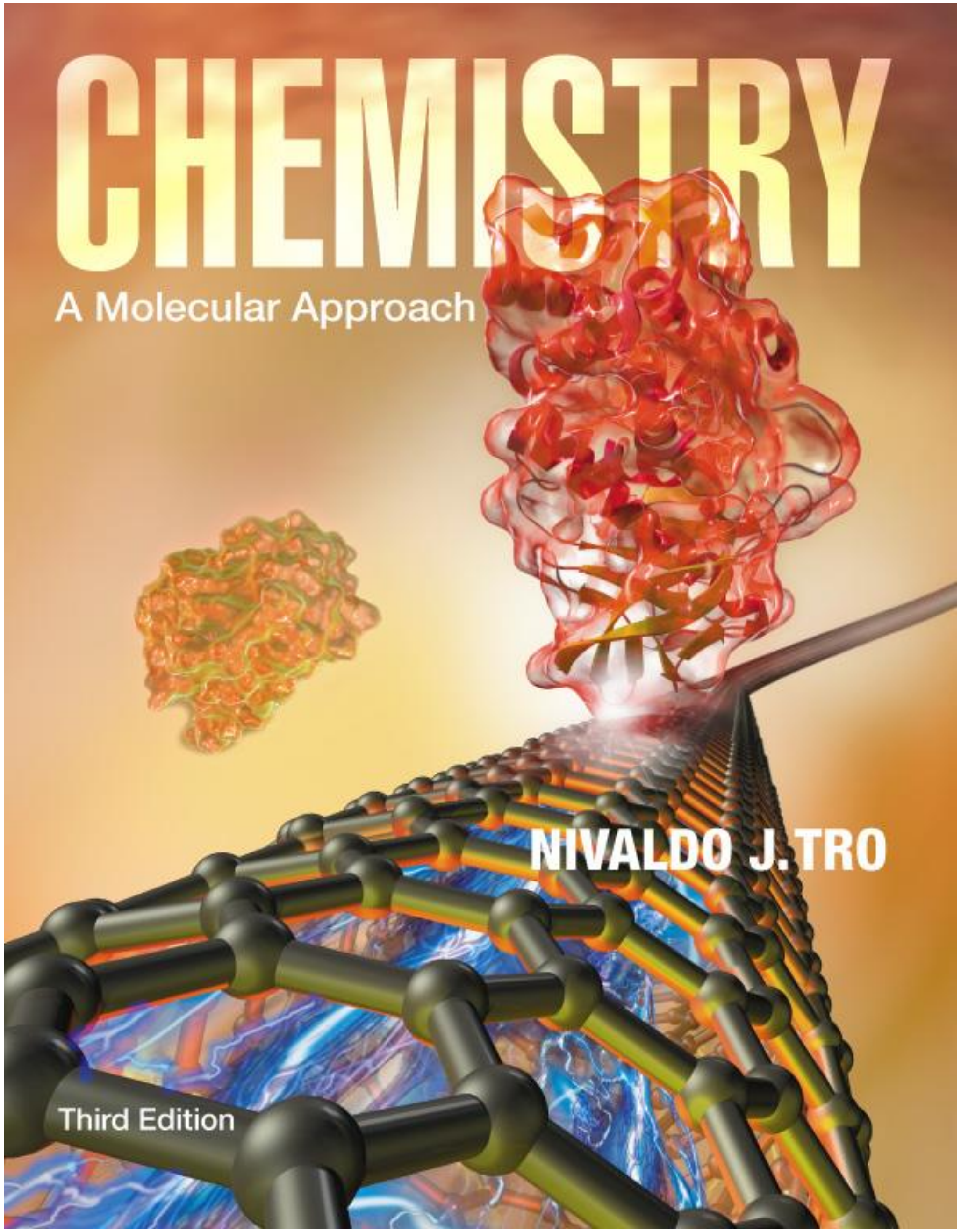


Lecture Presentation

Chapter 8

Periodic Properties of the Element

Sherril Soman
Grand Valley State University

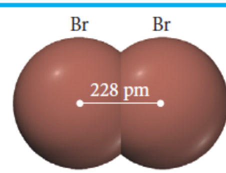
Pg. 373-374

Key Equations and Relationships

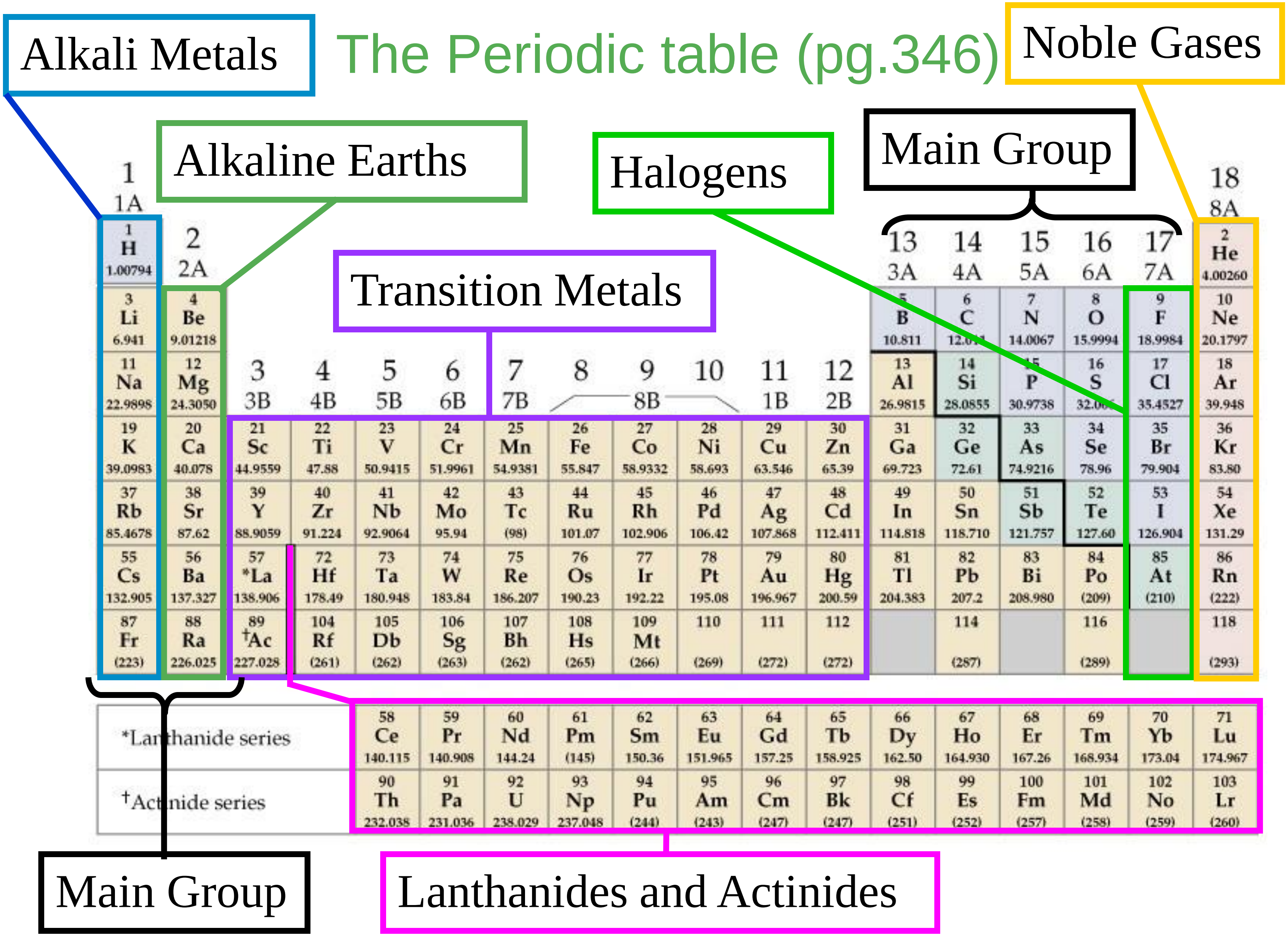
Order of Filling Quantum-Mechanical Orbitals (8.3)

$$1s\ 2s\ 2p\ 3s\ 3p\ 4s\ 3d\ 4p\ 5s\ 4d\ 5p\ 6s$$

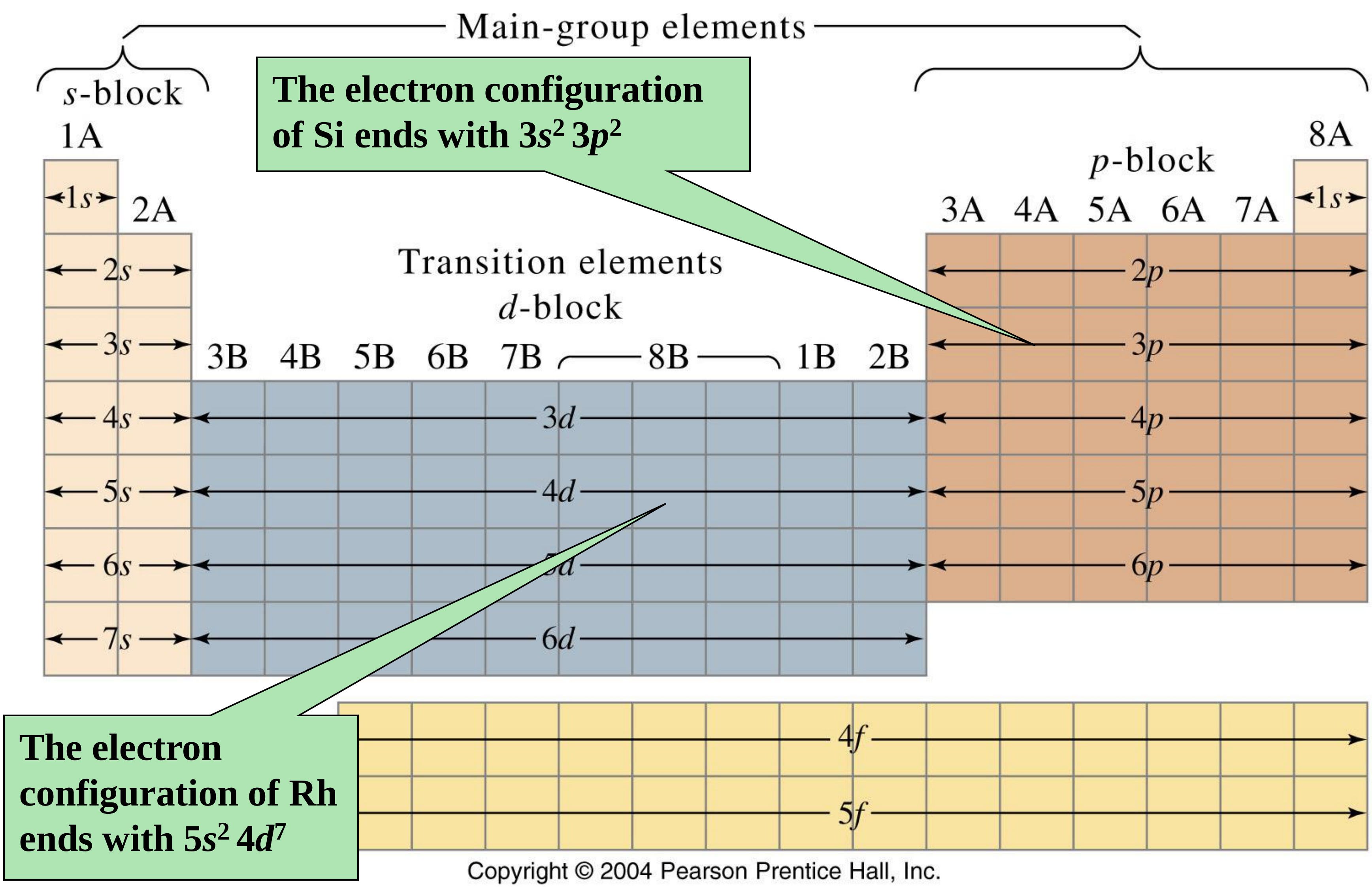
Key Learning Outcomes

Chapter Objectives	Assessment
Writing Electron Configurations (**8.3**)	Example 8.1 For Practice 8.1 Exercises 41, 42
Writing Orbital Diagrams (**8.3**) 	Example 8.2 For Practice 8.2 Exercises 43, 44
Valence Electrons and Core Electrons (**8.4**)	Example 8.3 For Practice 8.3 Exercises 51, 52
Electron Configurations from the Periodic Table (**8.4**)	Example 8.4 For Practice 8.4 For More Practice 8.4 Exercises 45, 46
Using Periodic Trends to Predict Atomic Size (**8.6**) 	Example 8.5 For Practice 8.5 For More Practice 8.5 Exercises 61–64
Writing Electron Configurations for Ions (**8.7**)	Example 8.6 For Practice 8.6 Exercises 65, 66
Using Periodic Trends to Predict Ion Size (**8.7**)	Example 8.7 For Practice 8.7 For More Practice 8.7 Exercises 69–72
Using Periodic Trends to Predict Relative Ionization Energies (**8.7**)	Example 8.8 For Practice 8.8 For More Practice 8.8 Exercises 73–76
Predicting Metallic Character Based on Periodic Trends (**8.8**) 	Example 8.9 For Practice 8.9 For More Practice 8.9 Exercises 81–84
Writing Reactions for Alkali Metal and Halogen Reactions (**8.9**)	Example 8.10 For Practice 8.10 Exercises 85–90

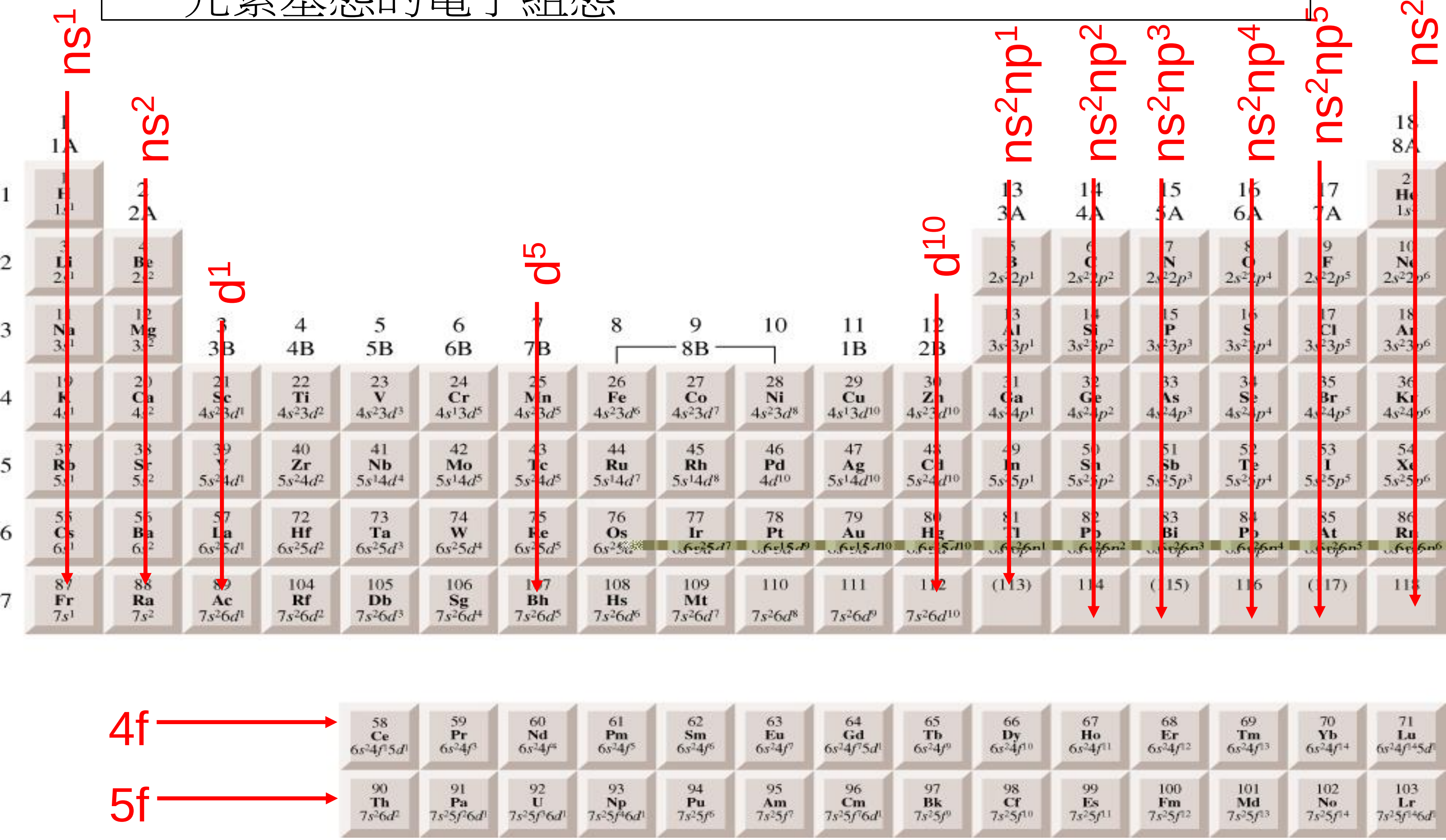
The Periodic table (pg.346)



Using the Periodic Table to Write Electron Configurations

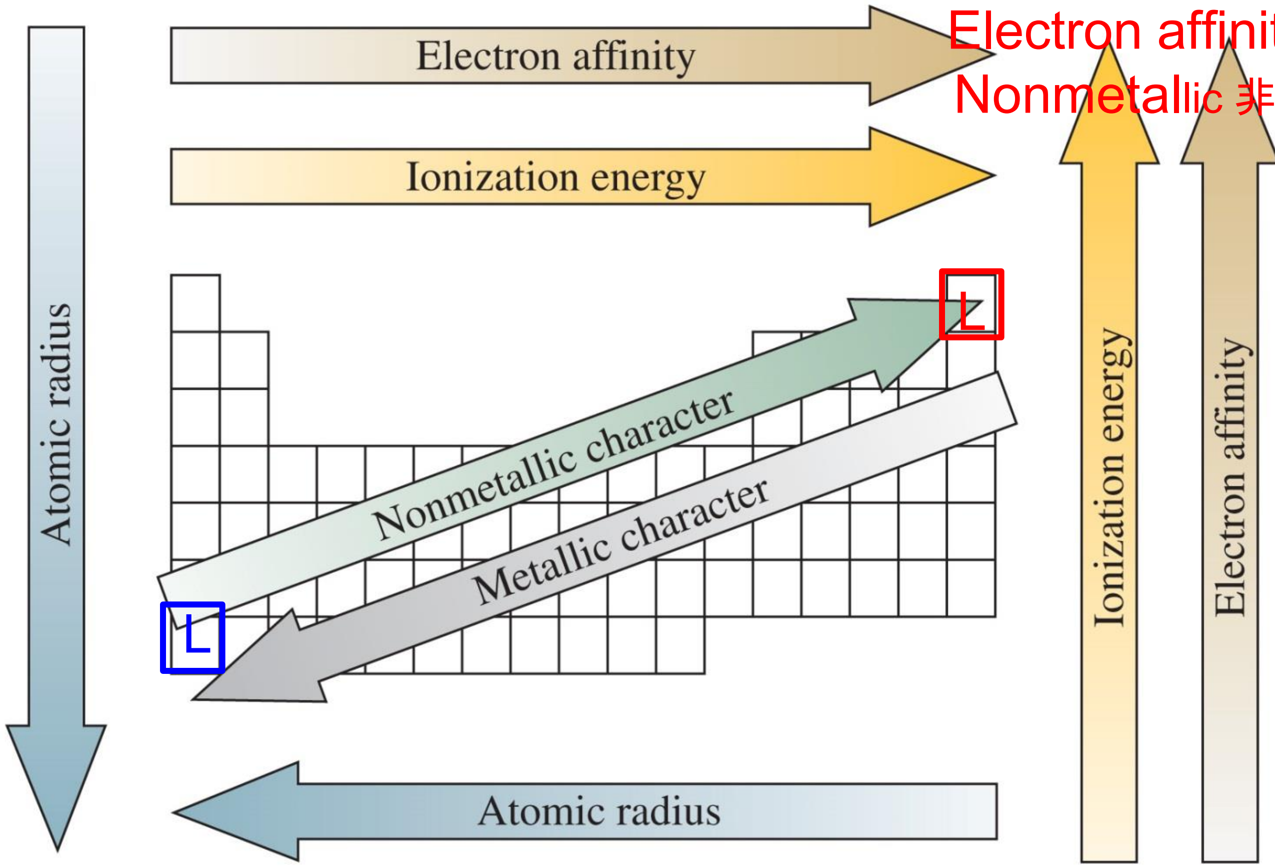


Ground State Electron Configurations of the Elements
元素基態的電子組態



Periodic Properties of the Elements 元素的週期性質

Large: Ionization energy游離能
Electron affinity電子親和力
Nonmetallic 非金屬



Large: atomic radius（原子半徑）
Metallic (金屬)

FIGURE 9-12 **Atomic properties and the periodic table – a summary**

Example 8.1 Electron Configurations

Write electron configurations for each element.

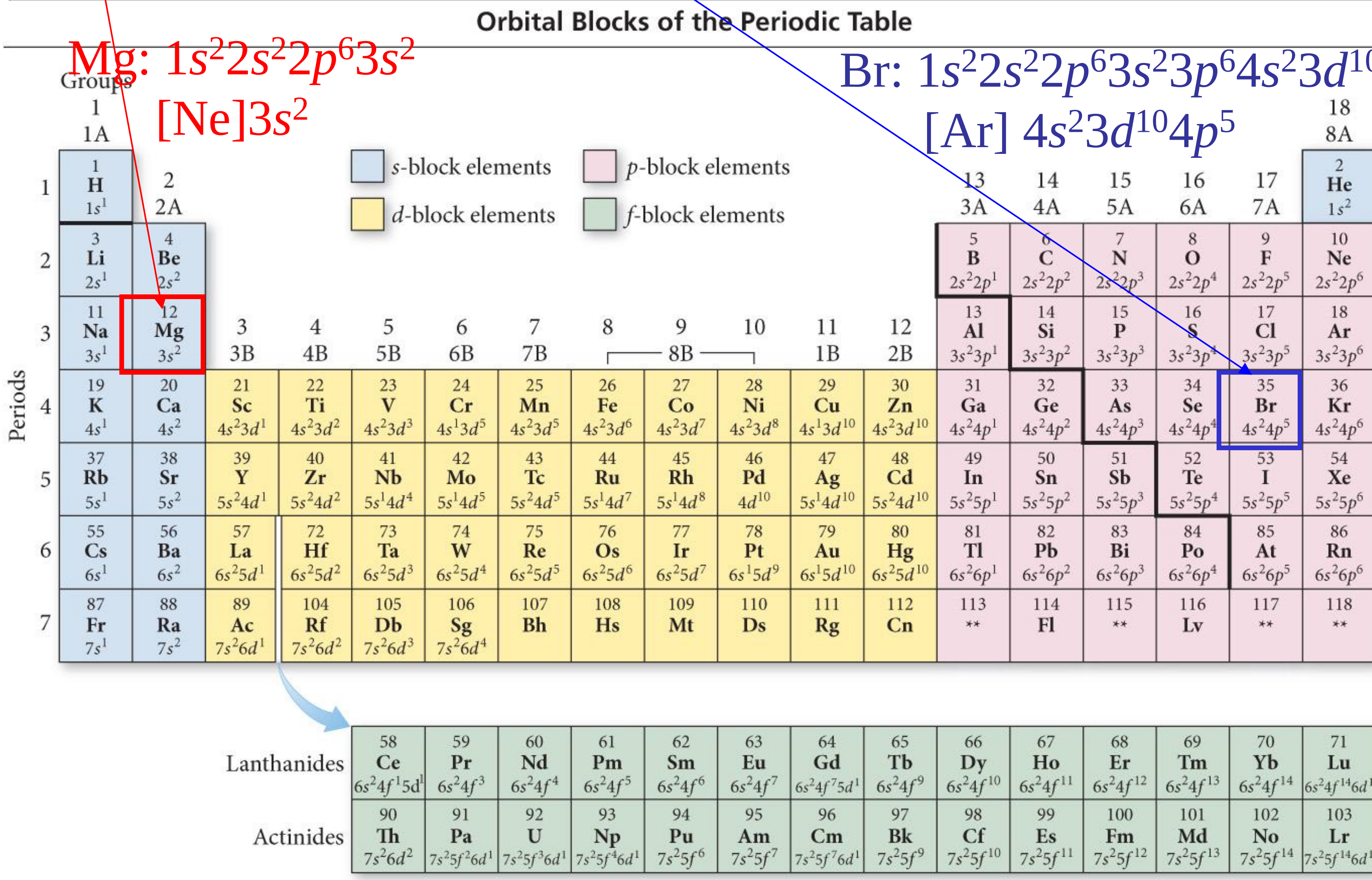
a. Mg **b.** P **c.** Br **d.** Al

Mg: $1s^22s^22p^63s^2$

$[Ne]3s^2$

Br: $1s^22s^22p^63s^23p^64s^23d^{10}4p^5$

$[Ar]\ 4s^23d^{10}4p^5$



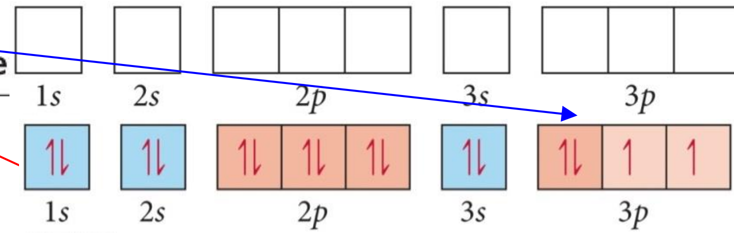
Example 8.2 Writing Orbital Diagrams 未成對電子＝2

Write the orbital diagram for sulfur and determine the number of unpaired electrons.

Solution

S: $1s^22s^22p^63s^23p^4$

$[Ne]3s^23p^4$



Orbital Blocks of the

s-block elements · p-block elements · d-block elements · f-block elements

Periods	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
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	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 **	114 Fl	115 **	116 Lv	117 **	118 **

Lanthanides	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 $6s^24f^{14}6d^1$
Actinides	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 $7s^25f^{14}6d^1$

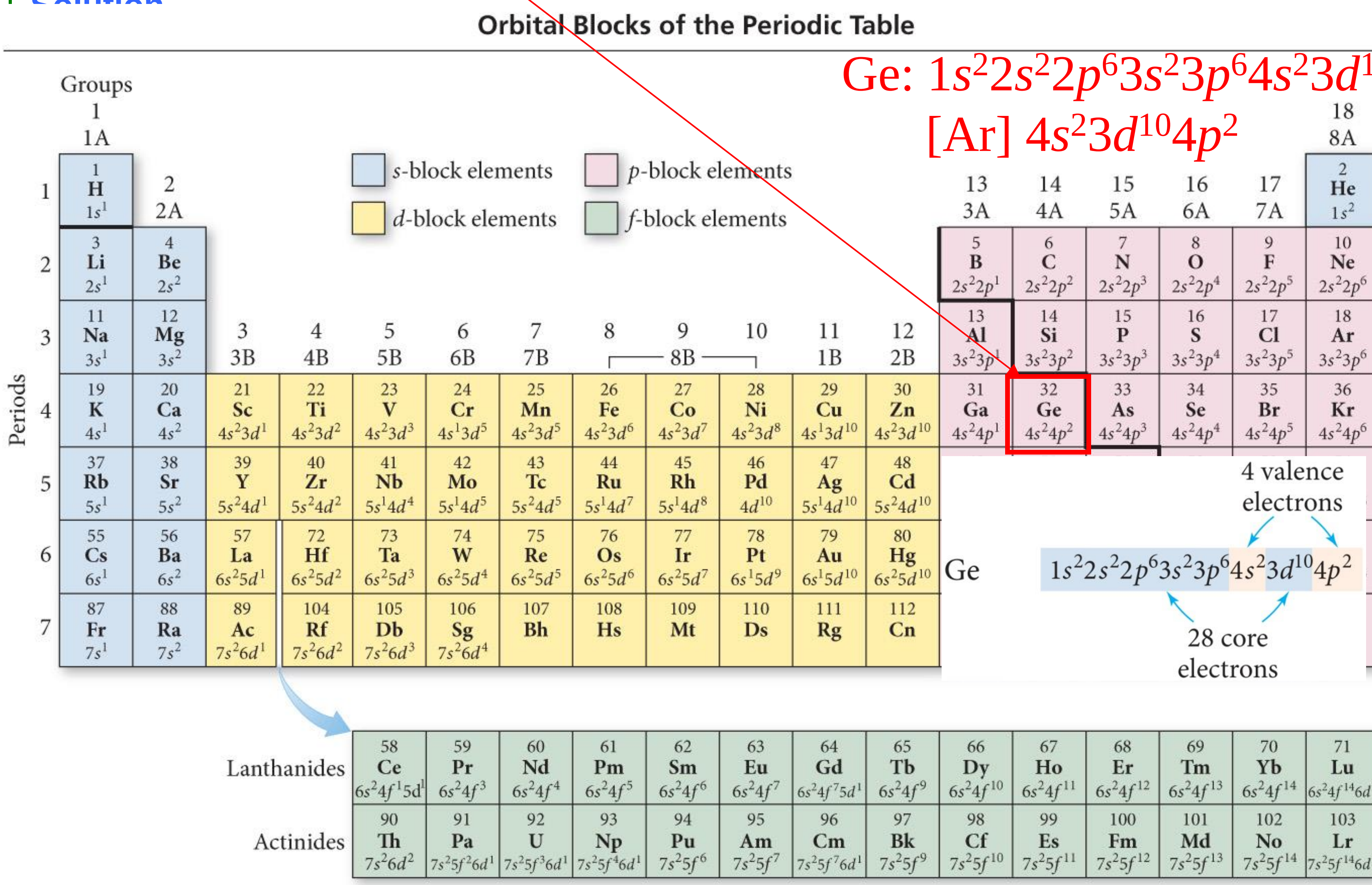
Example 8.3 Valence Electrons and Core Electrons

Write the electron configuration for Ge. Identify the valence electrons and the core electrons.

Solution

Ge: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$

$[Ar]\ 4s^2 3d^{10} 4p^2$

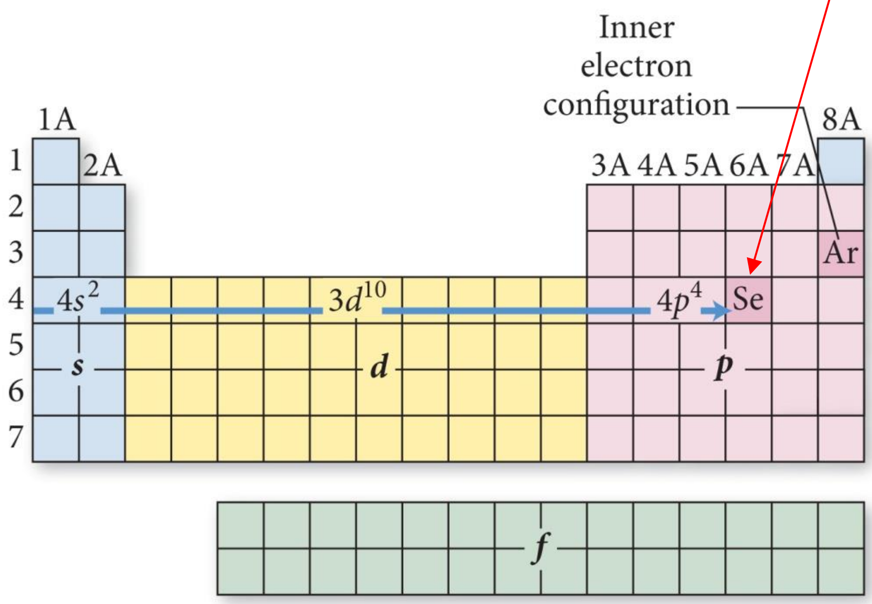


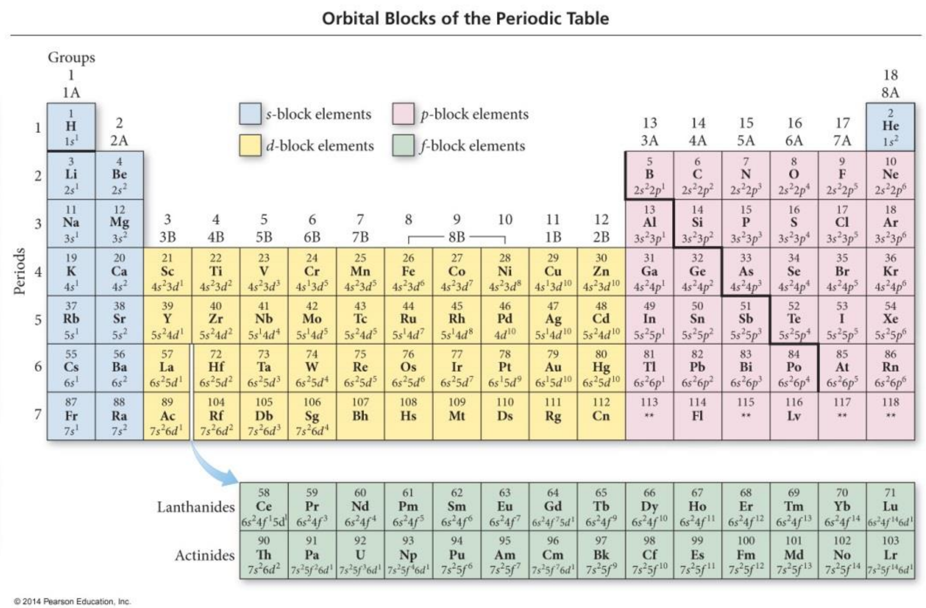
Example 8.4 Writing Electron Configurations from the Periodic Table

Use the periodic table to write the electron configuration for selenium (Se).

Solution

The atomic number of Se is 34. The noble gas that precedes Se in the periodic table is argon, so the inner electron configuration is [Ar]. Obtain the outer electron configuration by tracing the elements between Ar and Se and assigning electrons to the appropriate orbitals. Begin with [Ar]. Because Se is in row 4, add two 4*s* electrons as you trace across the *s* block (*n* = row number). Next, add ten 3*d* electrons as you trace across the *d* block (*n* = row number − 1). Lastly, add four 4*p* electrons as you trace across the *p* block to Se, which is in the fourth column of the *p* block (*n* = row number).





Se [Ar] $4s^2\,3d^{10}\,4p^4$

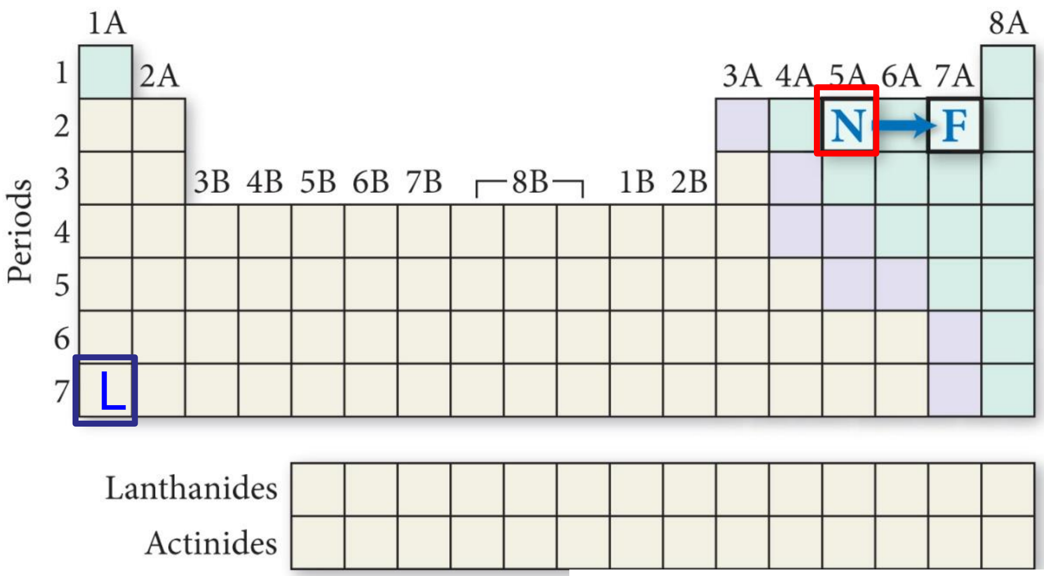
Example 8.5 Atomic Size

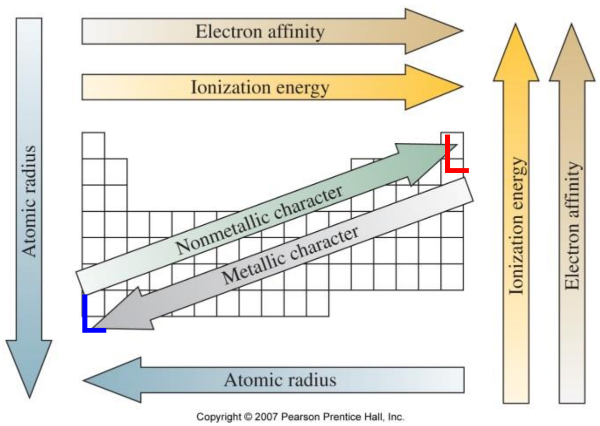
On the basis of periodic trends, choose the larger atom in each pair (if possible). Explain your choices.

a. N or F **b.** C or Ge **c.** N or Al **d.** Al or Ge

Solution

a. N atoms are larger than F atoms because as you trace the path between N and F on the periodic table, you move to the right within the same period. As you move to the right across a period, the effective nuclear charge experienced by the outermost electrons increases, resulting in a smaller radius.



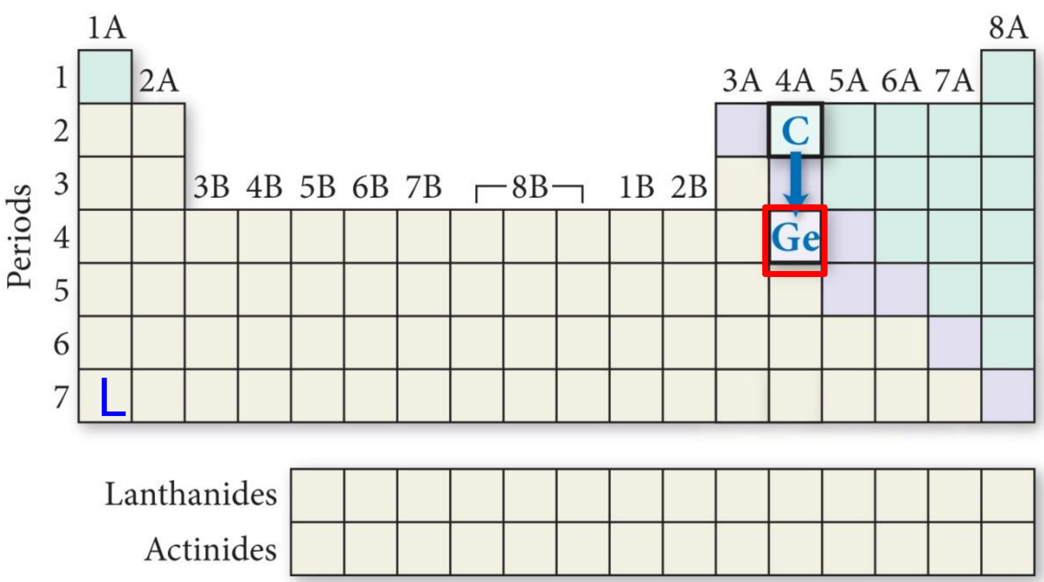


Large: atomic radius（原子半徑）
Metallic (金屬)

Example 8.5 Atomic Size

Continued

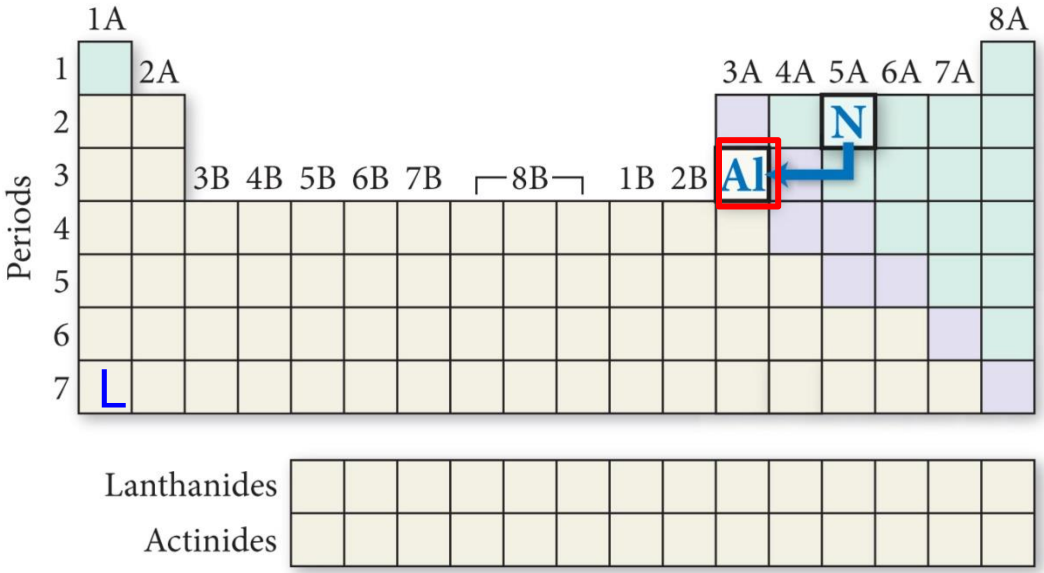
b. Ge atoms are larger than C atoms because as you trace the path between C and Ge on the periodic table, you move down a column. Atomic size increases as you move down a column because the outermost electrons occupy orbitals with a higher principal quantum number that are therefore larger, resulting in a larger atom.



Example 8.5 Atomic Size

Continued

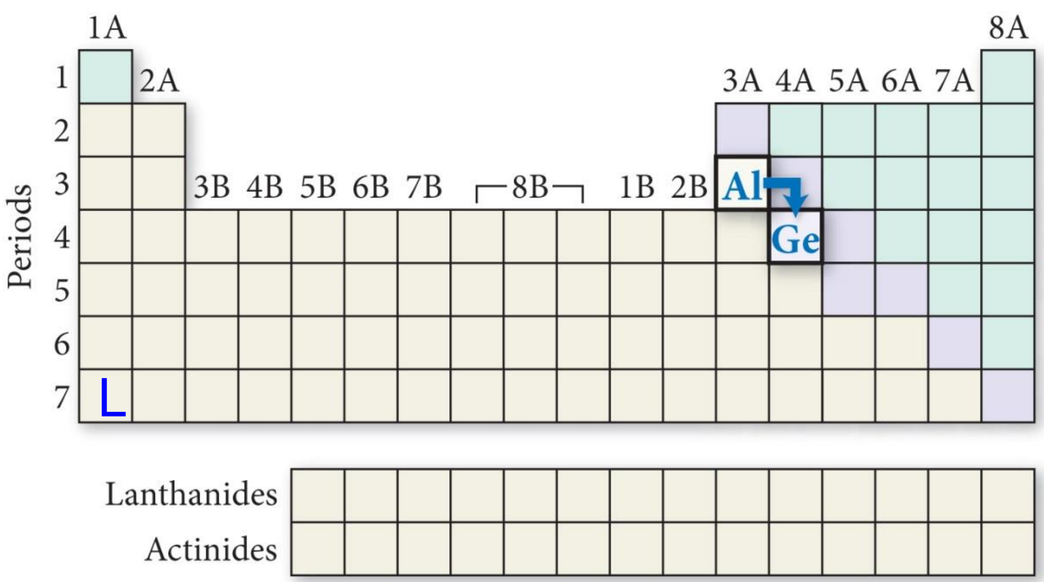
c. Al atoms are larger than N atoms because as you trace the path between N and Al on the periodic table, you move down a column (atomic size increases) and then to the left across a period (atomic size increases). These effects add together for an overall increase.



Example 8.5 Atomic Size

Continued

d. Based on periodic trends alone, you cannot tell which atom is larger, because as you trace the path between Al and Ge you go to the right across a period (atomic size decreases) and then down a column (atomic size increases). These effects tend to counter each other, and it is not easy to tell which will predominate.



無法判斷

Example 8.6 Electron Configurations and Magnetic Properties for Ions

Write the electron configuration and orbital diagram for each ion and determine whether each is diamagnetic or paramagnetic.

a. Al^{3+} **b.** S^{2-} **c.** Fe^{3+}

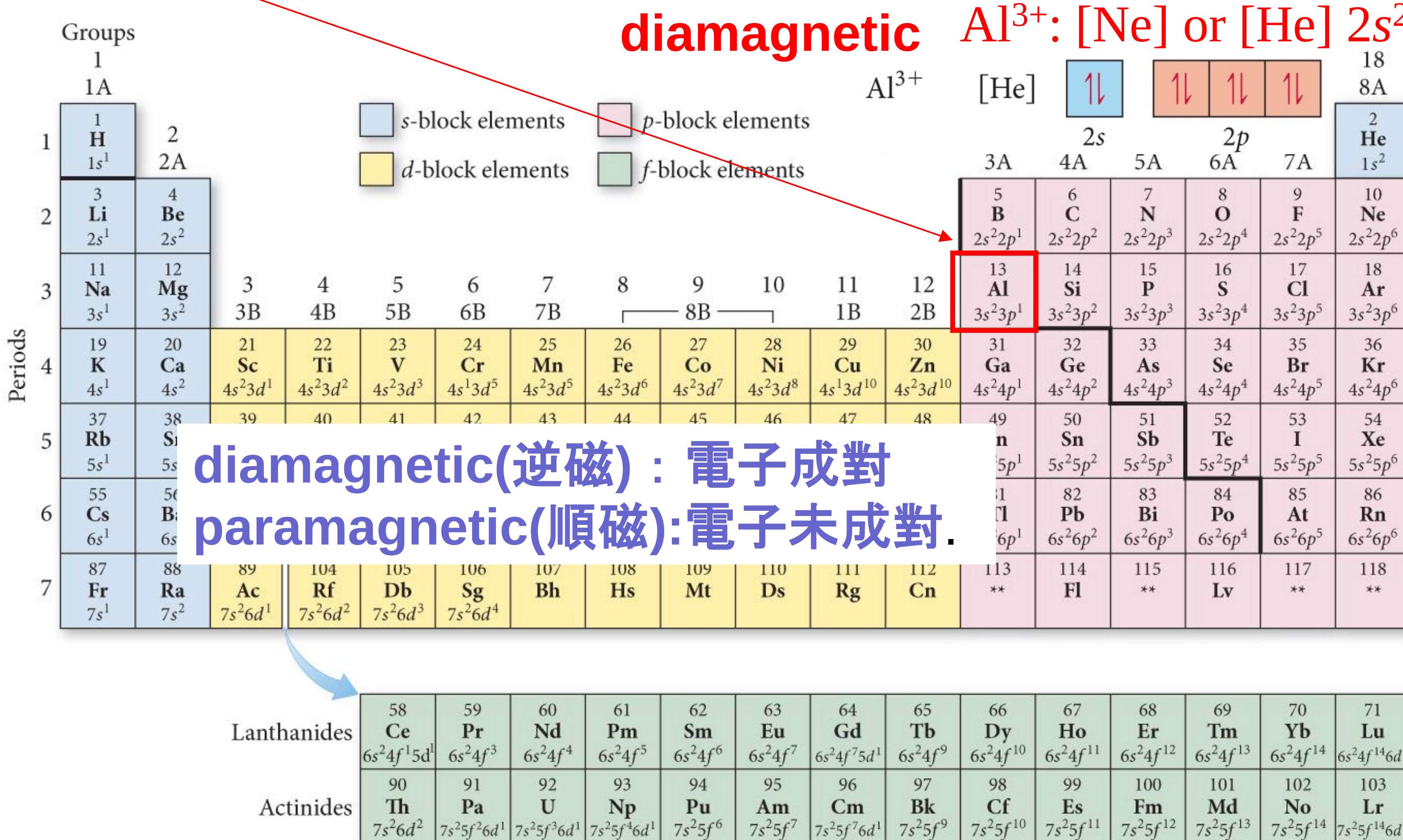
Al: $1s^22s^22p^63s^23p^1$

[Ne] $3s^23p^1$

Al^{3+}: [Ne] or [He] $2s^22p^6$

diamagnetic

Orbital Blocks of the Periodic Table



diamagnetic(逆磁)：電子成對

paramagnetic(順磁):電子未成對.

Example 8.6 Electron Configurations and Magnetic Properties for Ions

Continued

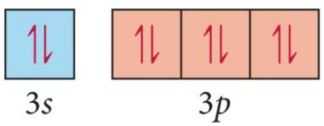
b. S^{2-}

Begin by writing the electron configuration of the neutral atom. Since this ion has a 2– charge, add two electrons to write the electron configuration of the ion. Write the orbital diagram by drawing half-arrows to represent each electron in boxes representing the orbitals. Because there are no unpaired electrons, S^{2-} is diamagnetic.

S [Ne] $3s^2\ 3p^4$

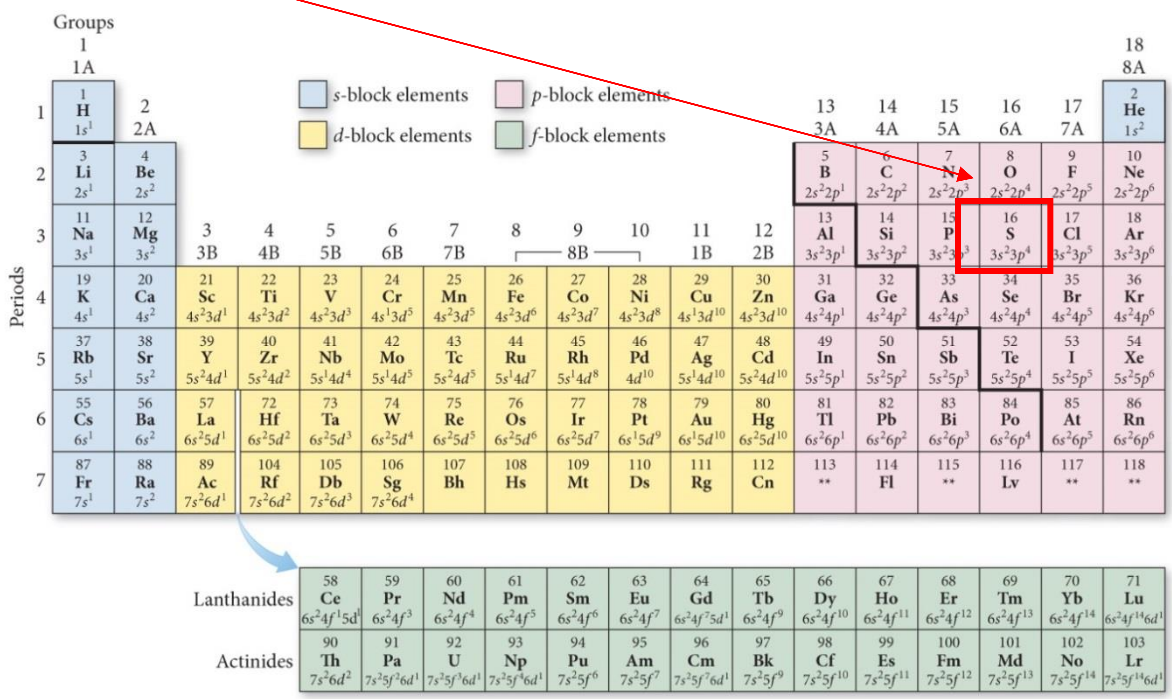
S^{2-} [Ne] $3s^2\ 3p^6$

S^{2-} [Ne]



Diamagnetic

Orbital Blocks of the Periodic Table



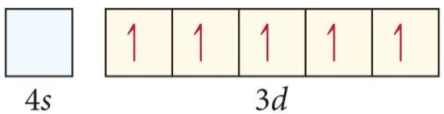
Example 8.6 Electron Configurations and Magnetic Properties for Ions

Continued

c. Fe^{3+}

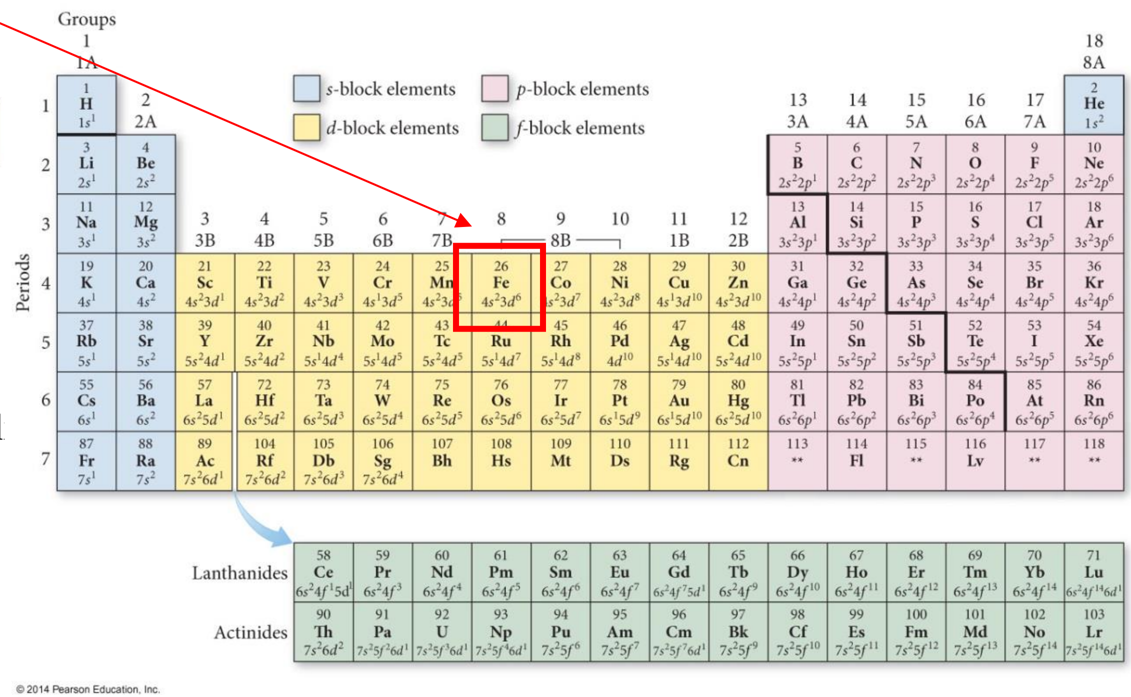
Begin by writing the electron configuration of the neutral atom. Since this ion has a 3+ charge, remove three electrons to write the electron configuration of the ion. Since it is a transition metal, remove the electrons from the 4*s* orbital before removing electrons from the 3*d* orbitals. Write the orbital diagram by drawing half-arrows to represent each electron in boxes representing the orbitals. There are unpaired electrons, so Fe^{3+} is paramagnetic.

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

Fe^{3+} [Ar] 

Paramagnetic

Orbital Blocks of the Periodic Table



For Practice 8.6

Write the electron configuration and orbital d
or diamagnetic.

a. Co^{2+} **b.** N^{3-}

Example 8.7 Ion Size

陰離子>中性原子>陽離子

Choose the larger atom or ion from each pair.

a. S or $\boxed{S^{2-}}$ **b.** $\boxed{Ca}$ or Ca^{2+} **c.** $\boxed{Br^{-}}$ or Kr

Solution

a. The S^{2-} ion is larger than an S atom because anions are larger than the atoms from which they are formed.

b. A Ca atom is larger than Ca^{2+} because cations are smaller than the atoms from which they are formed.

c. A Br^{-} ion is larger than a Kr atom because, although they are isoelectronic, Br^{-} has one fewer proton than Kr, resulting in a lesser pull on the electrons and therefore a larger radius.

For Practice 8.7

Choose the larger atom or ion from each pair.

a. K or K^{+} **b.** F or F^{-} **c.** Ca^{2+} or Cl^{-}

For More Practice 8.7

Arrange the following in order of decreasing radius: Ca^{2+}, Ar, Cl^{-}.

Example 8.8 Ionization Energy

On the basis of periodic trends, determine which element in each pair has the higher first ionization energy (if possible).

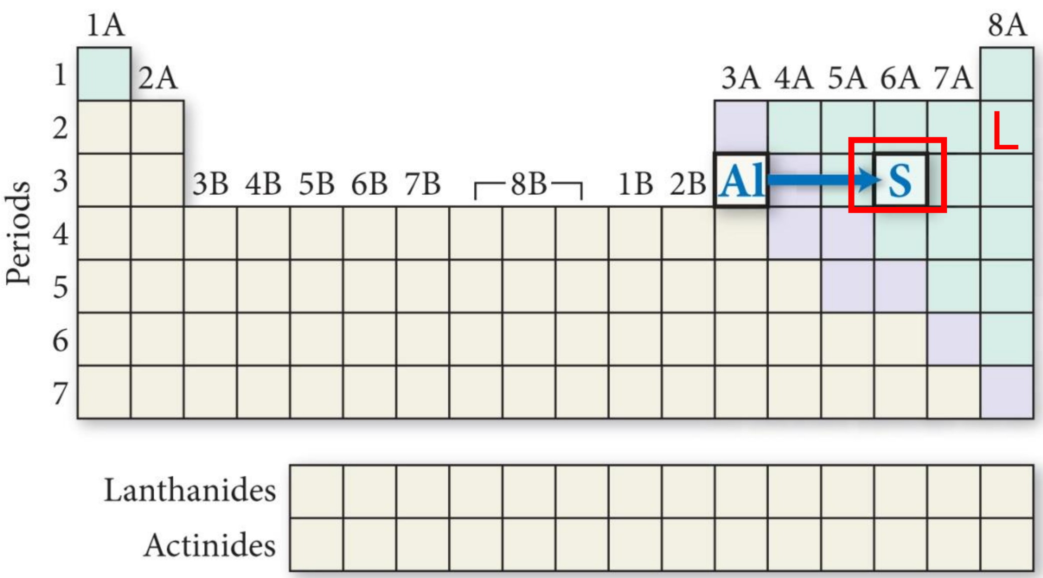
a. Al or S
b. As or Sb
c. N or Si
d. O or Cl

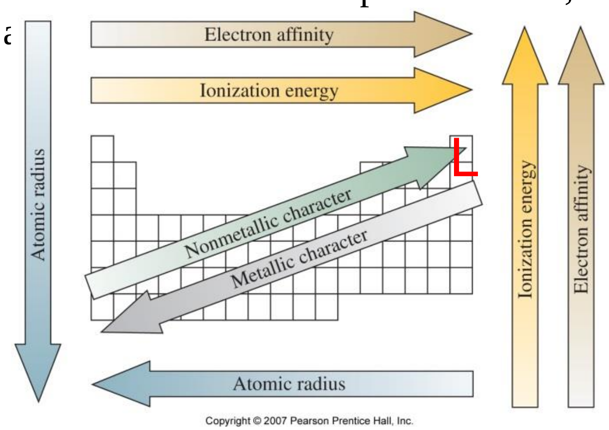
Solution

Ionization energy游離

a. Al or S

S has a higher ionization energy than Al because as you trace the path between Al and S on the periodic table, you move to the right within the same row. Ionization energy increases increasing effective nuclear charge.



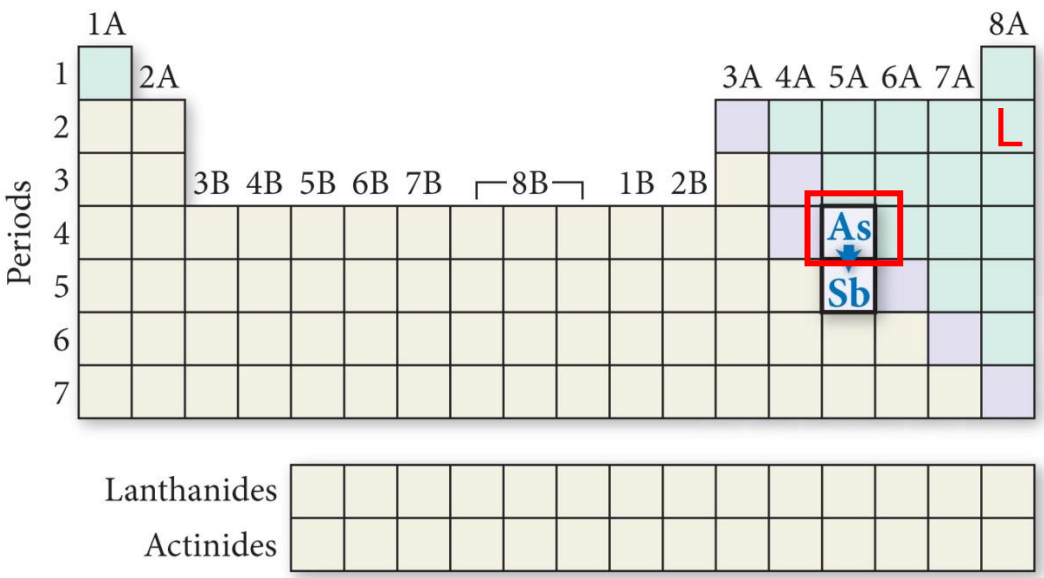


Example 8.8 Ionization Energy

Continued

b. As or Sb

As has a higher ionization energy than Sb because as you trace the path between As and Sb on the periodic table, you move down a column. Ionization energy decreases as you go down a column as a result of the increasing size of orbitals with increasing *n*.

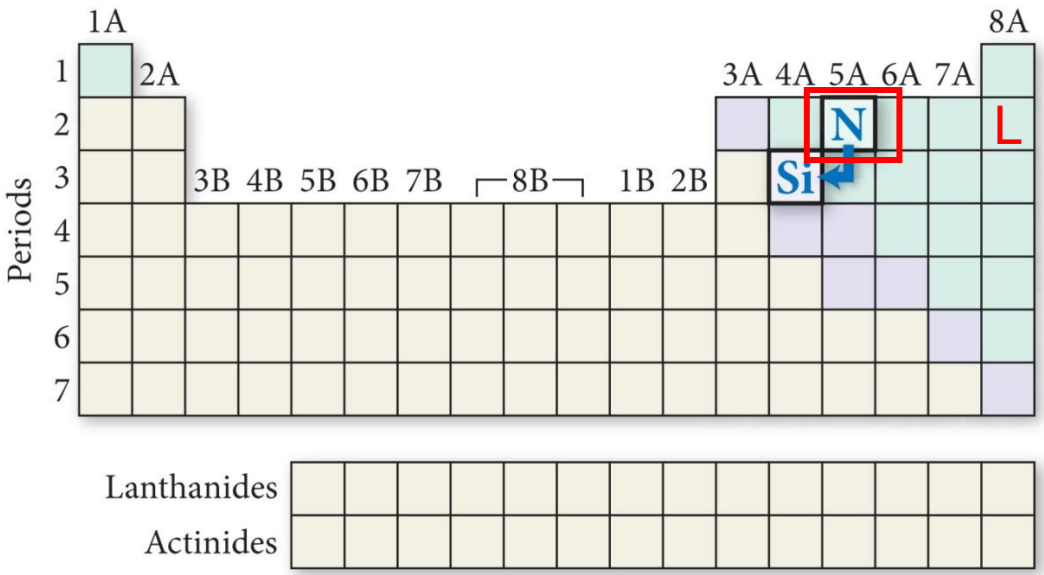


Example 8.8 Ionization Energy

Continued

c. N or Si

N has a higher ionization energy than Si because as you trace the path between N and Si on the periodic table, you move down a column (ionization energy decreases) and then to the left across a row (ionization energy decreases). These effects sum together for an overall decrease.

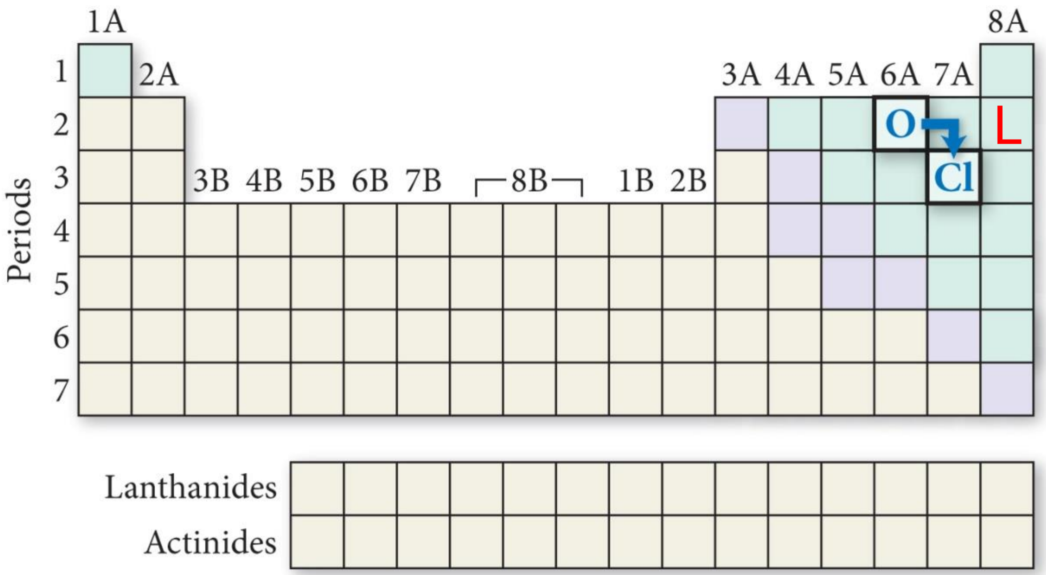


Example 8.8 Ionization Energy

Continued

d. O or Cl

Based on periodic trends alone, it is impossible to tell which has a higher ionization energy because, as you trace the path between O and Cl, you go to the right across a row (ionization energy increases) and then down a column (ionization energy decreases). These effects tend to counter each other, and it is not obvious which will dominate.



無法判斷

Example 8.9 Metallic Character

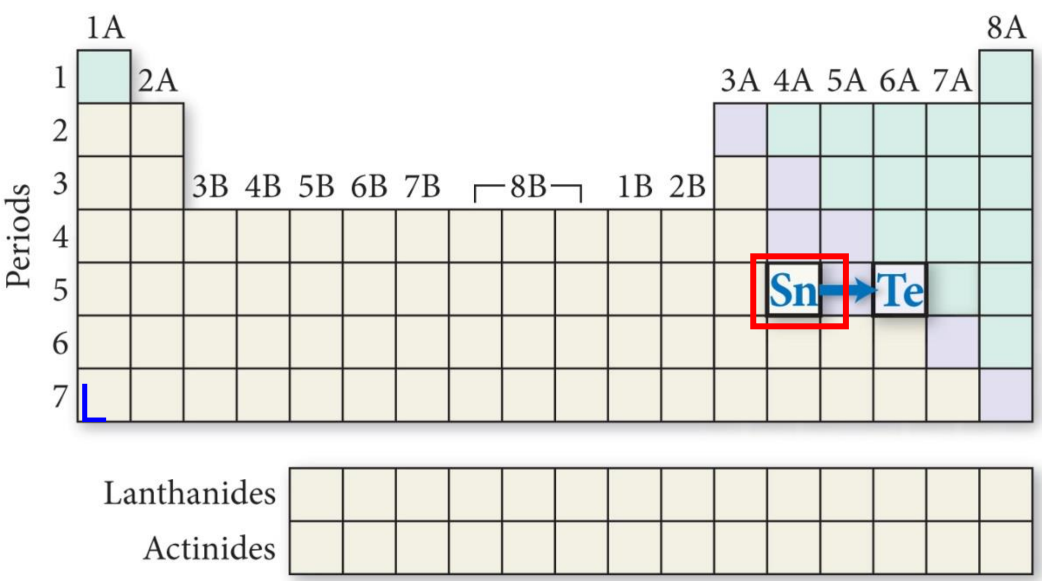
On the basis of periodic trends, choose the more metallic element from each pair (if possible).

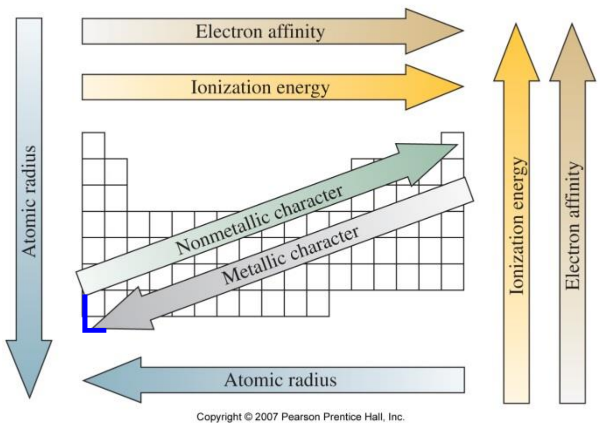
a. Sn or Te **b.** P or Sb **c.** Ge or In **d.** S or Br

Solution

a. Sn or Te

Sn is more metallic than Te because as you trace the path between Sn and Te on the periodic table, you move to the right within the same period. Metallic character decreases as we go to the right.





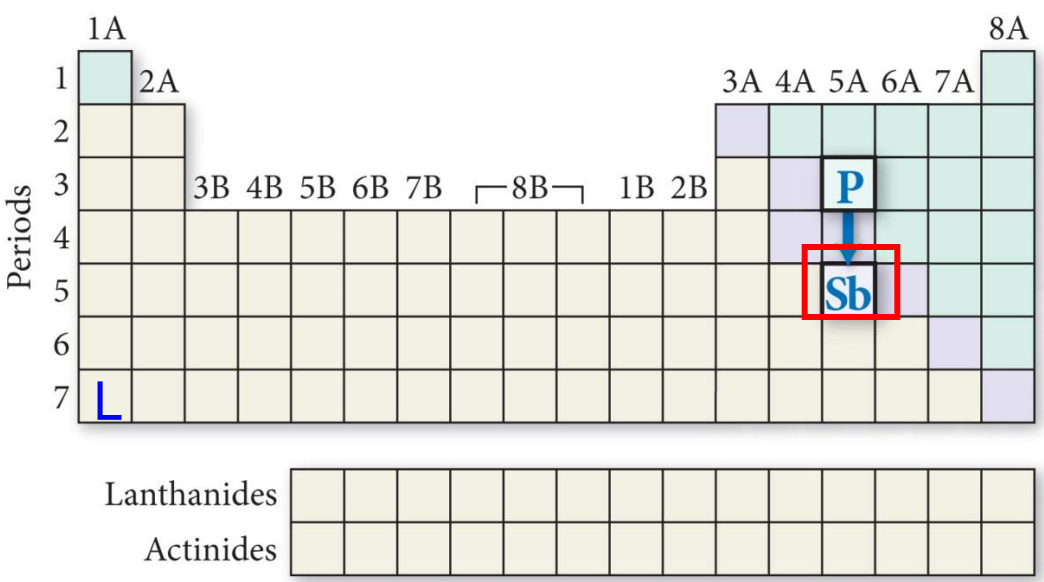
Metallic (金屬)

Example 8.9 Metallic Character

Continued

b. P or Sb

Sb is more metallic than P because as you trace the path between P and Sb on the periodic table, you move down a column. Metallic character increases as we go down a column.

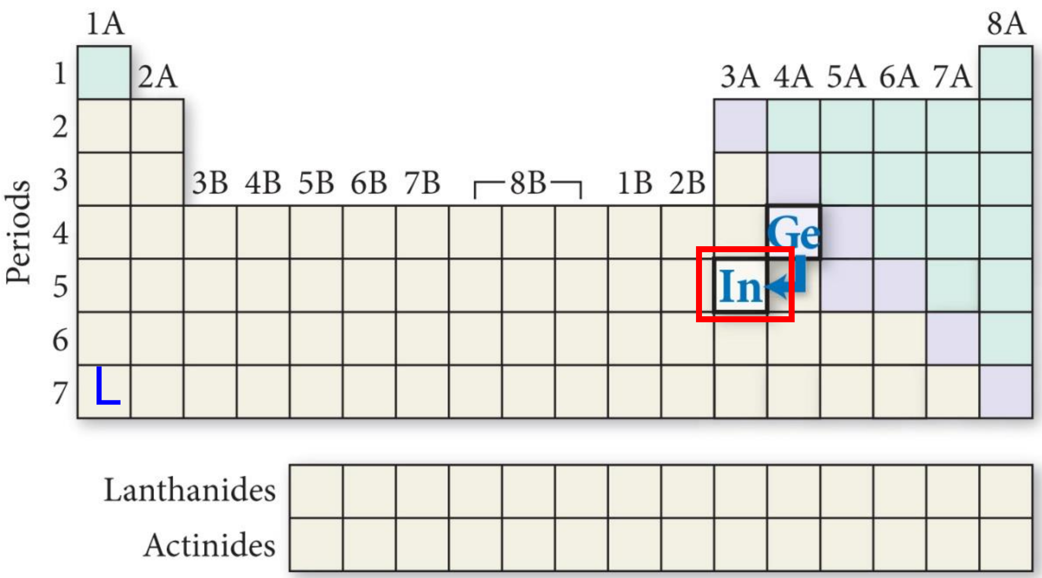


Example 8.9 Metallic Character

Continued

c. Ge or In

In is more metallic than Ge because as you trace the path between Ge and In on the periodic table, you move down a column (metallic character increases) and then to the left across a period (metallic character increases). These effects add together for an overall increase.

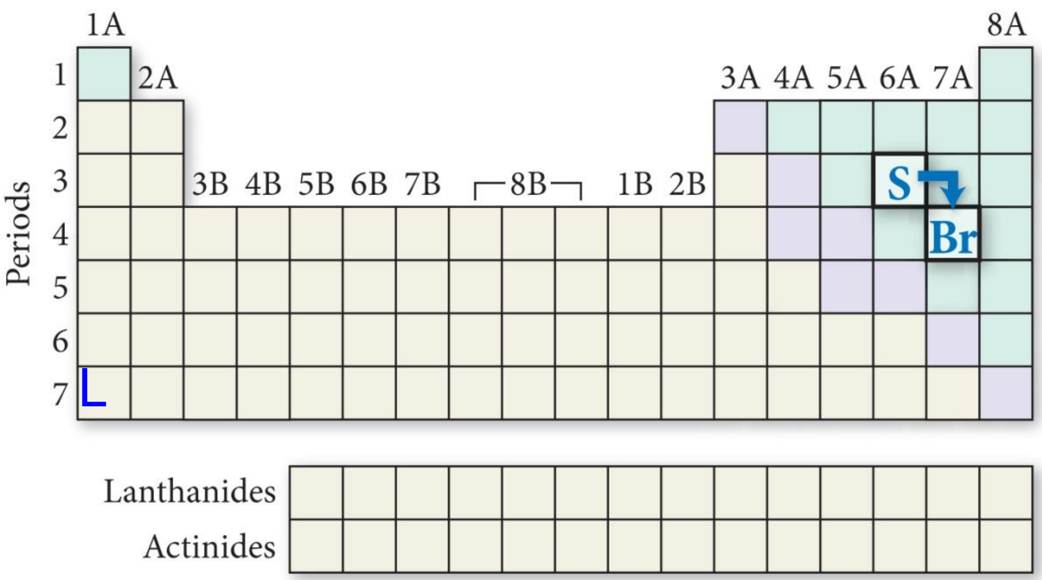


Example 8.9 Metallic Character

Continued

d. S or Br

Based on periodic trends alone, we cannot tell which is more metallic because as you trace the path between S and Br, you go to the right across a period (metallic character decreases) and then down a column (metallic character increases). These effects tend to counter each other, and it is not easy to tell which will predominate.



無法判斷

The Periodic table (pg.346)

Alkali Metals

Alkaline Earths

Transition Metals

Halogens

Main Group

Noble Gases

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
1 **H** 1.00794																	2 **He** 4.00260
3 **Li** 6.941	4 **Be** 9.01218											5 **B** 10.811	6 **C** 12.011	7 **N** 14.0067	8 **O** 15.9994	9 **F** 18.9984	10 **Ne** 20.1797
11 **Na** 22.9898	12 **Mg** 24.3050											13 **Al** 26.9815	14 **Si** 28.0855	15 **P** 30.9738	16 **S** 32.06	17 **Cl** 35.4527	18 **Ar** 39.948
19 **K** 39.0983	20 **Ca** 40.078	21 **Sc** 44.9559	22 **Ti** 47.88	23 **V** 50.9415	24 **Cr** 51.9961	25 **Mn** 54.9381	26 **Fe** 55.847	27 **Co** 58.9332	28 **Ni** 58.693	29 **Cu** 63.546	30 **Zn** 65.39	31 **Ga** 69.723	32 **Ge** 72.61	33 **As** 74.9216	34 **Se** 78.96	35 **Br** 79.904	36 **Kr** 83.80
37 **Rb** 85.4678	38 **Sr** 87.62	39 **Y** 88.9059	40 **Zr** 91.224	41 **Nb** 92.9064	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.906	46 **Pd** 106.42	47 **Ag** 107.868	48 **Cd** 112.411	49 **In** 114.818	50 **Sn** 118.710	51 **Sb** 121.757	52 **Te** 127.60	53 **I** 126.904	54 **Xe** 131.29
55 **Cs** 132.905	56 **Ba** 137.327	57 ***La** 138.906	72 **Hf** 178.49	73 **Ta** 180.948	74 **W** 183.84	75 **Re** 186.207	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.967	80 **Hg** 200.59	81 **Tl** 204.383	82 **Pb** 207.2	83 **Bi** 208.980	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** 226.025	89 **†Ac** 227.028	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (263)	107 **Bh** (262)	108 **Hs** (265)	109 **Mt** (266)	110 (269)	111 (272)	112 (272)		114 (287)		116 (289)		118 (293)

*Lanthanide series	58 **Ce** 140.115	59 **Pr** 140.908	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.965	64 **Gd** 157.25	65 **Tb** 158.925	66 **Dy** 162.50	67 **Ho** 164.930	68 **Er** 167.26	69 **Tm** 168.934	70 **Yb** 173.04	71 **Lu** 174.967
†Actinide series	90 **Th** 232.038	91 **Pa** 231.036	92 **U** 238.029	93 **Np** 237.048	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (260)

Main Group

Lanthanides and Actinides

Example 8.10 Alkali Metal and Halogen Reactions

Write a balanced chemical equation for each reaction.

a. the reaction between potassium metal and bromine gas

b. the reaction between rubidium metal and liquid water

c. the reaction between gaseous chlorine and solid iodine

Solution

a. Alkali metals react with halogens to form metal halides. Write the formulas for the reactants and the metal halide product (making sure to write the correct ionic chemical formula for the metal halide, as outlined in Section 3.5), and then balance the equation.

$$2\ K(s) + Br_2(g) \rightarrow 2\ KBr(s)$$

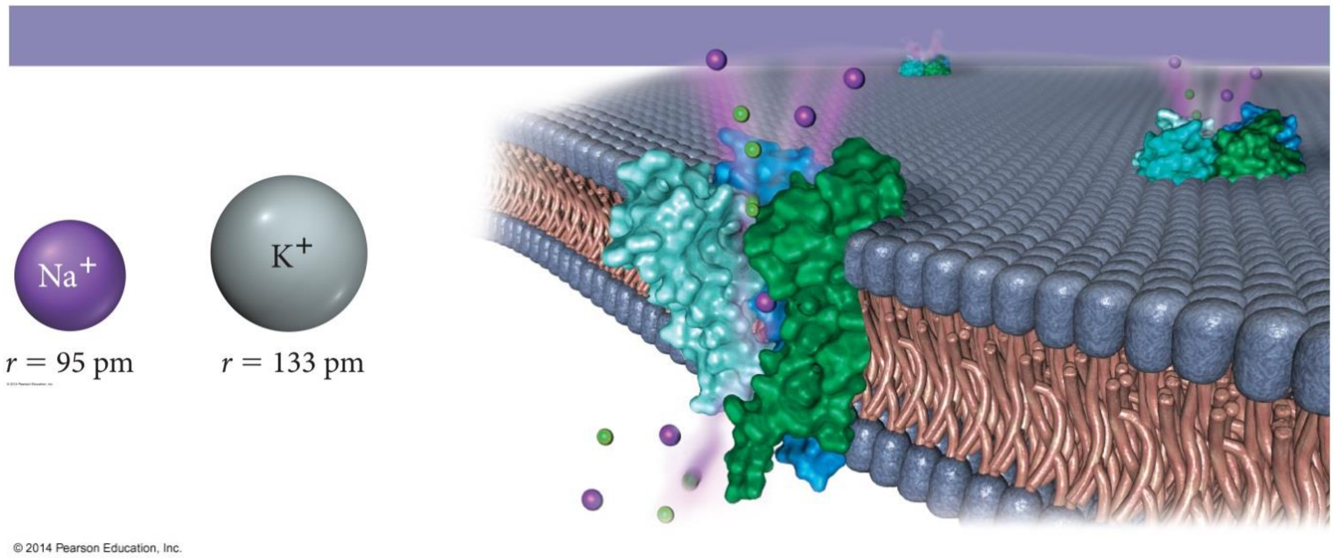
b. Alkali metals react with water to form the dissolved metal ion, the hydroxide ion, and hydrogen gas. Write the skeletal equation including each of these and then balance it.

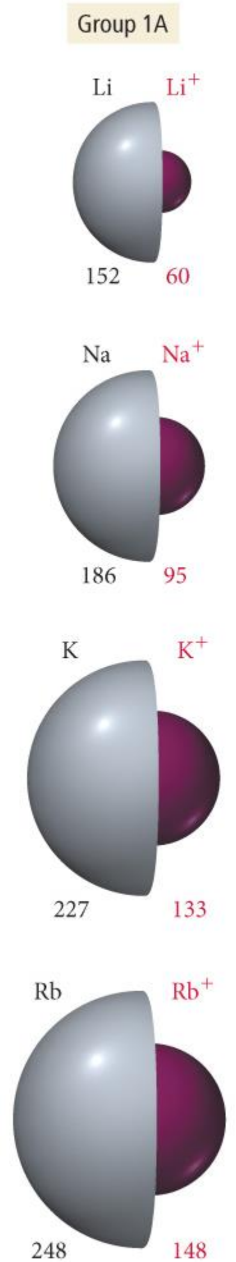
$$2\ Rb(s) + 2\ H_2O(l) \rightarrow 2\ Rb+(aq) + 2\ OH^-(aq) + H_2(g)$$

c. Halogens react with each other to form interhalogen compounds. Write the skeletal equation with each of the halogens as the reactants and the interhalogen compound as the product and balance the equation.

$$Cl_2(g) + I_2(s) \rightarrow 2\ ICl(g)$$

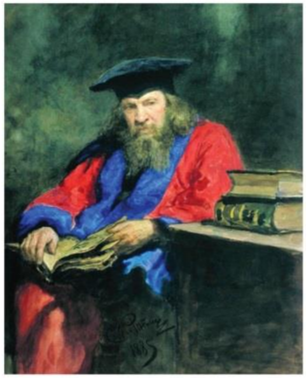
7.1 Nerve Transmission





- Movement of ions across cell membranes is the basis for the transmission of nerve signals.
- Na^+ and K^+ ions are pumped across membranes in opposite directions through ion channels.
  - Na^+ out and K^+ in
- The ion channels can differentiate Na^+ from K^+ by their difference in size.
- Ion size and other properties of atoms are **periodic properties**—properties whose values can be predicted based on the element's position on the periodic table.

8.2 The Development of the Periodic Table週期表



- Mendeleev’ s Periodic Law allows us to predict **what** the properties of an element will be based on its position on the table.
- It doesn’ t explain why the pattern exists.
- Quantum mechanics is a theory that explains **why** the periodic trends in the properties exist.
  - Knowing *why* allows us to predict *what.*

Mendeleev (1834–1907)

Gallium (eka-aluminum)

	Mendeleev’s predicted properties	Actual properties
Atomic mass	About 68 amu	69.72 amu
Melting point	Low	29.8 °C
Density	5.9 g/cm^3	5.90 g/cm^3
Formula of oxide	X_2O_3	Ga_2O_3
Formula of chloride	XCl_3	$GaCl_3$

Germanium (eka-silicon)

	Mendeleev’s predicted properties	Actual properties
Atomic mass	About 72 amu	72.64 amu
Density	5.5 g/cm^3	5.35 g/cm^3
Formula of oxide	XO_2	GeO_2
Formula of chloride	XCl_4	$GeCl_4$

8.3 Electron Configurations: How Electrons Occupy Orbitals

- Quantum-mechanical theory describes the behavior of electrons in atoms.
- The electrons in atoms exist in orbitals.
- A description of the orbitals occupied by electrons is called an **electron configuration(電子組態)**.

H $1s^1$ Number of electrons in orbital

Orbital

Schrodinger Wave Equation $\Psi = \mathrm{fn}(n, l, m_l), \boldsymbol{m_s}$

spin quantum number(自旋量子數) , m_s.

-Not in the Schrödinger equation

-Schrödinger equation(n, l, m_l)

Quantum Numbers and Electron Orbitals

$$\psi(x, y, z) = \psi(r, \theta, \phi)$$

$$\boxed{\Psi_{n,l,m_l}(r,\theta,\phi) = R_{n,l}(r) \cdot \Theta_{l,m_l}(\theta) \cdot \Phi_{m_l}(\phi)}$$

Principle quantum number, n = 1, 2, 3…

Angular momentum quantum number,
ℓ = 0, 1, 2…(n-1)

ℓ = 0, s

ℓ = 1, p

ℓ = 2, d

ℓ = 3, f

Magnetic quantum number,
m_l = - ℓ…-2, -1, 0, 1, 2…+ℓ

The Property of Electron Spin

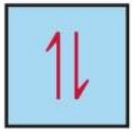
Stern–Gerlach experiment:
https://www.youtube.com/watch?v=rg4Fnag4V-E

- Spin is a fundamental property of all electrons.
- All electrons have the same amount of spin.
- The orientation of the electron spin is quantized, it can only be in one direction or its opposite.
  - Spin up or spin down
- **spin quantum number(自旋量子數) , m_s.**
- m_s can have values of +½ or +½ −½.
  - Not in the Schrödinger equation
  - Schrödinger equation(n, l, m_l)

Pauli Exclusion Principle不相容原理

- **No two electrons in an atom may have the same set of four quantum numbers.**
- Therefore, no orbital may have more than two electrons, and they must have opposite spins.

Allowed Quantum Numbers

	Electron configuration	Orbital diagram
He	$1s^2$	 $1s$

Helium has two electrons.

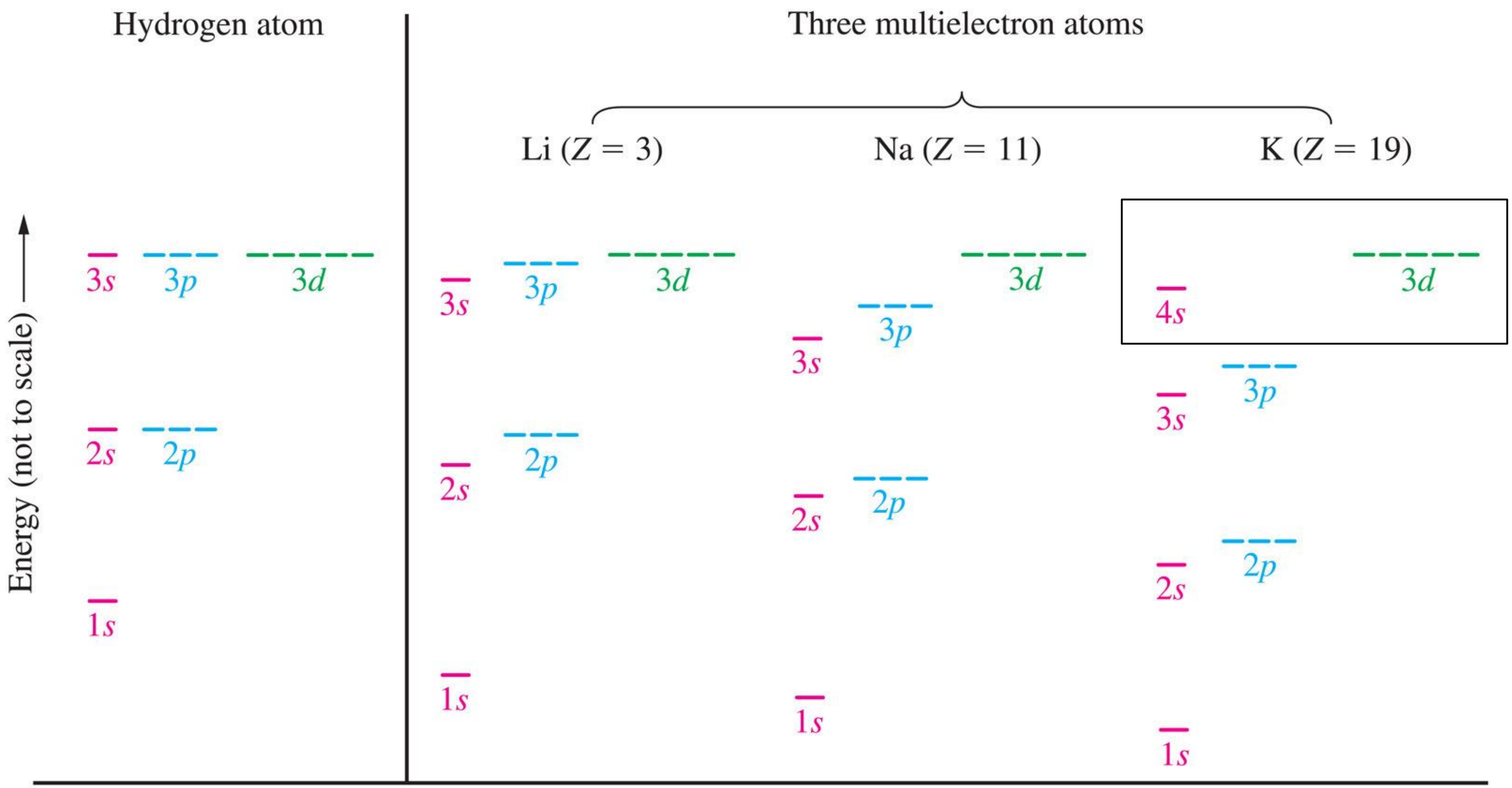
Both electrons are in the first energy level.

Both electrons are in the *s* orbital of the first energy level.

Because they are in the same orbital, they must have opposite spins

n	l	m_l	m_s
1	0	0	$+\frac{1}{2}$
1	0	0	$-\frac{1}{2}$

Sublevel Splitting in Multielectron Atoms



Orbital energy-level diagram for the first three electronic shells

We call orbitals with the same energy **degenerate.(等能量）**

$s\ (l = 0) < p\ (l = 1) < d\ (l = 2) < f\ (l = 3)$

Multi-electron Atoms (多電子原子)

Schrödinger equation was for only one e^-.

Electron-electron repulsion in multi-electron atoms.

Hydrogen-like orbitals (by approximation).

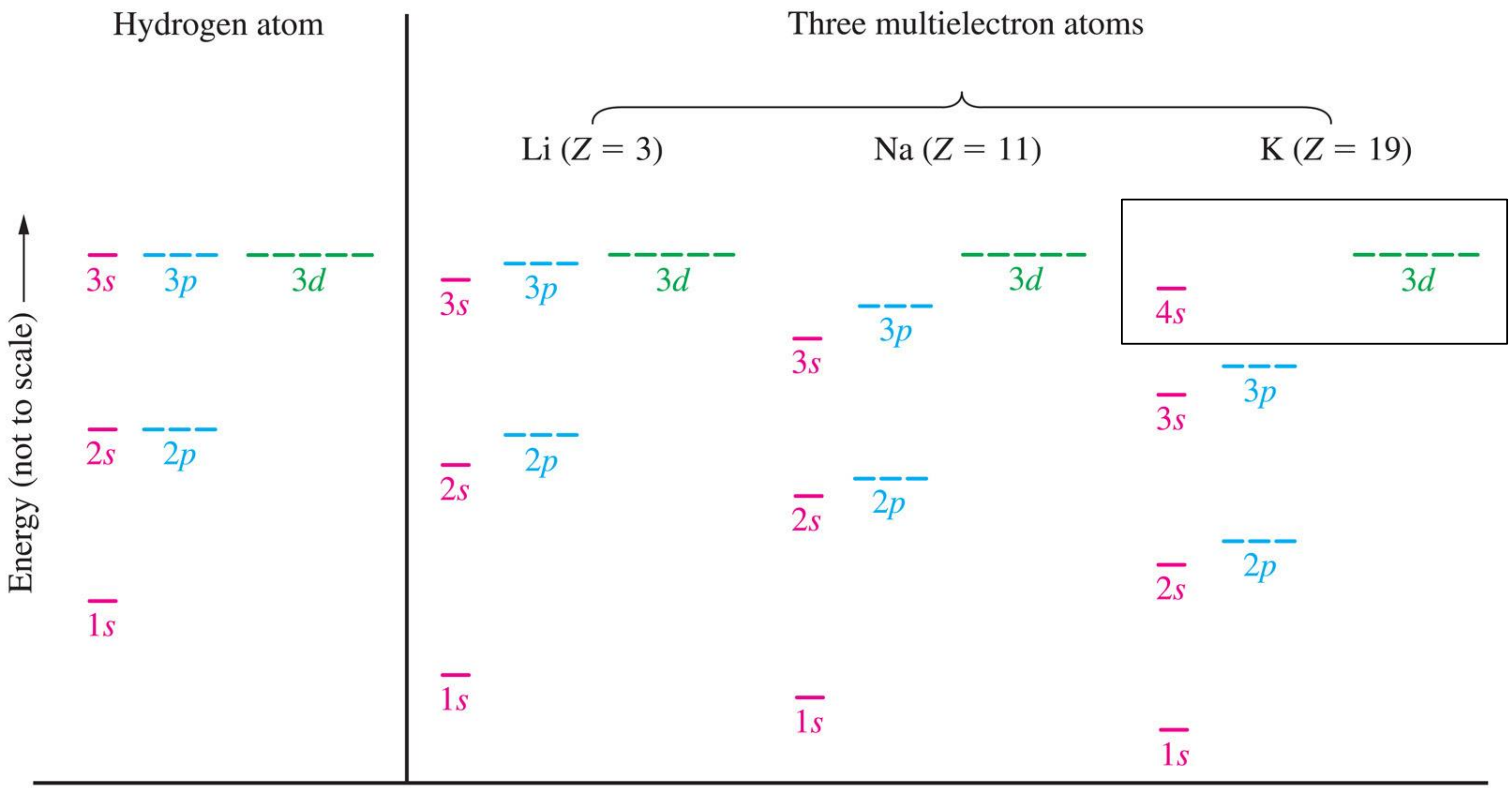
- 多電子原子的波函數, 只能求得數值上的近似解
- 利用量子力學, 可準確地計算出多電子原子的能階等性質。
- 多電子原子與氫原子的結構有一主要分別:

➔氫原子的軌域能階由主量子數 n 決定;

氫原子能階: $1s < 2s = 2p < 3s = 3p = 3d < 4s = 4p = 4d$

➔多電子原子的軌域能階由主量子數 n 和角動量子數 l 決定。

鉀原子能階: $1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p$



▲ FIGURE 8-36

Orbital energy-level diagram for the first three electronic shells

多電子原子的能階(1)

要考慮核對各電子的吸引力及電子間的互斥作用。

$E=-2.179*10^{-18}*(Z-\sigma)^2/n^2$

E：每一個電子的能量(j/個)

Z:核電荷

σ：遮蔽常數

n：主量子數

多電子原子的能階(2)

(1)隨主量子數n增大，主階能階變大，K＜L ＜M ＜N…，同一主層各副層能階順序隨l值增大而變大，所以產生能階分裂:

$E_{ns} < E_{np} < E_{nd} < E_{nf}$

(2)因為屏蔽效應及穿透效應, 所以產生能階交錯：

$E_{ns} < E_{(n-2)f} < E_{(n-1)d} < E_{np}$

屏蔽效應

(1)電子間的互斥而使核對外層電子的吸引被減弱的作用稱屏蔽效應。

(2)內層電子對外層電子的屏蔽作用較大，外層電子對內層電子可看做不產生屏蔽。

(3)離核愈近的主層上的電子被其他電子遮蔽的效應愈小，所以各電子層遮蔽作用大小為K ＞ L ＞ M ＞ N...

穿透效應

(1)外層電子的電子雲具有鑽到內部空間而更靠近核的現象較電子穿透，電子穿透降低了其他電子對他的屏蔽作用，而使軌域能量降低，這種效應稱為穿透效應。

(2)由電子雲徑向分佈圖發現：ns比np易穿透，np比nd易穿透。

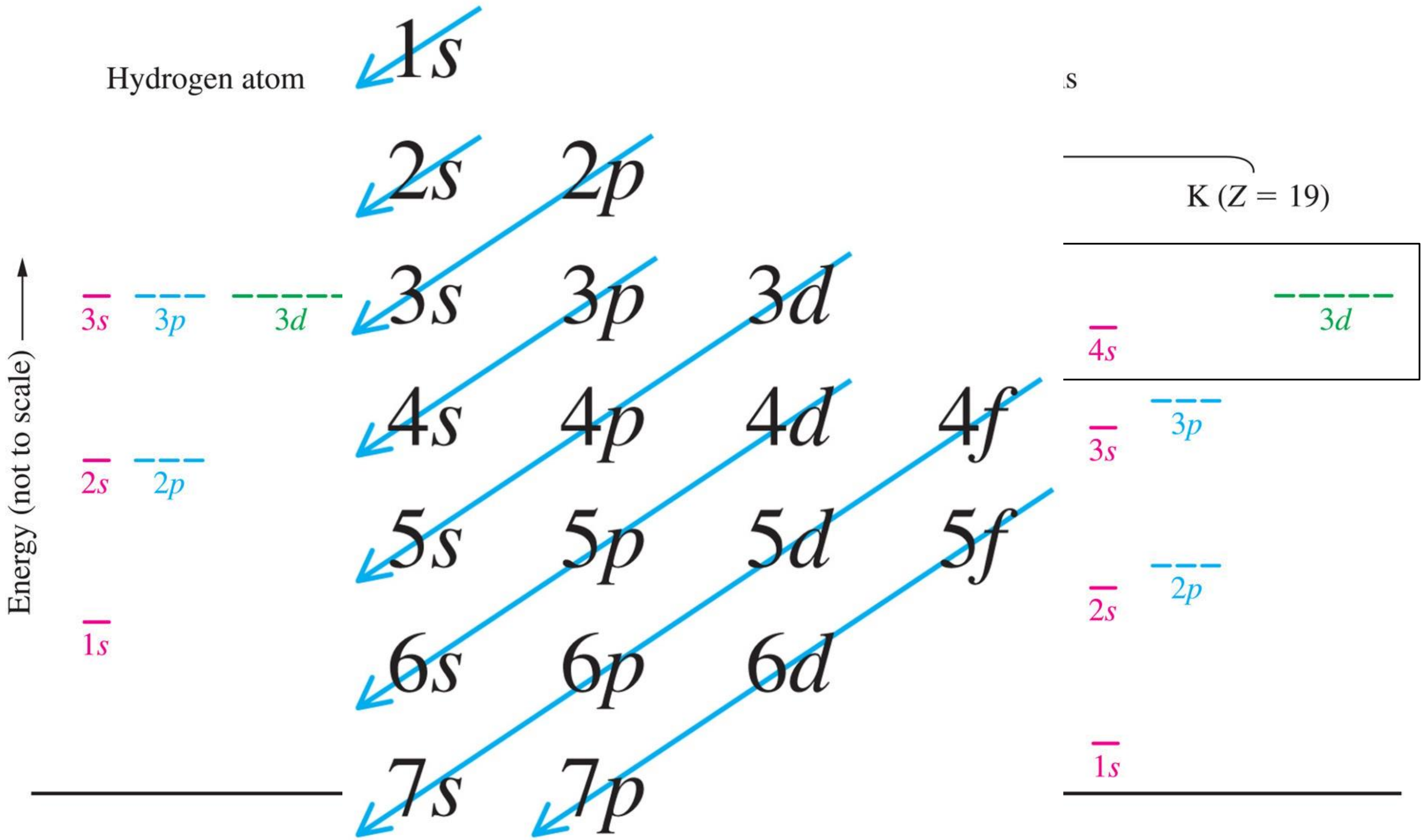
能級交錯舉例說明

$E_{4s} < E_{3d}$：

4s: n=4, s → l=0

3d: n=3, d → l=2

因為4s主量子數比3d大1，但l角動量量子數比3d小2，即s orbital 大於 d orbital 穿透效應增大而使軌域能量降低的作用超過主量子數增加對軌域能量的升高作用。



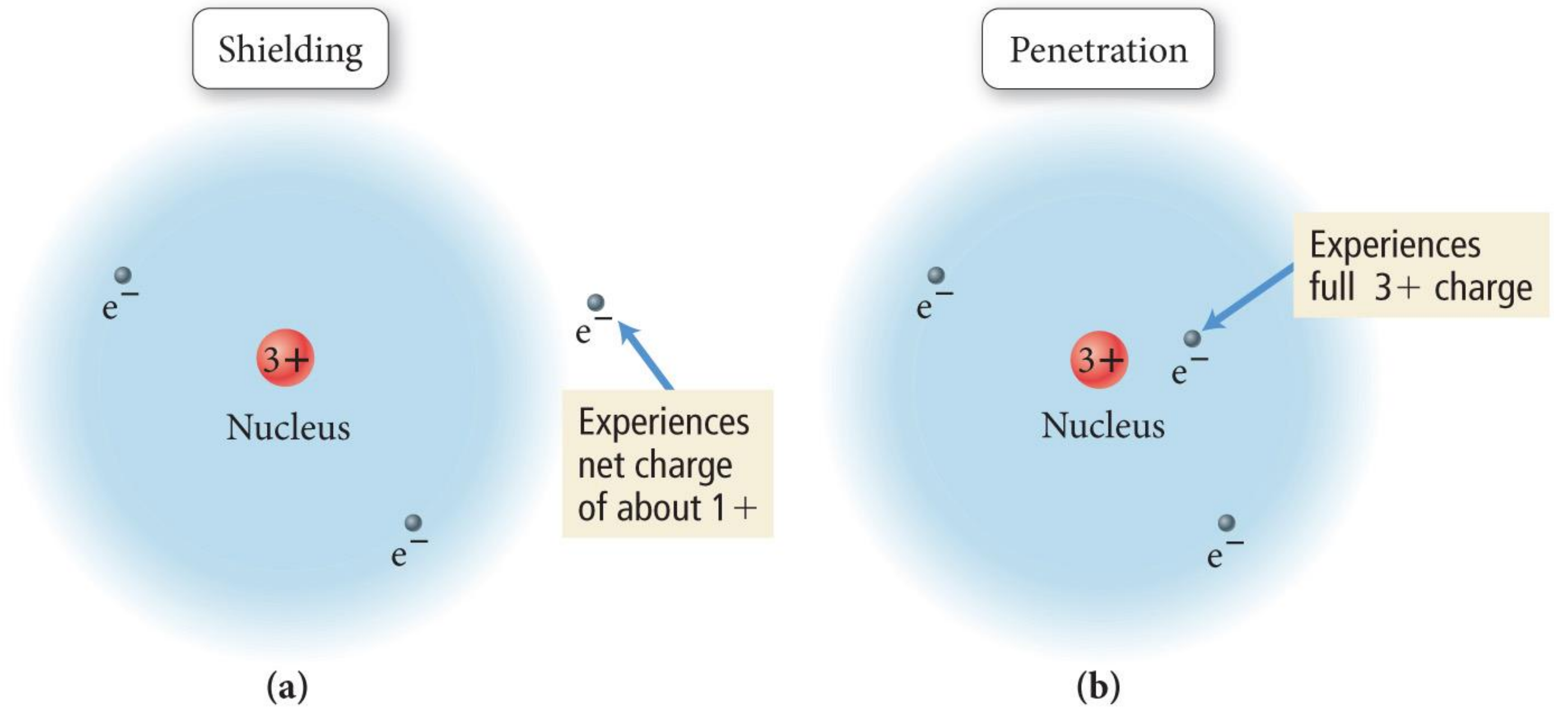
▲ FIGURE 8-36

▲ FIGURE 8-37

Orbital energy-level diagrams for the first three electronic shells

The order of filling of electronic subshells

Shielding(遮蔽) and Penetration(穿透)

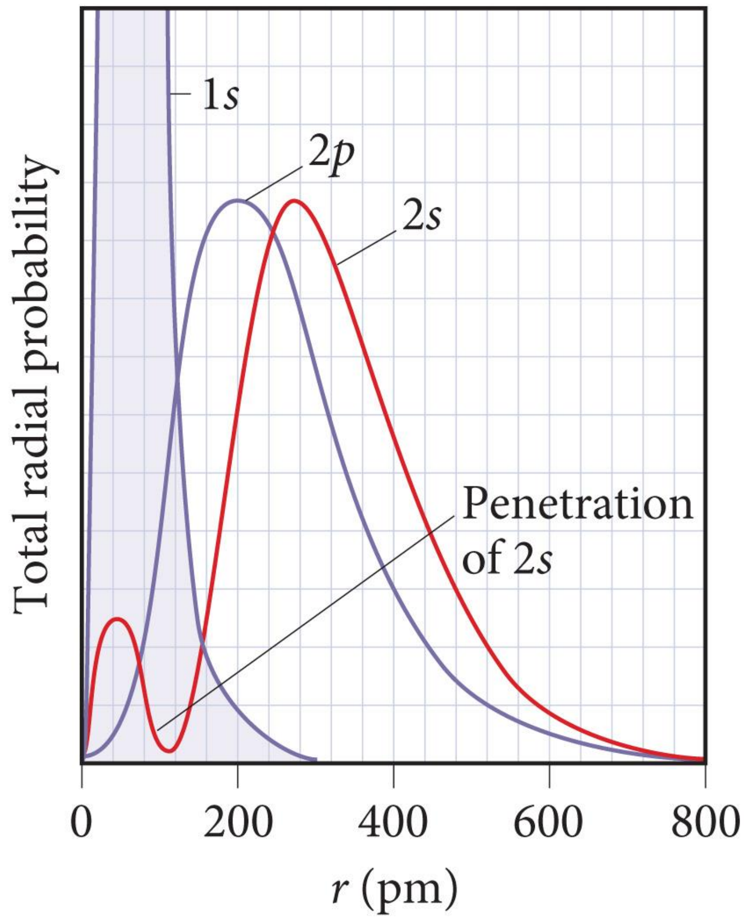


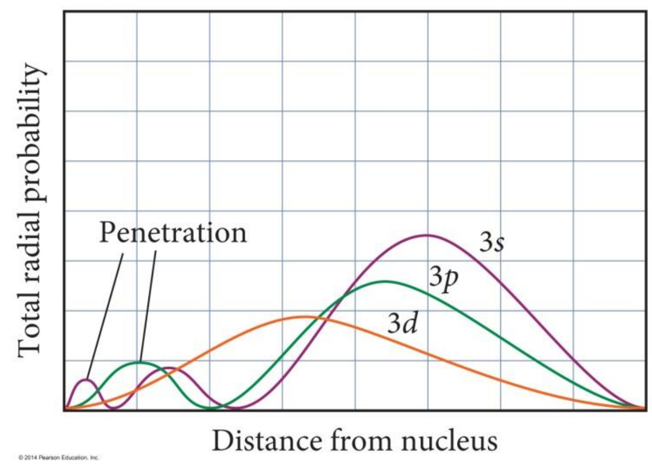
- Each electron in a multielectron atom experiences both the attraction to the nucleus and repulsion by other electrons in the atom.
- These repulsions cause the electron to have a net reduced attraction to the nucleus; it is **shielded** from the nucleus.
- The total amount of attraction that an electron feels for the nucleus is called the **effective nuclear charge** of the electron.

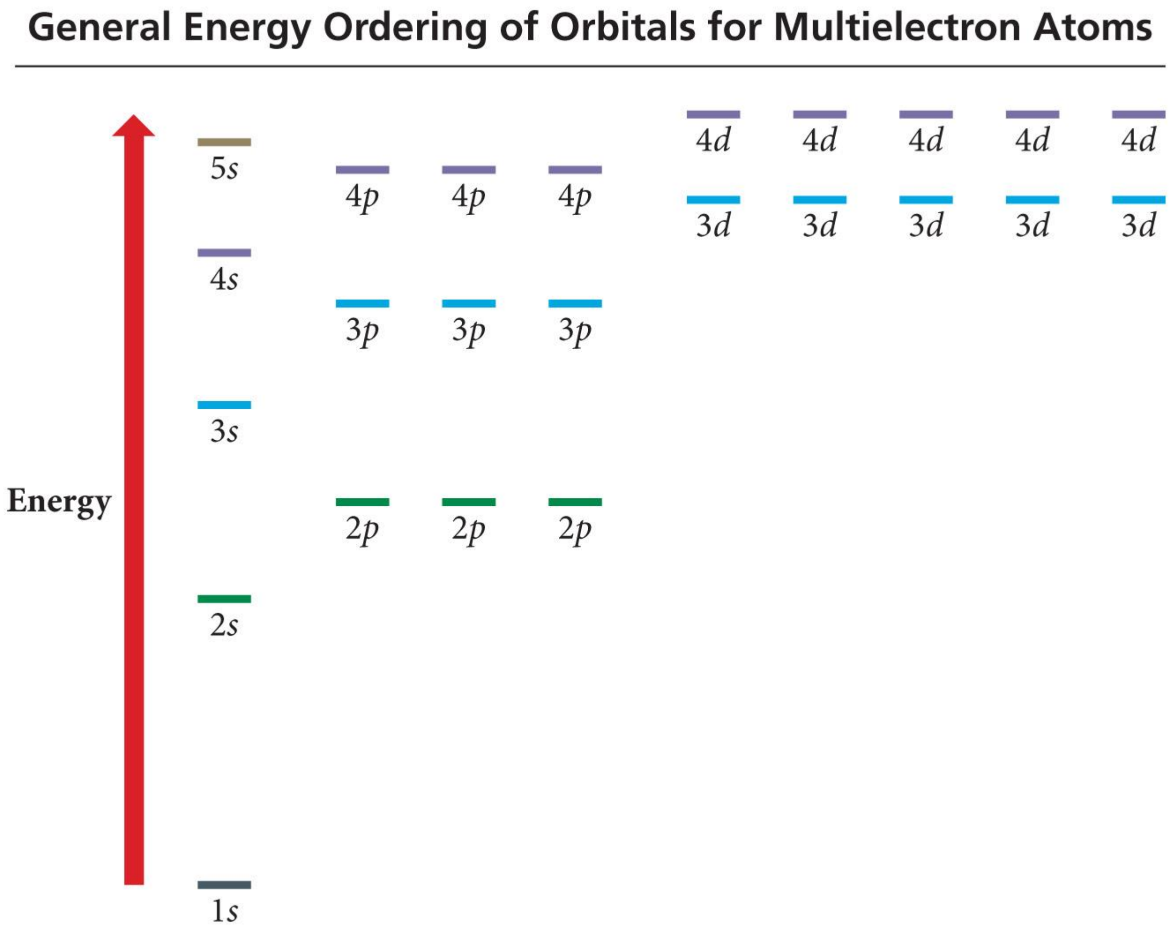
- The closer an electron is to the nucleus, the more attraction it experiences.
- The better an outer electron is at **penetrating** through the electron cloud of inner electrons, the more attraction it will have for the nucleus.
- The degree of penetration is related to the orbital’ s radial distribution function.
  - In particular, the distance the maxima of the function are from the nucleus

Penetration and Shielding

- 2s>2p
- 3s > 3p > 3d







▲ FIGURE 8.5 General Energy Ordering of Orbitals for Multielectron Atoms.

Multielectron Electron Configurations

Aufbau process

Electrons occupy orbitals in a way that minimizes the energy of the atom.

Pauli exclusion principle

No two electrons can have all four quantum numbers alike.

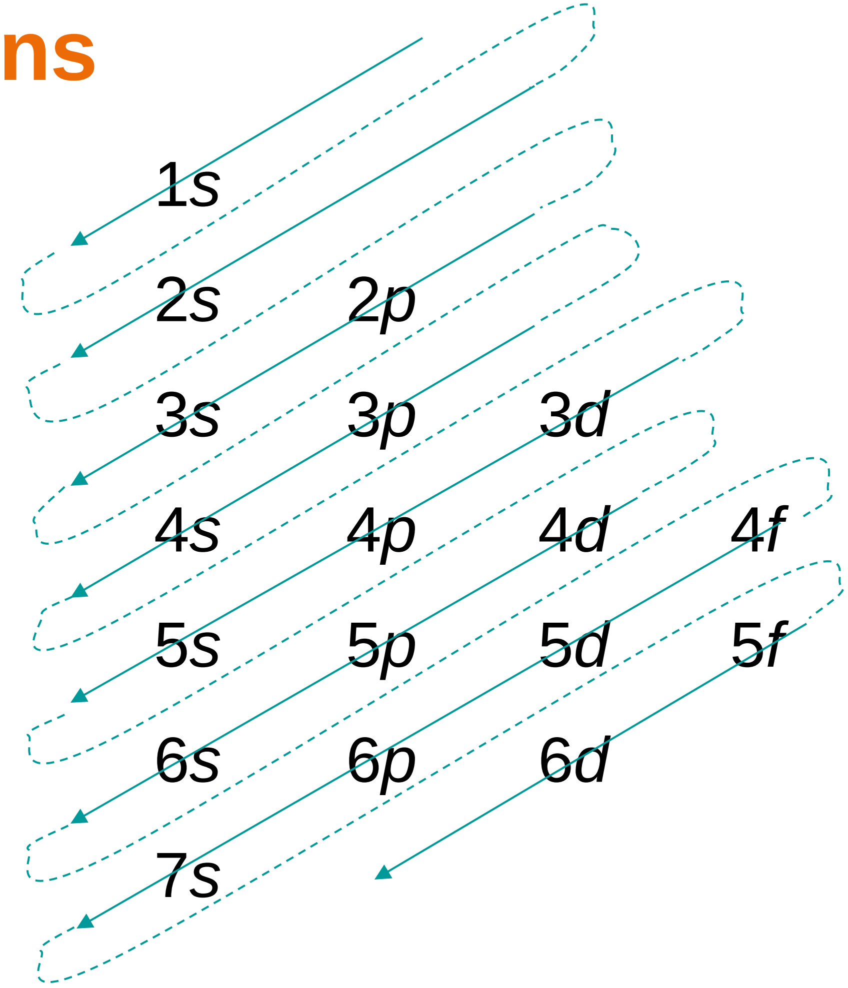
Hund's rule

When orbitals of identical energy (degenerate orbitals) are available, electrons initially occupy these orbitals singly.

Order of Sublevel Filling in Ground State Electron Configurations

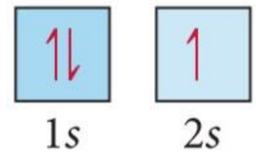
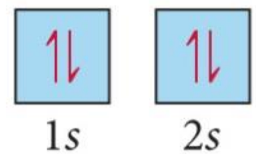
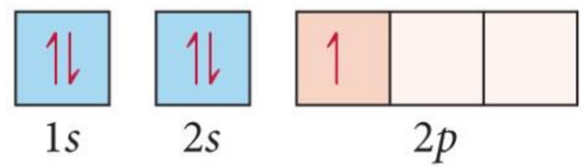
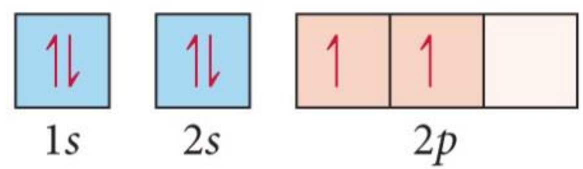
Start by drawing a diagram, putting each energy shell on a row and listing the sublevels (*s, p, d, f*) for that shell in order of energy (from left to right).

Next, draw arrows through the diagonals, looping back to the next diagonal each time.



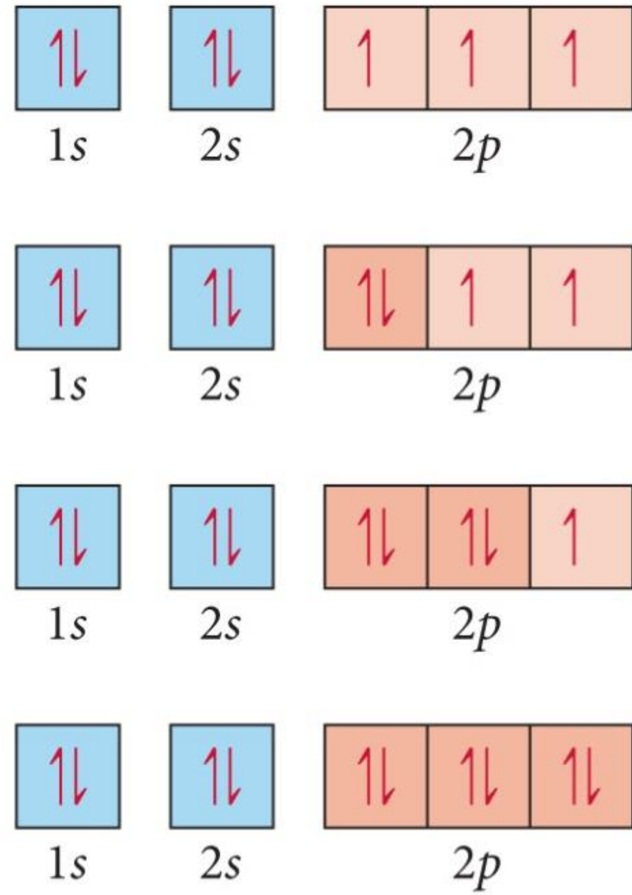
1s <2s< 2p<3s<3p<4s<3d<4p<5s<4d<5p<6s<4f<5d<6p<6d<7s

Electron Configuration of Atoms in Their Ground State

Symbol	Number of electrons	Electron configuration	Orbital diagram
Li	3	$1s^2 2s^1$	
Be	4	$1s^2 2s^2$	
B	5	$1s^2 2s^2 2p^1$	
C	6	$1s^2 2s^2 2p^2$	

Electron Configurations

N	7	$1s^2 2s^2 2p^3$
O	8	$1s^2 2s^2 2p^4$
F	9	$1s^2 2s^2 2p^5$
Ne	10	$1s^2 2s^2 2p^6$

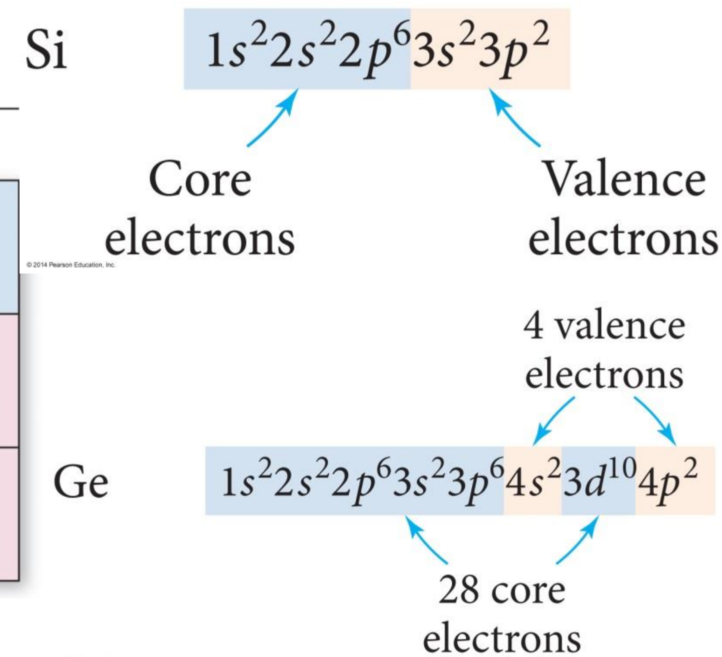


© 2014 Pearson Education, Inc.

8.4 Electron Configurations, Valence Electrons, and the Periodic Table

Outer Electron Configurations of Elements 1–18

1A	2A	3A	4A	5A	6A	7A	8A
1 **H** $1s^1$							2 **He** $1s^2$
3 **Li** $2s^1$	4 **Be** $2s^2$	5 **B** $2s^2 2p^1$	6 **C** $2s^2 2p^2$	7 **N** $2s^2 2p^3$	8 **O** $2s^2 2p^4$	9 **F** $2s^2 2p^5$	10 **Ne** $2s^2 2p^6$
11 **Na** $3s^1$	12 **Mg** $3s^2$	13 **Al** $3s^2 3p^1$	14 **Si** $3s^2 3p^2$	15 **P** $3s^2 3p^3$	16 **S** $3s^2 3p^4$	17 **Cl** $3s^2 3p^5$	18 **Ar** $3s^2 3p^6$



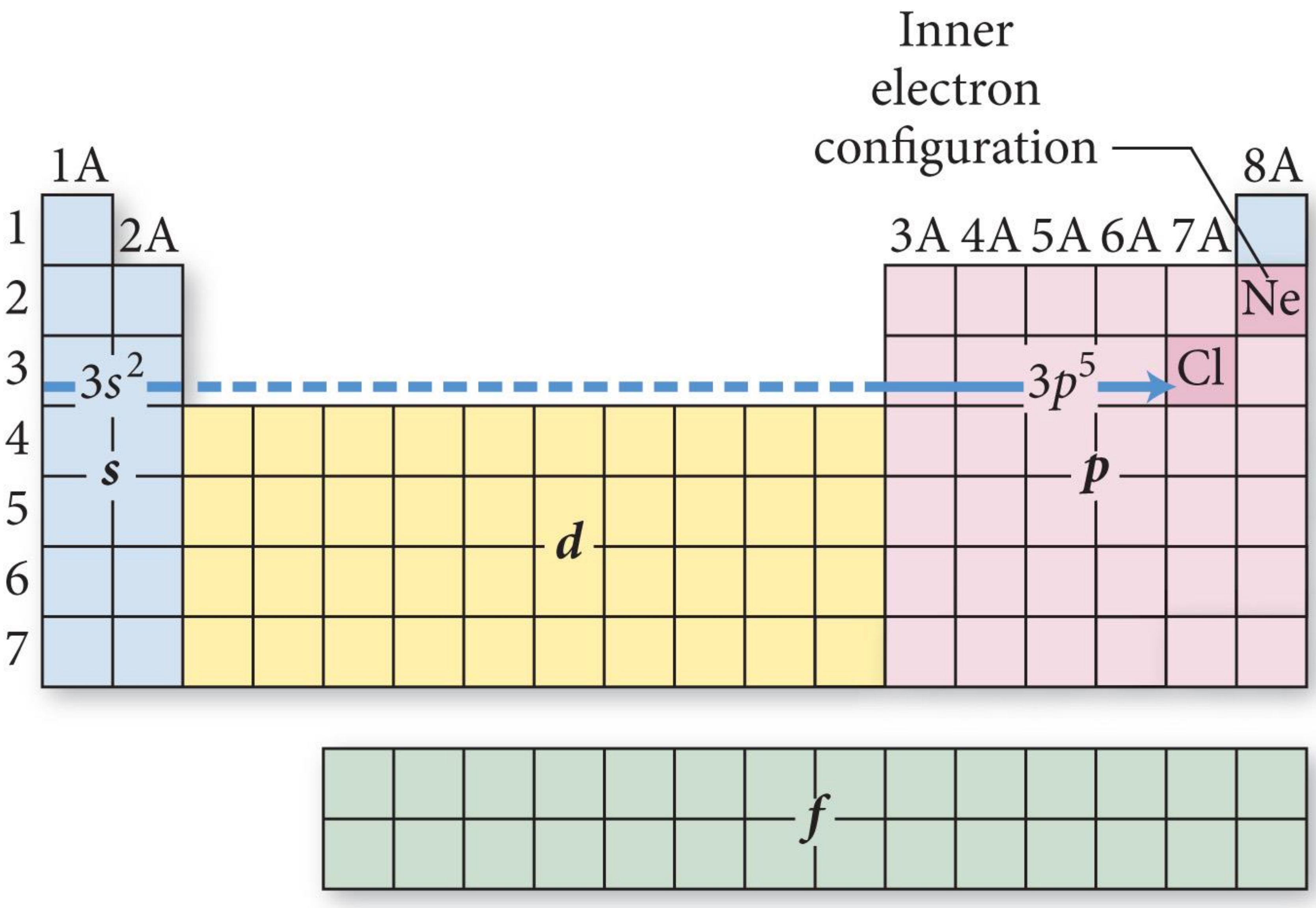
- The electrons in all the sublevels with the **highest principal energy shell** are called the **valence electrons(價電子)**.
- Electrons in lower energy shells are called **core electrons (內層電子)**. .

Orbital Blocks of the Periodic Table

- *s*-block elements
- *p*-block elements
- *d*-block elements
- *f*-block elements

Periods \ Groups	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
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**	108 **Hs**	109 **Mt**	110 **Ds**	111 **Rg**	112 **Cn**	113 **	114 **Fl**	115 **	116 **Lv**	117 **	118 **

Lanthanides	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** $6s^24f^{14}6d^1$
Actinides	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** $7s^25f^{14}6d^1$


Inner electron configuration
1A
2A
3A 4A 5A 6A 7A
8A
1
2
3
4
5
6
7
Ne
Cl
3s2
3p5
s
p
d
f

Irregular Electron Configurations

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** $4s^23d^{10}$
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** $5s^24d^{10}$

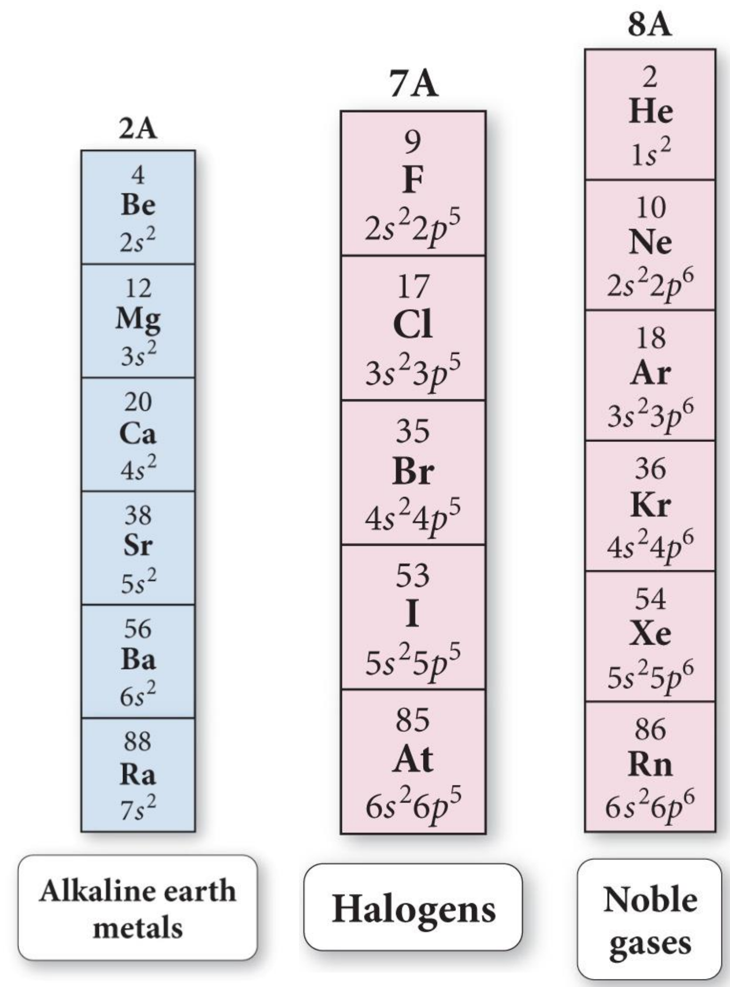
- Expected
- Cr = [Ar]$4s^23d^4$
- Cu = [Ar]$4s^23d^9$
- Mo = [Kr]$5s^24d^4$
- Ru = [Kr]$5s^24d^6$
- Pd = [Kr]$5s^24d^8$

- Found experimentally
- Cr = [Ar]$4s^13d^5$
- Cu = [Ar]$4s^13d^{10}$
- Mo = [Kr]$5s^14d^5$
- Ru = [Kr]$5s^14d^7$
- Pd = [Kr]$5s^04d^{10}$

8.5 The Explanatory Power of the Quantum-Mechanical Model

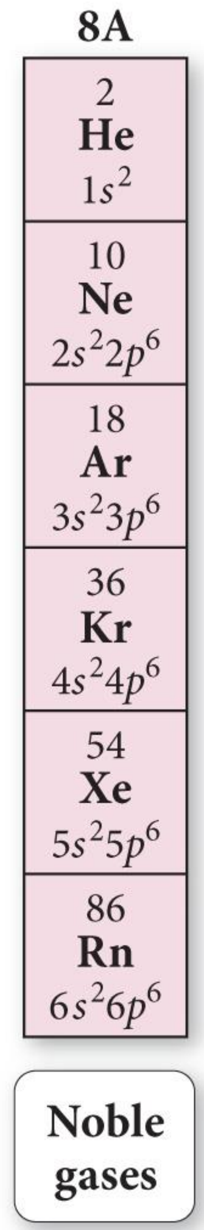
Properties and Electron Configuration

- The properties of the elements follow a periodic pattern.
  - Elements in the same column have similar properties.
  - The elements in a period show a pattern that repeats.
- The quantum-mechanical model explains this because the number of valence electrons and the types of orbitals they occupy are also periodic.



The Noble Gas（鈍氣） Electron Configuration

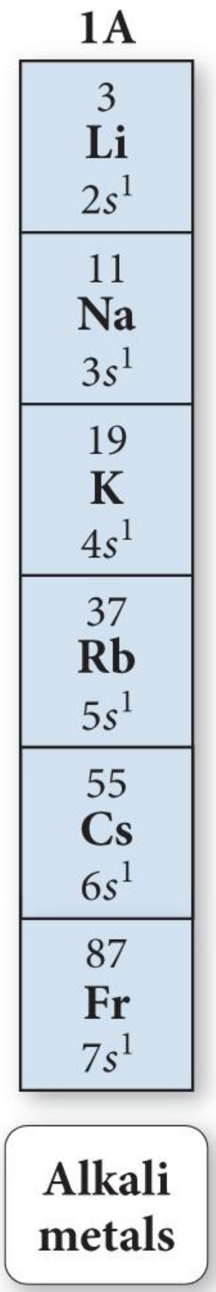
- The noble gases have eight valence electrons.
  - Except for He, which has only two electrons
- They are especially nonreactive.
  - He and Ne are practically inert.
- The reason the noble gases are so nonreactive is that the electron configuration of the noble gases is especially stable.



The Alkali Metals（鹼金屬）

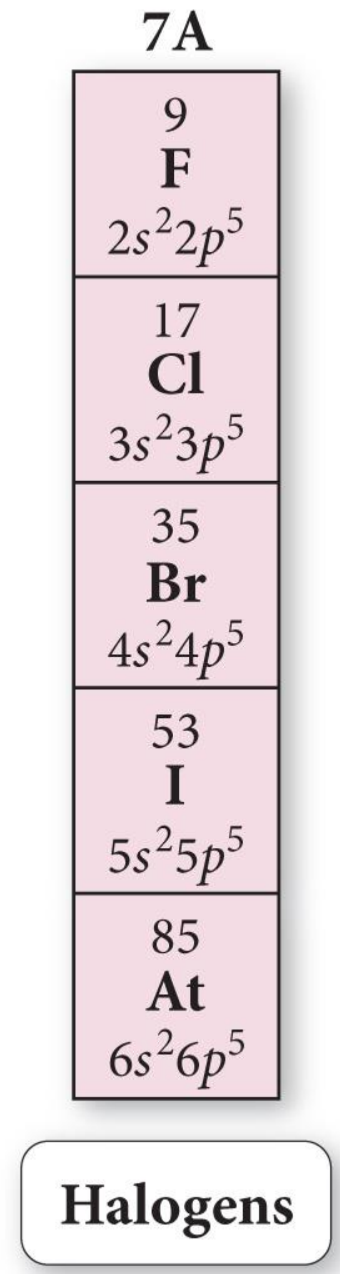
- The **alkali** metals have one more electron than the previous noble gas.

- In their reactions, the alkali metals tend to lose one electron, resulting in the same electron configuration as a noble gas.
  - Forming a cation with a **1+** charge



The Halogens(鹵素)

- Have one fewer electron than the next noble gas
- In their reactions with metals, the halogens tend to gain an electron and attain the electron configuration of the next noble gas, forming an anion with charge **1−**.
- In their reactions with nonmetals, they tend to share electrons with the other nonmetal so that each attains the electron configuration of a noble gas.



Electron Configuration and Ion Charge

Mg atom = $1s^2 2s^2 2p^6 3s^2$

Mg^{2+} cation = $1s^2 2s^2 2p^6$

Cation 陽離子

Elements That Form Ions with Predictable Charges

	1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
1	Li^+														N^{3-}	O^{2-}	F^-	
2	Na^+	Mg^{2+}											Al^{3+}			S^{2-}	Cl^-	
3	K^+	Ca^{2+}														Se^{2-}	Br^-	
4	Rb^+	Sr^{2+}														Te^{2-}	I^-	
5	Cs^+	Ba^{2+}																

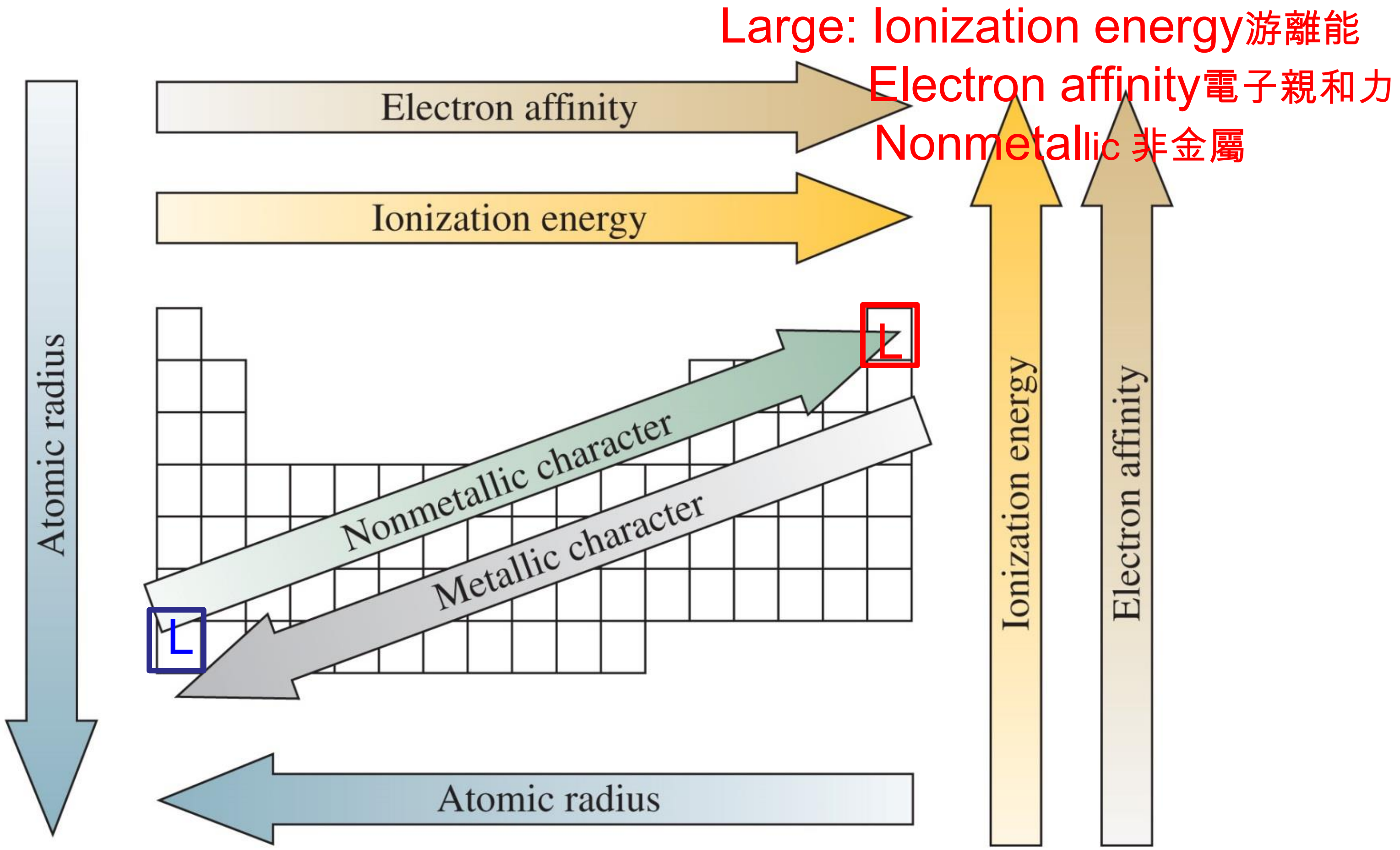
Mg^{2+} cation = $1s^2 2s^2 2p^6$

Anions 陰離子

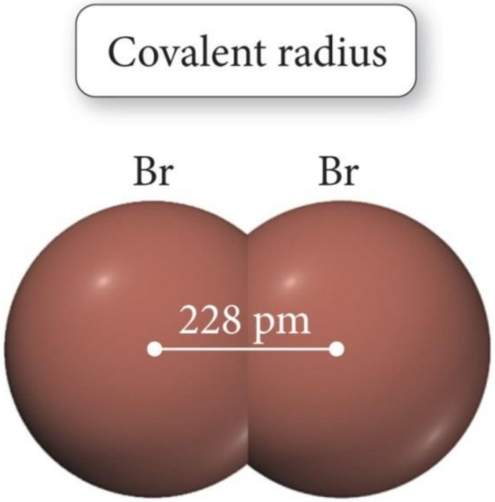
S atom = $1s^2 2s^2 2p^6 3s^2 3p^4$

S atom = $1s^2 2s^2 2p^6 3s^2 3p^4$

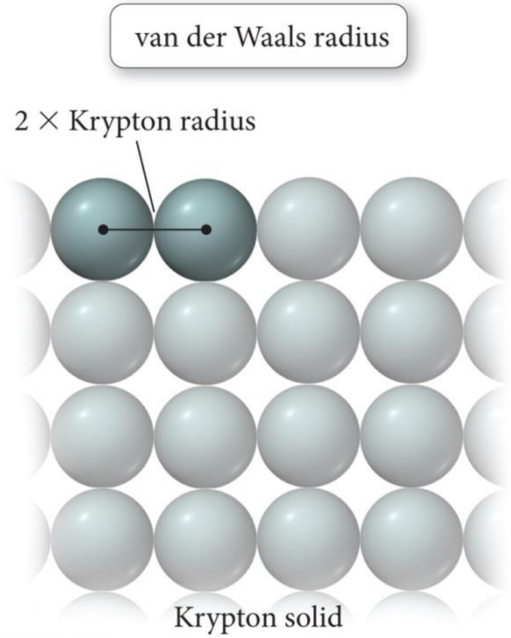
8.6 Periodic Trends in the Size of Atoms and Effective Nuclear Charge

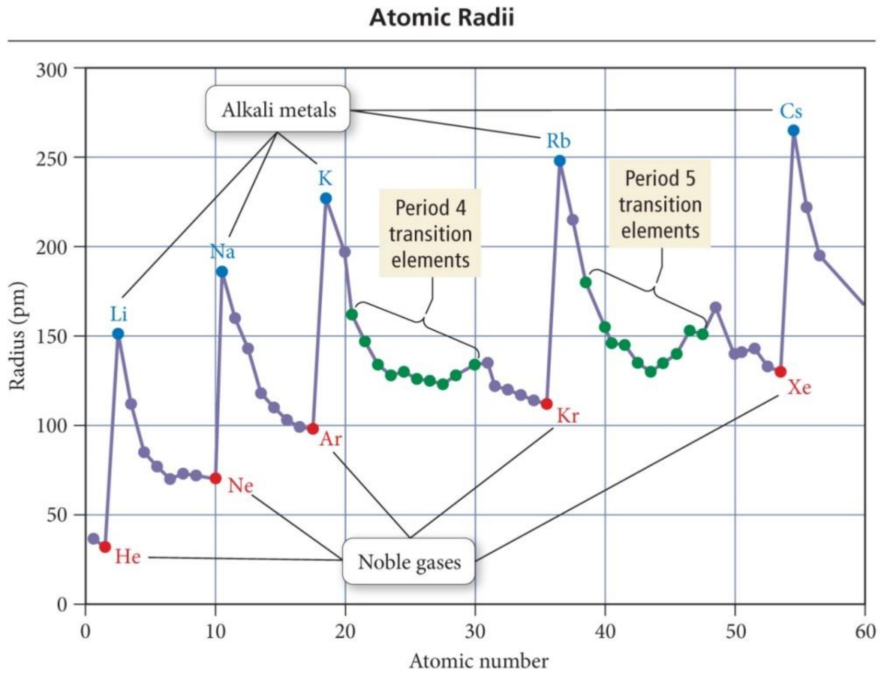


Large: atomic radius（原子半徑）
Metallic (金屬)

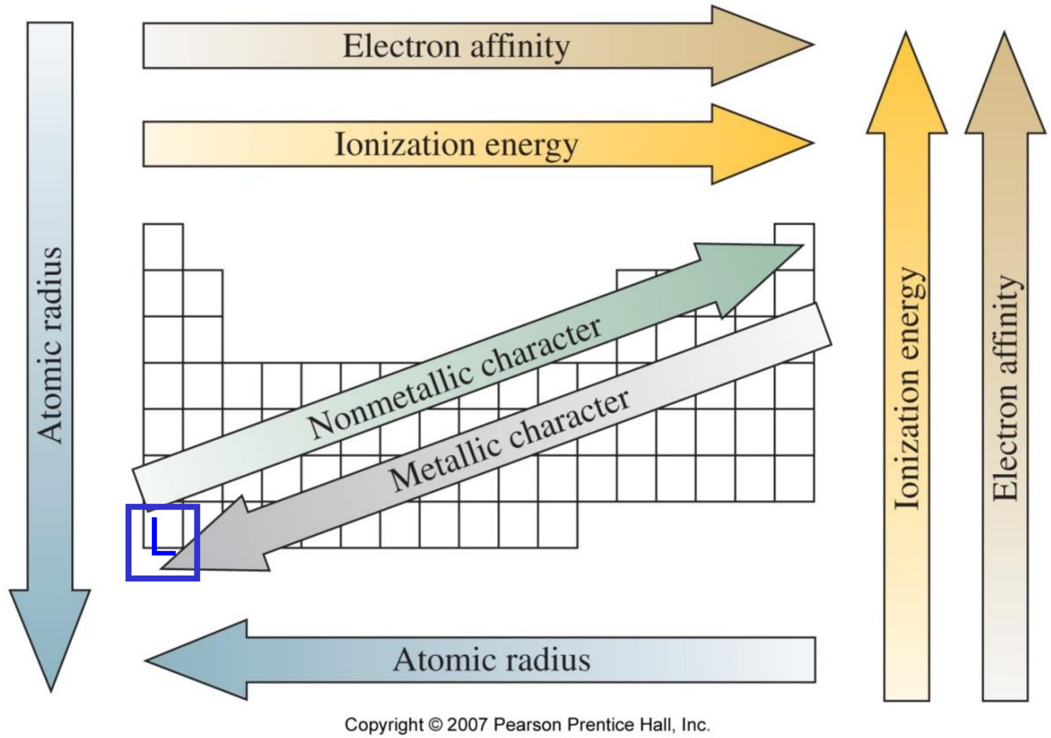


$$\text{Br radius} = \frac{228\text{ pm}}{2} = 114\text{ pm}$$





- Atomic radius increases down group
  - ✓ Valence shell farther from nucleus
  - ✓ Effective nuclear charge fairly close



- Atomic radius decreases across period (left to right)
  - ✓ Adding electrons to same valence shell
  - ✓ Effective nuclear charge increases
  - ✓ Valence shell held closer

Periodic Trends in Atomic Radius

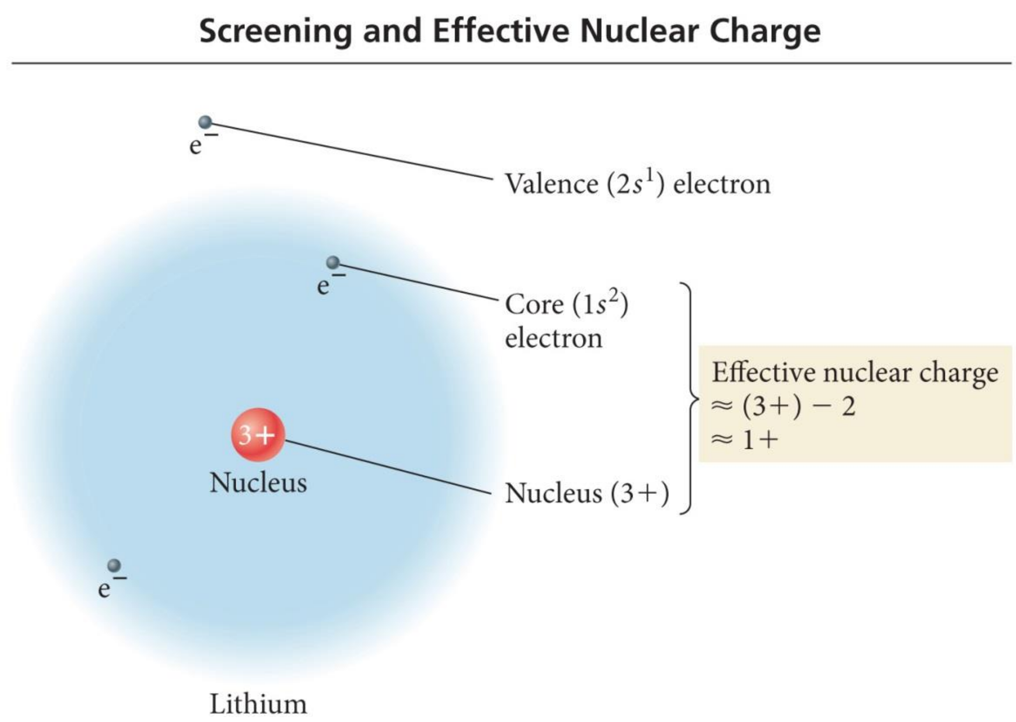
Gallium (eka-aluminum)

	Mendeleev's predicted properties	Actual properties
Atomic mass	About 68 amu	69.72 amu
Melting point	Low	29.8 °C
Density	5.9 g/cm^3	5.90 g/cm^3
Formula of oxide	X_2O_3	Ga_2O_3
Formula of chloride	XCl_3	$GaCl_3$

Germanium (eka-silicon)

	Mendeleev's predicted properties	Actual properties
Atomic mass	About 72 amu	72.64 amu
Density	5.5 g/cm^3	5.35 g/cm^3
Formula of oxide	XO_2	GeO_2
Formula of chloride	XCl_4	$GeCl_4$

Screening and Effective Nuclear Charge有效核電荷



Li $1s^2 2s^1$

Z_{eff} = 3-2 =1

Be $1s^2 2s^2$

Z_{eff} = 4-2 =2

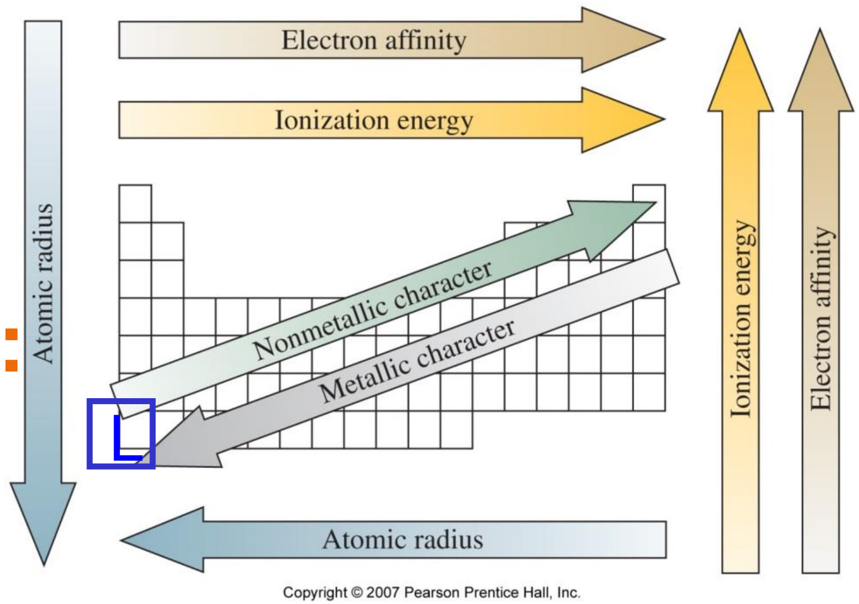
$Z_{effective} = Z - S$

Z is the nuclear charge

S is the number of electrons in lower energy levels

Quantum-Mechanical Explanation for the Group Trend in Atomic Radius

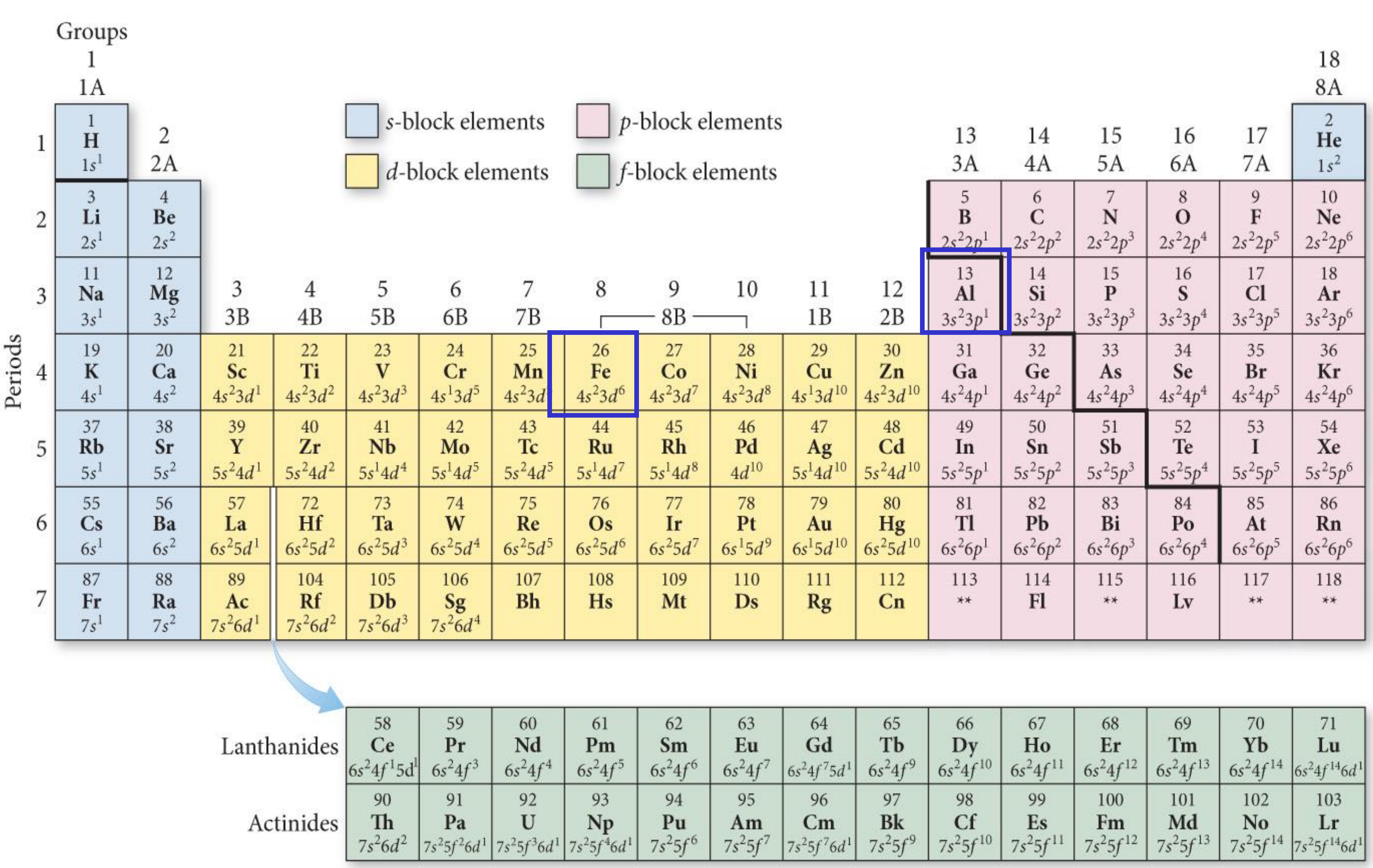
Trends in Atomic Radius: Transition Metals



- Atomic radii of transition metals are roughly the **same size** across the *d* block.
  - Much less difference than across main group elements
  - Valence shell ns^2, not the $(n-1)d$ electrons
  - **Effective nuclear charge on the ns^2** electrons approximately the same

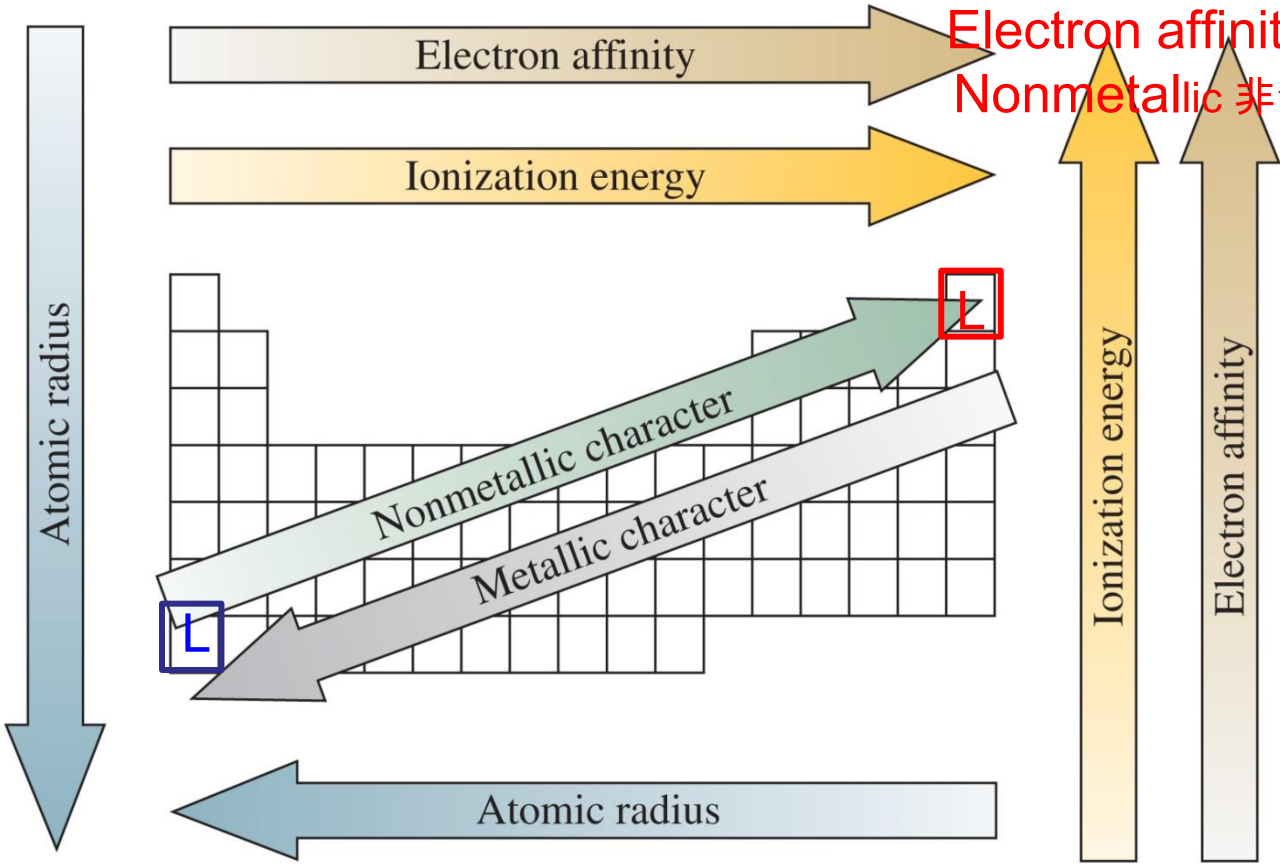
8.7 Ions: Electron Configurations, Magnetic Properties, Ionic Radii, and Ionization Energy

Orbital Blocks of the Periodic Table



8.7 **Ions: Electron Configurations, Magnetic Properties, Ionic Radii, and Ionization Energy**

Large: Ionization energy游離能
Electron affinity電子親和力
Nonmetallic 非金屬



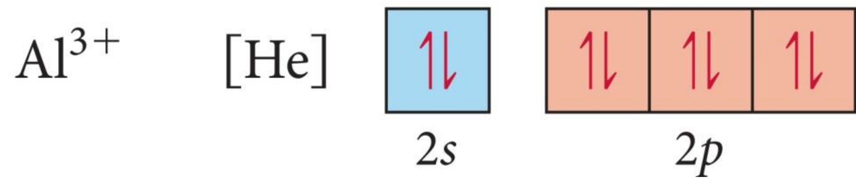
Large: atomic radius（原子半徑）
Metallic (金屬)

paramagnetism(順磁).電子都成對

-Will be attracted to a magnetic field

Al atom = $1s^22s^22p^63s^23p^1$

Al^{3+} ion = $1s^22s^22p^6$



- **diamagnetism(逆磁).有不成對電子**
  – Slightly repelled by a magnetic field

- The iron atom has two valence electrons:

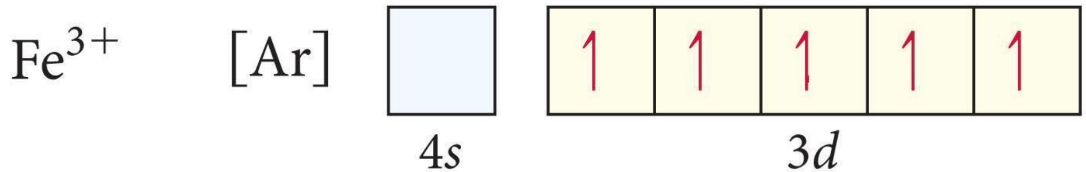
  Fe atom = $1s^22s^22p^63s^23p^64s^23d^6$

- When iron forms a cation, it first loses its valence electrons:

  Fe^{2+} cation = $1s^22s^22p^63s^23p^63d^6$

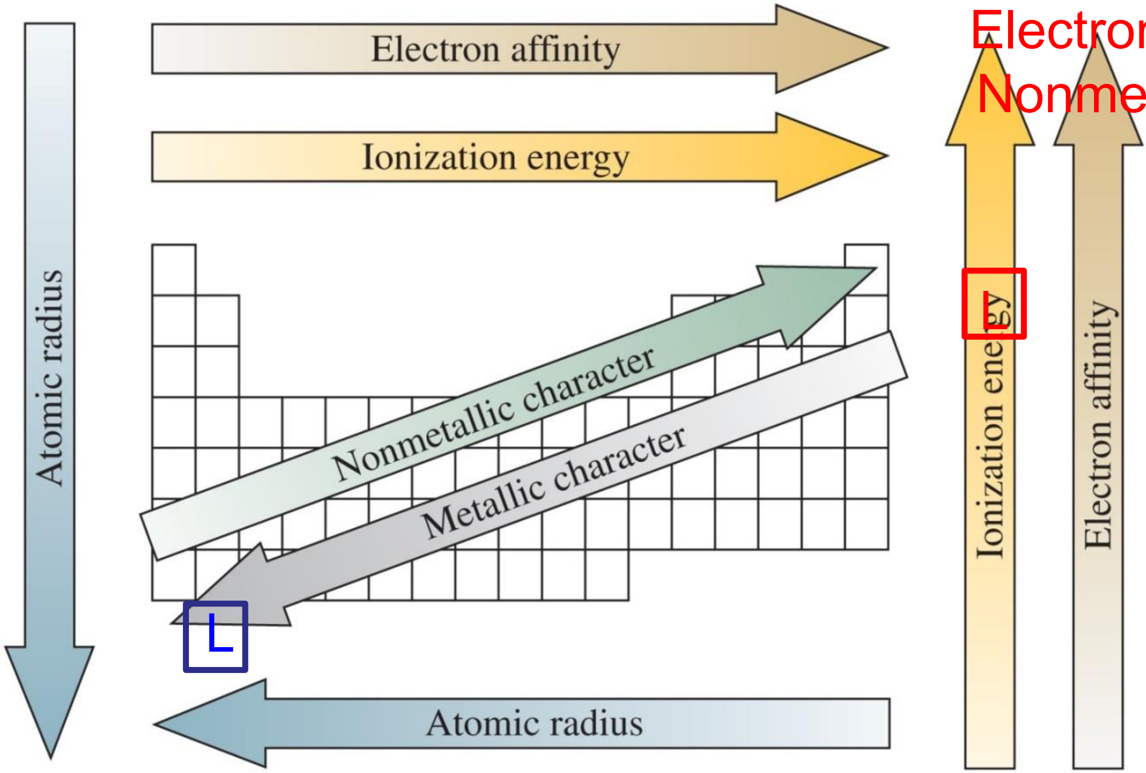
- It can then lose 3*d* electrons:

  Fe^{3+} cation = $1s^22s^22p^63s^23p^63d^5$



Trends in Ionic Radius

Large: Ionization energy游離能
Electron affinity電子親和力
Nonmetallic 非金屬



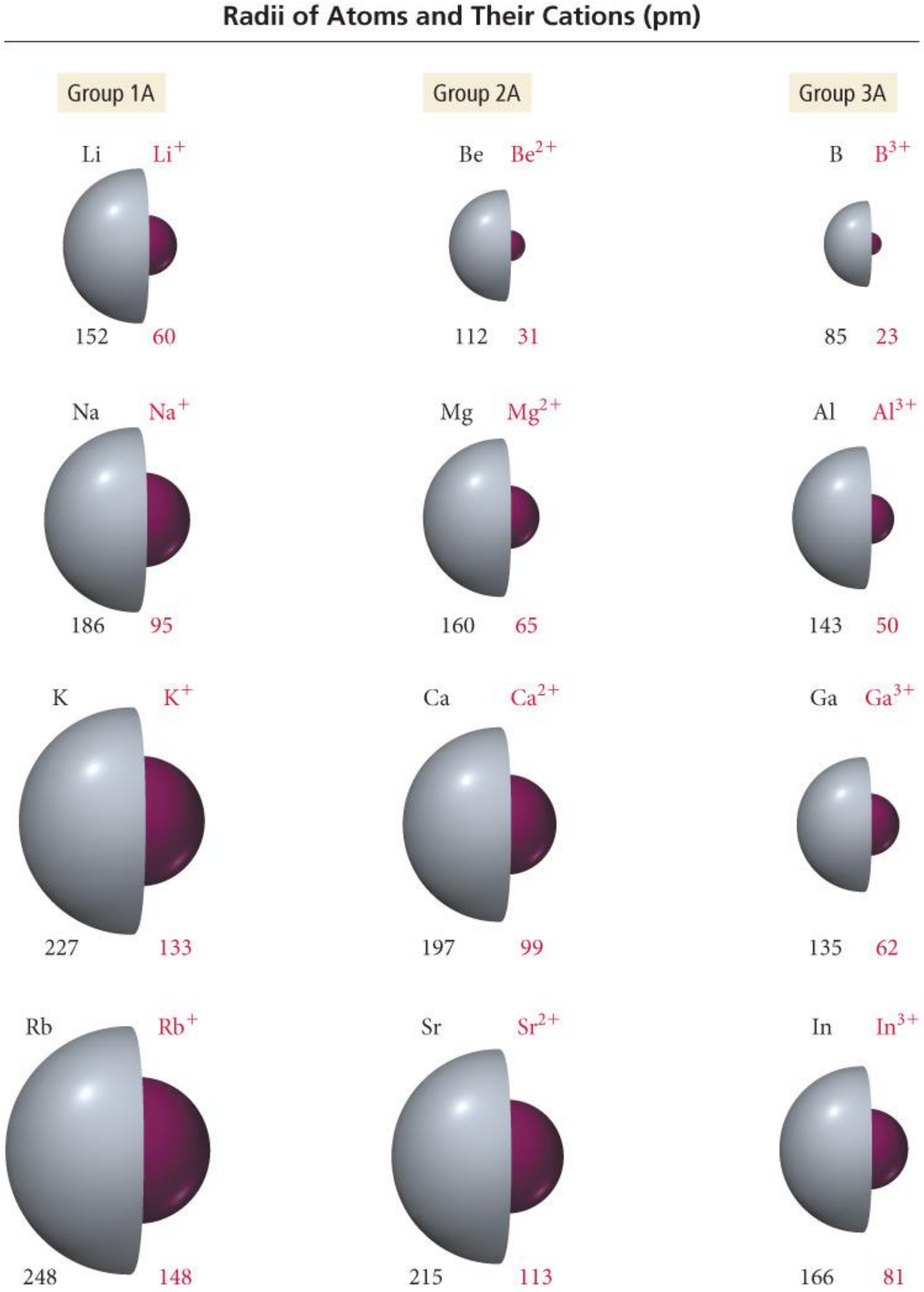
Large: atomic radius（原子半徑）
Metallic (金屬)

$O^{2-} > F^{-} > Na^{+} > Mg^{2+} > Al^{3+}$

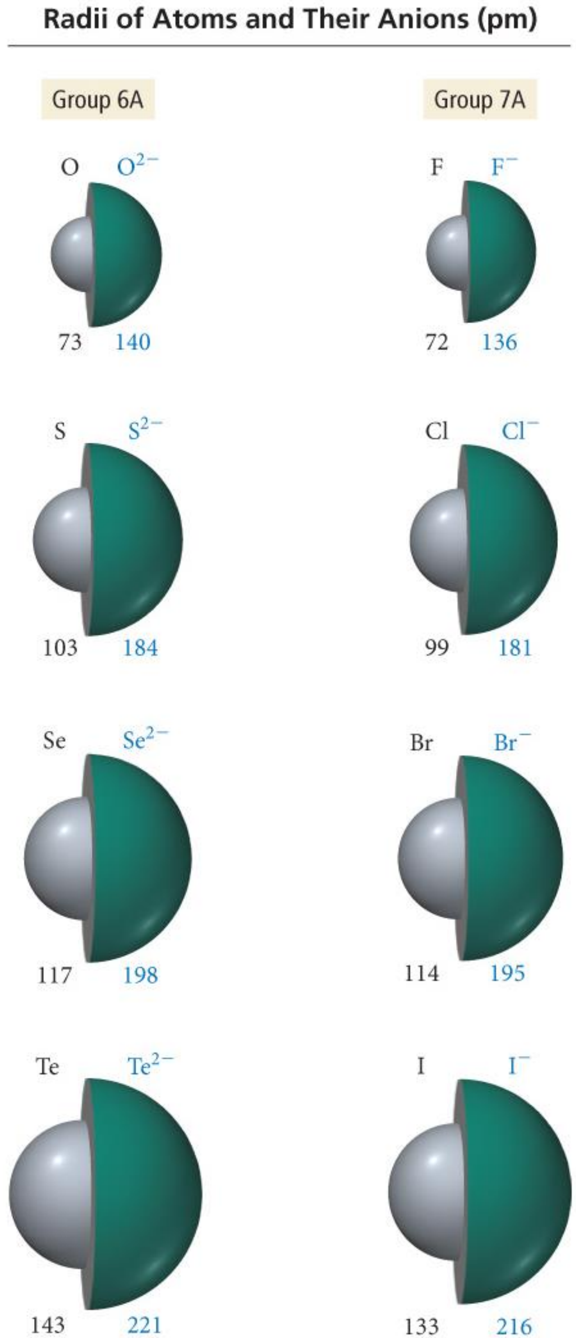
✓ Isoelectronic = same electron configuration

anions陰離子 > neutral atoms中性原子 > Cations陽離子
越負愈大 越正越小

Periodic Trends in Ionic Radius

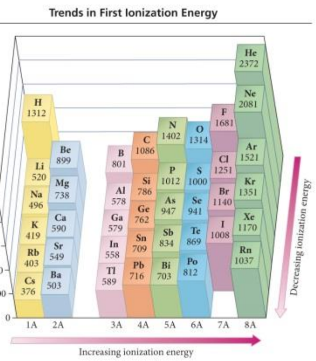


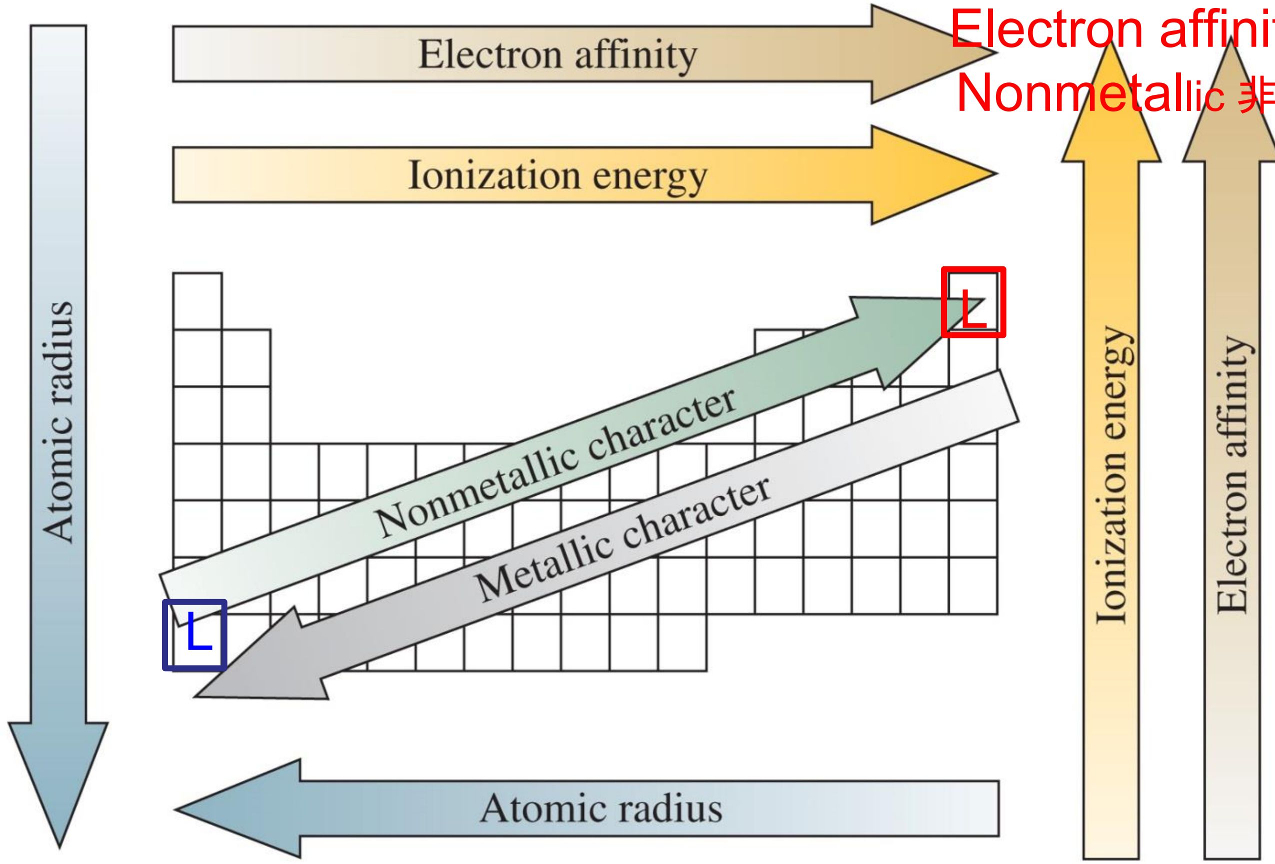
Periodic Trends in Ionic Radius



Ionization Energy (IE)

Large: Ionization energy游離能
Electron affinity電子親和力
Nonmetallic 非金屬

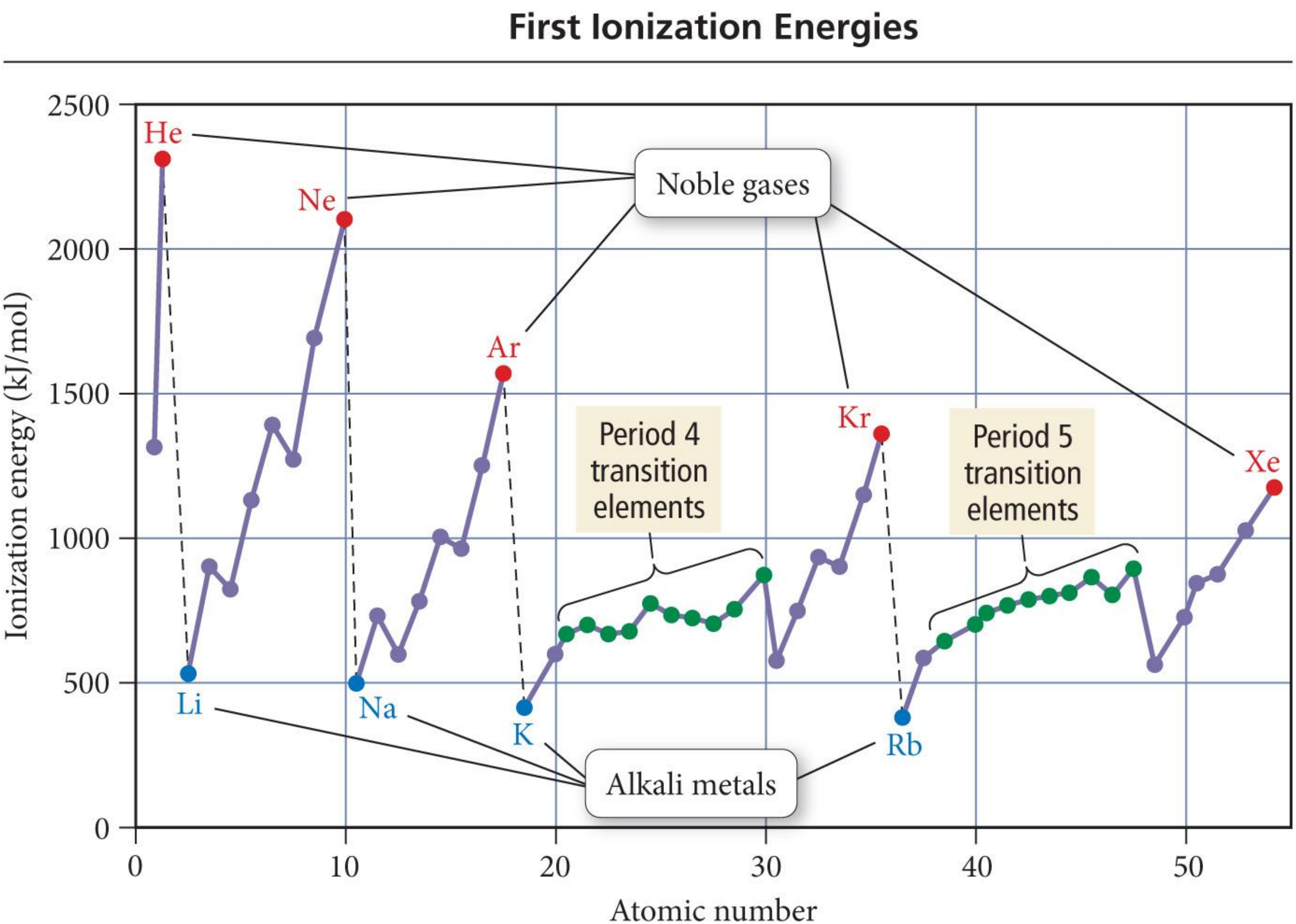




Large: atomic radius（原子半徑）
Metallic (金屬)

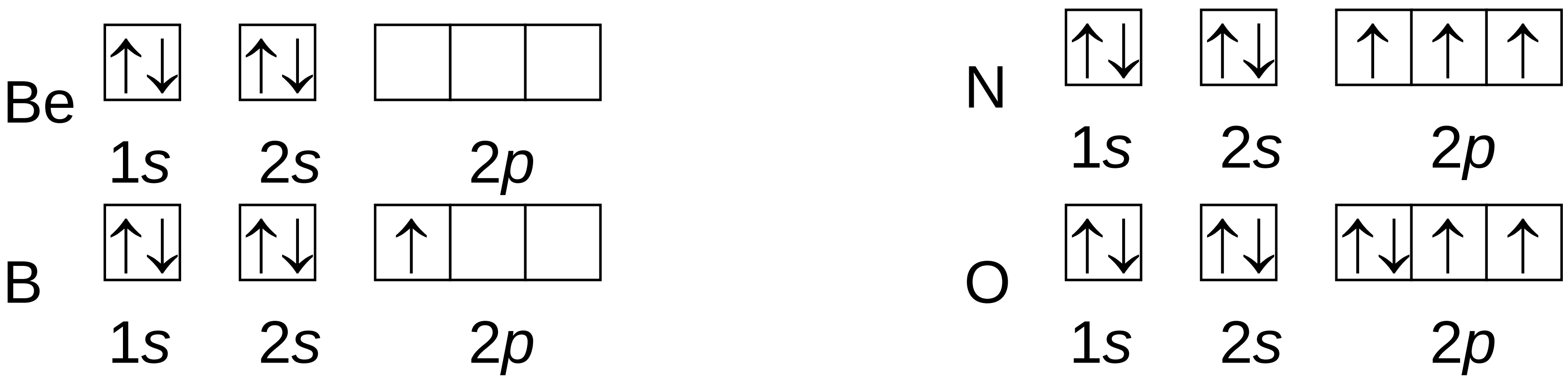
Ionization Energy (IE)

- Minimum energy needed to remove an electron from an atom or ion
  - ✓ Gas state
  - ✓ Endothermic process
  - ✓ Valence electron easiest to remove, lowest IE
  - ✓ $M(g) + IE_1 \rightarrow M^{1+}(g) + 1\ e^{-}$
  - ✓ $M^{+1}(g) + IE_2 \rightarrow M^{2+}(g) + 1\ e^{-}$
    - ➢ First ionization energy = energy to remove electron from neutral atom, second IE = energy to remove from 1+ ion, etc.

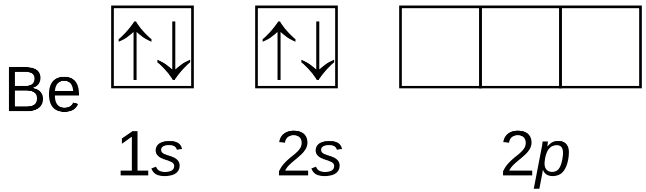

First Ionization Energies
Ionization energy (kJ/mol)
2500
2000
1500
1000
500
0
Atomic number
0
10
20
30
40
50
He
Ne
Ar
Kr
Xe
Noble gases
Li
Na
K
Rb
Alkali metals
Period 4 transition elements
Period 5 transition elements

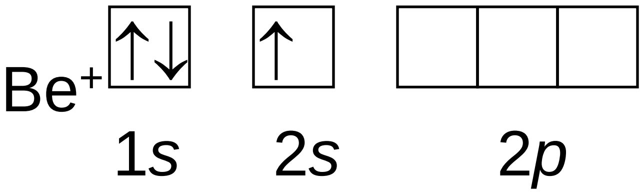
Exceptions in the First IE Trends

- First ionization energy generally increases from left to right across a period
- Except from 2A to 3A, 5A to 6A

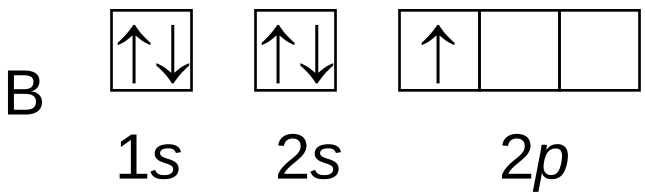


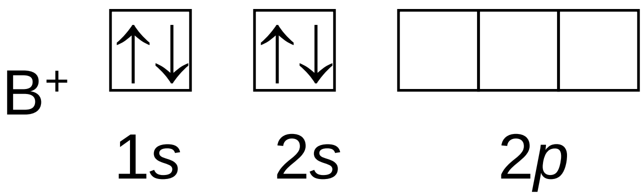
Exceptions in the First Ionization Energy Trends, Be and B





To ionize Be, you must break up a full sublevel, which costs extra energy.





When you ionize B, you get a full sublevel, which costs less energy.

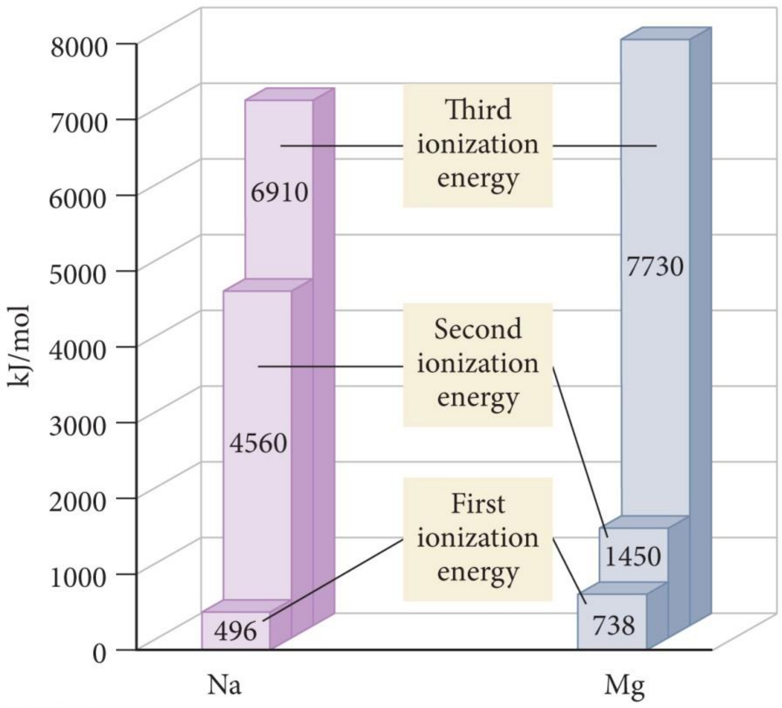
Exceptions in the First Ionization Energy Trends, N and O

To ionize N, you must break up a half-full sublevel, which costs extra energy.

When you ionize O, you get a half-full sublevel, which costs less energy.

Trends in Successive Ionization Energies

- Removal of each successive electron costs more energy.
  – Shrinkage in size due to having more protons than electrons
  – Outer electrons closer to the nucleus; therefore harder to remove
- There's a regular increase in energy for each successive valence electron.
- There's a large increase in energy when core electrons are removed.



Trends in Second and Successive Ionization Energies

TABLE 8.1 Successive Values of Ionization Energies for the Elements Sodium through Argon (kJ/mol)

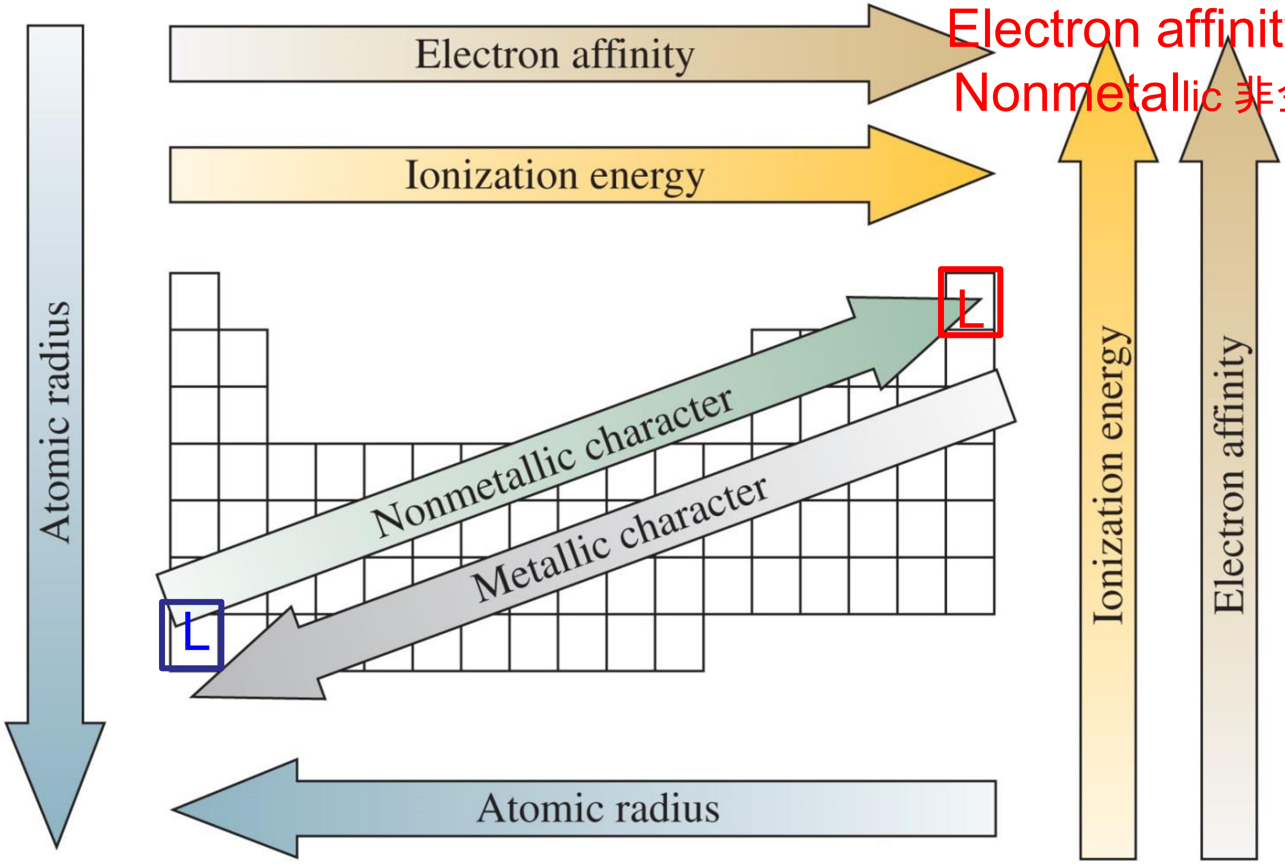
Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7
Na	496	4560					
Mg	738	1450	7730				
Al	578	1820	2750	11,600			
Si	786	1580	3230	4360	16,100		
P	1012	1900	2910	4960	6270	22,200	
S	1000	2250	3360	4560	7010	8500	27,100
Cl	1251	2300	3820	5160	6540	9460	11,000
Ar	1521	2670	3930	5770	7240	8780	12,000

Core electrons

8.8 Electron Affinities and Metallic Character

(EA)電子親和力 ✓M(g) + 1e⁻ → M^{1-}(g) + EA

Large: Ionization energy游離能
Electron affinity電子親和力
Nonmetallic 非金屬



Large: atomic radius（原子半徑）
Metallic (金屬)

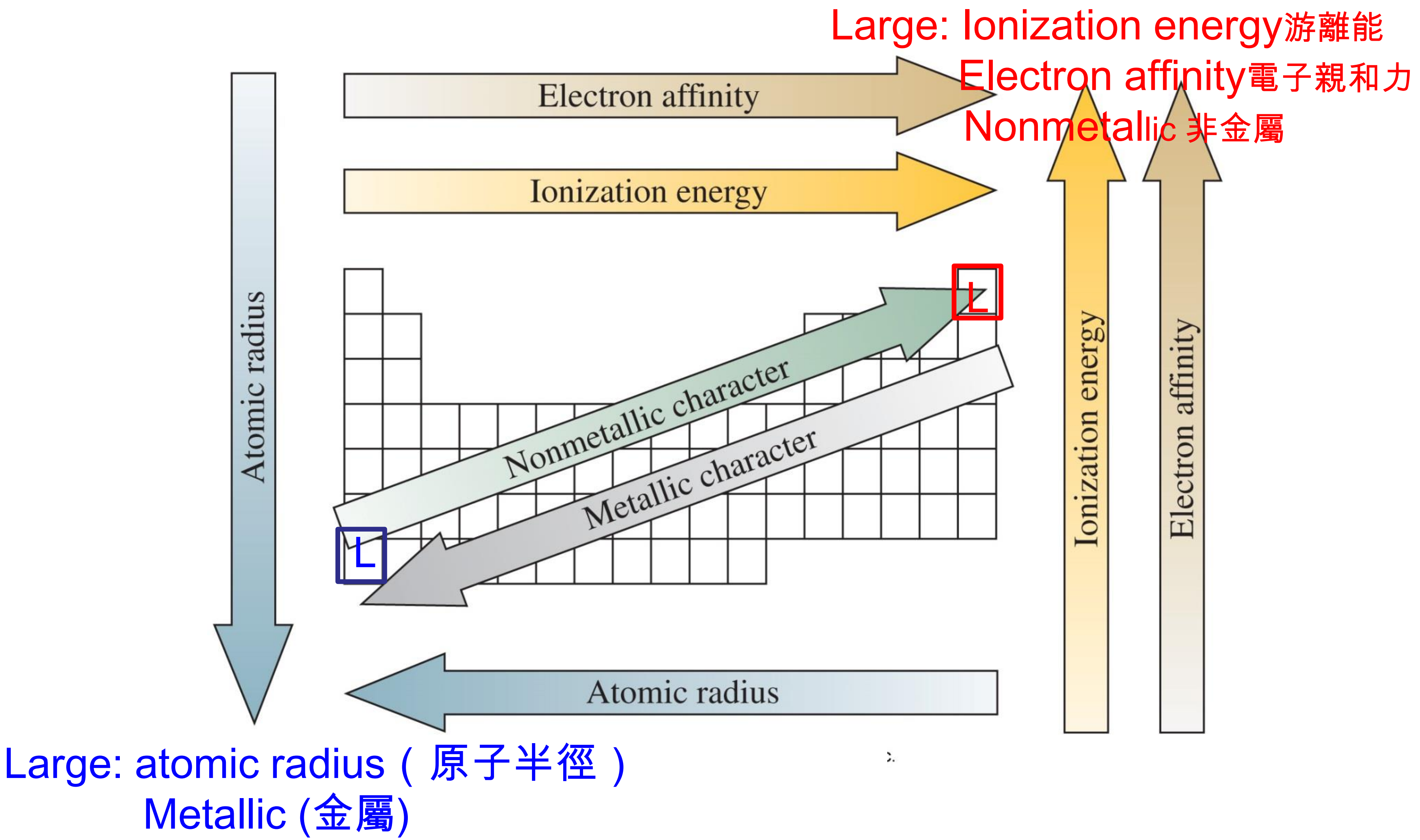
Electron Affinities (kJ/mol)

1A	2A	3A	4A	5A	6A	7A	8A
H −73							**He** >0
Li −60	**Be** >0	**B** −27	**C** −122	**N** >0	**O** −141	**F** −328	**Ne** >0
Na −53	**Mg** >0	**Al** −43	**Si** −134	**P** −72	**S** −200	**Cl** −349	**Ar** >0
K −48	**Ca** −2	**Ga** −30	**Ge** −119	**As** −78	**Se** −195	**Br** −325	**Kr** >0
Rb −47	**Sr** −5	**In** −30	**Sn** −107	**Sb** −103	**Te** −190	**I** −295	**Xe** >0

Trends in Electron Affinity

- Alkali metals decrease electron affinity down the column.
  - But not all groups do
  - Generally irregular increase in EA from second period to third period
- "Generally" increases across period
  - Becomes more negative from left to right
  - Not absolute
  - Group 5A generally lower EA than expected because extra electron must pair
  - Groups 2A and 8A generally very low EA because added electron goes into higher energy level or sublevel
- Highest EA in any period = halogen

Metallic Character



Properties of Metals and Nonmetals

- Metals
  - ✓Malleable and ductile
  - ✓Shiny, lustrous, reflect light
  - ✓Conduct heat and electricity
  - ✓Most oxides basic and ionic
  - ✓Form cations in solution
  - ✓Lose electrons in reactions – **oxidized**
- Nonmetals
  - ✓Brittle in solid state
  - ✓Dull, nonreflective solid surface
  - ✓Electrical and thermal insulators
  - ✓Most oxides are acidic and molecular
  - ✓Form anions and polyatomic anions
  - ✓Gain electrons in reactions – **reduced**

8.9 Some Examples of Periodic Chemical Behavior: The Alkali Metals, the Halogens, and the Noble Gases

Trends in the Alkali Metals,1A

TABLE 8.2 Properties of the Alkali Metals*

Element	Electron Configuration	Atomic Radius (pm)	IE_1 (kJ/mol)	Density at 25 °C (g/cm^3)	Melting Point (°C)
Li	[He] $2s^1$	152	520	0.535	181
Na	[Ne] $3s^1$	186	496	0.968	102
K	[Ar] $4s^1$	227	419	0.856	98
Rb	[Kr] $5s^1$	248	403	1.532	39
Cs	[Xe] $6s^1$	265	376	1.879	29

*Francium is omitted because it has no stable isotopes.

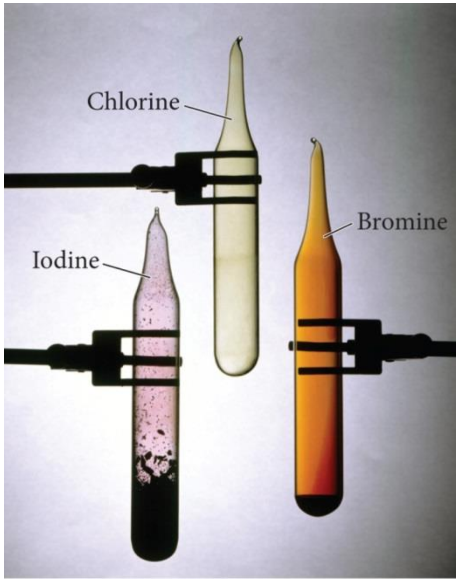
- Very low ionization energies
  - Good reducing agents; easy to oxidize
  - Very reactive; not found uncombined in nature
  - React with nonmetals to form salts
  - Compounds generally soluble in water ∴ found in seawater

Trends in the Halogens, 7A

TABLE 8.3 Properties of the Halogens*

Element	Electron Configuration	Atomic Radius (pm)	EA (kJ/mol)	Melting Point (°C)	Boiling Point (°C)	Density of Liquid (g/cm^3)
F	[He] $2s^2 2p^5$	72	−328	−219	−188	1.51
Cl	[Ne] $3s^2 3p^5$	99	−349	−101	−34	2.03
Br	[Ar] $4s^2 4p^5$	114	−325	−7	59	3.19
I	[Kr] $5s^2 5p^5$	133	−295	114	184	3.96

*At is omitted because it is rare and radioactive.

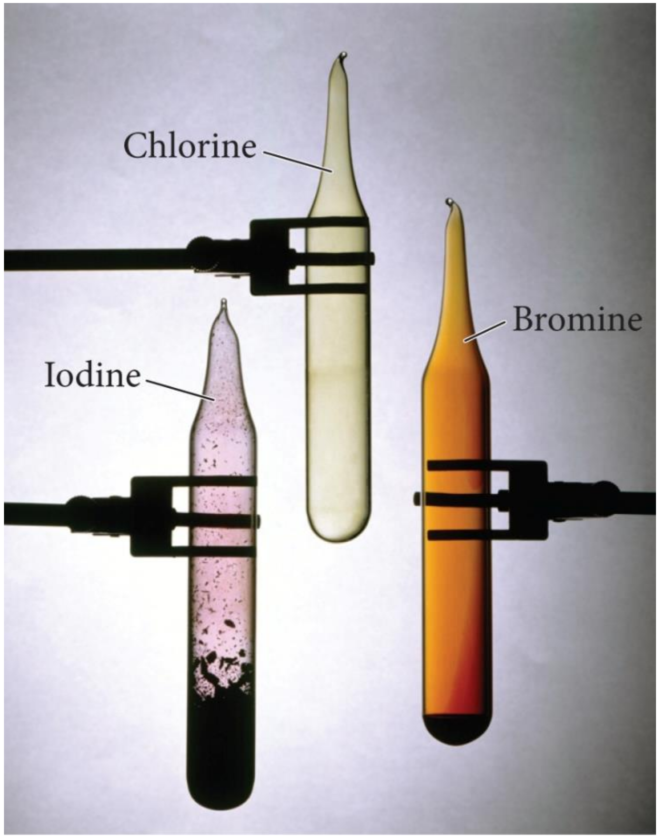


- Very high electron affinities
  - Good oxidizing agents; easy to reduce
  - Very reactive; not found uncombined in nature
  - React with metals to form salts
  - Compounds generally soluble in water ∴ found in seawater

Halogens

TABLE 8.3 Properties of the Halogens*						
Element	Electron Configuration	Atomic Radius (pm)	EA (kJ/mol)	Melting Point (°C)	Boiling Point (°C)	Density of Liquid (g/cm^3)
F	[He] $2s^2 2p^5$	72	−328	−219	−188	1.51
Cl	[Ne] $3s^2 3p^5$	99	−349	−101	−34	2.03
Br	[Ar] $4s^2 4p^5$	114	−325	−7	59	3.19
I	[Kr] $5s^2 5p^5$	133	−295	114	184	3.96

*At is omitted because it is rare and radioactive.



Reactions of Alkali Metals with Halogens

$$2\ Na\,(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$$

- Alkali metals are oxidized to the 1+ ion.
- Halogens are reduced to the 1− ion.
- The ions then attach together by ionic bonds.
- The reaction is exothermic.

Reactions of Alkali Metals with Water

$$2\ M(s) + 2\ H_2O(l) \rightarrow M^+(aq) + 2\ OH^-(aq) + H_2(g)$$

- Alkali metals are oxidized to the 1+ ion.
- H_2O is split into $H_2(g)$ and OH^- ion.
- The Li, Na, and K are less dense than the water, so they float on top.
- The ions then attach together by ionic bonds.
- The reaction is exothermic, and often the heat released ignites the $H_2(g)$.

Reactions of the Alkali Metals with Water

Lithium Sodium Potassium

Trends in the Noble Gases, 8A

TABLE 8.4 Properties of the Noble Gases*

Element	Electron Configuration	Atomic Radius (pm)**	IE_1 (kJ/mol)	Boiling Point (K)	Density of Gas (g/L at STP)
He	$1s^2$	32	2372	4.2	0.18
Ne	$[He]2s^2 2p^6$	70	2081	27.1	0.90
Ar	$[Ne]3s^2 3p^6$	98	1521	87.3	1.78
Kr	$[Ar]4s^2 4p^6$	112	1351	119.9	3.74
Xe	$[Kr]5s^2 5p^6$	130	1170	165.1	5.86

*Radon is omitted because it is radioactive.

**Since only the heavier noble gases form compounds, covalent radii for the smaller noble gases are estimated.

Liquid helium, a cryogenic liquid, cools substances to temperatures as low as 1.2 K.

- Very unreactive
  - Only found uncombined in nature
  - Used as "inert" atmosphere when reactions with other gases would be undesirable.
- Melting point and boiling point increase down the column.
  - All gases at room temperature
  - Very low boiling points

Xxpg339 Coulomb's Law

$$E = \frac{1}{4\pi\varepsilon_0}\frac{q_1 q_2}{r}$$

- Coulomb's law describes the attractions and repulsions between charged particles.
- For like charges, the potential energy (E) is positive and decreases as the particles get farther apart as *r* increases.
- For opposite charges, the potential energy is negative and becomes more negative as the particles get closer together.
- The strength of the interaction increases as the size of the charges increases.
  - Electrons are more strongly attracted to a nucleus with a 2+ charge than a nucleus with a 1+ charge.