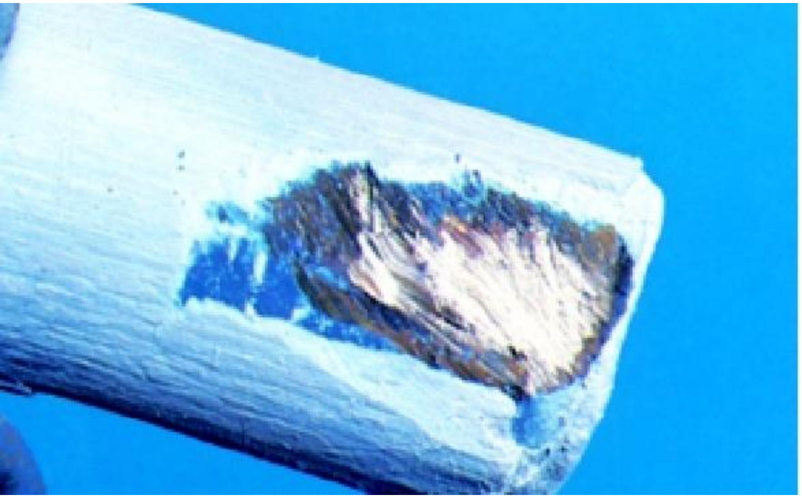
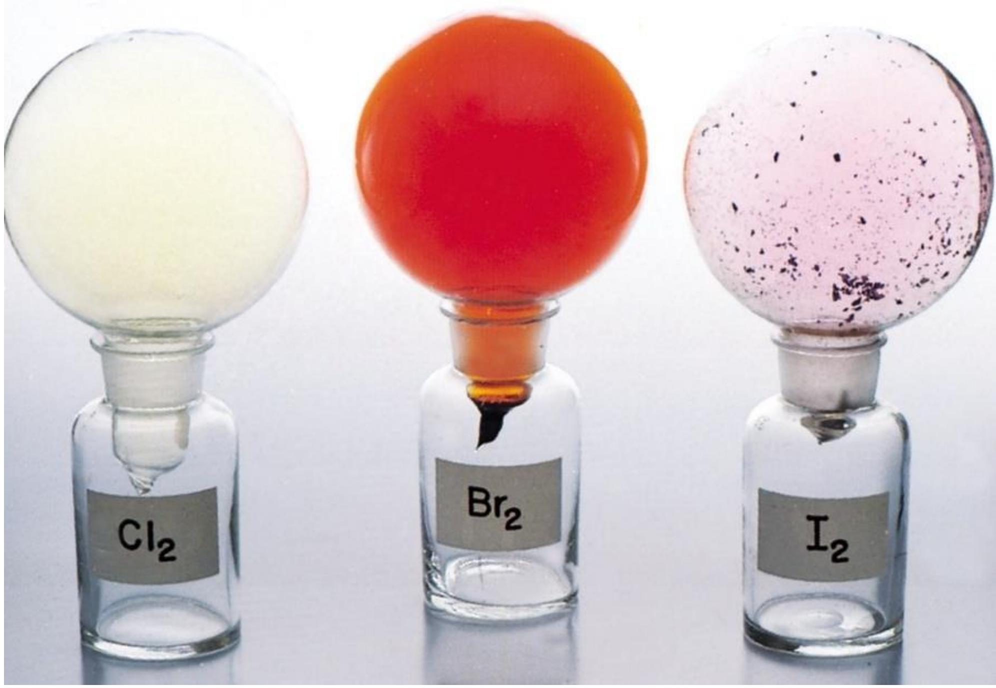




Periodic Relationships Among the Elements

Chapter 8

Development of the Periodic Table

- The Periodic Table
  - During the time of Robert Boyle (1661) only 13 elements were known
  - (Sb, As, Bi, C, Cu, Au, Fe, Pb, Hg, Ag, S, Sn and Zn)

- By the end of the eighteenth century Lavoisier added another 11 elements
  - (Cl, Co, H, Mn, Mo, Ni, N, O, P, Pt and W)

- Since that time, a new element has been discovered on the average of every two and one-half years.

Development of the Periodic Table

- With the increase of element number, Scientists began to find patterns in their properties. In 1829, J.W. Dobereiner discovered the existence of families of three elements with similar chemical properties and called them **triads**.
  - Li, Na, K   Ca, Sr, Ba   S, Se, Te   Cl. Br, I   and Mn, Cr, Fe

- In 1865, Newlands arranged the elements in order of increasing atomic weights. He placed seven elements in a group and every eighth element had similar properties. Known as **law of octaves**.

  But he left no room for any newer elements.

Development of the Periodic Table

- In 1871 Mendeleeff proposed his version of the periodic table in which elements were arranged in order of increasing atomic weights.

  He grouped the elements on the basis of similarities in their chemical properties.

  His table had **blank spaces** where he predicted the discovery of 10 other elements.

- More than 700 versions of the periodic table were proposed since then.

Mendeleev's Periodic Table

1871

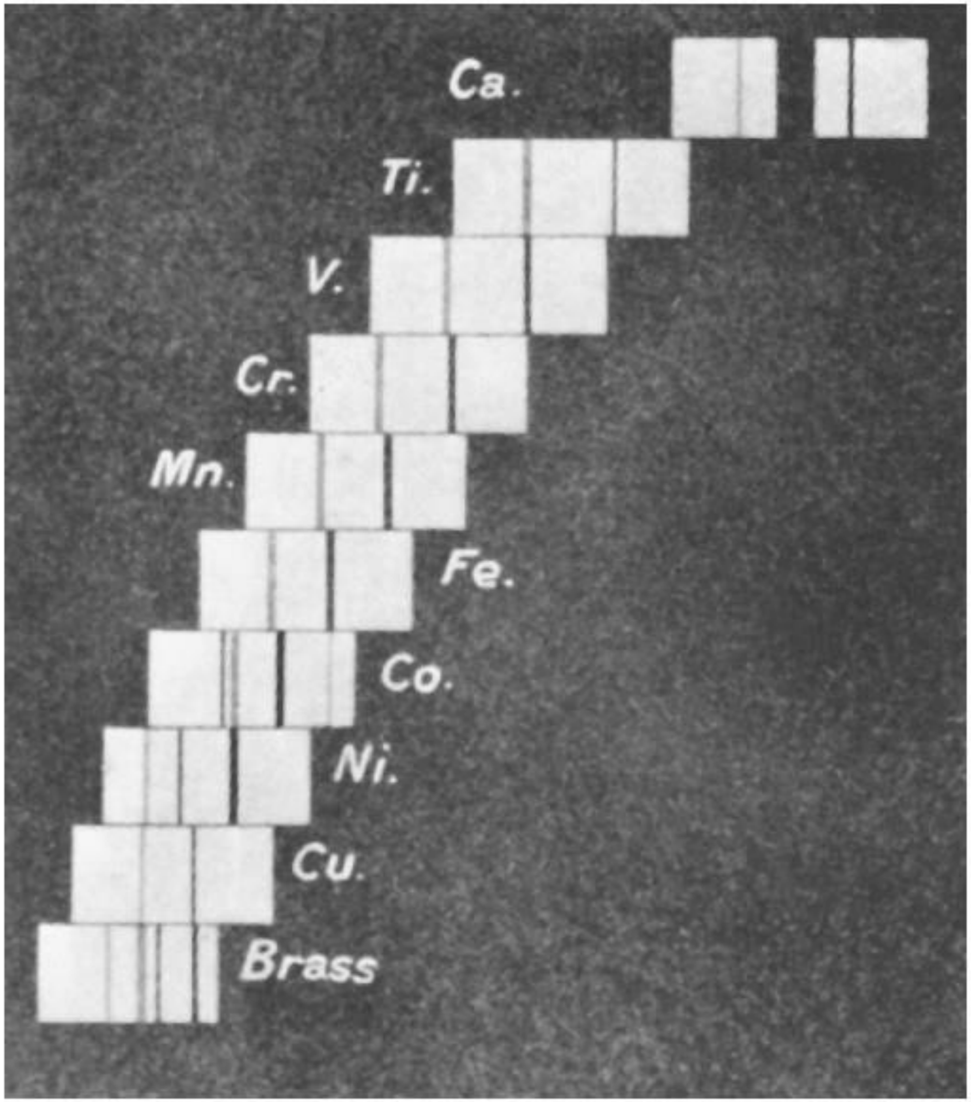
Reihen	Gruppe I. — R^2O	Gruppe II. — RO	Gruppe III. — R^2O^3	Gruppe IV. RH^4 RO^2	Gruppe V. RH^3 R^2O^5	Gruppe VI. RH^2 RO^3	Gruppe VII. RH R^2O^7	Gruppe VIII. — RO^4
1	H = 1							
2	Li = 7	Be = 9,4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27,3	Si = 28	P = 31	S = 32	Cl = 35,5	
4	K = 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		

Predicted Elements were Found

TABLE 10.1 Properties of Germanium: Predicted and Observed

Property	Predicted Eka-silicon (1871)	Observed Germanium (1886)
Atomic mass	72	72.6
Density, g/cm^3	5.5	5.47
Color	dirty gray	grayish white
Density of oxide, g/cm^3	EsO_2: 4.7	GeO_2: 4.703
Boiling point of chloride	$EsCl_4$: below 100 °C	$GeCl_4$: 86 °C
Density of chloride, g/cm^3	$EsCl_4$: 1.9	$GeCl_4$: 1.887

X-Ray Spectra



- Moseley 1913
  - X-ray emission is explained in terms of transitions in which e^- drop into orbits close to the atomic nucleus.
  - Correlated frequencies to nuclear charges.
- $\nu = A\,(Z - b)^2$
  - Used to predict new elements (43, 61, 75) later discovered.

Development of the Periodic Table

- The most popular version of the periodic table shows the elements arranged in order of **increasing atomic number** and grouped into **periods** and **groups**.

- Each group contains elements with **similar chemical properties**.

- In most cases, the elements in a column also have **similar electronic configuration**.

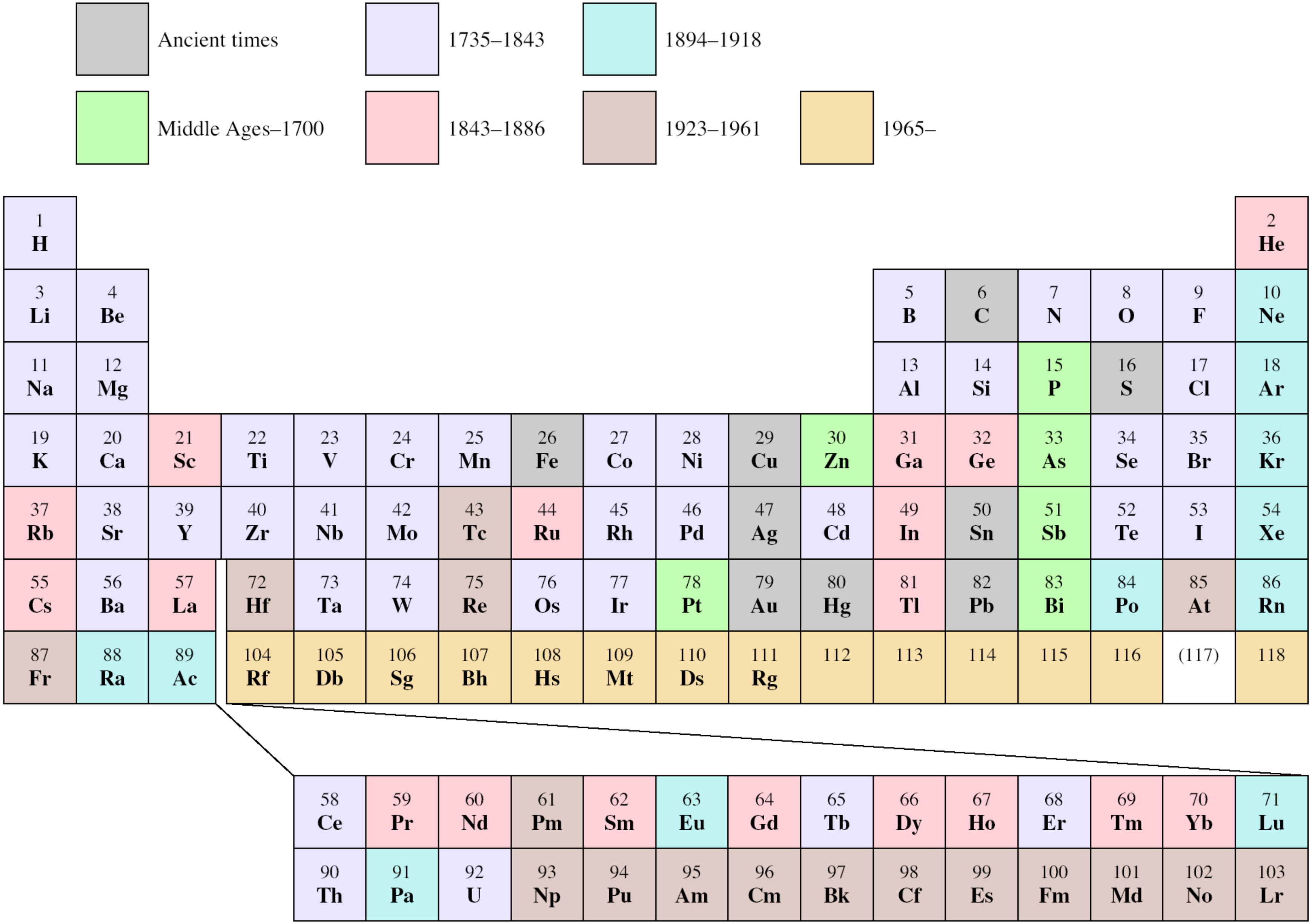
- The two most important exceptions are **hydrogen** and **helium**.

Development of the Periodic Table

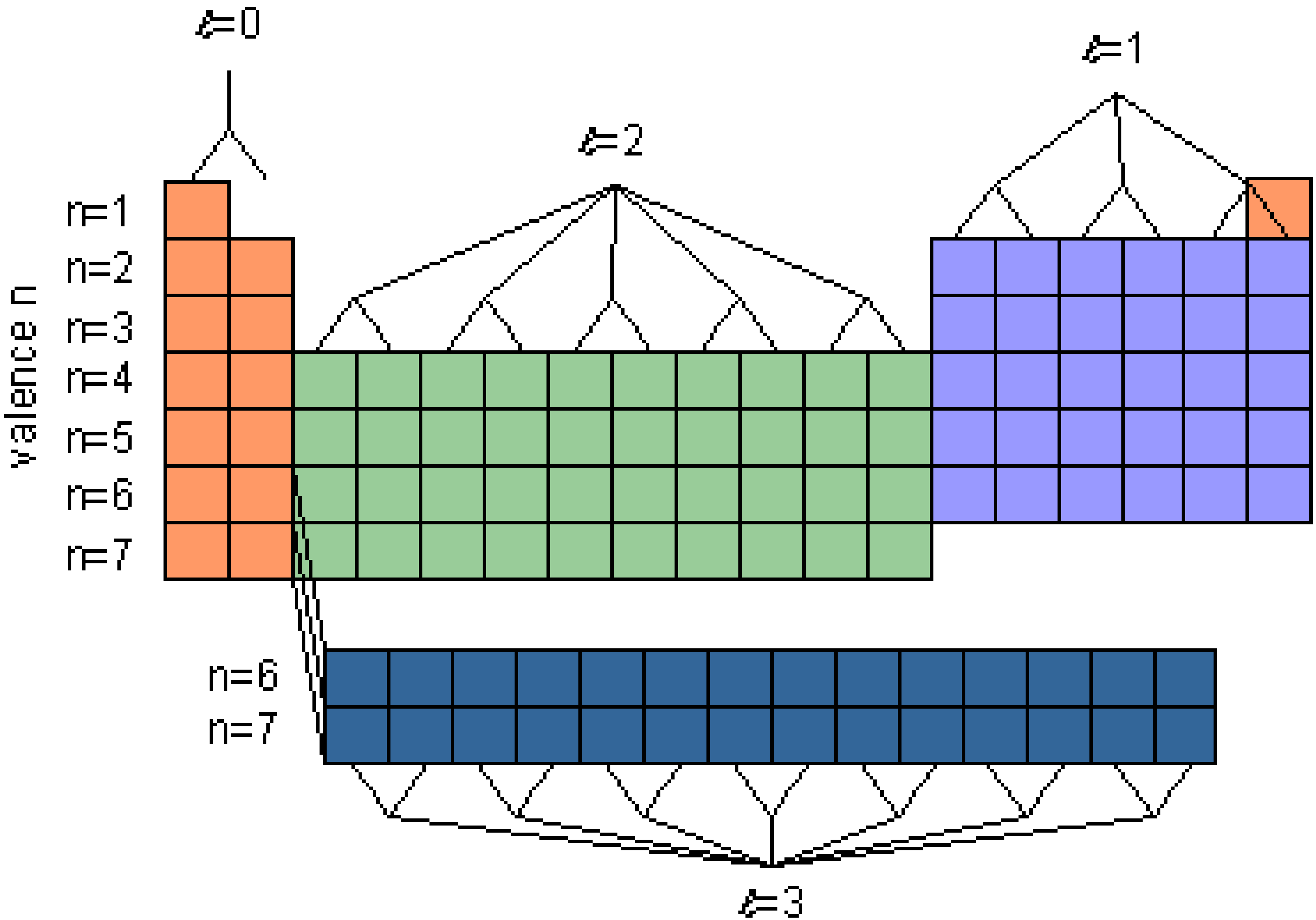
- The periodic table provides a basis for predicting general trends in a number of physical properties, such as:
  - the size of the atoms or the ions,
  - ionization energies (**IE**),
  - electron affinity (**EA**),
  - electronegativities etc.

profoundly influence the **chemical behavior** of elements.

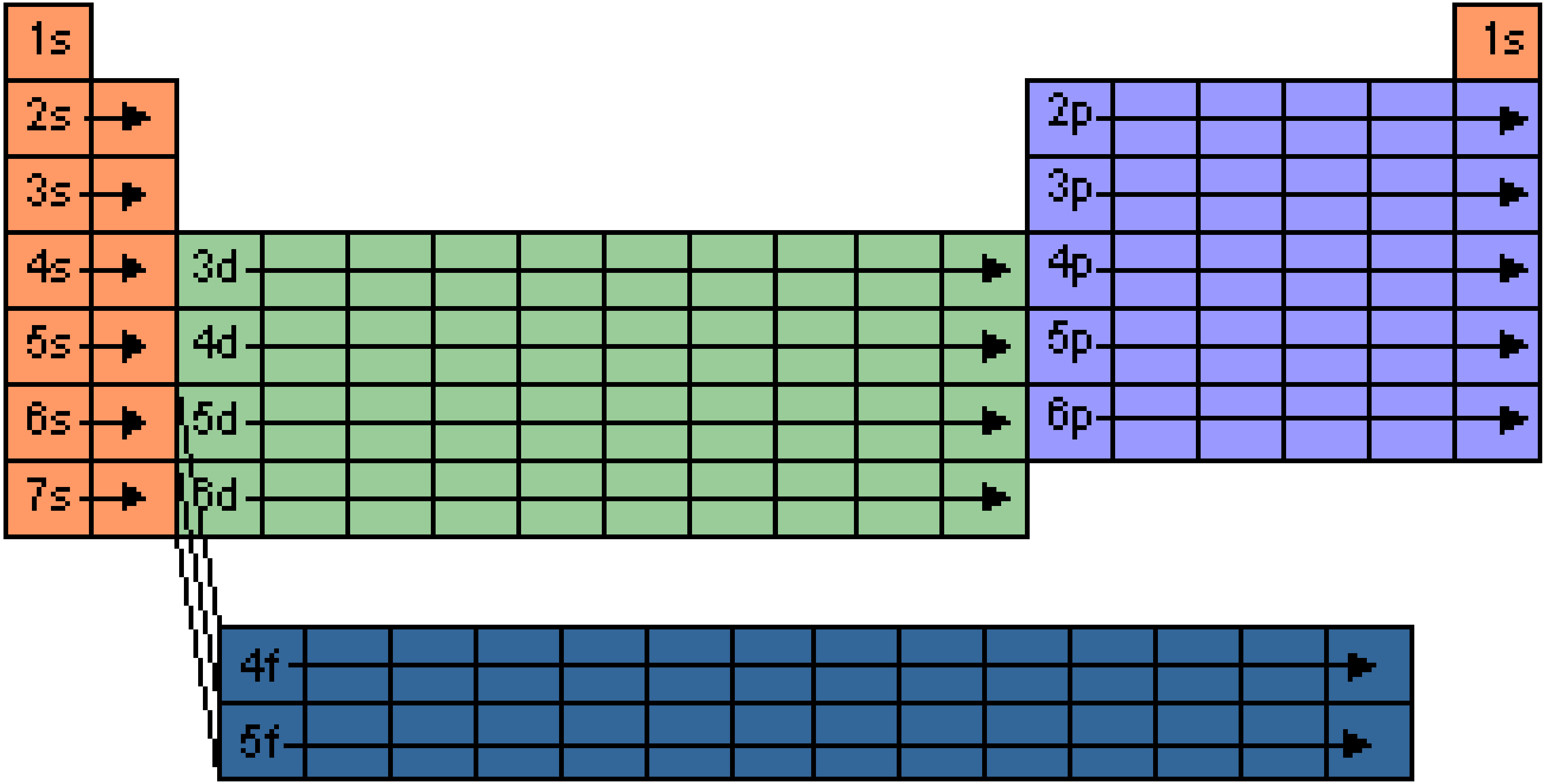
When the Elements Were Discovered



Outermost subshell electrons quantum numbers



Outermost subshell being filled with electrons



Ground State Electron Configurations of the Elements

	1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
	ns^1	ns^2	d^1				d^5					d^{10}	ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6
1	1 **H** $1s^1$																	2 **He** $1s^2$
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** $2s^22p^6$
3	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** $3s^23p^6$
4	19 **K** $4s^1$	20 **Ca** $4s^2$	21 **Sc** $4s^23d^1$	22 **Ti** $4s^23d^2$	23 **V** $4s^23d^3$	24 **Cr** $4s^13d^5$	25 **Mn** $4s^23d^5$	26 **Fe** $4s^23d^6$	27 **Co** $4s^23d^7$	28 **Ni** $4s^23d^8$	29 **Cu** $4s^13d^{10}$	30 **Zn** $4s^23d^{10}$	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** $4s^24p^6$
5	37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $5s^24d^1$	40 **Zr** $5s^24d^2$	41 **Nb** $5s^14d^4$	42 **Mo** $5s^14d^5$	43 **Tc** $5s^24d^5$	44 **Ru** $5s^14d^7$	45 **Rh** $5s^14d^8$	46 **Pd** $4d^{10}$	47 **Ag** $5s^14d^{10}$	48 **Cd** $5s^24d^{10}$	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** $5s^25p^6$
6	55 **Cs** $6s^1$	56 **Ba** $6s^2$	57 **La** $6s^25d^1$	72 **Hf** $6s^25d^2$	73 **Ta** $6s^25d^3$	74 **W** $6s^25d^4$	75 **Re** $6s^25d^5$	76 **Os** $6s^25d^6$	77 **Ir** $6s^25d^7$	78 **Pt** $6s^15d^9$	79 **Au** $6s^15d^{10}$	80 **Hg** $6s^25d^{10}$	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** $6s^26p^6$
7	87 **Fr** $7s^1$	88 **Ra** $7s^2$	89 **Ac** $7s^26d^1$	104 **Rf** $7s^26d^2$	105 **Db** $7s^26d^3$	106 **Sg** $7s^26d^4$	107 **Bh** $7s^26d^5$	108 **Hs** $7s^26d^6$	109 **Mt** $7s^26d^7$	110 **Ds** $7s^26d^8$	111 **Rg** $7s^26d^9$	112 $7s^26d^{10}$	113 $7s^27p^1$	114 $7s^27p^2$	115 $7s^27p^3$	116 $7s^27p^4$	(117)	118 $7s^27p^6$

4f	58 **Ce** $6s^24f^15d^1$	59 **Pr** $6s^24f^3$	60 **Nd** $6s^24f^4$	61 **Pm** $6s^24f^5$	62 **Sm** $6s^24f^6$	63 **Eu** $6s^24f^7$	64 **Gd** $6s^24f^75d^1$	65 **Tb** $6s^24f^9$	66 **Dy** $6s^24f^{10}$	67 **Ho** $6s^24f^{11}$	68 **Er** $6s^24f^{12}$	69 **Tm** $6s^24f^{13}$	70 **Yb** $6s^24f^{14}$	71 **Lu** $6s^24f^{14}5d^1$
5f	90 **Th** $7s^26d^2$	91 **Pa** $7s^25f^26d^1$	92 **U** $7s^25f^36d^1$	93 **Np** $7s^25f^46d^1$	94 **Pu** $7s^25f^6$	95 **Am** $7s^25f^7$	96 **Cm** $7s^25f^76d^1$	97 **Bk** $7s^25f^9$	98 **Cf** $7s^25f^{10}$	99 **Es** $7s^25f^{11}$	100 **Fm** $7s^25f^{12}$	101 **Md** $7s^25f^{13}$	102 **No** $7s^25f^{14}$	103 **Lr** $7s^25f^{14}6d^1$

Classification of the Elements

- Representative elements
- Noble gases
- Transition metals
- Zinc Cadmium Mercury
- Lanthanides
- Actinides

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112	113	114	115	116	(117)	118

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Representing Free Elements in Chemical Equations

- How metals, metalloids, and nonmetals are represented as free elements in chemical equations?
- Metals do not exist in discrete molecular units, empirical formulas are used.
  - Empirical formula = Symbols (Mg, Ca, Fe, Cu, Zn, etc.)
- For nonmetals no single rule.
  - Carbon exists as an extensive three-dimensional network of atoms, empirical formula (C) is used.
  - Hydrogen, nitrogen, oxygen, & halogens exists as diatomic molecule, molecular formula, H_2, N_2, O_2, & F_2, Cl_2, Br_2, I_2 are used.
  - Stable form of molecular phosphorus P_4 is used.

Representing Free Elements in Chemical Equations

- For nonmetals no single rule. (contd.)
  - Stable form of molecular sulfur, S_8 but (S) is used. Combustion of sulfur:

$$\cancel{S_8(s) + 8O_2(g) \rightarrow 8SO_2(g)}$$

$$S(s) + O_2(g) \rightarrow SO_2(g)$$

Identical stoichiometry

- All noble gases are monatomic, symbols are used.
  - He, Ne, Ar, Kr, Xe, & Rn
- Metalloids have complex three-dimensional networks, empirical formulas are used,
  - B, Si, Ge, As, Sb, Po, & At

Valence & Core Electrons

- the electrons in all the subshells with the highest principal energy shell are called the **valence electrons**.
- electrons in lower energy shells are called **core electrons**.
- chemists have observed that one of the most important factors in the way an atom behaves, both chemically and physically, is the **number of valence electrons**.

Valence & Core Electrons

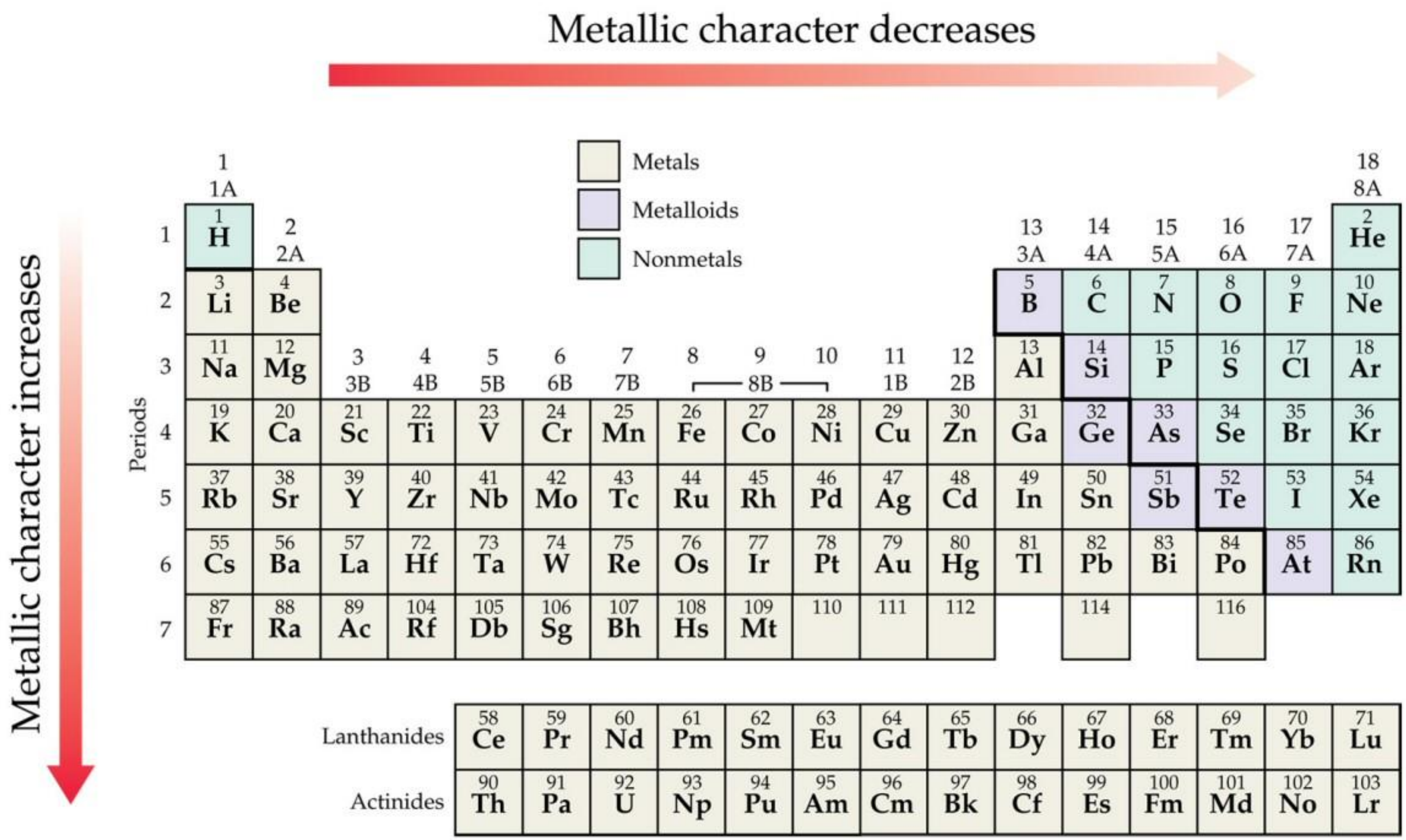
Rb = 37 electrons = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$

- the highest principal energy shell of Rb that contains electrons is the 5th, therefore Rb has 1 valence electron and 36 core electrons

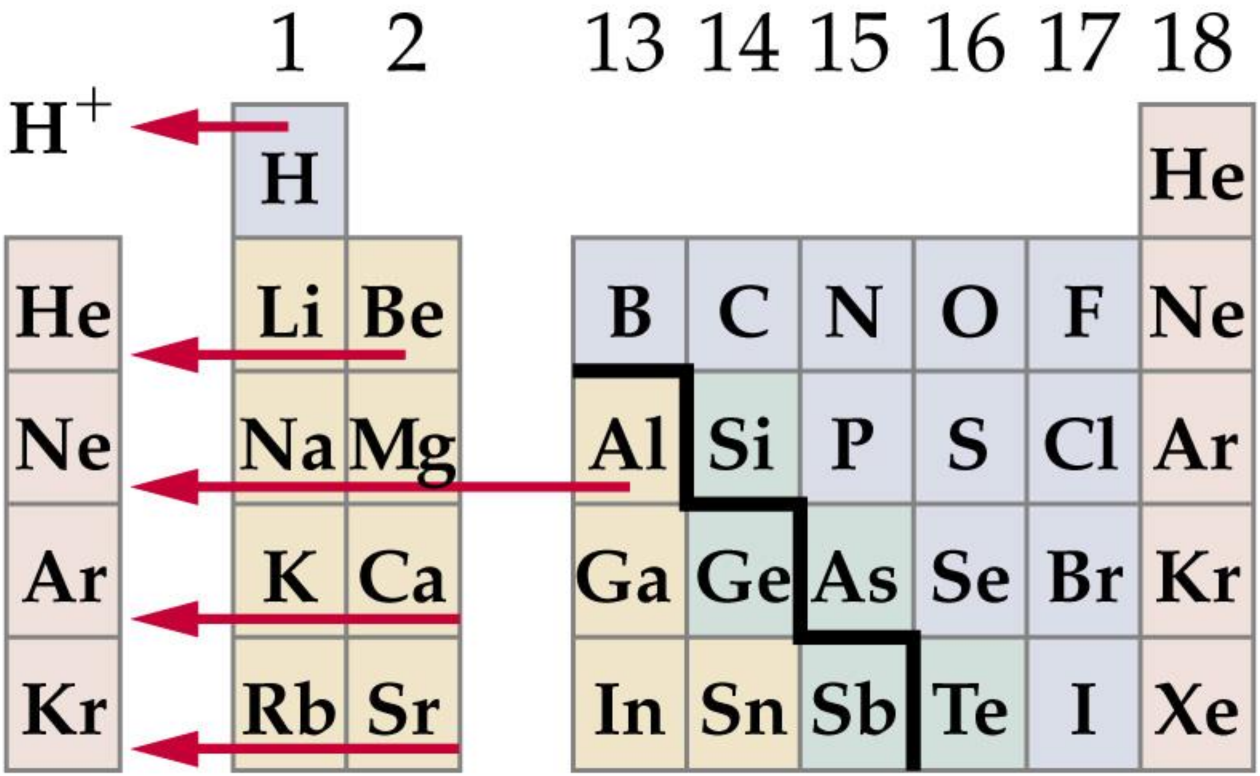
Kr = 36 electrons = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

- the highest principal energy shell of Kr that contains electrons is the 4th, therefore Kr has 8 valence electrons and 28 core electrons

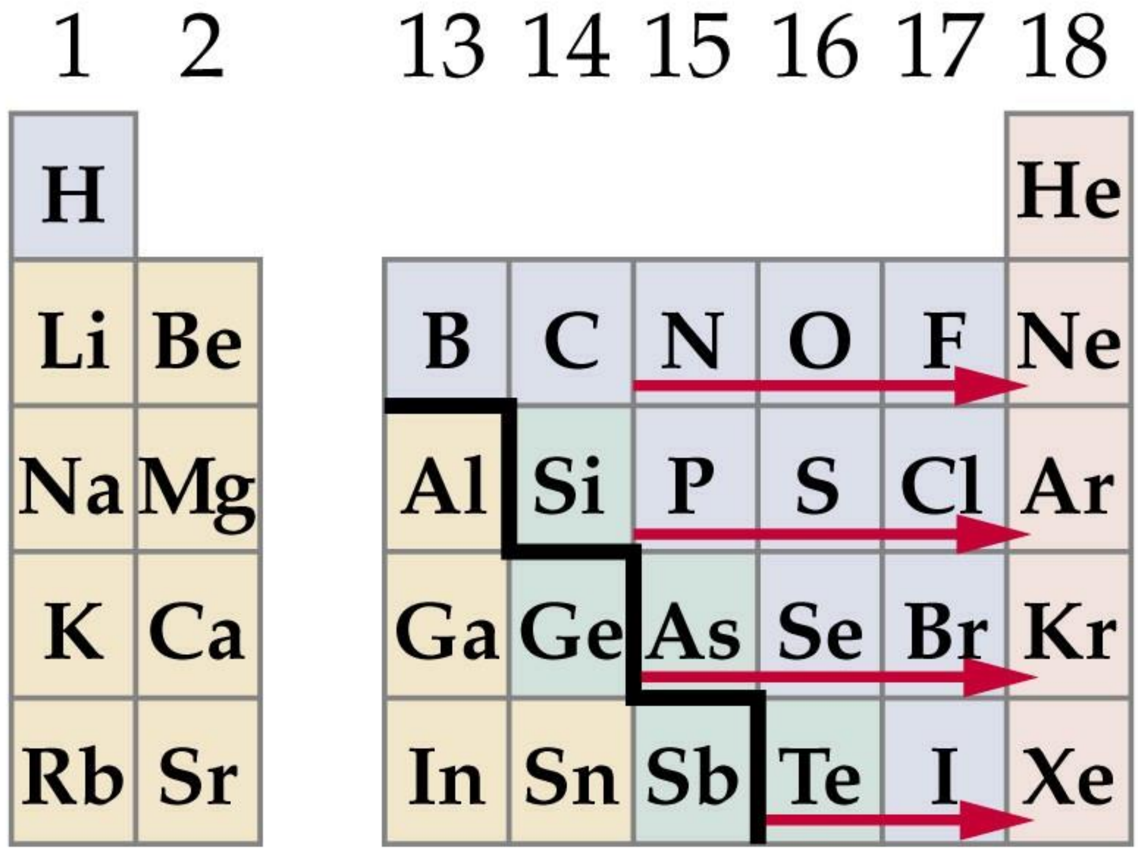
Trends in Metallic Character



Metals Tend to Lose Electrons



Nonmetals Tend to Gain Electrons



Electron Configurations of Cations and Anions of Representative Elements

Na $[Ne]3s^1$ — Na^+ $[Ne]$

Ca $[Ar]4s^2$ — Ca^{2+} $[Ar]$

Al $[Ne]3s^23p^1$ — Al^{3+} $[Ne]$

Atoms lose electrons so that cation has a noble-gas outer electron configuration.

H $1s^1$ — H^- $1s^2$ or $[He]$

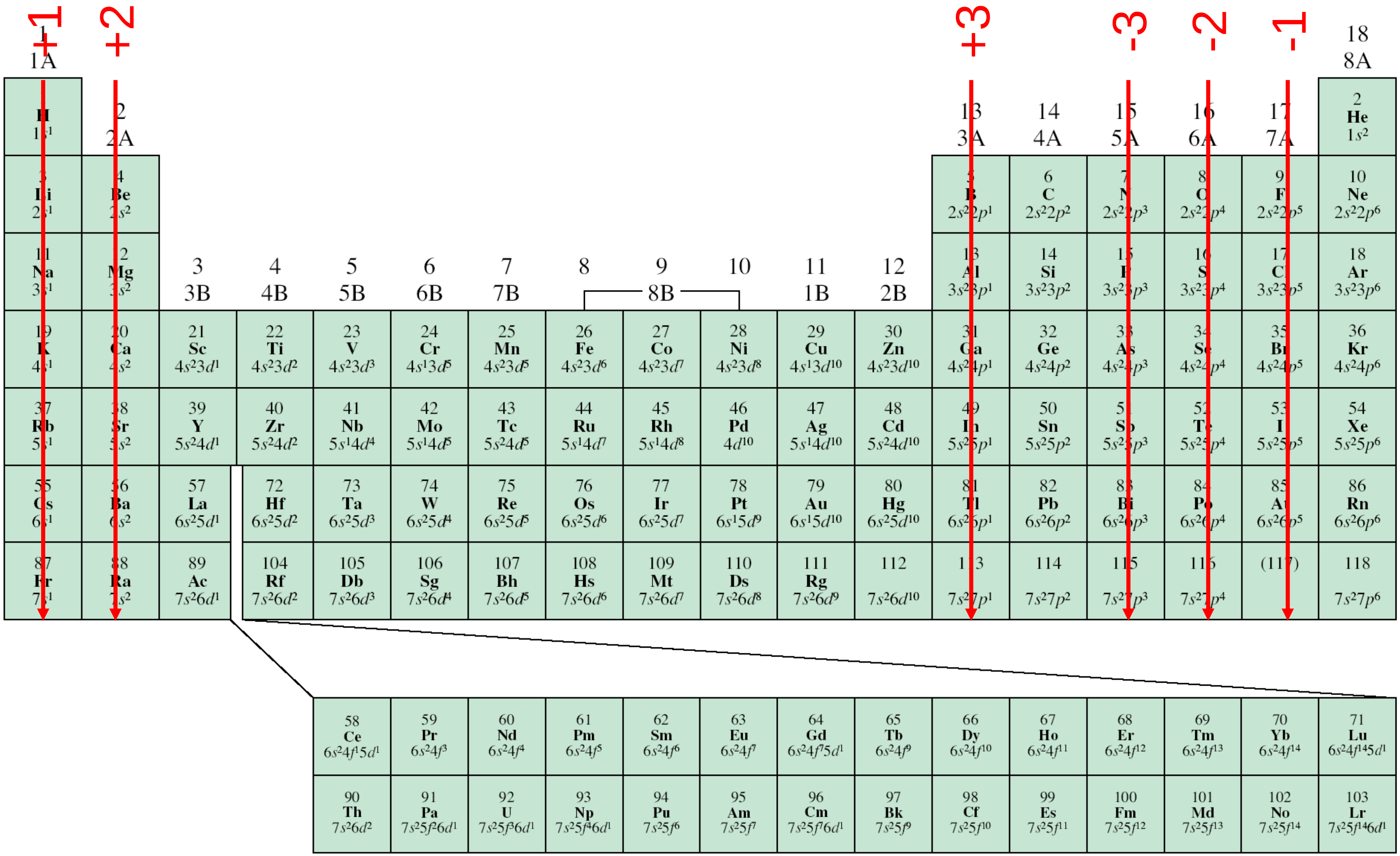
F $1s^22s^22p^5$ — F^- $1s^22s^22p^6$ or $[Ne]$

O $1s^22s^22p^4$ — O^{2-} $1s^22s^22p^6$ or $[Ne]$

N $1s^22s^22p^3$ — N^{3-} $1s^22s^22p^6$ or $[Ne]$

Atoms gain electrons so that anion has a noble-gas outer electron configuration.

Cations and Anions of Representative Elements



Electron Configurations of Cations of Transition Metals

When a cation is formed from an atom of a transition metal, electrons are always removed first from the *ns* orbital and then from the $(n - 1)d$ orbitals.

Fe: $[Ar]4s^2 3d^6$

Fe^{2+}: $[Ar]4s^0 3d^6$ or $[Ar]3d^6$

Fe^{3+}: $[Ar]4s^0 3d^5$ or $[Ar]3d^5$

Mn: $[Ar]4s^2 3d^5$

Mn^{2+}: $[Ar]4s^0 3d^5$ or $[Ar]3d^5$

Effective Nuclear Charge (Z_{eff})

Effective nuclear charge (Z_{eff}) is the "positive charge" felt by an electron.

$Z_{eff} = Z - \sigma \qquad 0 < \sigma < Z$ (σ = shielding constant)

$Z_{eff} \approx Z$ – number of inner or core electrons

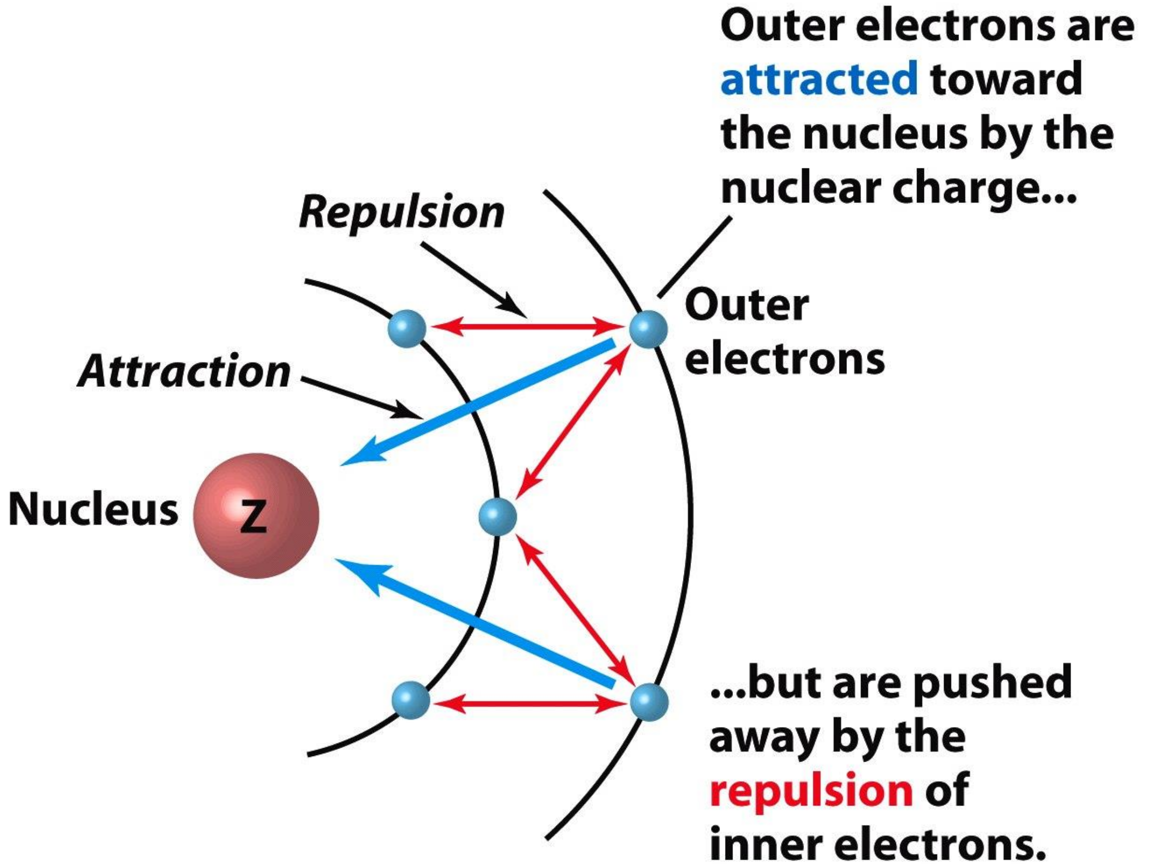


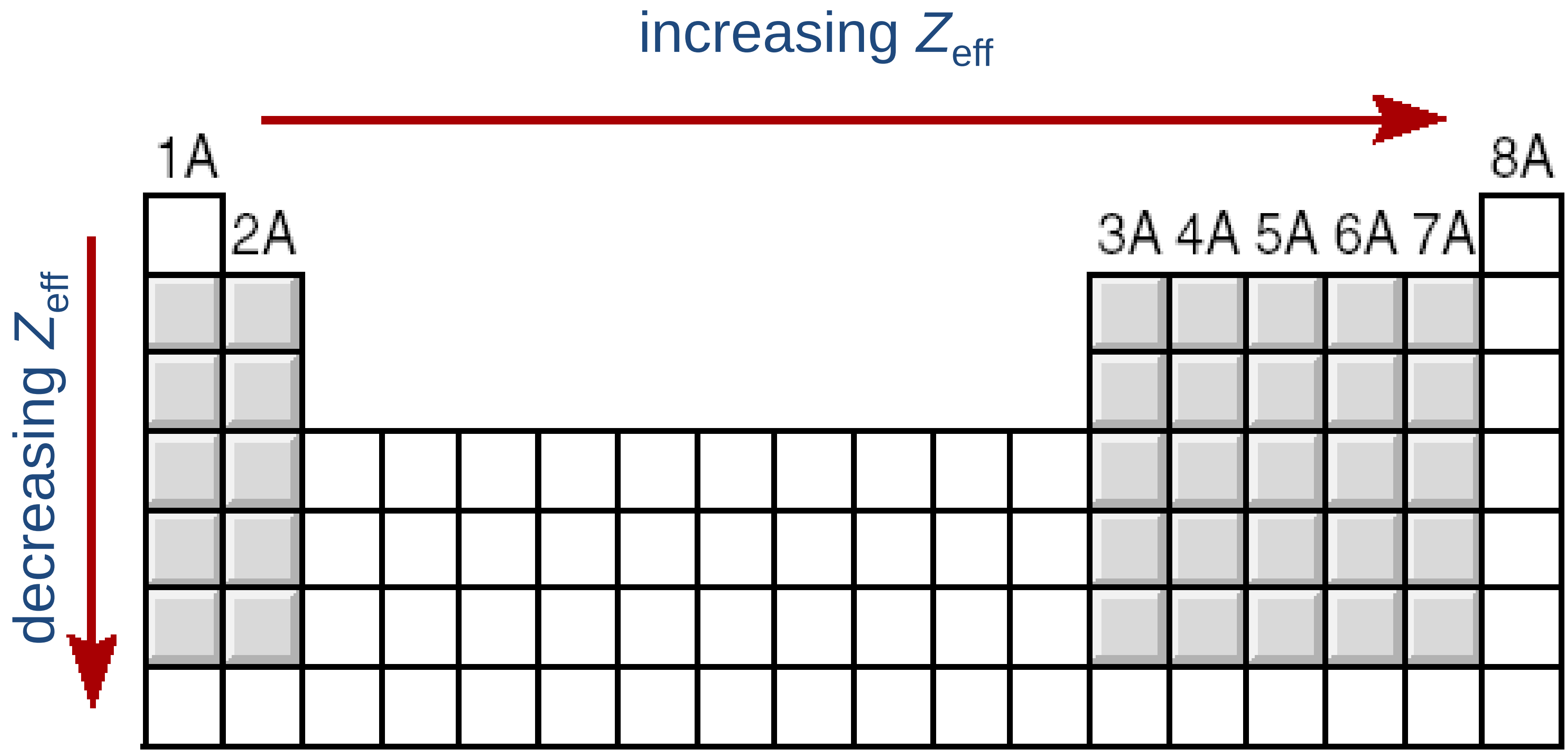
Figure 5-16 Chemistry, 5/e

Effective Nuclear Charge (Z_{eff})

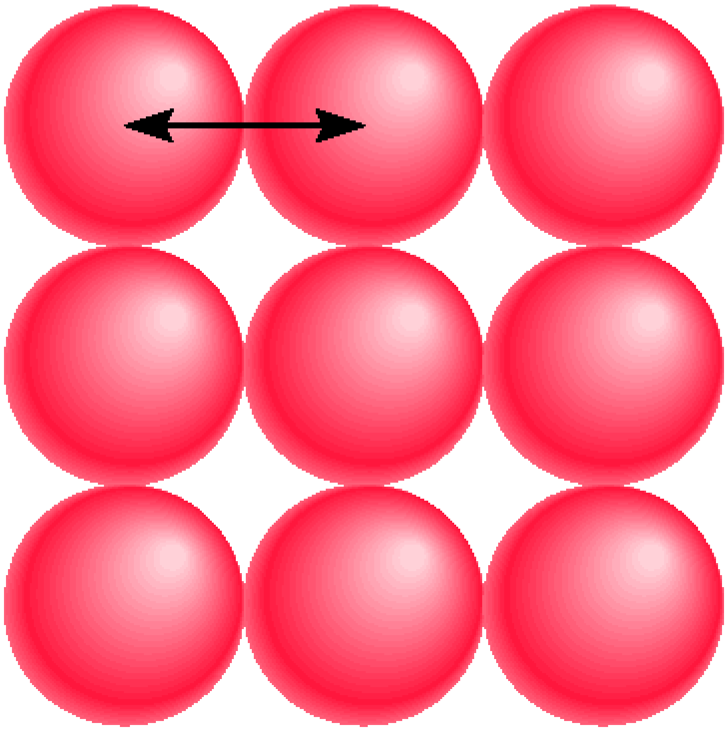
$Z_{eff} = Z_{actual}$ – No. of Core Electrons **or** Electron shielding

	Z	Core	Z_{eff}	Radius (pm)
Na	11	10	1	186
Mg	12	10	2	160
Al	13	10	3	143
Si	14	10	4	132

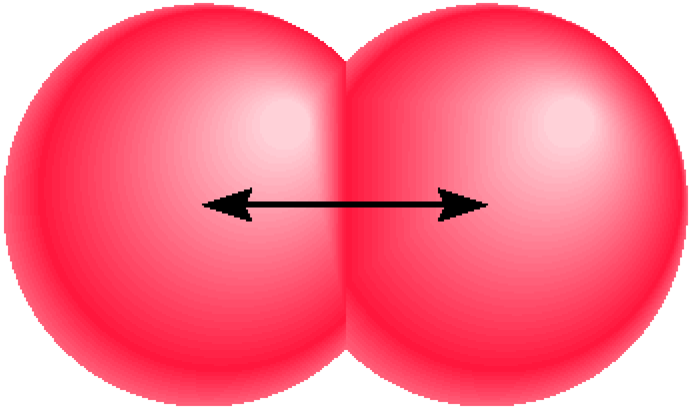
Effective Nuclear Charge (Z_{eff})



Atomic Radii

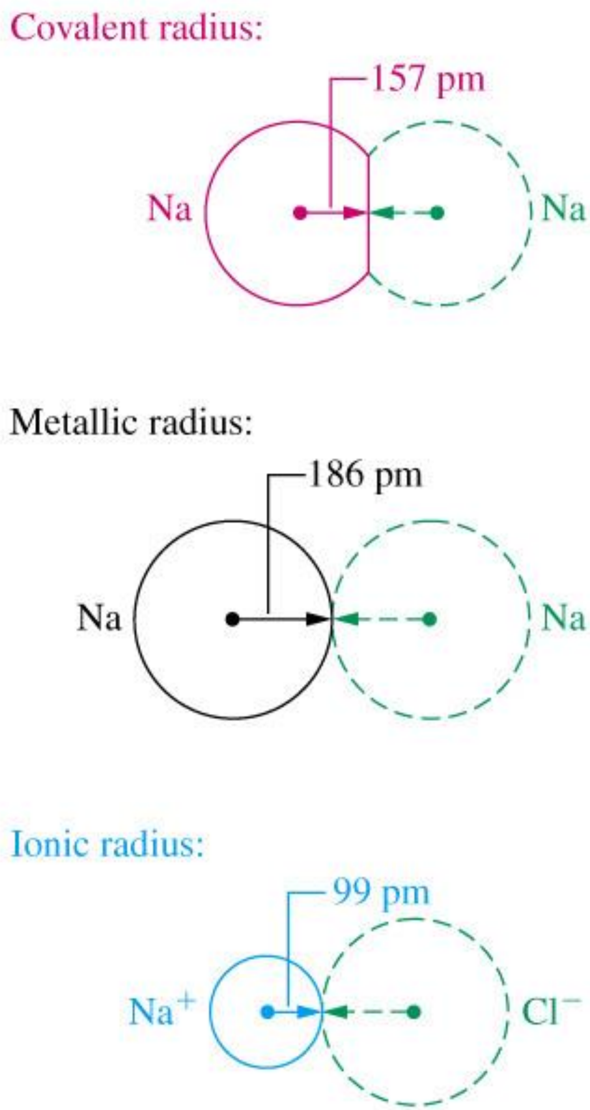


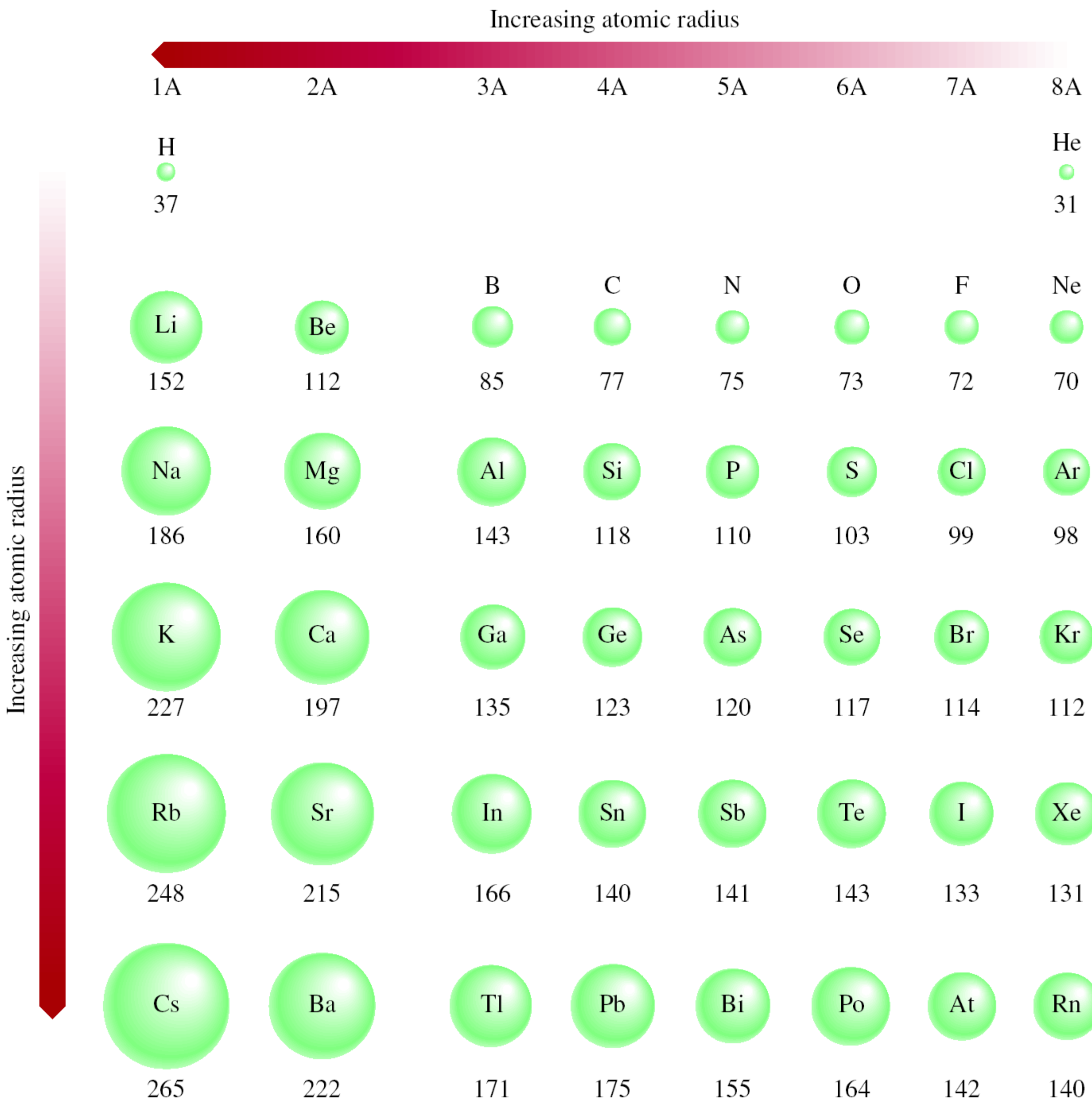
metallic radius



covalent radius

The Sizes of Atoms and Ions



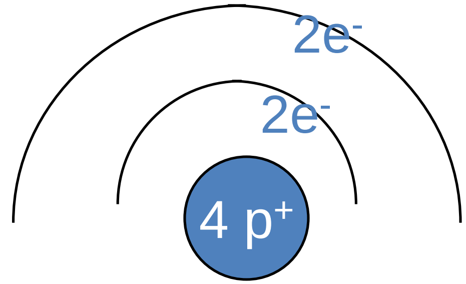

Increasing atomic radius
1A 2A 3A 4A 5A 6A 7A 8A
H 37
He 31
Li 152
Be 112
B 85
C 77
N 75
O 73
F 72
Ne 70
Na 186
Mg 160
Al 143
Si 118
P 110
S 103
Cl 99
Ar 98
K 227
Ca 197
Ga 135
Ge 123
As 120
Se 117
Br 114
Kr 112
Rb 248
Sr 215
In 166
Sn 140
Sb 141
Te 143
I 133
Xe 131
Cs 265
Ba 222
Tl 171
Pb 175
Bi 155
Po 164
At 142
Rn 140
Increasing atomic radius

Atomic Radii Increases Down the Group

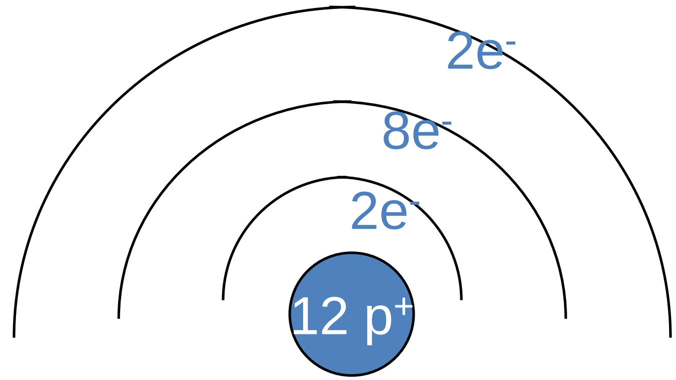
Group IIA

Be
Mg
Ca
Sr
Ba
Ra

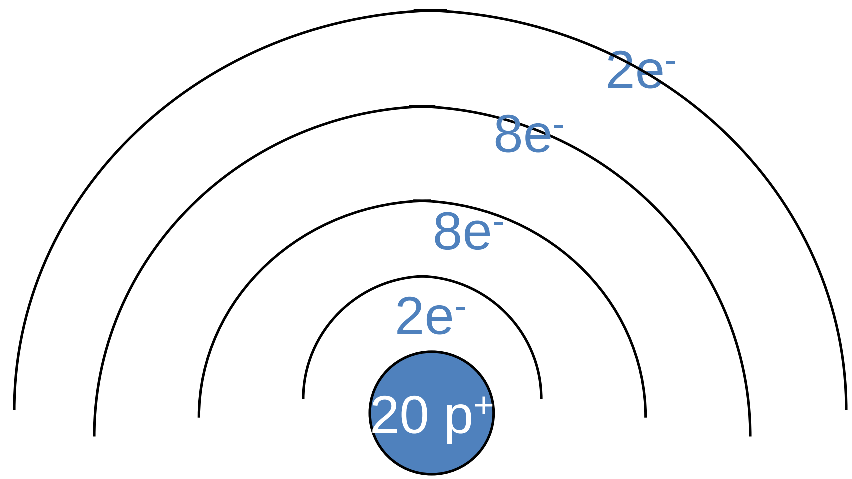
Be ($4p^+$ & $4e^-$)



Mg ($12p^+$ & $12e^-$)



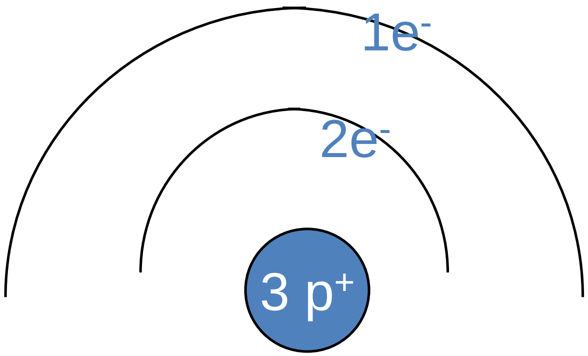
Ca ($20p^+$ & $20e^-$)



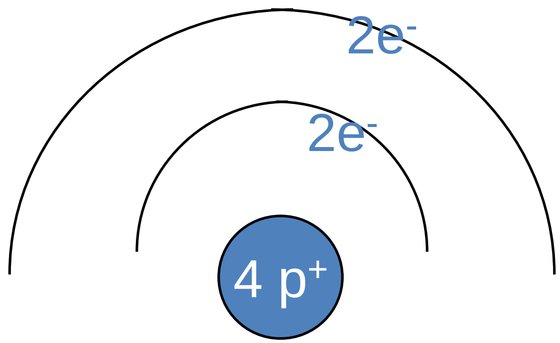
Atomic Radii Decreases Across the Period

Period 2

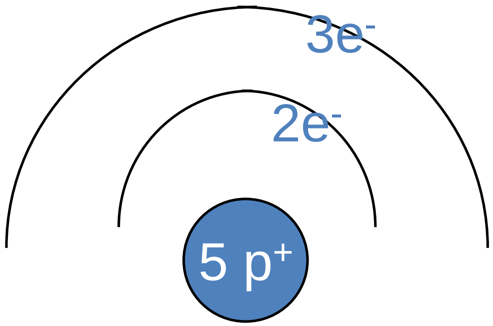
Li	Be	B	C	O	Ne



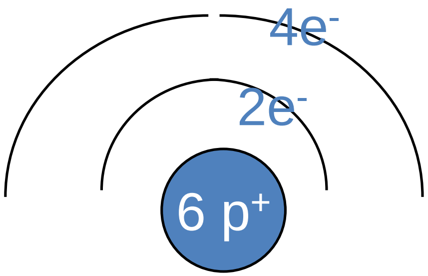
Li ($3p^+$ & $3e^-$)



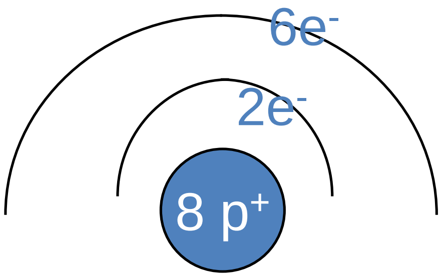
Be ($4p^+$ & $4e^-$)



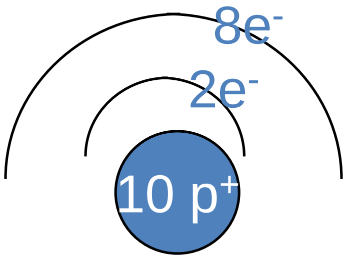
B ($5p^+$ & $5e^-$)



C ($6p^+$ & $6e^-$)

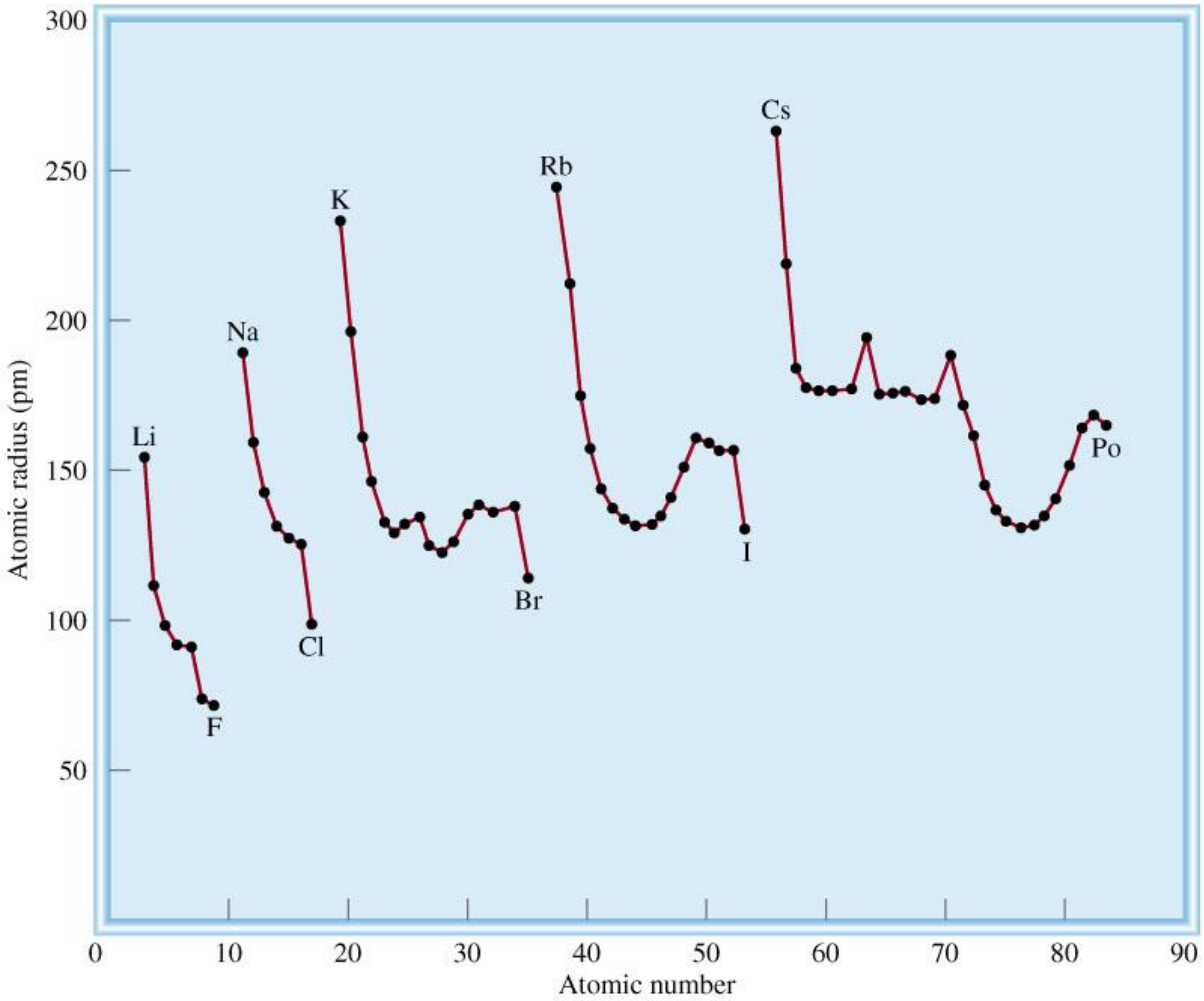


O ($8p^+$ & $8e^-$)

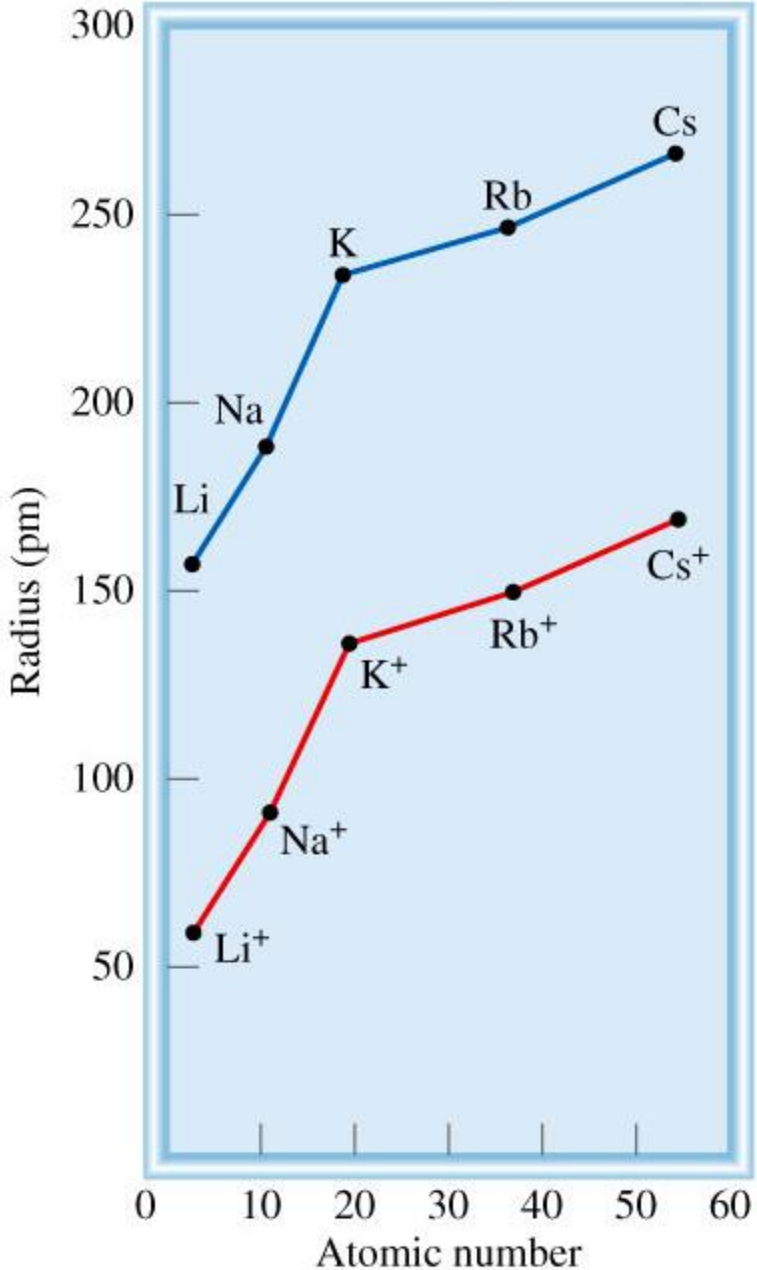


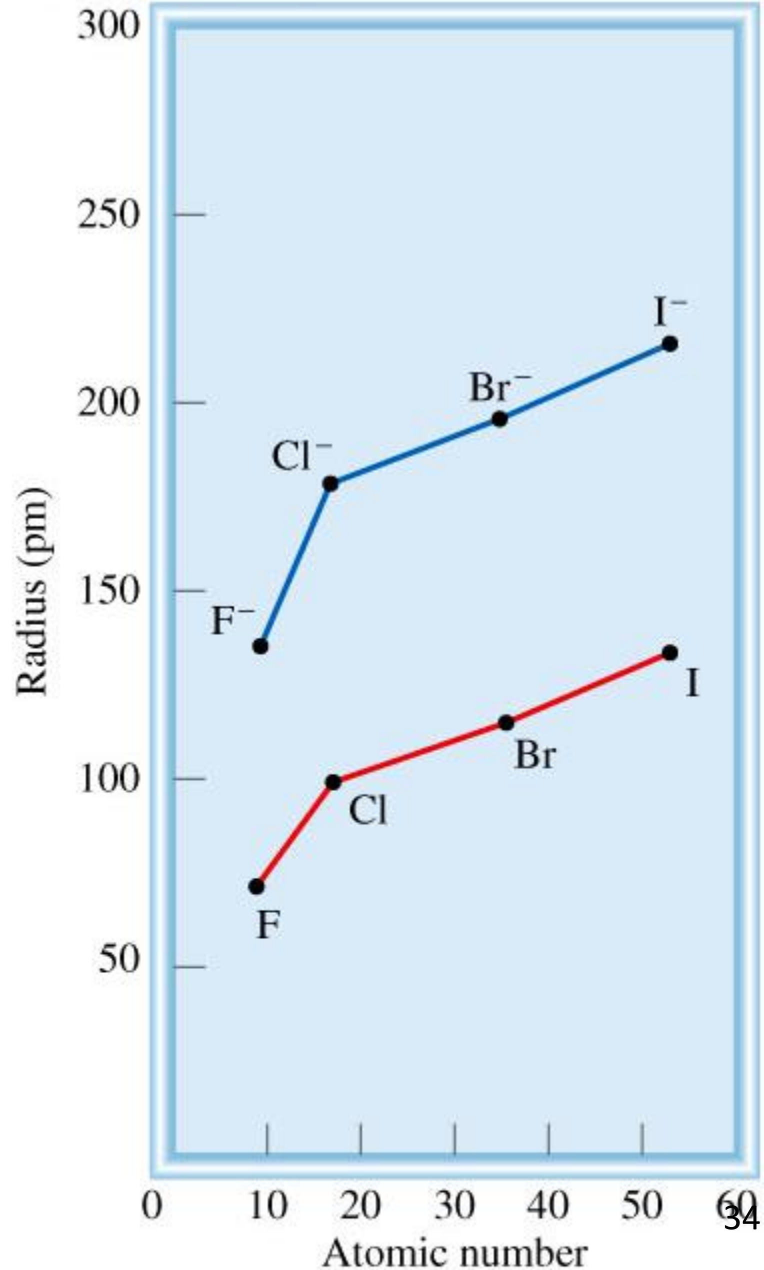
Ne ($10p^+$ & $10e^-$)

Trends in Atomic Radii

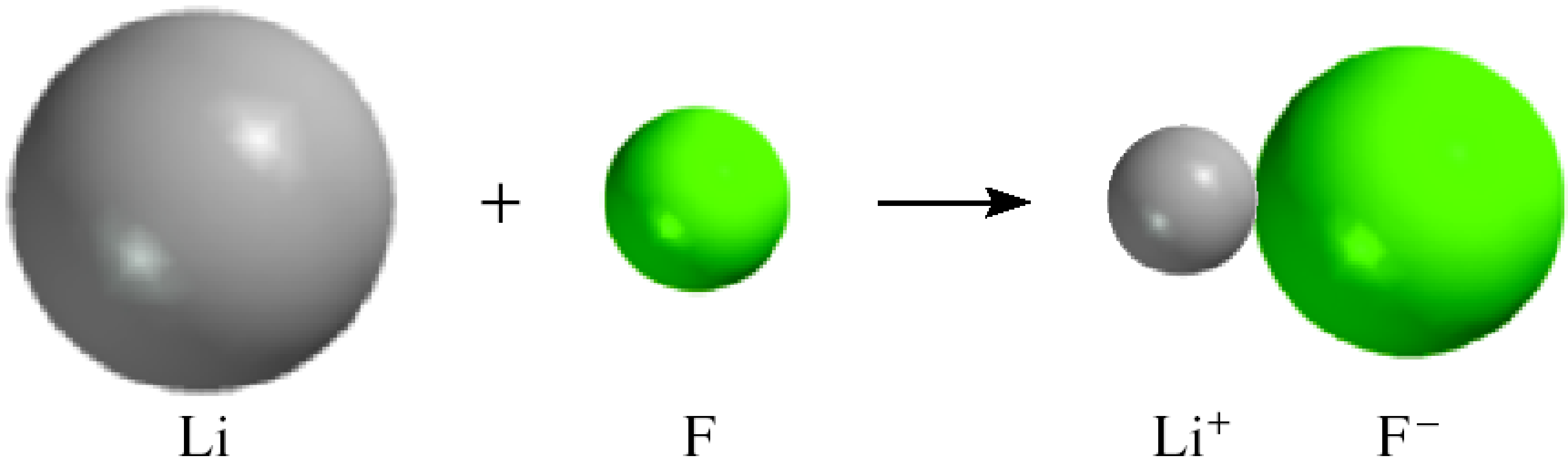


Comparison of Atomic Radii with Ionic Radii





Atomic Radii & Ionic Radii



Cation is always **smaller** than atom from which it is formed.
Anion is always **larger** than atom from which it is formed.

Radii for Isoelectronic Atoms or Ions

Isoelectronic: have the same number of electrons, and hence the same ground-state electron configuration

Na^+: [Ne] Al^{3+}: [Ne] F^-: $1s^22s^22p^6$ or [Ne]

O^{2-}: $1s^22s^22p^6$ or [Ne] N^{3-}: $1s^22s^22p^6$ or [Ne]

Na^+, Al^{3+}, F^-, O^{2-}, and N^{3-} are all ***isoelectronic*** with Ne

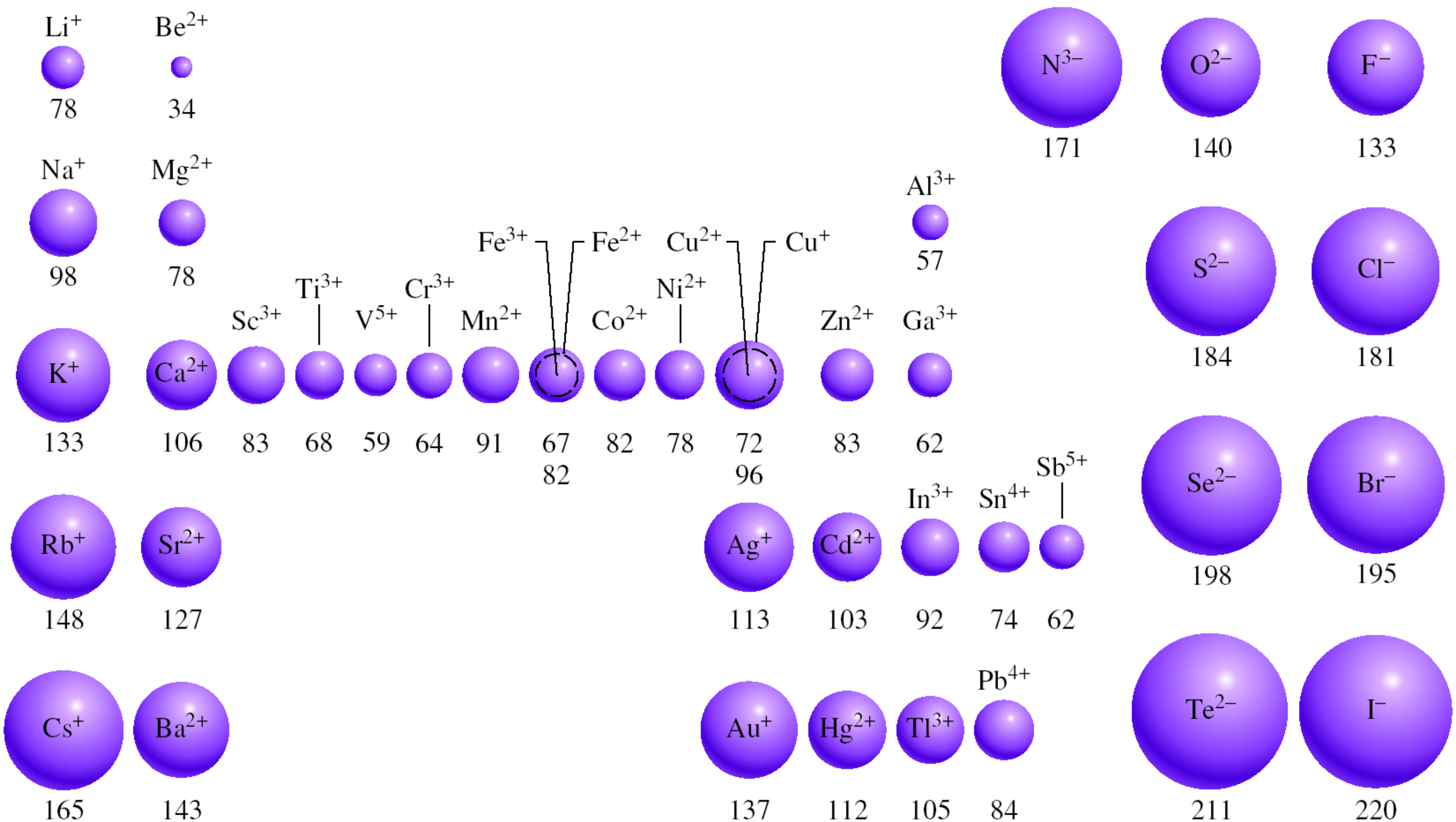
What neutral atom is isoelectronic with H^- ?

H^- : $1s^2$ same electron configuration as He

Radii for Isoelectronic Atoms or Ions contd.

Atom or Ion	Radius (nm)	Electron config.
C^{4-}	0.260	$1s^2 2s^2 2p^6$
N^{3-}	0.171	$1s^2 2s^2 2p^6$
O^{-2}	0.140	$1s^2 2s^2 2p^6$
F^{-}	0.136	$1s^2 2s^2 2p^6$
Ne	0.112	$1s^2 2s^2 2p^6$
Na^{+}	0.095	$1s^2 2s^2 2p^6$
Mg^{2+}	0.065	$1s^2 2s^2 2p^6$
Al^{3+}	0.050	$1s^2 2s^2 2p^6$

The Radii (in pm) of Ions of Familiar Elements



Ionization Energy

Ionization energy is the minimum energy (kJ/mol) required to remove an electron from a gaseous atom in its ground state.

$\boldsymbol{I}_1 + X_{(g)} \longrightarrow X^{+}_{(g)} + e^-$ $\boldsymbol{I}_1$ **first** ionization energy

$\boldsymbol{I}_2 + X^{+}_{(g)} \longrightarrow X^{2+}_{(g)} + e^-$ $\boldsymbol{I}_2$ **second** ionization energy

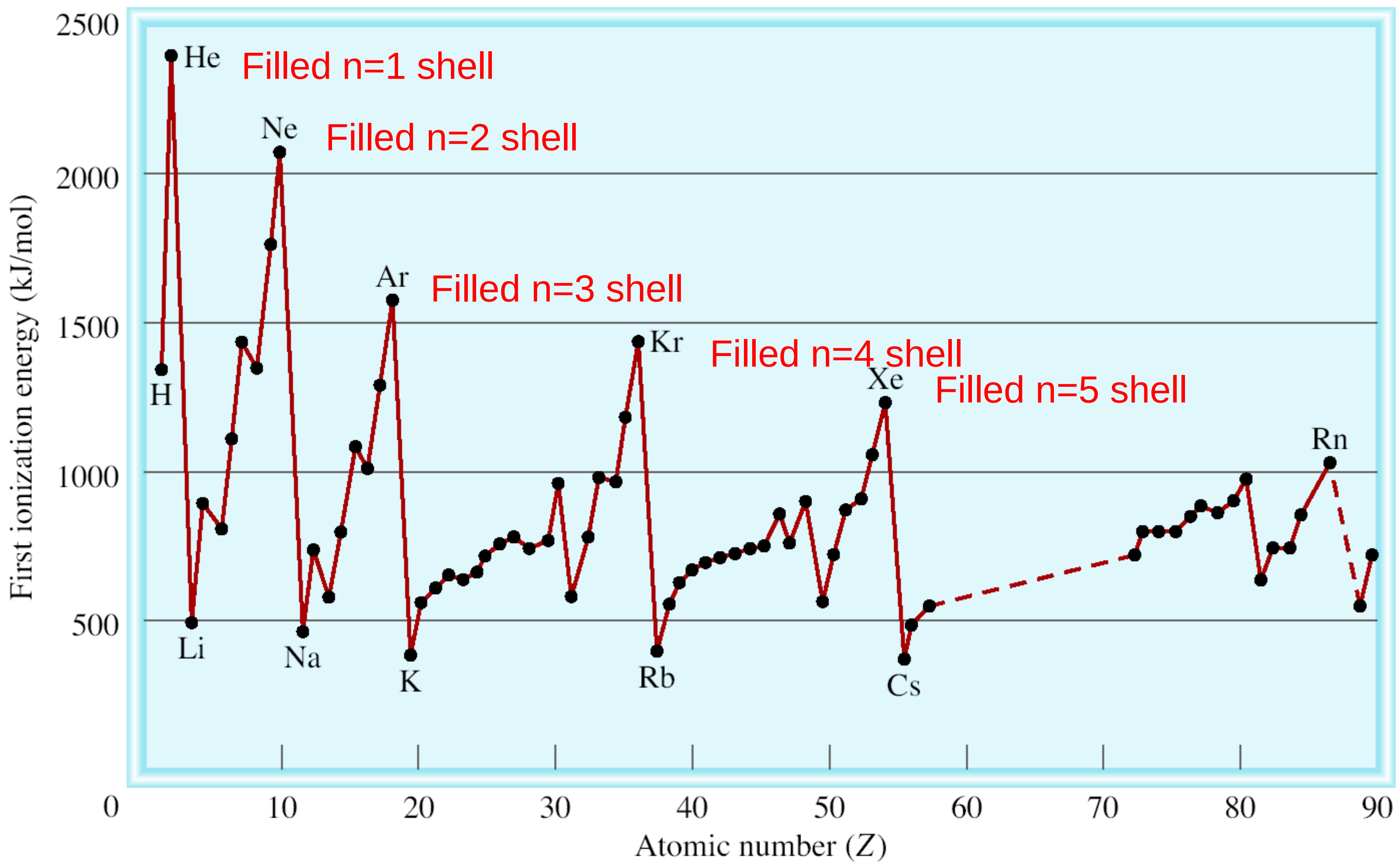
$\boldsymbol{I}_3 + X^{2+}_{(g)} \longrightarrow X^{3+}_{(g)} + e^-$ $\boldsymbol{I}_3$ **third** ionization energy

$$\boxed{\boldsymbol{I}_1 < \boldsymbol{I}_2 < \boldsymbol{I}_3}$$

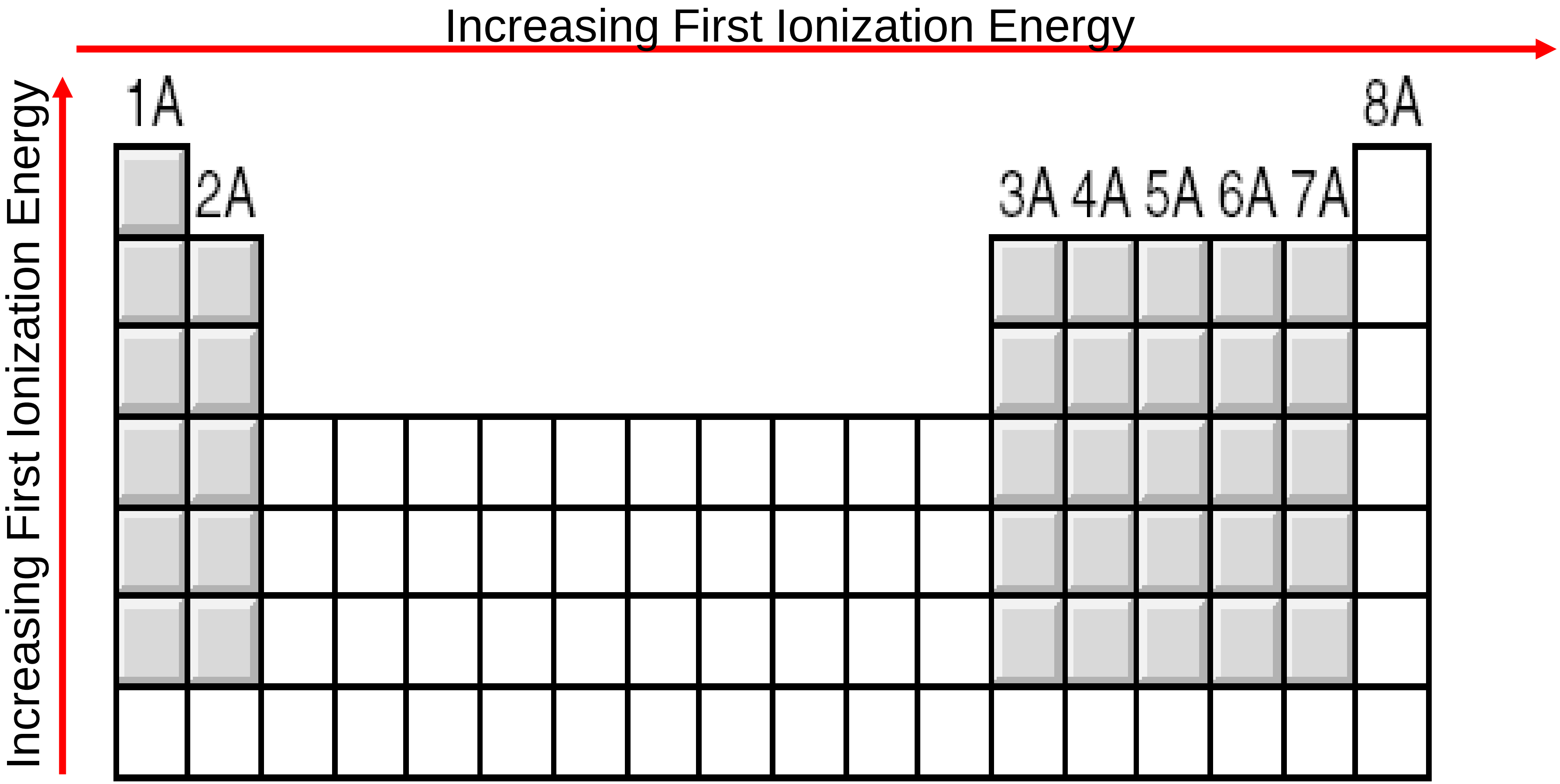
TABLE 8.2 The Ionization Energies (kJ/mol) of the First 20 Elements

Z	Element	First	Second	Third	Fourth	Fifth	Sixth
1	H	1,312					
2	He	2,373	5,251				
3	Li	520	7,300	11,815			
4	Be	899	1,757	14,850	21,005		
5	B	801	2,430	3,660	25,000	32,820	
6	C	1,086	2,350	4,620	6,220	38,000	47,261
7	N	1,400	2,860	4,580	7,500	9,400	53,000
8	O	1,314	3,390	5,300	7,470	11,000	13,000
9	F	1,680	3,370	6,050	8,400	11,000	15,200
10	Ne	2,080	3,950	6,120	9,370	12,200	15,000
11	Na	495.9	4,560	6,900	9,540	13,400	16,600
12	Mg	738.1	1,450	7,730	10,500	13,600	18,000
13	Al	577.9	1,820	2,750	11,600	14,800	18,400
14	Si	786.3	1,580	3,230	4,360	16,000	20,000
15	P	1,012	1,904	2,910	4,960	6,240	21,000
16	S	999.5	2,250	3,360	4,660	6,990	8,500
17	Cl	1,251	2,297	3,820	5,160	6,540	9,300
18	Ar	1,521	2,666	3,900	5,770	7,240	8,800
19	K	418.7	3,052	4,410	5,900	8,000	9,600
20	Ca	589.5	1,145	4,900	6,500	8,100	11,000

Variation of the First Ionization Energy with Atomic Number



General Trends in First Ionization Energies



Electron Affinity

Electron affinity is the negative of the energy change that occurs when an electron is accepted by an atom in the gaseous state to form an anion.

$$X_{(g)} + e^- \longrightarrow X^-_{(g)}$$

$F_{(g)} + e^- \longrightarrow F^-_{(g)}$ $\Delta H = -328$ kJ/mol   EA = +328 kJ/mol

$O_{(g)} + e^- \longrightarrow O^-_{(g)}$ $\Delta H = -141$ kJ/mol   EA = +141 kJ/mol

$O^-_{(g)} + e^- \longrightarrow O^{2-}_{(g)}$ $\Delta H = +780$ kJ/mol   EA = −780 kJ/mol**?**

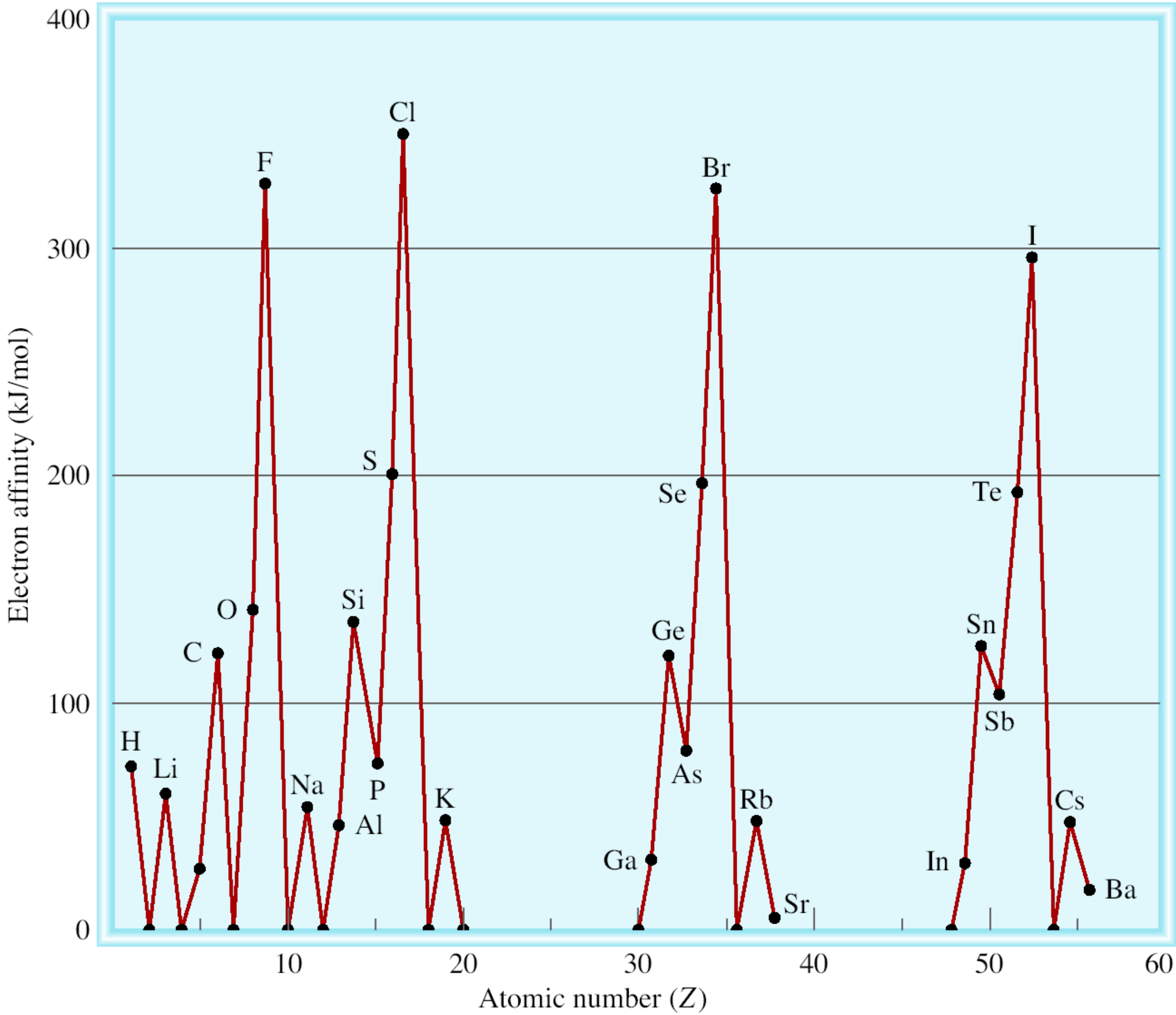
- EA of O^{2-} is highly –ve; *i.e.* O^{2-} formation is unfavorable,
- electron-electron repulsion (*g*) > stability from noble gas configuration;
- O^{2-} ion stabilized by neighboring cation, *e.g.* Li_2O, Mg_2O

TABLE 8.3 **Electron Affinities (kJ/mol) of Some Representative Elements and the Noble Gases***

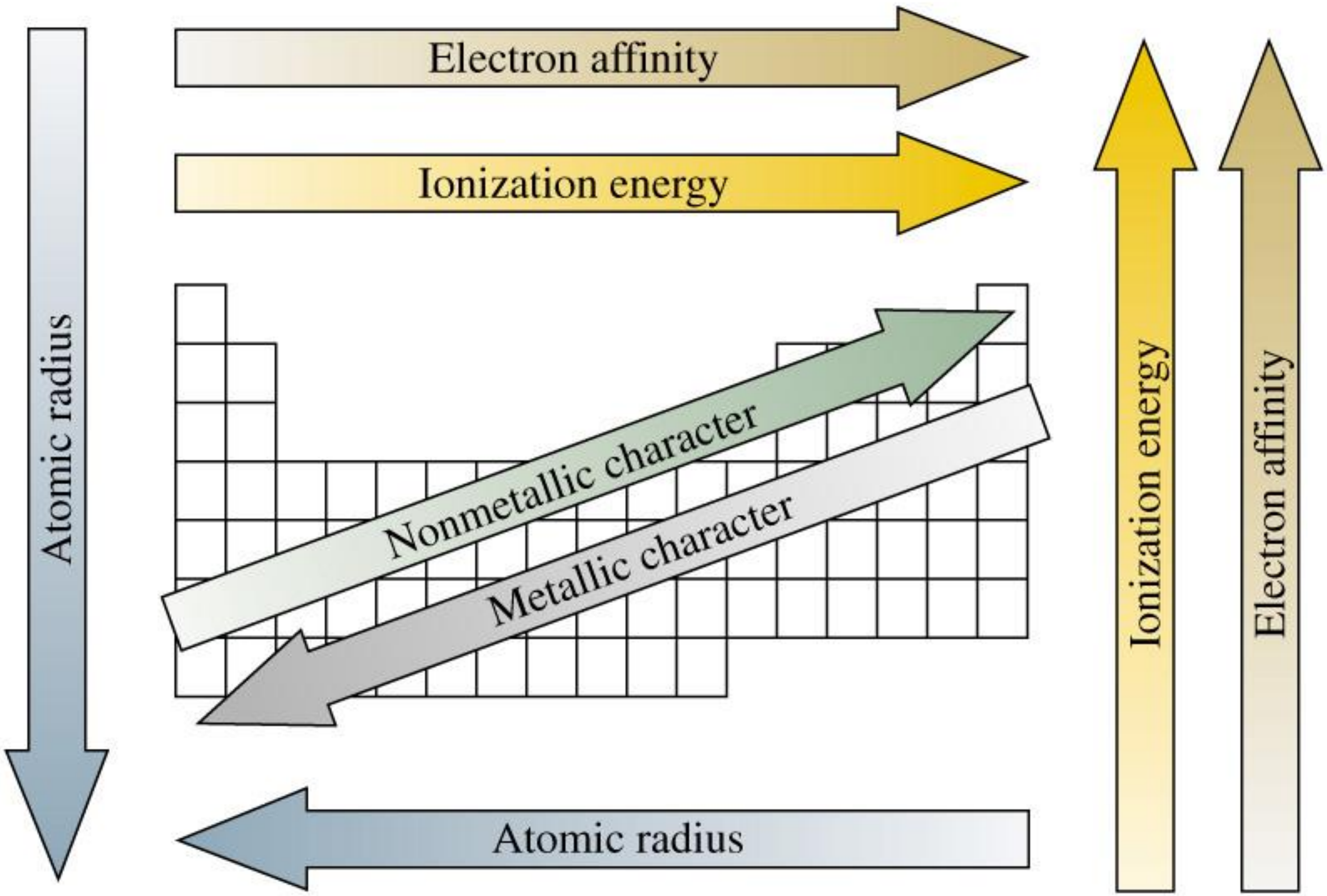
1A	2A	3A	4A	5A	6A	7A	8A
H							He
73							< 0
Li	Be	B	C	N	O	F	Ne
60	≤ 0	27	122	0	141	328	< 0
Na	Mg	Al	Si	P	S	Cl	Ar
53	≤ 0	44	134	72	200	349	< 0
K	Ca	Ga	Ge	As	Se	Br	Kr
48	2.4	29	118	77	195	325	< 0
Rb	Sr	In	Sn	Sb	Te	I	Xe
47	4.7	29	121	101	190	295	< 0
Cs	Ba	Tl	Pb	Bi	Po	At	Rn
45	14	30	110	110	?	?	< 0

*The electron affinities of the noble gases, Be, and Mg have not been determined experimentally, but are believed to be close to zero or negative.

Variation of Electron Affinity With Atomic Number (H – Ba)



Periodic Properties of the Elements



Variation in Chemical Properties of the Representative Elements

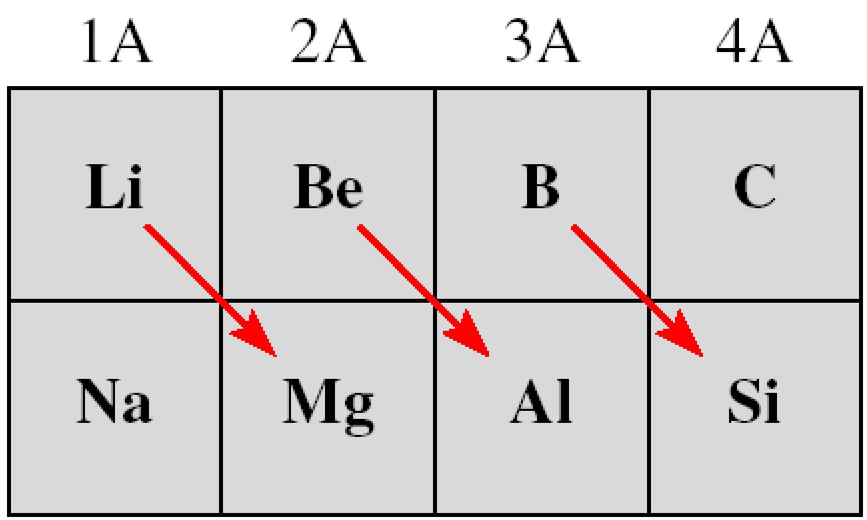
- Ionization energy (**IE**) and electron affinity (**EA**) help us to understand:
  - the **types** of reactions an element undergo
  - the **nature** of the elements' compounds formed
- **IE** measures the attraction of an atom for its **own** electrons
- **EA** expresses the attraction of an atom for **additional** electrons
  - **IE** and **EA** together gives the insight into the **general attraction** of an atom for electron
- With these concepts we can survey the chemical behavior of the elements systematically.

Variation in Chemical Properties of the Representative Elements

- The **metallic character** of the elements
  - **decreases** from left to right across a period, and
  - **increases** from top to bottom within a group
- Metals usually have **low** ionization energies, and Nonmetals usually have **high** electron affinity
- On the basis of these trends we can predict the outcome of a reaction involving some of these elements

General Trends in Chemical Properties

- First member of each group **differ** from the rest of the member of the same group.
  - *Li & Be* ***do not*** *show all the properties of Group 1A & 2A, because of unusually small size.*
- ***Diagonal relationship:*** similarities between pairs of elements in different groups and periods.
  - *Reason is the closeness of* ***charge densities*** *of their cations.*



Hydrogen ($1s^1$)

- No suitable position for hydrogen in the periodic table, it really could be a class by itself
- Like alkali metals, has a single *s* electron and forms H^+ ion, which is hydrated in solution (H_3O^+)
- Also forms hydride ion H^- in ionic compounds, NaH, CaH_2
- Hydrides reacts with water:

$$2NaH(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

$$CaH_2(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + 2H_2(g)$$

- Most important compound of hydrogen is water,

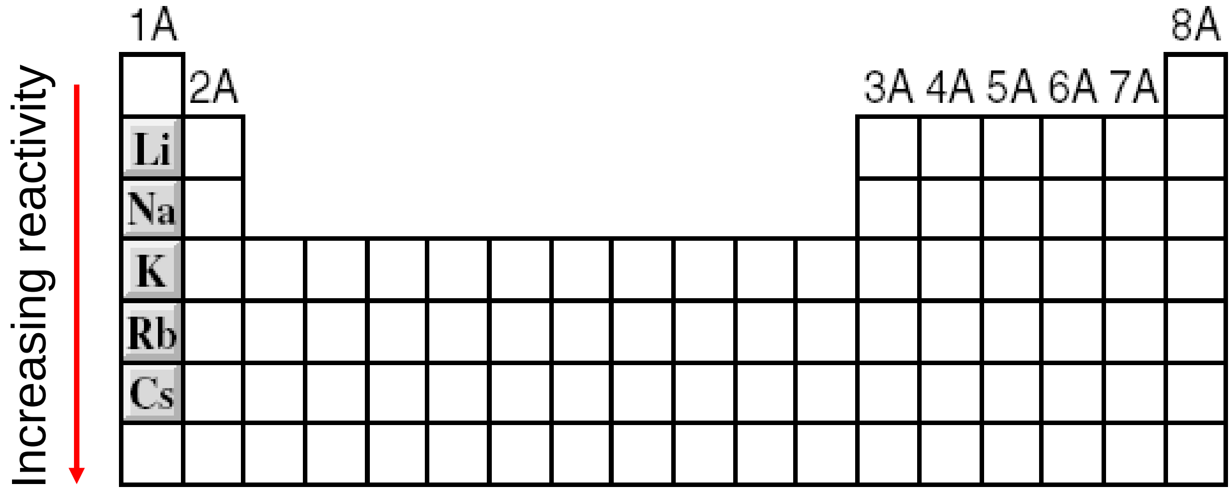
$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

Group 1A Elements (ns^1, $n \geq 2$)

$$M \longrightarrow M^{+1} + 1e^-$$

So reactive, never found in pure state in nature.

$$2M_{(s)} + 2H_2O_{(l)} \longrightarrow 2MOH_{(aq)} + H_{2(g)}$$



Group 1A Elements (ns^1, $n \geq 2$)

In air lose shiny appearance by combining with O_2 to form oxides:

- Li forms *oxides*
- Na, K, Rb, & Cs forms *oxides* and *peroxides*
- K, Rb, & Cs also forms *superoxides*

$4Li_{(s)} + O_{2(g)} \longrightarrow 2Li_2O_{(s)}$; Lithium oxide, contg. O^{2-}

$2Na_{(s)} + O_{2(g)} \longrightarrow Na_2O_{2(s)}$; peroxide, contg. O_2^{2-}

$K_{(s)} + O_{2(g)} \longrightarrow KO_{2(s)}$; superoxide, contg. O_2^{-}

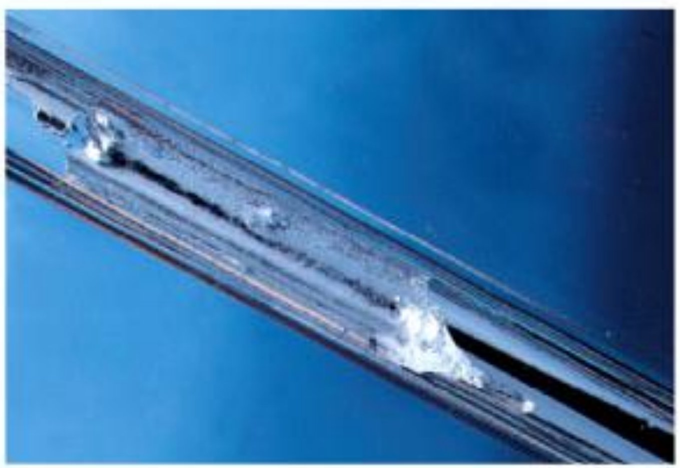
Stability of oxides depends on how strongly the cations and anions attracts one another in the solid states.

Group 1A Elements (ns^1, $n \geq 2$)

Lithium (Li)

Sodium (Na)

Potassium (K)



Rubidium (Rb)

Cesium (Cs)

Group 2A Elements (ns^2, $n \geq 2$)

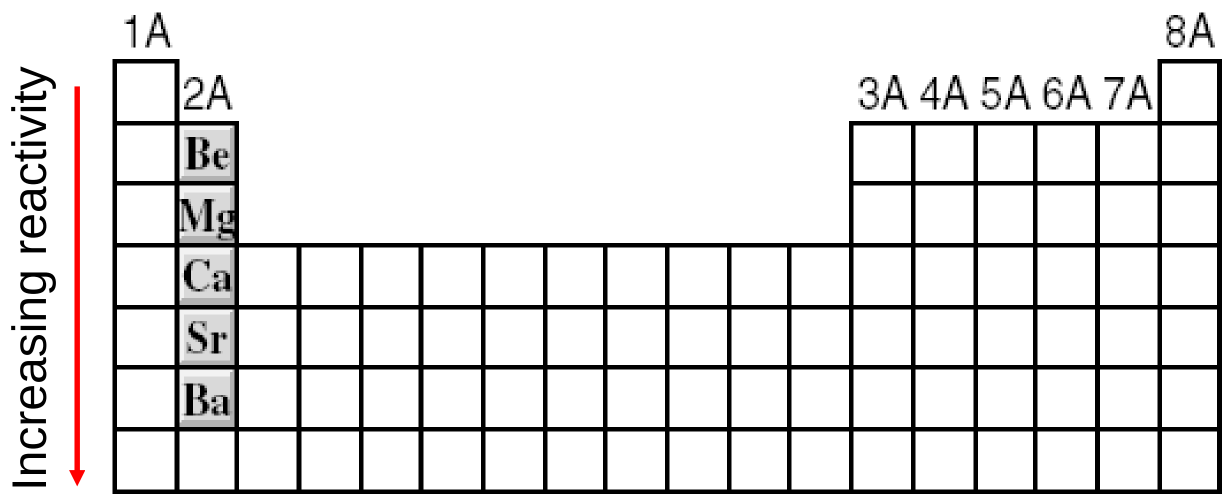
$M \longrightarrow M^{+2} + 2e^-$; but BeH_2, $BeCl_2$ & MgH_2 are molecular

$Be_{(s)} + 2H_2O_{(l)} \longrightarrow$ No Reaction

$Mg_{(s)} + 2H_2O_{(g)} \longrightarrow Mg(OH)_{2(aq)} + H_{2(g)}$

$M_{(s)} + 2H_2O_{(l)} \longrightarrow M(OH)_{2(aq)} + H_{2(g)}$ M = Ca, Sr, or Ba

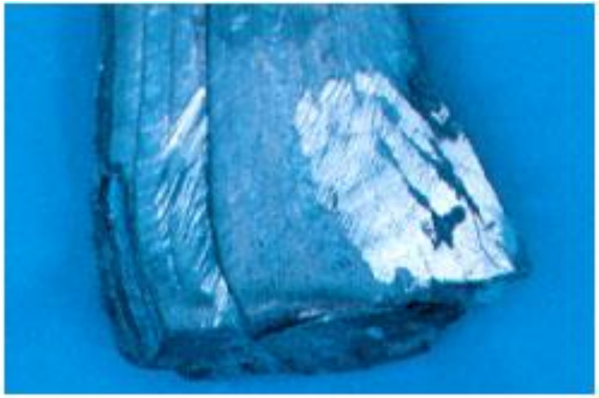
BeO & MgO form at high temp., CaO, SrO & BaO form at room temp.



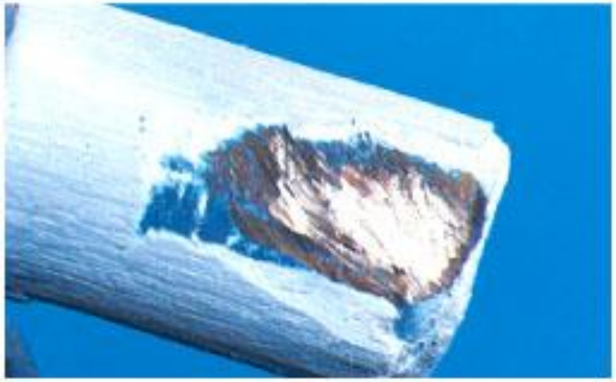
Group 2A Elements (ns^2, $n \geq 2$)

Beryllium (Be)

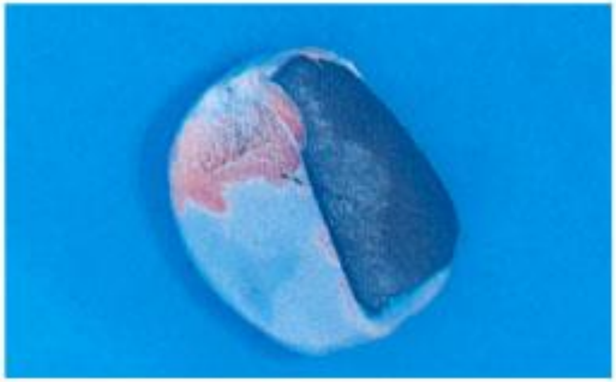
Magnesium (Mg)



Calcium (Ca)



Strontium (Sr)



Barium (Ba)

Radium (Ra)

Group 3A Elements (ns^2np^1, $n \geq 2$)

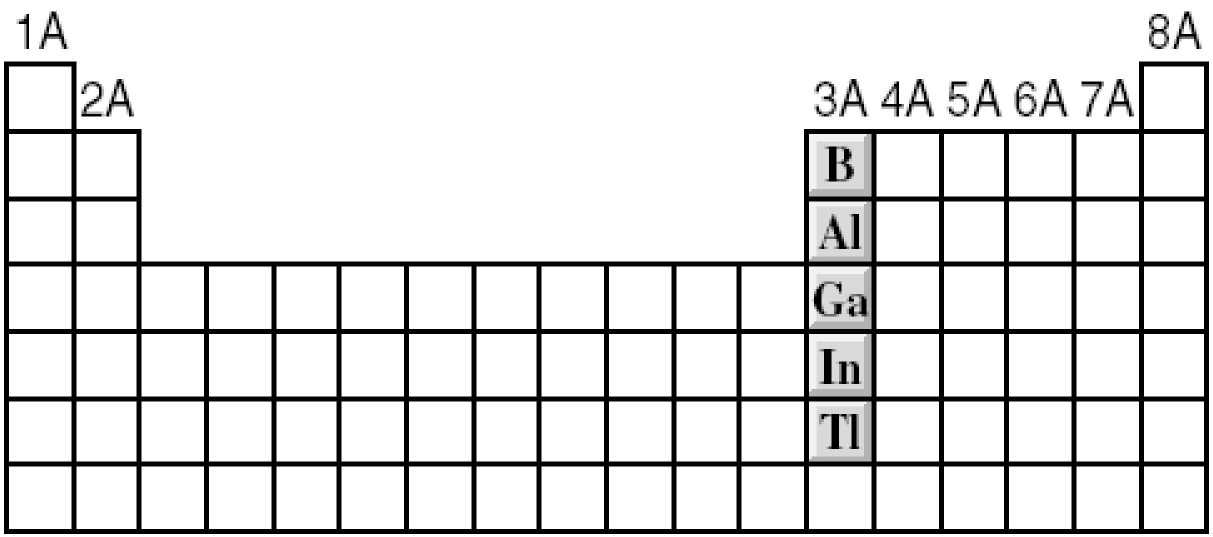
B does not form binary ionic comp. and unreactive toward O_2 & H_2O

$$4Al_{(s)} + 3O_{2(g)} \longrightarrow 2Al_2O_{3(s)}$$; only forms Al^{3+}

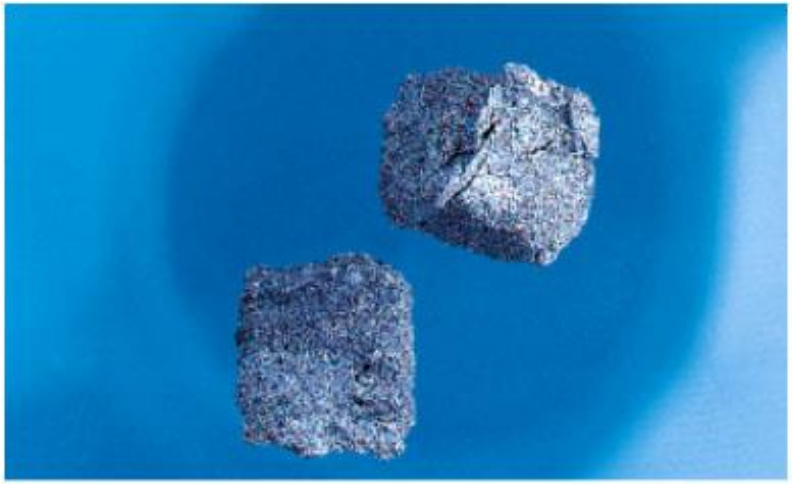
$$2Al_{(s)} + 6H^+_{(aq)} \longrightarrow 2Al^{3+}_{(aq)} + 3H_{2(g)}$$

Ga, In & Tl forms +1 & +3 ions, down the group +1 more stable

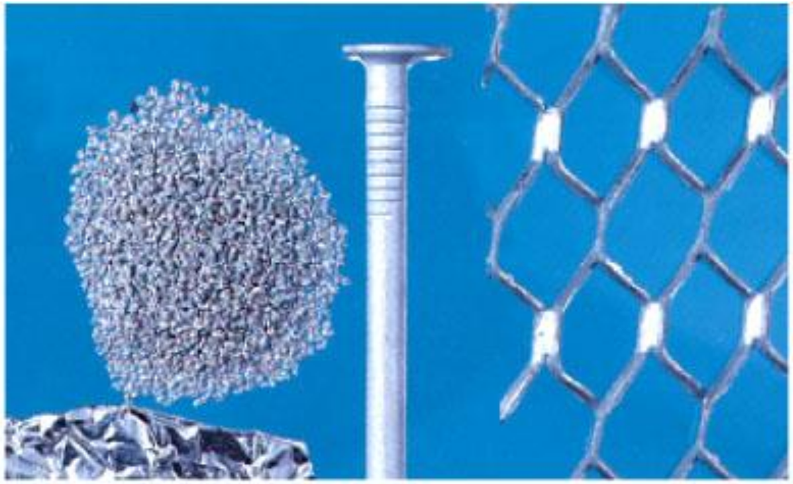
AlH_3 is molecular, resembles BeH_2, example of diagonal relatn.



Group 3A Elements (ns^2np^1, $n \geq 2$)



Boron (B)



Aluminum (Al)

Gallium (Ga)

Indium (In)

Group 4A Elements (ns^2np^2, $n \geq 2$)

Sn & Pb do not react with H_2O, but reacts with acids

$$Sn_{(s)} + 2H^+_{(aq)} \longrightarrow Sn^{2+}_{(aq)} + H_{2\ (g)}$$

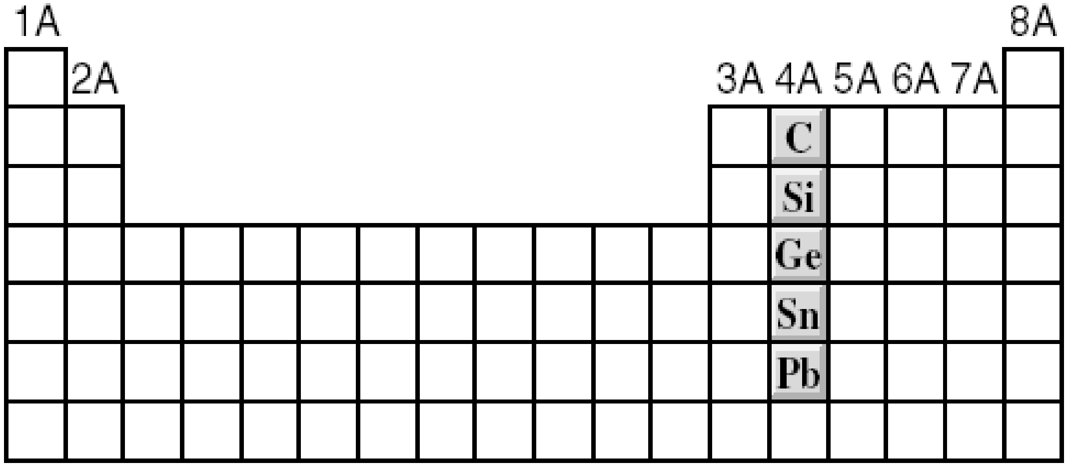
$$Pb_{(s)} + 2H^+_{(aq)} \longrightarrow Pb^{2+}_{(aq)} + H_{2\ (g)}$$

Forms +2 & +4 oxidation states,

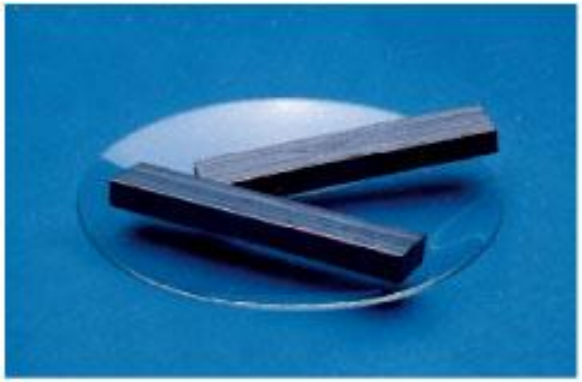
for C, +4 is more stable, CO_2 more stable than CO,

for Si, SiO_2 is stable but SiO is unstable,

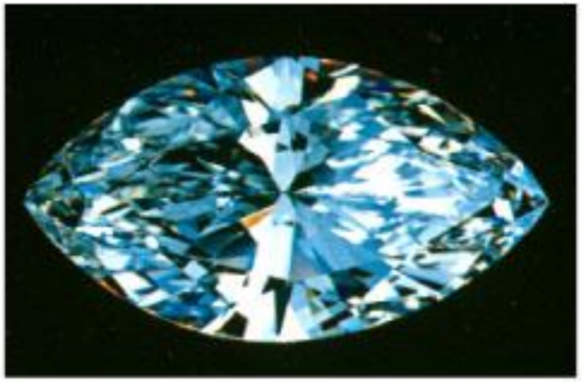
down the group trend is reversed, $Sn^{+4} \approx Sn^{+2}$, $Pb^{+2} > Pb^{+4}$



Group 4A Elements (ns^2np^2, $n \geq 2$)

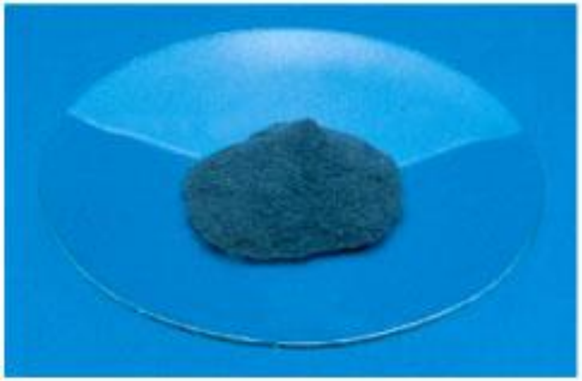


Carbon (graphite)



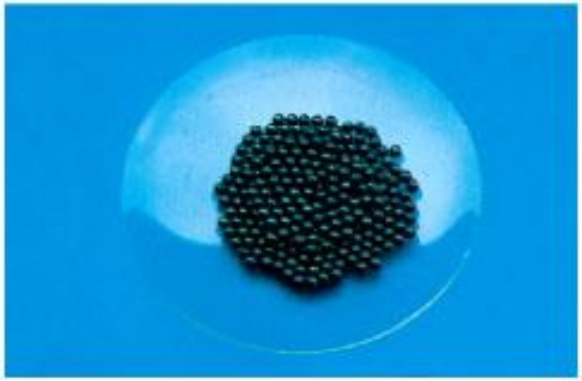
Carbon (diamond)

Silicon (Si)



Germanium (Ge)

Tin (Sn)



Lead (Pb)

Group 5A Elements (ns^2np^3, $n \geq 2$)

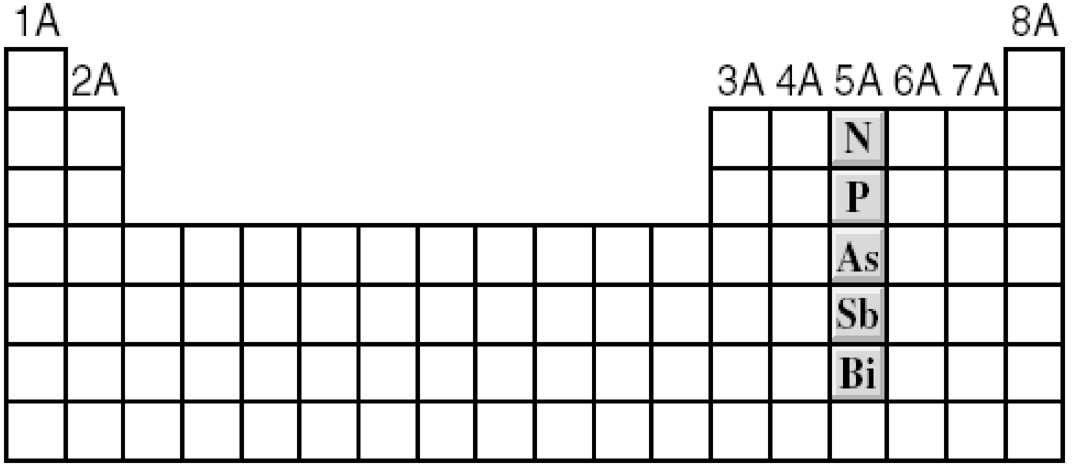
Nitrogen is $N_2(g)$, forms $NO(g)$, $N_2O(g)$, $NO_2(g)$, $N_2O_4(g)$ & $N_2O_5(s)$, form nitride ion N^{3-} (isoelectronic with Ne), Li_3N, Mg_3N_2 are ionic

Phosphorus is P_4 molecule, forms $P_4O_6(s)$ & $P_4O_{10}(s)$

$$N_2O_{5(s)} + H_2O_{(l)} \longrightarrow 2HNO_{3(aq)}$$

$$P_4O_{10(s)} + 6H_2O_{(l)} \longrightarrow 4H_3PO_{4(aq)}$$

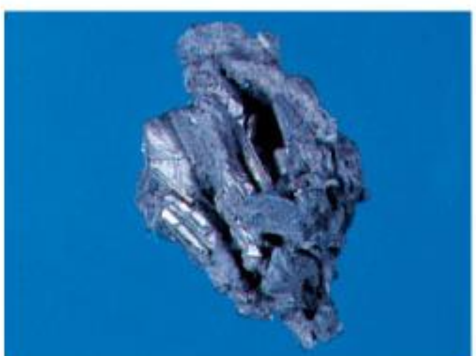
As, Sb & Bi have three dimensional structures, Bi is less reactive



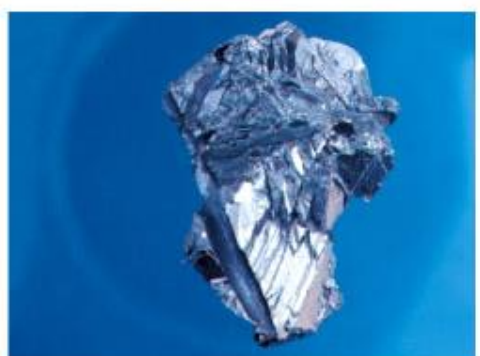
Group 5A Elements (ns^2np^3, $n \geq 2$)

Nitrogen (N_2)

White and red phosphorus (P)



Arsenic (As)



Antimony (Sb)

Bismuth (Bi)

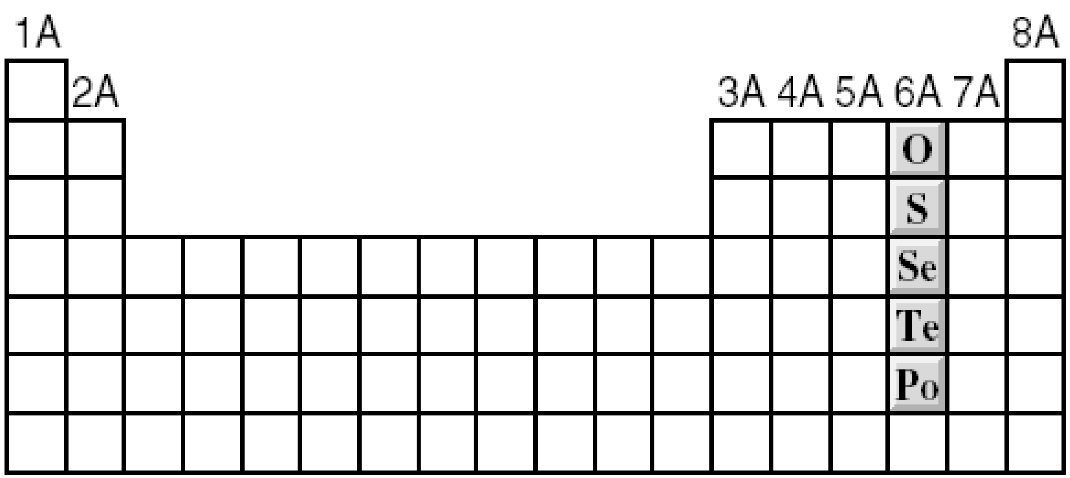
Group 6A Elements (ns^2np^4, $n \geq 2$)

Oxygen is O_2, Sulfur is S_8 & Selenium is Se_8,

Te & Po have three dimensional structure, Po is radioactive,

Forms O^{2-}, S^{2-}, Se^{2-} & Te^{2-} anions, and many molecular compounds with nonmetals, *e.g.* SO_2, SO_3 and H_2S

$$SO_{3(g)} + H_2O_{(l)} \longrightarrow H_2SO_{4(aq)}$$



Group 6A Elements (ns^2np^4, $n \geq 2$)

Sulfur (S_8)

Selenium (Se_8)

Tellurium (Te)

Group 7A Elements (ns^2np^5, $n \geq 2$)

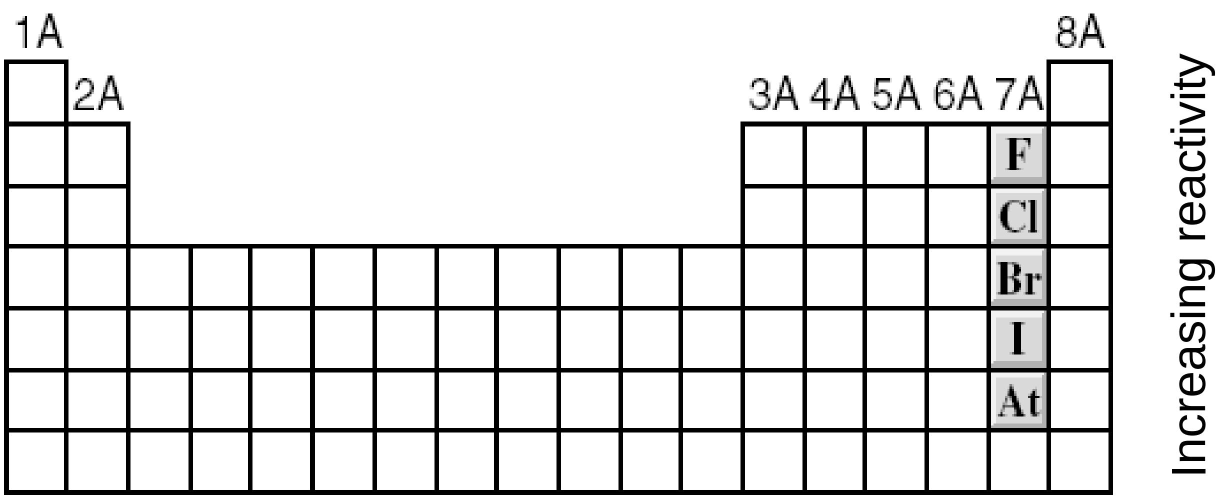
Halogens are X_2, very reactive, never found as elements in nature, At is radioactive

$X + 1e^- \longrightarrow X^{-1}$, have high IE and large positive EA

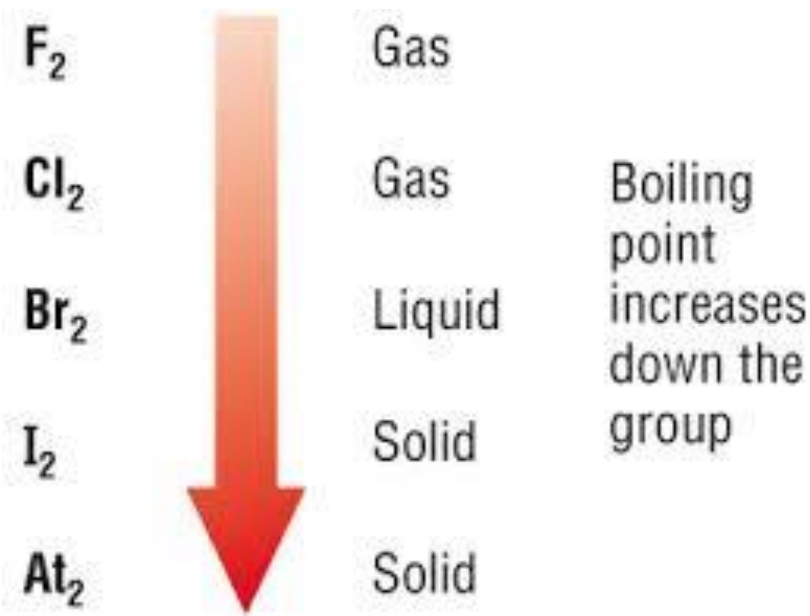
$$2F_{2(g)} + 2H_2O_{(l)} \longrightarrow 4HF_{(aq)} + O_{2(g)}$$

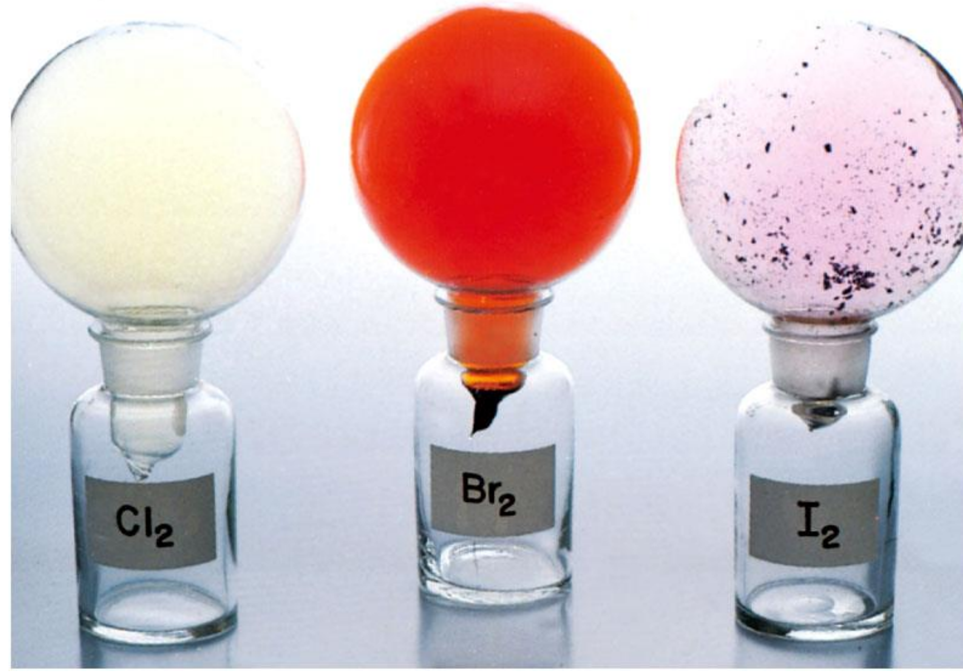
Gr. 1A & 2A halides are ionic, forms molecular compounds, *e.g.*, ICl & BrF_3 and NF_3, PCl_5 & SF_6

$X_{2(g)} + H_{2(g)} \longrightarrow 2HX_{(g)}$ HF*(aq)* is weak and other are strong acid



Group 7A Elements (ns^2np^5, $n \geq 2$)





Fluorine is very pale **yellow**, chlorine is **yellow-green**, and bromine is **red-brown**. Iodine crystals are shiny **purple**.

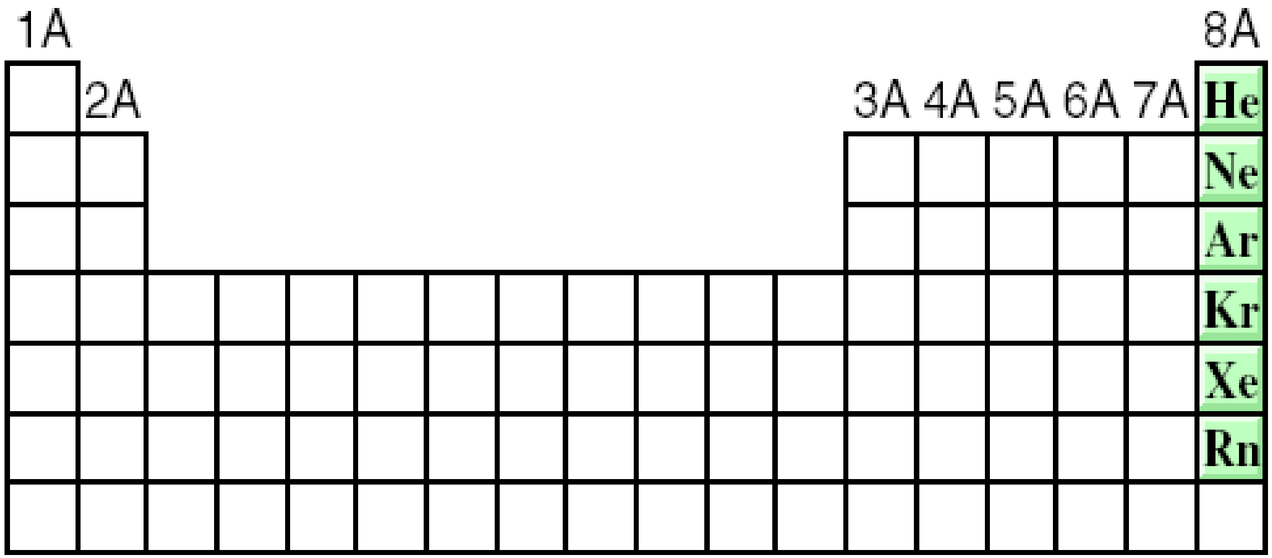
Group 8A Elements (ns^2np^6, $n \geq 2$)

All noble gases exists as monatomic species.

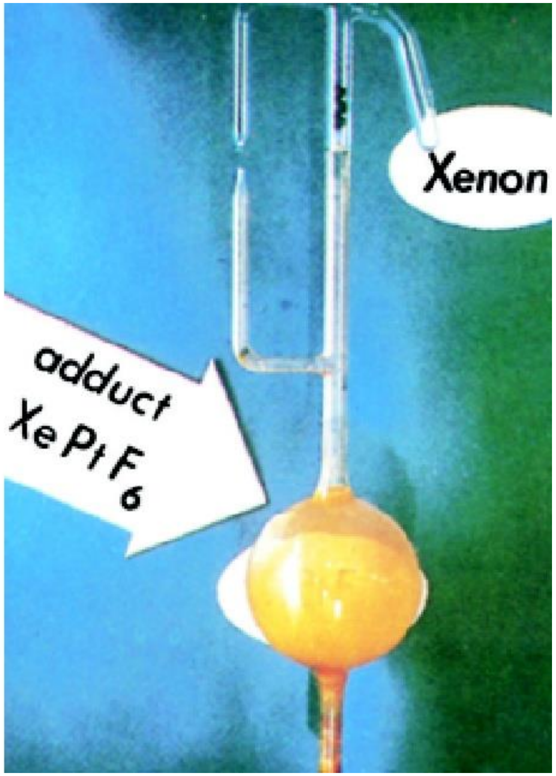
Completely filled *ns* and *np* subshells.

Highest ionization energy (IE) of all elements.

No tendency to accept extra electrons.



Compounds of the Noble Gases



A number of **xenon** compounds XeF_4, XeO_3, XeO_4, $XeOF_4$ exist.

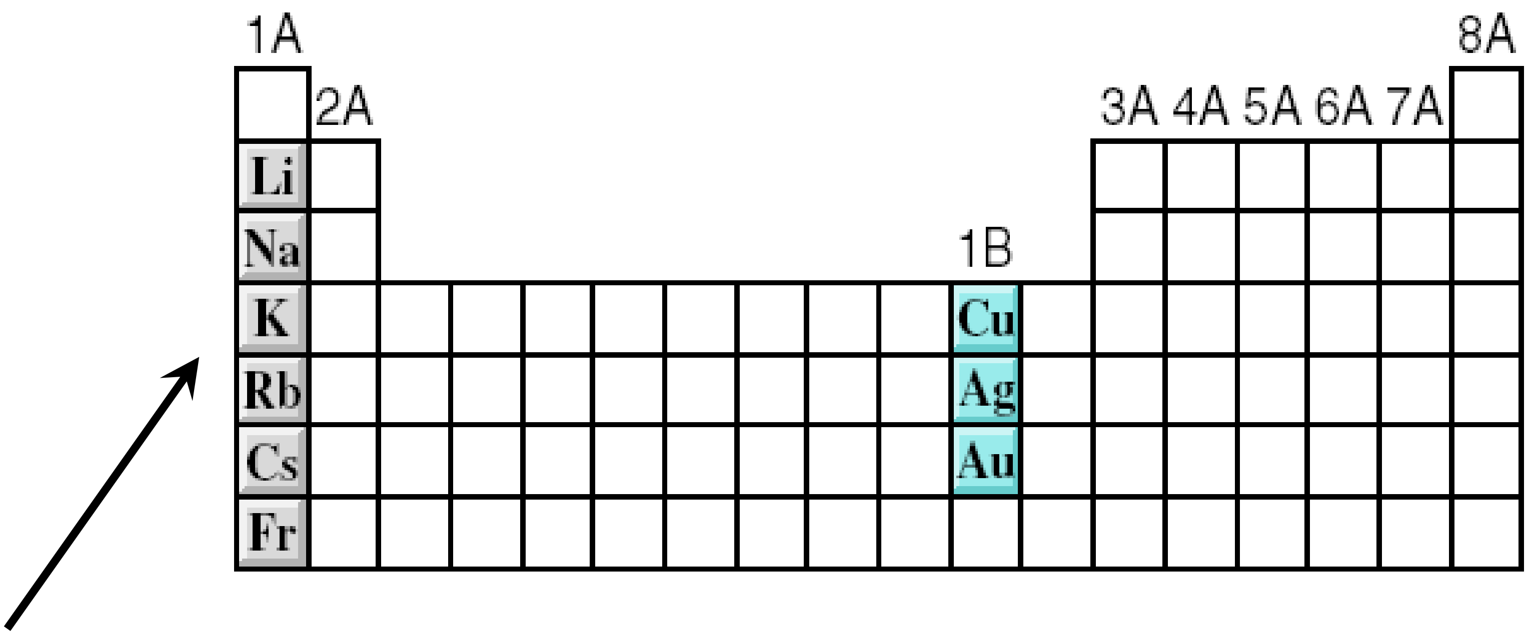
A few **krypton** compounds (*e.g.*, KrF_2) have been prepared.

No compounds of **helium** and **neon** are known.

Comparison of Group 1A and 1B

The metals in these two groups have similar outer electron configurations, with one electron in the outermost *s* orbital.

Chemical properties are quite different due to difference in the ionization energy.



Lower I_1, more reactive

Properties of Oxides Across a Period

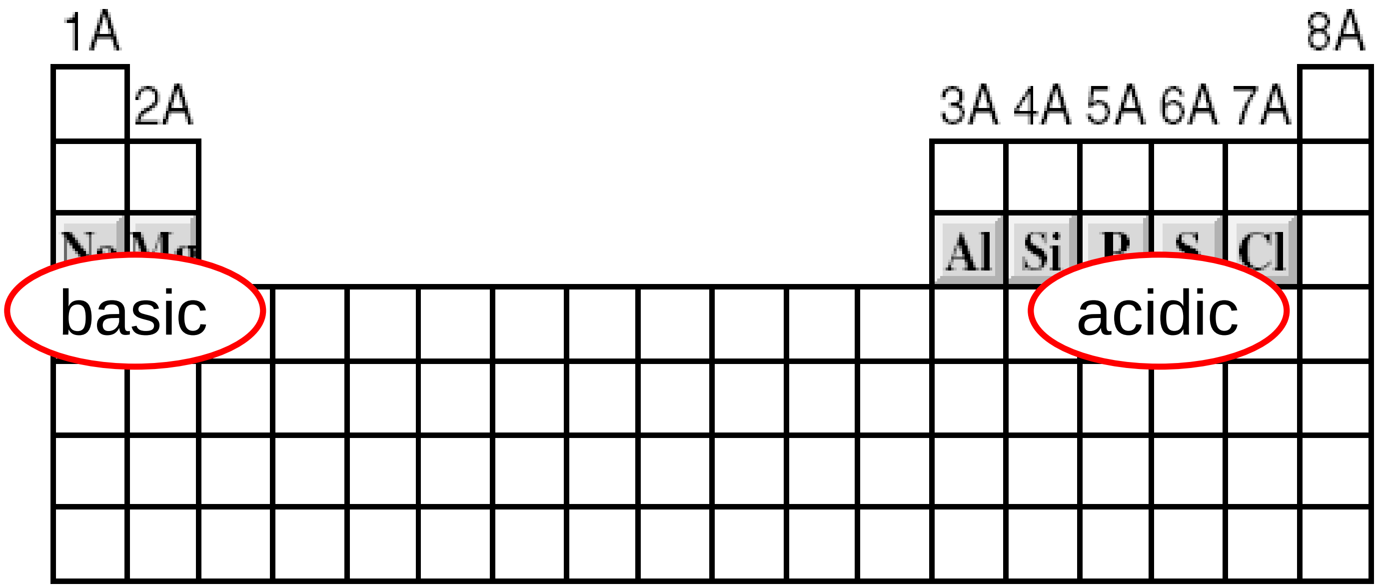


TABLE 8.4 Some Properties of Oxides of the Third-Period Elements

	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
Type of compound	⟵ Ionic ⟶			⟵ Molecular ⟶			
Structure	⟵ Extensive three-dimensional ⟶				⟵ Discrete molecular units ⟶		
Melting point (°C)	1275	2800	2045	1610	580	16.8	−91.5
Boiling point (°C)	?	3600	2980	2230	?	44.8	82
Acid-base nature	Basic	Basic	Amphoteric	⟵ Acidic ⟶			

Properties of Oxides Across a Period

- **Basic oxides:**

Na_2O and MgO are **basic** oxides,

$$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$$

$$MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$$

- **Amphoteric oxide:**

Al_2O_3 is an **amphoteric** oxide,

$$Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O(l)$$

$$Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq)$$

Other amphoteric oxides are, ZnO, BeO, and Bi_2O_3

Properties of Oxides Across a Period

- **Acidic oxides:**

SiO_2, P_4O_{10}, SO_3 and Cl_2O_7 are **acidic** oxides,

$$SiO_2(s) + 2NaOH(aq) \rightarrow Na_2SiO_3(aq) + H_2O(l)$$

$$P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$$

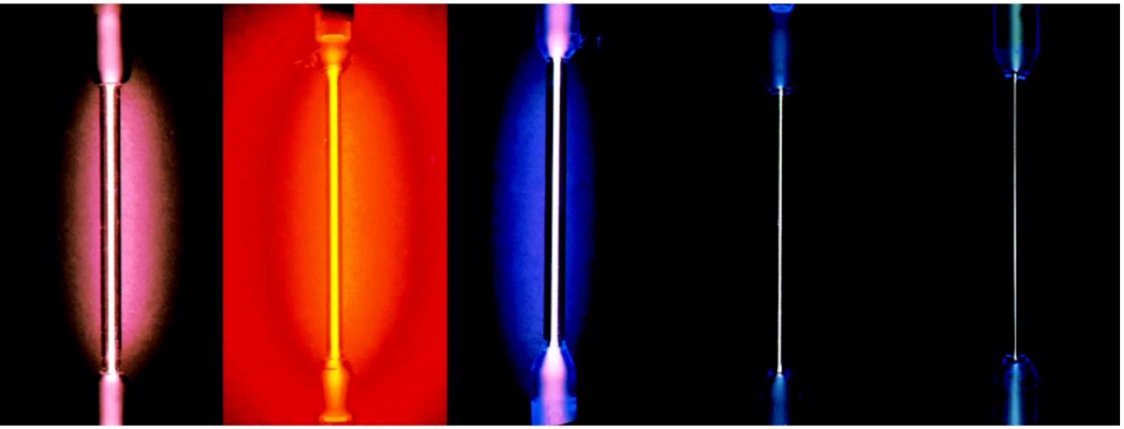
$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

$$Cl_2O_7(l) + H_2O(l) \rightarrow 2HClO_4(aq)$$

- **Neutral oxides:**

Some oxides are **neutral**, CO and NO,
Do **not** react with water to produce acidic or basic solution

Chemistry in Action: Discovery of the Noble Gases



Helium (He) Neon (Ne) Argon (Ar) Krypton (Kr) Xenon (Xe)

Sir William Ramsay