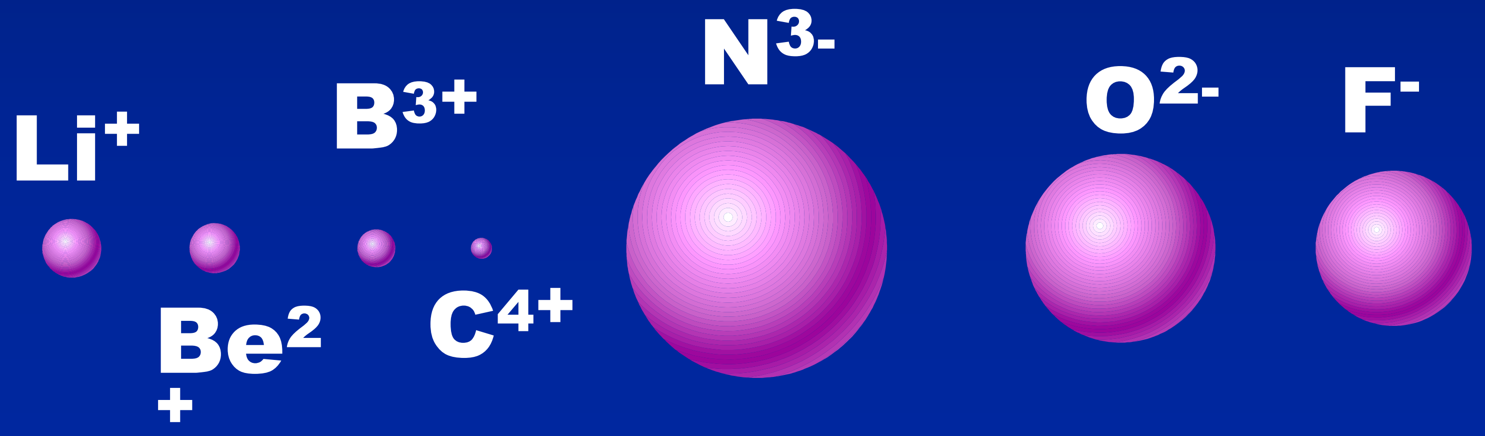
- When an atom gains or loses electrons, the radius changes
- Cations are always smaller than their parent atoms (often losing an energy level)
- Anions are always larger than their parent atoms (increased e^- repulsions)

Size of ions

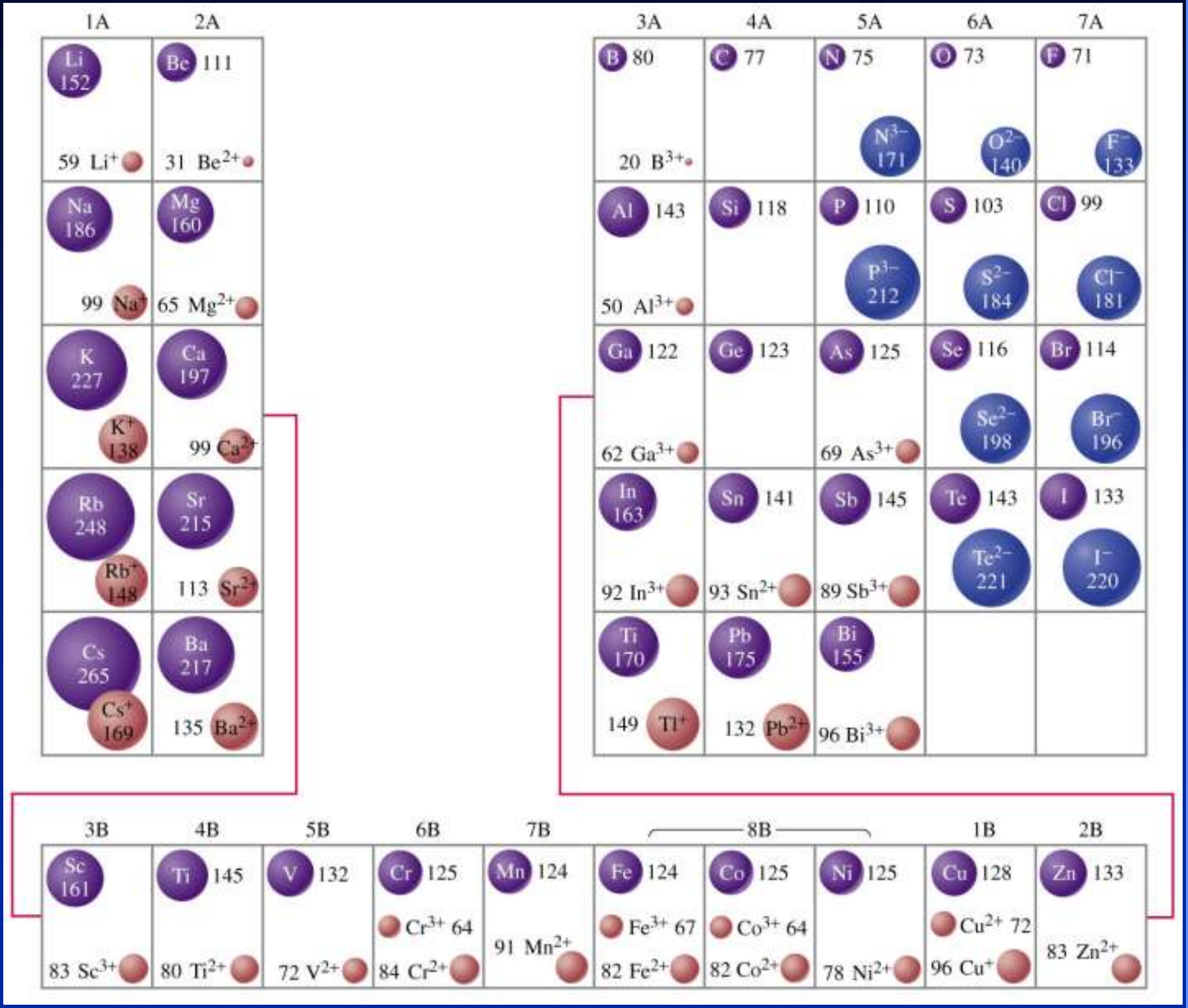
- **Ion size increases down a group.**
- **Cations are <u>smaller</u> than the atoms they came from.**
- **Anions are <u>larger</u> than atoms they came from.**
- **across a row ions get smaller, and then suddenly larger.**
- **First half are cations.**
- **Second half are anions.**

Periodic Trends

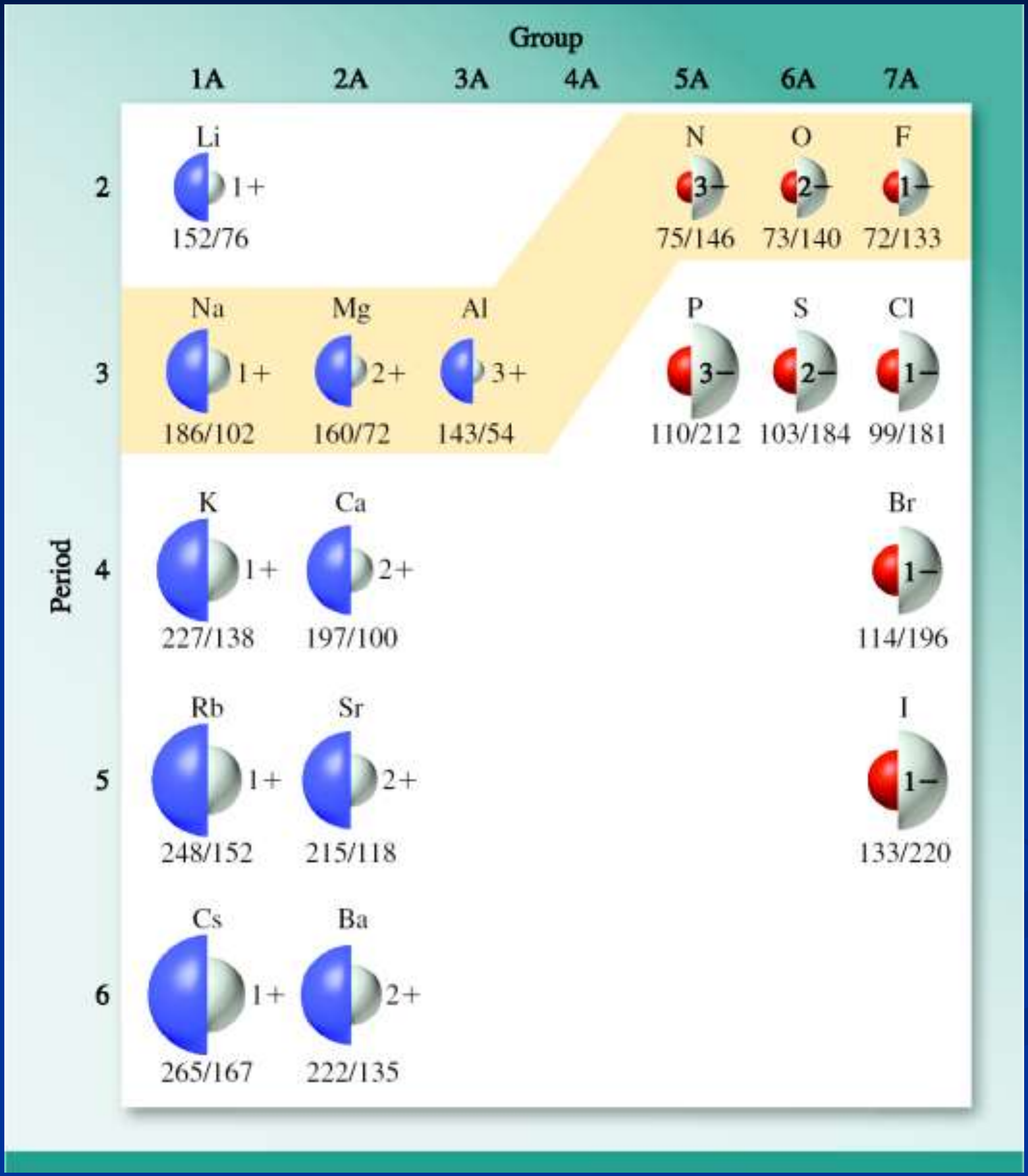
- Across the period, the change is complicated because of the change from predominantly **metals** on the left to **nonmetals** on the right.



Atomic and Ionic Radii

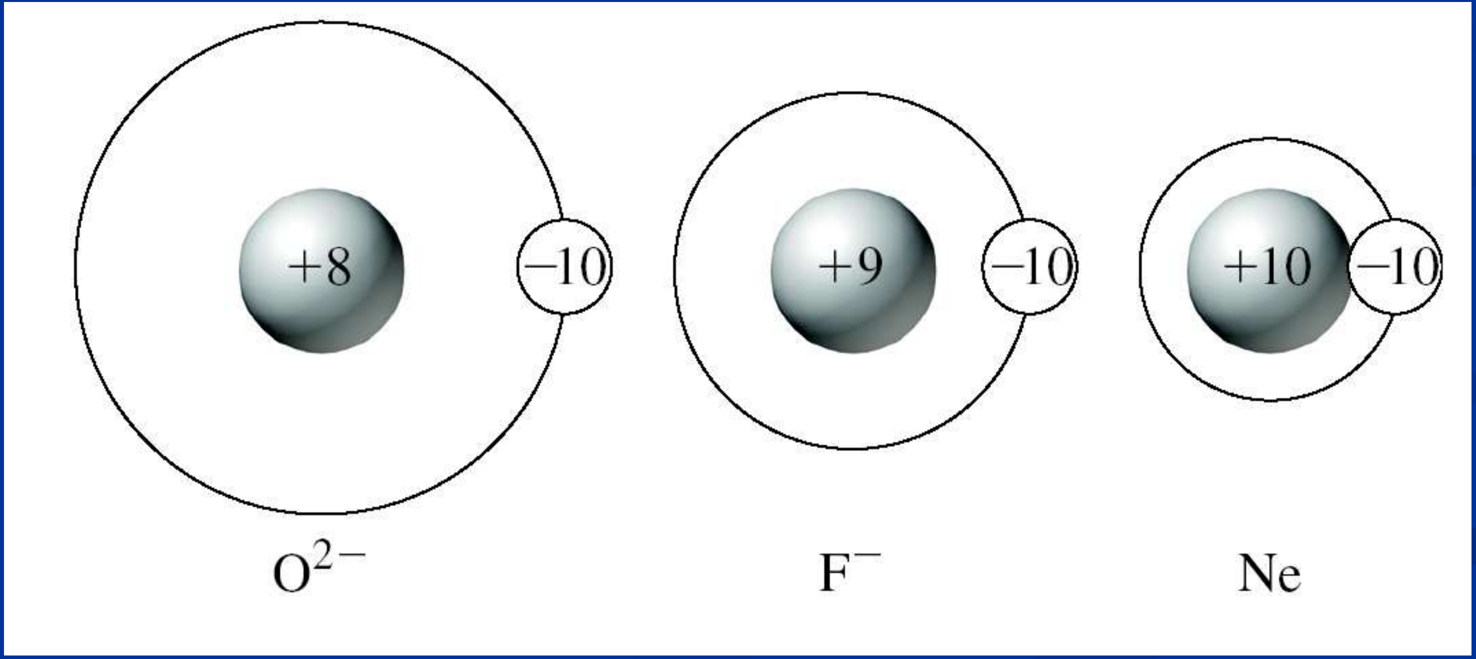


Comparison of Atomic and Ionic Radii



Isoelectronic Series

- Two or more species having the same electron configuration but different nuclear charges
- Size varies significantly

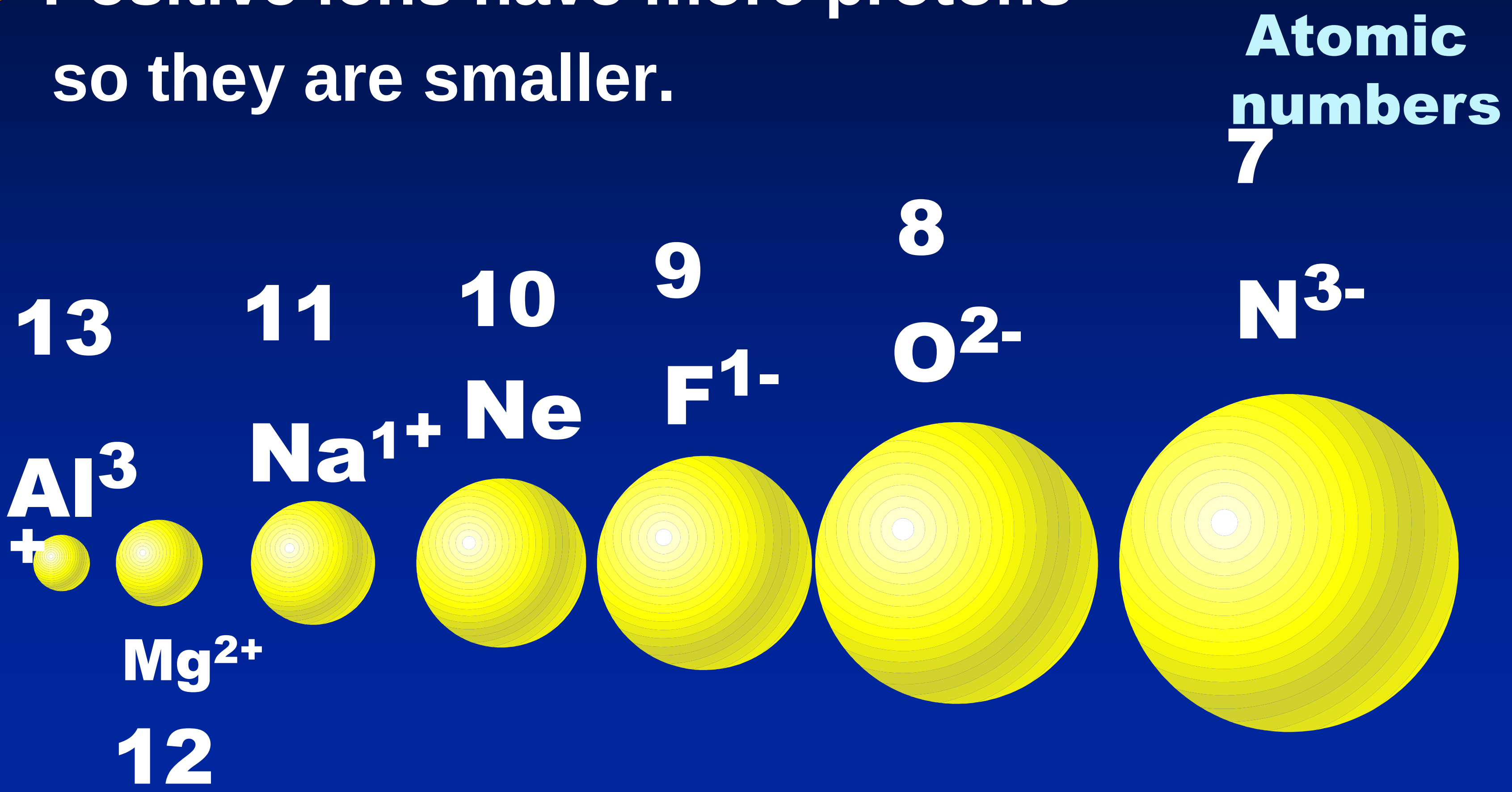


Isoelectronic ions

- Iso - same
- Iso electronic ions have the same # of electrons in the outermost level
- Al^{3+} , Mg^{2+}, Na^{+}, Ne, F^{-}, O^{2-} and N^{3-}
- All have ***10*** electrons.
- All have the same configuration $1s^2 2s^2 2p^6$

Size of Isoelectronic ions

- Positive ions have more protons so they are smaller.



Key Points

- Development of the periodic table
- Modern table and its arrangement
- Main group elements
- Valence electrons
- Effective nuclear charge and relationship to periodic trends
- Atomic radius (ionic radii, covalent radii, metallic radii)

Key Points

- Ionization energy (*IE*) - trends of 1st and multiple *IE*'s
- Electron affinity (*EA*) - trends
- Properties of metals, metalloids and nonmetals
- Isoelectronic - predict charges of ions and electron configurations of ions
- Write and/or recognize an isoelectronic series