Chapter 8

Electron Configuration and Chemical Periodicity

Electron Configuration and Chemical Periodicity

8.1 Characteristics of Many-Electron Atoms

8.2 The Quantum-Mechanical Model and the Periodic Table

8.3 Trends in Three Atomic Properties

8.4 Atomic Properties and Chemical Reactivity

Figure 8.1 The effect of electron spin.

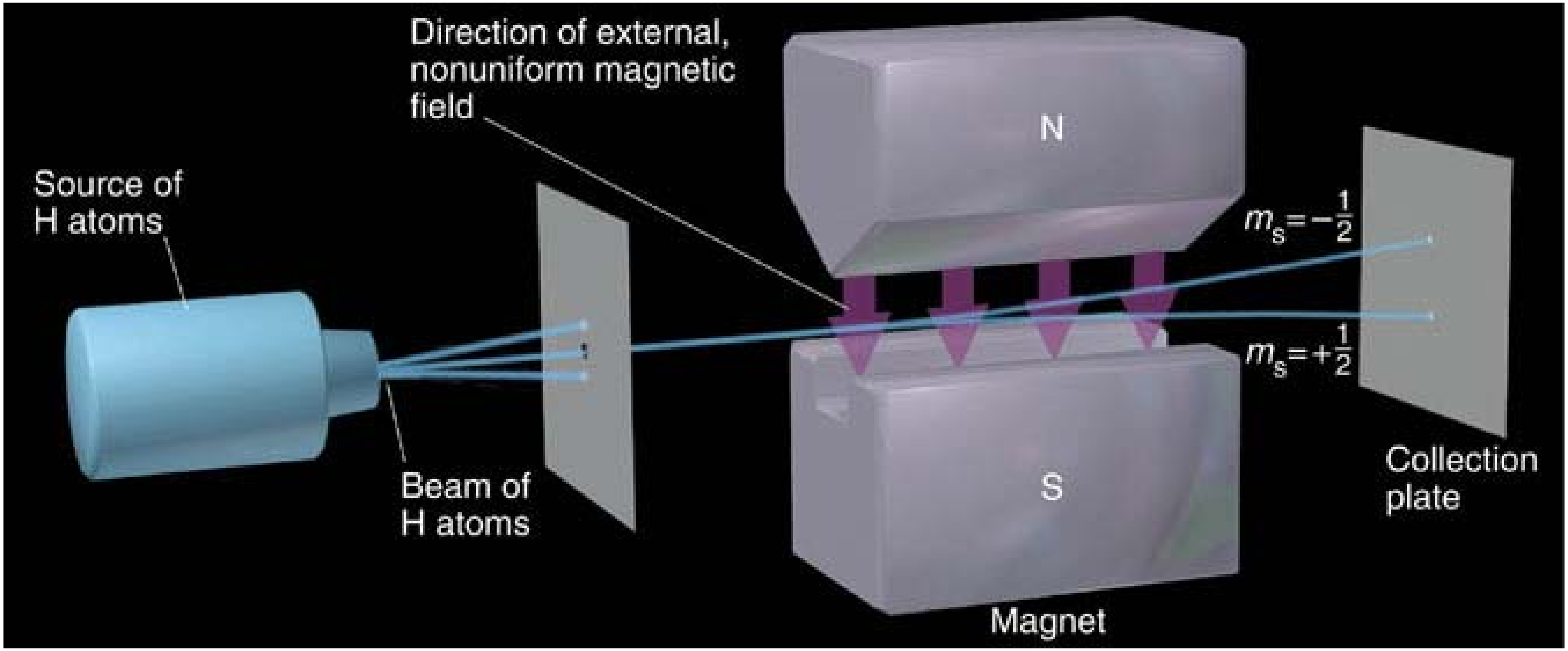


Table 8.1 Summary of Quantum Numbers of Electrons in Atoms

Name	Symbol	Permitted Values	Property
principal	n	positive integers (1, 2, 3, …)	orbital energy (size)
angular momentum	l	integers from 0 to n-1	orbital shape (The l values 0, 1, 2, and 3 correspond to *s*, *p*, *d*, and *f* orbitals, respectively.)
magnetic	m_l	integers from $-l$ to 0 to $+l$	orbital orientation
spin	m_s	+½ or -½	direction of e^- spin

Quantum Numbers and The Exclusion Principle

Each electron in any atom is described completely by a set of ***four*** quantum numbers.

The first three quantum numbers describe the orbital, while the fourth quantum number describes electron spin.

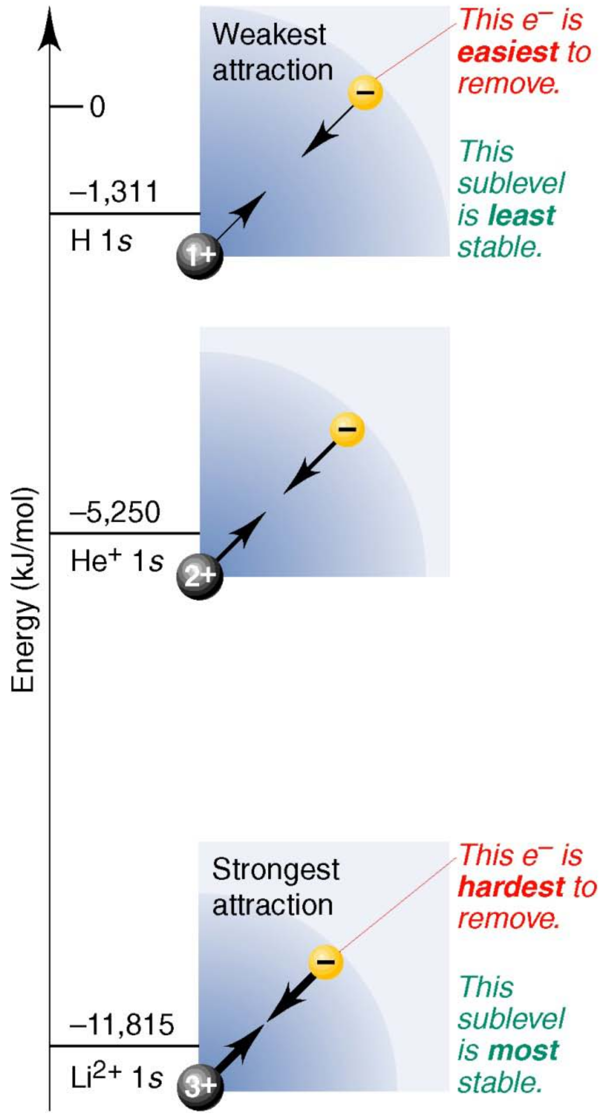
Pauli's **exclusion principle** states that *no two electrons in the same atom can have the same four quantum numbers*.

An atomic orbital can hold a ***maximum of two electrons*** and they must have ***opposing spins***.

Factors Affecting Atomic Orbital Energies

- The energies of atomic orbitals are affected by
  - nuclear charge (*Z*) and
  - shielding by other electrons.
- A **higher nuclear charge** increases nucleus-electron interactions and lowers sublevel energy.
- **Shielding** by other electrons reduces the full nuclear charge to an ***effective nuclear charge*** ($\boldsymbol{Z_{\mathrm{eff}}}$).
  - Z_{eff} is the nuclear charge an electron actually experiences.
- **Orbital shape** also affects sublevel energy.

Figure 8.2 The effect of nuclear charge on sublevel energy.



Greater nuclear charge lowers sublevel energy.

It takes *more energy* to remove the 1*s* electron from He^+ than from H.

Shielding and Orbital Energy

- Electrons in the ***same*** energy level shield each other to some extent.
- Electrons in ***inner*** energy levels shield the outer electrons very effectively.
  - The further from the nucleus an electron is, the lower the Z_{eff} for that particular electron.

Figure 8.3 Shielding and energy levels.

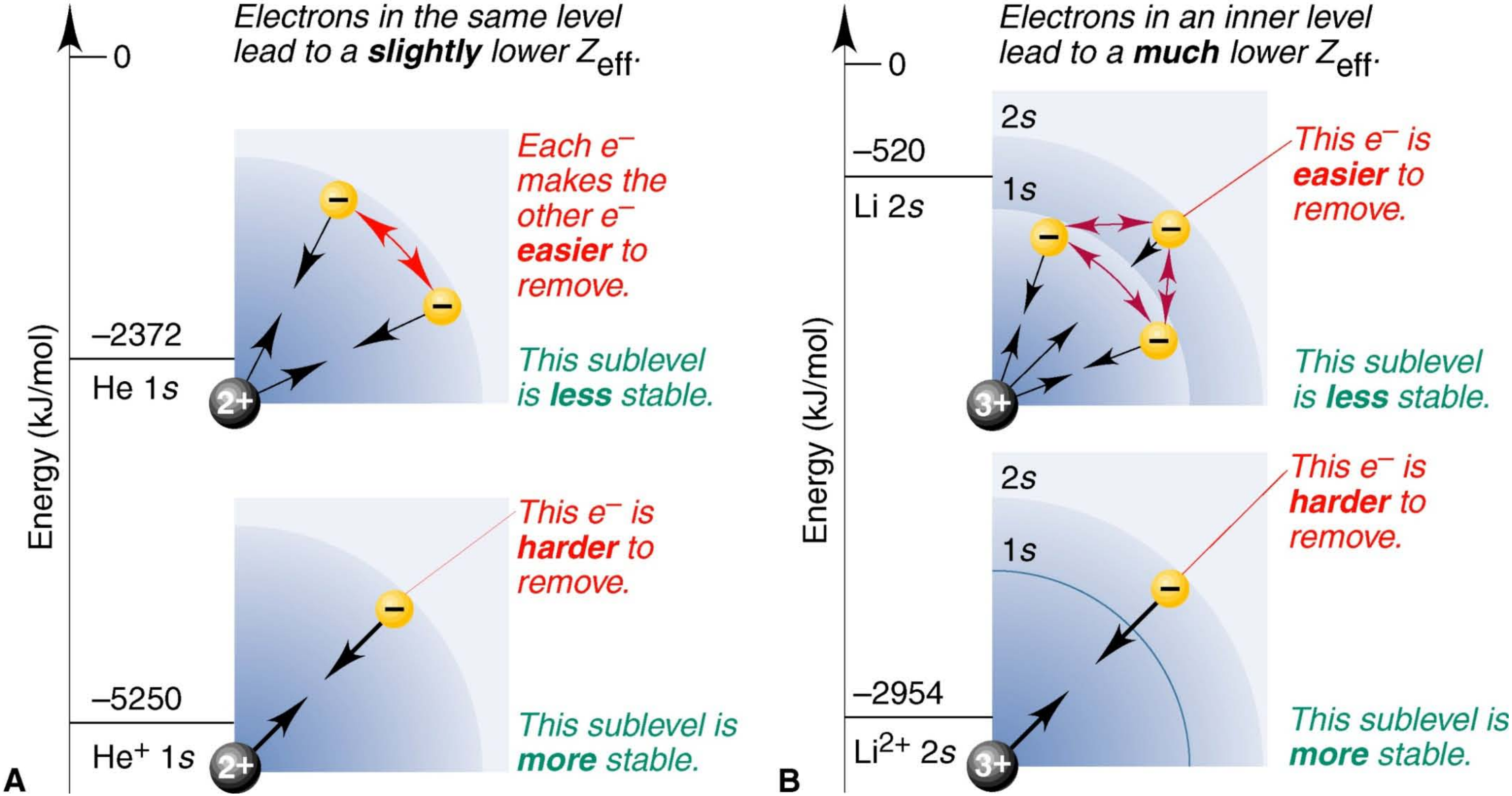
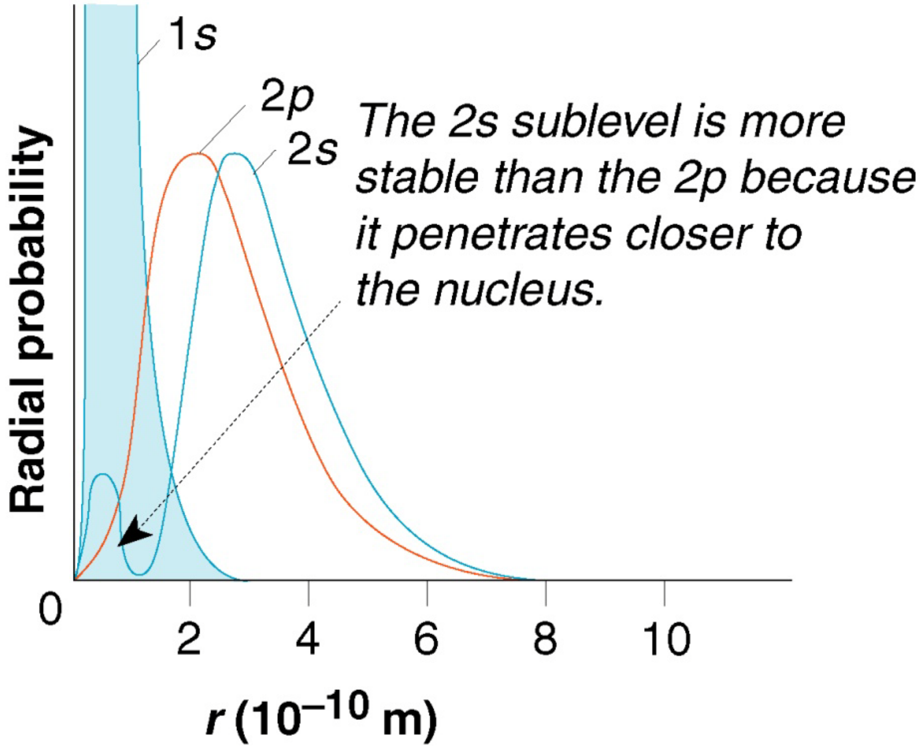


Figure 8.4 Penetration and sublevel energy.



Orbital shape causes electrons in some orbitals to "penetrate" close to the nucleus.

Penetration increases nuclear attraction and decreases shielding.

Splitting of Levels into Sublevels

Each energy level is split into ***sublevels*** of differing energy.

Splitting is caused by penetration and its effect on shielding.

For a given n value, a lower l value indicates a lower energy sublevel.

Order of sublevel energies: $\boldsymbol{s < p < d < f}$

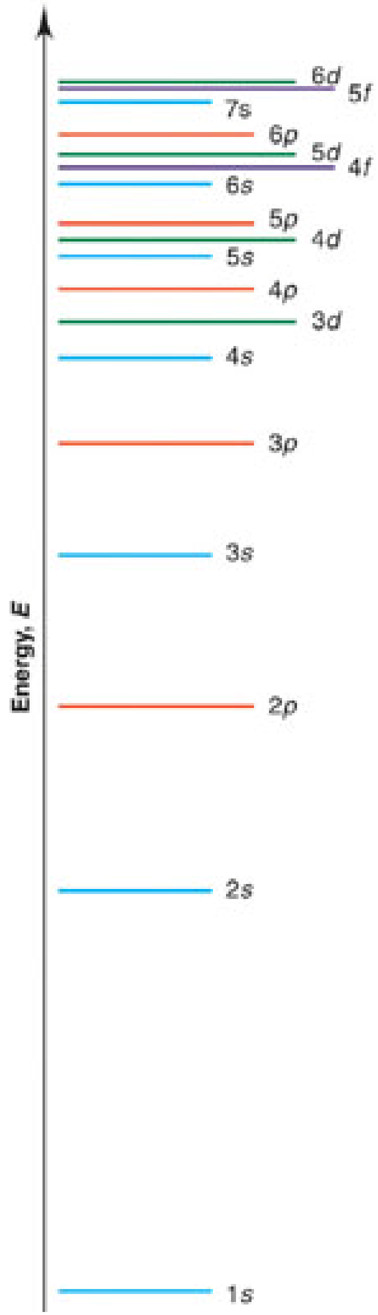


Figure 8.5

Order for filling energy sublevels with electrons.

In general, energies of sublevels increase as *n* increases (1 < 2 < 3, etc.)
and as *l* increases (*s* < *p* < *d* < *f*).

As *n* increases, some sublevels overlap.

Electron Configurations and Orbital Diagrams

Electron configuration is indicated by a shorthand notation:

$$n\,l^{\#}$$

# of electrons in the sublevel

as *s*, *p*, *d*, *f*

Orbital diagrams make use of a box, circle, or line for each orbital in the energy level. An arrow is used to represent an electron *and* its spin.

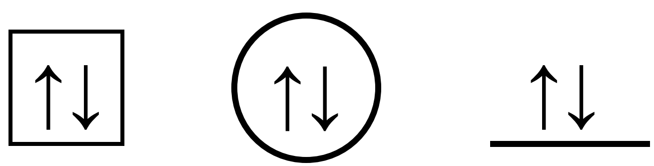
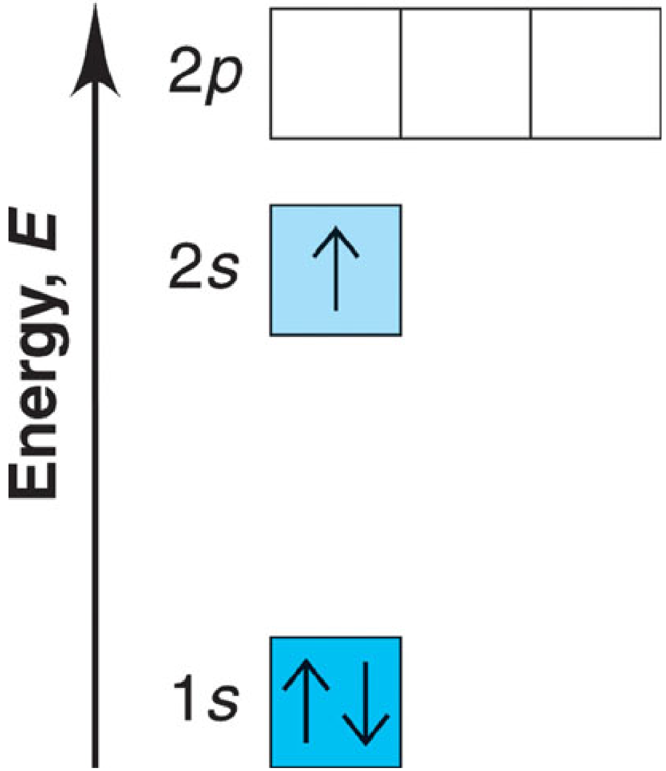


Figure 8.6 **A vertical orbital diagram for the Li ground state.**



Building Orbital Diagrams

The **aufbau principle** is applied – electrons are always placed in the lowest energy sublevel available.

H (Z = 1) $1s^1$ $\boxed{\uparrow}$

$1s$

The **exclusion principle** states that each orbital may contain a maximum of 2 electrons, which must have opposite spins.

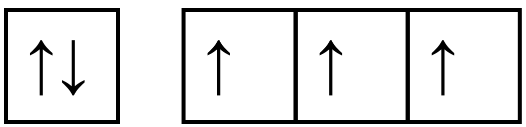
He (Z = 2) $1s^2$ $\boxed{\uparrow\downarrow}$

$1s$

Building Orbital Diagrams

Hund's rule specifies that when orbitals of equal energy are available, the lowest energy electron configuration has the maximum number of unpaired electrons with parallel spins.

N (Z = 7) $1s^22s^22p^3$



$2s$ $2p$

Sample Problem 8.1

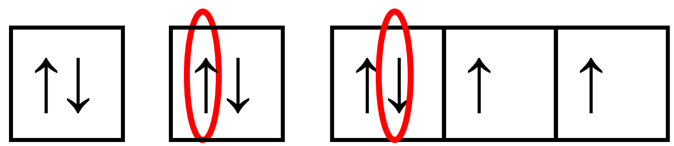
Determining Quantum Numbers from Orbital Diagrams

PROBLEM: Write a set of quantum numbers for the third electron and a set for the eighth electron of the F atom.

PLAN: Identify the electron of interest and note its level (n), sublevel, (l), orbital (m_l) and spin (m_s). Count the electrons in the order in which they are placed in the diagram.

SOLUTION:

F (Z = 9) $1s^2 2s^2 2p^3$

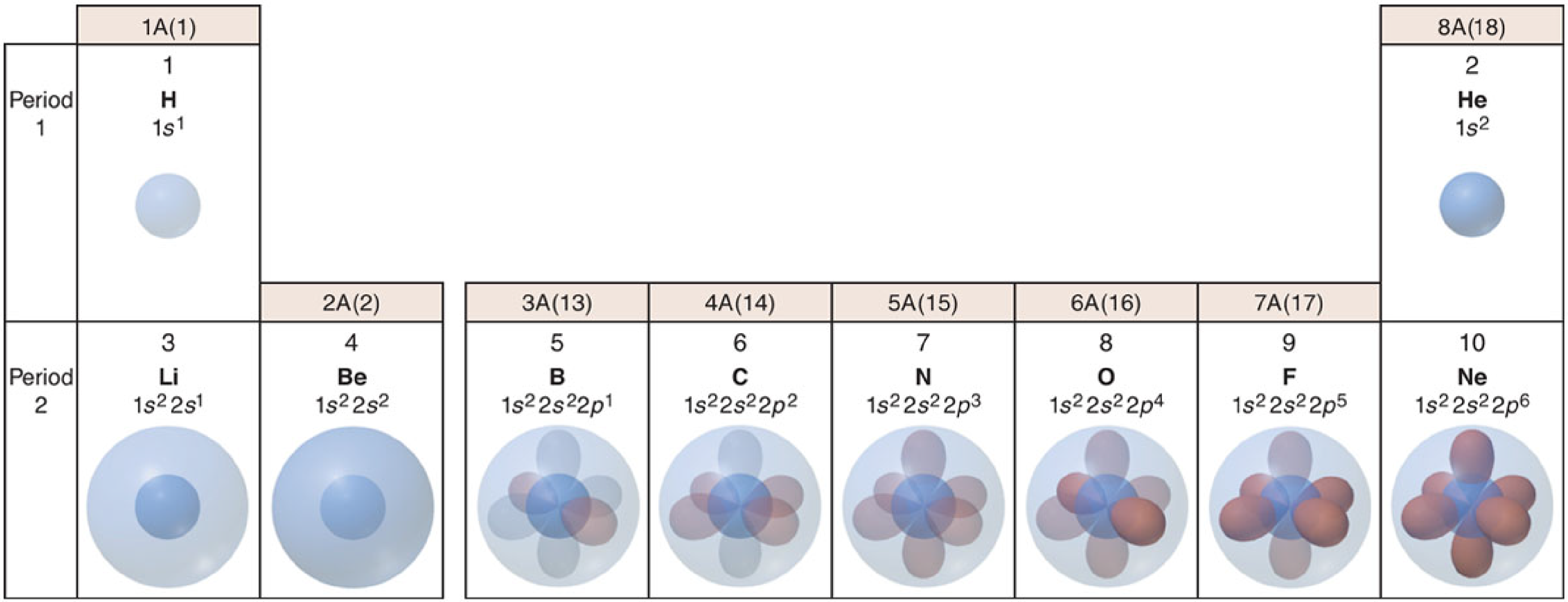


1s 2s 2p

For the 3rd electron: n = 2, l = 0, m_l = 0, m_s = +½

For the 8th electron: n = 2, l = 1, m_l = -1, m_s = -½

Figure 8.7 Depicting orbital occupancy for the first 10 elements.



Partial Orbital Diagrams and Condensed Configurations

A ***partial orbital diagram*** shows only the highest energy sublevels being filled.

Al (Z = 13) $1s^2 2s^2 2p^6 3s^2 3p^1$

↑↓		↑		
$3s$			$3p$	

A ***condensed electron configuration*** has the element symbol of the ***previous*** noble gas in square brackets.

Al has the condensed configuration [Ne]$3s^2 3p^1$

Table 8.2 Partial Orbital Diagrams and Electron Configurations* for the Elements in Period 3.

Atomic Number	Element	Partial Orbital Diagram (3*s* and 3*p* Sublevels Only)		Full Electron Configuration†	Condensed Electron Configuration
		3*s*	3*p*		
11	Na	↑	_ _ _	$[1s^2 2s^2 2p^6]\ 3s^1$	$[Ne]\ 3s^1$
12	Mg	↑↓	_ _ _	$[1s^2 2s^2 2p^6]\ 3s^2$	$[Ne]\ 3s^2$
13	Al	↑↓	↑ _ _	$[1s^2 2s^2 2p^6]\ 3s^2 3p^1$	$[Ne]\ 3s^2 3p^1$
14	Si	↑↓	↑ ↑ _	$[1s^2 2s^2 2p^6]\ 3s^2 3p^2$	$[Ne]\ 3s^2 3p^2$
15	P	↑↓	↑ ↑ ↑	$[1s^2 2s^2 2p^6]\ 3s^2 3p^3$	$[Ne]\ 3s^2 3p^3$
16	S	↑↓	↑↓ ↑ ↑	$[1s^2 2s^2 2p^6]\ 3s^2 3p^4$	$[Ne]\ 3s^2 3p^4$
17	Cl	↑↓	↑↓ ↑↓ ↑	$[1s^2 2s^2 2p^6]\ 3s^2 3p^5$	$[Ne]\ 3s^2 3p^5$
18	Ar	↑↓	↑↓ ↑↓ ↑↓	$[1s^2 2s^2 2p^6]\ 3s^2 3p^6$	$[Ne]\ 3s^2 3p^6$

*Colored type indicates the sublevel to which the last electron is added.

Electron Configuration and Group

Elements in the same group of the periodic table have the same outer electron configuration.

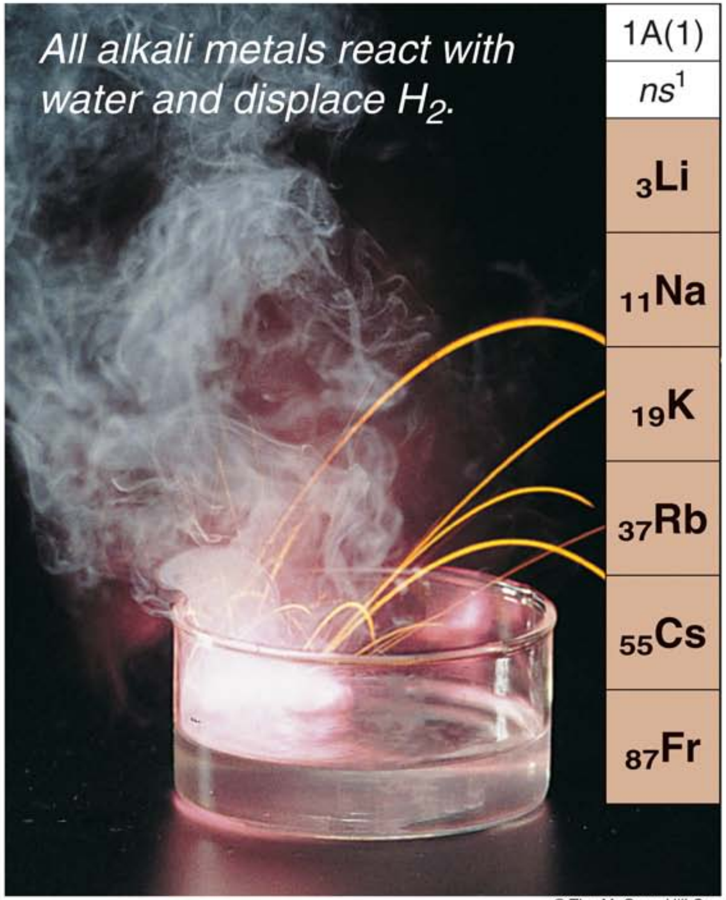
Elements in the same group of the periodic table exhibit similar chemical behavior.

Similar outer electron configurations correlate with similar chemical behavior.

Figure 8.8 **Condensed electron configurations in the first three periods.**

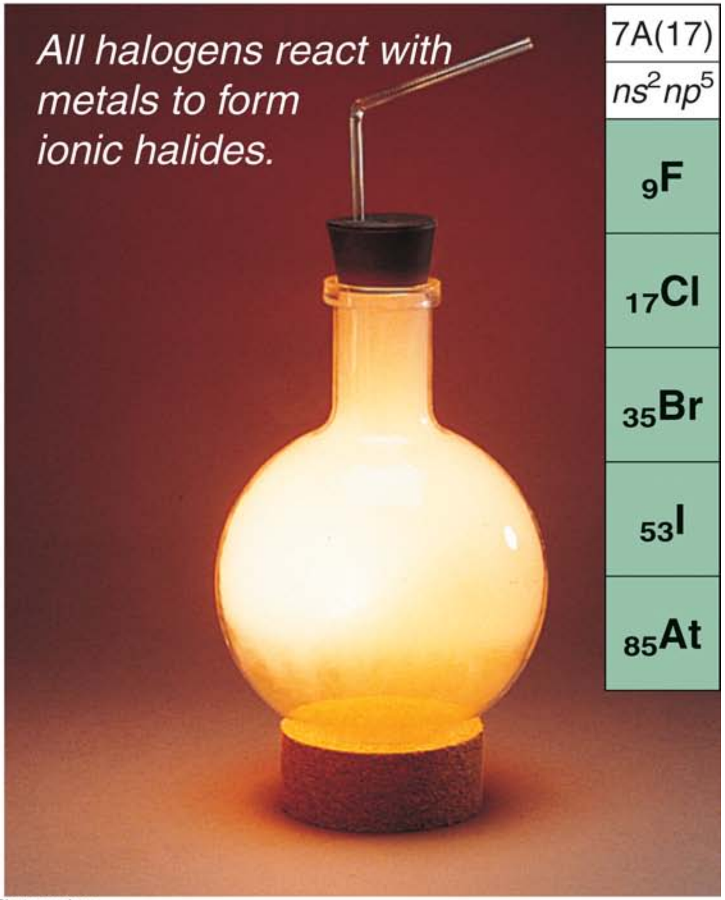
Period	1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
1	1 **H** $1s^1$							2 **He** $1s^2$
2	3 **Li** [He] $2s^1$	4 **Be** [He] $2s^2$	5 **B** [He] $2s^2 2p^1$	6 **C** [He] $2s^2 2p^2$	7 **N** [He] $2s^2 2p^3$	8 **O** [He] $2s^2 2p^4$	9 **F** [He] $2s^2 2p^5$	10 **Ne** [He] $2s^2 2p^6$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

Figure 8.9 Similar reactivities in a group.



A

Potassium reacting with water.



B

Chlorine reacting with potassium.

Table 8.3 Partial Orbital Diagrams and Electron Configurations* for the Elements in Period 4.

Atomic Number	Element	Partial Orbital Diagram (4*s*, 3*d*, and 4*p* Sublevels Only)									Full Electron Configuration	Condensed Electron Configuration
		4*s*	3*d*					4*p*				
19	K	↑									$1s^22s^22p^63s^23p^64s^1$	$[Ar]\ 4s^1$
20	Ca	↑↓									$1s^22s^22p^63s^23p^64s^2$	$[Ar]\ 4s^2$
21	Sc	↑↓	↑								$1s^22s^22p^63s^23p^64s^23d^1$	$[Ar]\ 4s^23d^1$
22	Ti	↑↓	↑	↑							$1s^22s^22p^63s^23p^64s^23d^2$	$[Ar]\ 4s^23d^2$
23	V	↑↓	↑	↑	↑						$1s^22s^22p^63s^23p^64s^23d^3$	$[Ar]\ 4s^23d^3$
24	Cr	↑	↑	↑	↑	↑	↑				$1s^22s^22p^63s^23p^64s^13d^5$	$[Ar]\ 4s^13d^5$
25	Mn	↑↓	↑	↑	↑	↑	↑				$1s^22s^22p^63s^23p^64s^23d^5$	$[Ar]\ 4s^23d^5$
26	Fe	↑↓	↑↓	↑	↑	↑	↑				$1s^22s^22p^63s^23p^64s^23d^6$	$[Ar]\ 4s^23d^6$
27	Co	↑↓	↑↓	↑↓	↑	↑	↑				$1s^22s^22p^63s^23p^64s^23d^7$	$[Ar]\ 4s^23d^7$

*Colored type indicates the sublevel to which the last electron is added.

Table 8.3 Partial Orbital Diagrams and Electron Configurations* for the Elements in Period 4.

Atomic Number	Element	Partial Orbital Diagram (4*s*, 3*d*, and 4*p* Sublevels Only)			Full Electron Configuration	Condensed Electron Configuration
		4s	3d	4p		
28	Ni	[↑↓]	[↑↓][↑↓][↑↓][↑][↑]	[ ][ ][ ]	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$	$[Ar]\ 4s^2 3d^8$
29	Cu	[↑]	[↑↓][↑↓][↑↓][↑↓][↑↓]	[ ][ ][ ]	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$	$[Ar]\ 4s^1 3d^{10}$
30	Zn	[↑↓]	[↑↓][↑↓][↑↓][↑↓][↑↓]	[ ][ ][ ]	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$	$[Ar]\ 4s^2 3d^{10}$
31	Ga	[↑↓]	[↑↓][↑↓][↑↓][↑↓][↑↓]	[↑][ ][ ]	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$	$[Ar]\ 4s^2 3d^{10} 4p^1$
32	Ge	[↑↓]	[↑↓][↑↓][↑↓][↑↓][↑↓]	[↑][↑][ ]	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$	$[Ar]\ 4s^2 3d^{10} 4p^2$
33	As	[↑↓]	[↑↓][↑↓][↑↓][↑↓][↑↓]	[↑][↑][↑]	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$	$[Ar]\ 4s^2 3d^{10} 4p^3$
34	Se	[↑↓]	[↑↓][↑↓][↑↓][↑↓][↑↓]	[↑↓][↑][↑]	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$	$[Ar]\ 4s^2 3d^{10} 4p^4$
35	Br	[↑↓]	[↑↓][↑↓][↑↓][↑↓][↑↓]	[↑↓][↑↓][↑]	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$	$[Ar]\ 4s^2 3d^{10} 4p^5$
36	Kr	[↑↓]	[↑↓][↑↓][↑↓][↑↓][↑↓]	[↑↓][↑↓][↑↓]	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$	$[Ar]\ 4s^2 3d^{10} 4p^6$

*Colored type indicates the sublevel to which the last electron is added.

Figure 8.10 A periodic table of partial ground-state electron configurations.

Main-Group Elements (*s* block) | Transition Elements (*d* block) | Main-Group Elements (*p* block)

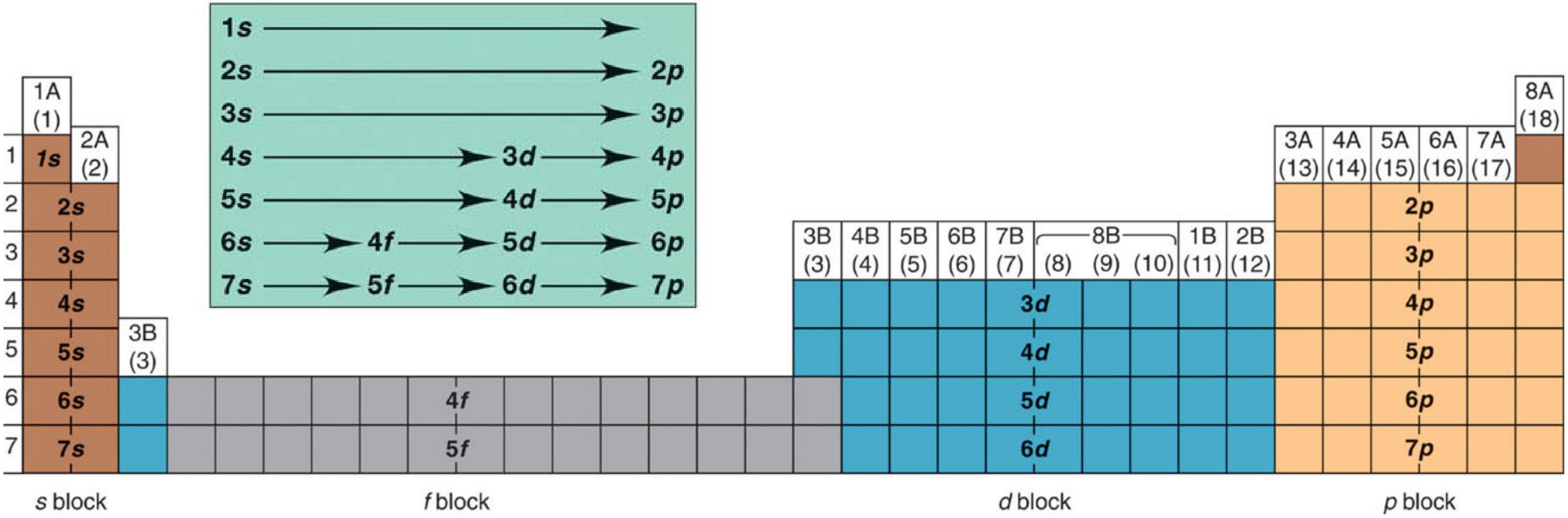
Period number: highest occupied energy level

Period	1A (1)	2A (2)	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	8B (9)	8B (10)	1B (11)	2B (12)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
	ns^1	ns^2											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 **Cn** $7s^26d^{10}$	113 $7s^27p^1$	114 $7s^27p^2$	115 $7s^27p^3$	116 $7s^27p^4$		118 $7s^27p^6$

Inner Transition Elements (*f* block)

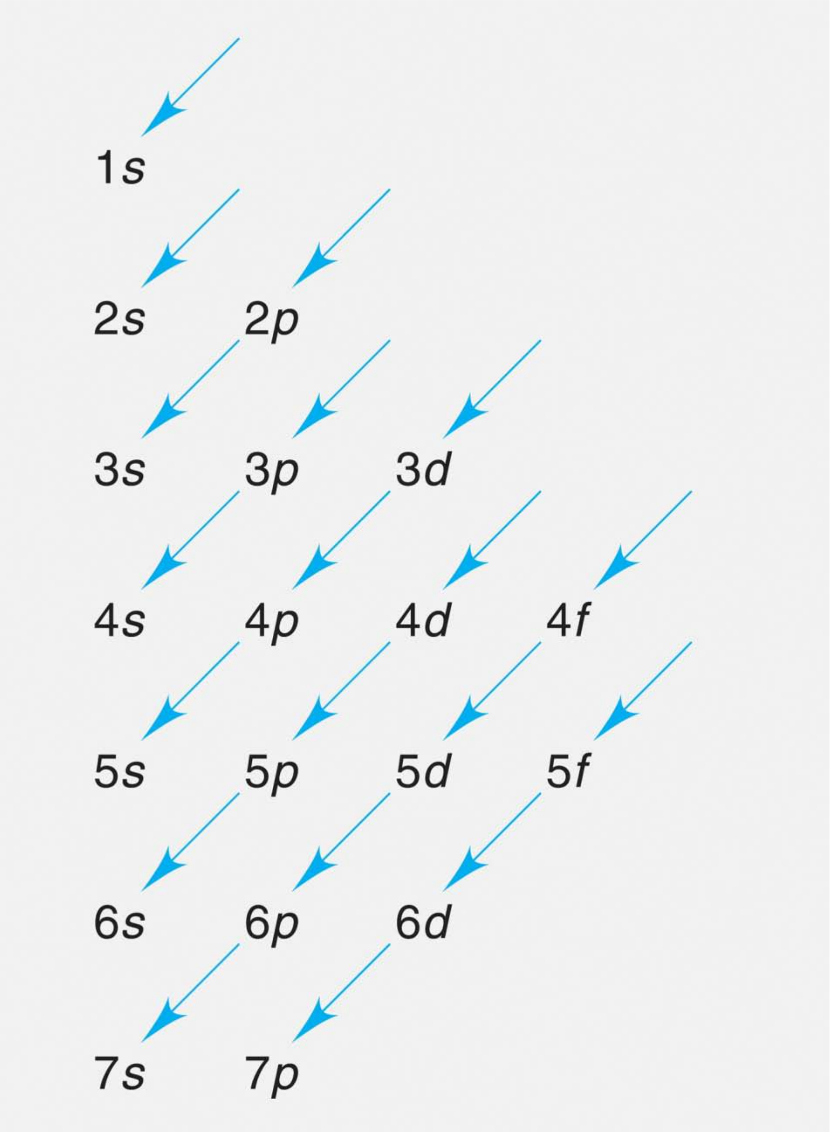
Period															
6	*Lanthanides	58 **Ce** $6s^24f^15d^1$	59 **Pr** $6s^24f^3$	60 **Nd** $6s^24f^4$	61 **Pm** $6s^24f^5$	62 **Sm** $6s^24f^6$	63 **Eu** $6s^24f^7$	64 **Gd** $6s^24f^75d^1$	65 **Tb** $6s^24f^9$	66 **Dy** $6s^24f^{10}$	67 **Ho** $6s^24f^{11}$	68 **Er** $6s^24f^{12}$	69 **Tm** $6s^24f^{13}$	70 **Yb** $6s^24f^{14}$	71 **Lu** $6s^24f^{14}5d^1$
7	**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** $7s^25f^{14}6d^1$

Figure 8.11 **Orbital filling and the periodic table.**



The order in which the orbitals are filled can be obtained directly from the periodic table.

Aid to memorizing sublevel filling order.



The *n* value is constant horizontally.
The *l* value is constant vertically.
n + *l* is constant diagonally.

Categories of Electrons

Inner (core) electrons are those an atom has in common with the pervious noble gas and any ***completed*** transition series.

Outer electrons are those in the ***highest*** energy level (highest *n* value).

Valence electrons are those involved in forming compounds.

For **main group** elements, the valence electrons ***are*** the outer electrons.

For **transition elements**, the valence electrons include the outer electrons and any (n-1)d electrons.

Sample Problem 8.2 **Determining Electron Configurations**

PROBLEM: Using the periodic table on the inside cover of the text (not Figure 8.10 or Table 8.3), give the full and condensed electron configurations, partial orbital diagrams showing valence electrons only, and number of inner electrons for the following elements:

(a) potassium (K; Z = 19) **(b)** technetium (Tc; Z = 43) **(c)** lead (Pb; Z = 82)

PLAN: The atomic number gives the number of electrons, and the periodic table shows the order for filling orbitals. The partial orbital diagram includes all electrons added after the previous noble gas except those in filled inner sublevels.

Sample Problem 8.2

SOLUTION:

(a) For K (Z = 19)

full configuration $1s^22s^22p^63s^23p^64s^1$

condensed configuration [Ar] $4s^1$

partial orbital diagram

There are 18 inner electrons.

Sample Problem 8.2

SOLUTION:

(b) For Tc ($Z = 43$)

full configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^5$

condensed configuration $[Kr]5s^2 4d^5$

partial orbital diagram

5s	4d					5p		
↑↓	↑	↑	↑	↑	↑			

There are 36 inner electrons.

Sample Problem 8.2

SOLUTION:

(a) For Pb (Z = 82)

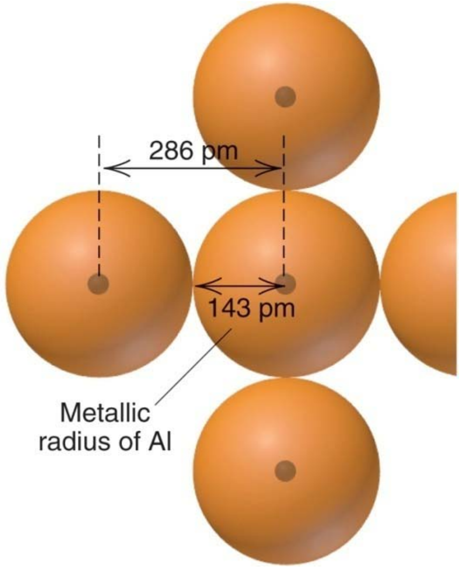
full configuration $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^2$

condensed configuration [Xe] $6s^24f^{14}5d^{10}6p^2$

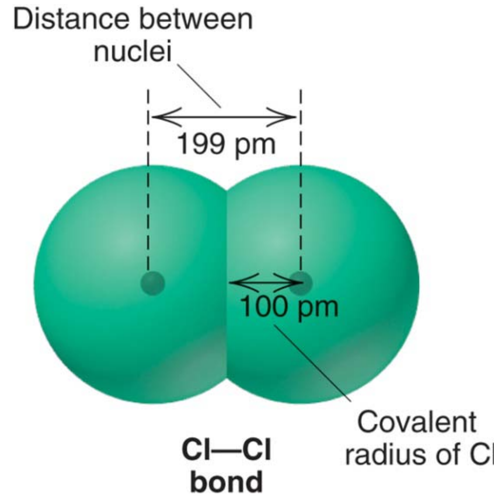
partial orbital diagram	↑↓	↑	↑	
	$6s$		$6p$	

There are 78 inner electrons.

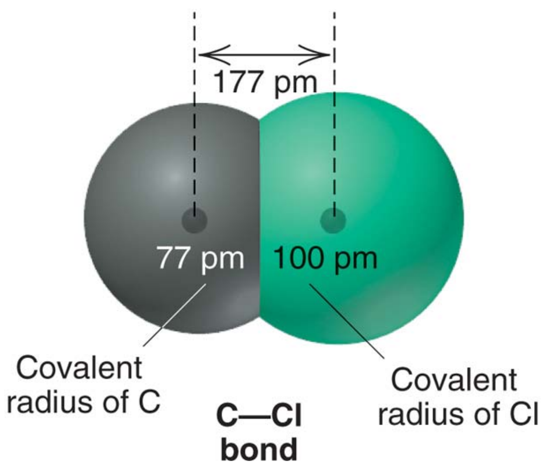
Figure 8.12 Defining atomic size.



A. The metallic radius of aluminum.



B. The covalent radius of chlorine.



C. Known covalent radii and distances between nuclei can be used to find unknown radii.

Trends in Atomic Size

Atomic size ***increases*** as the principal quantum number *n* ***increases***.

As *n* increases, the probability that the outer electrons will be further from the nucleus increases.

Atomic size ***decreases*** as the effective nuclear charge Z_{eff} ***increases***.

As Z_{eff} increases, the outer electrons are pulled closer to the nucleus.

For ***main group elements***:

atomic size ***increases*** down a group in the periodic table and ***decreases*** across a period.

Figure 8.13

Atomic radii of the main-group and transition elements.

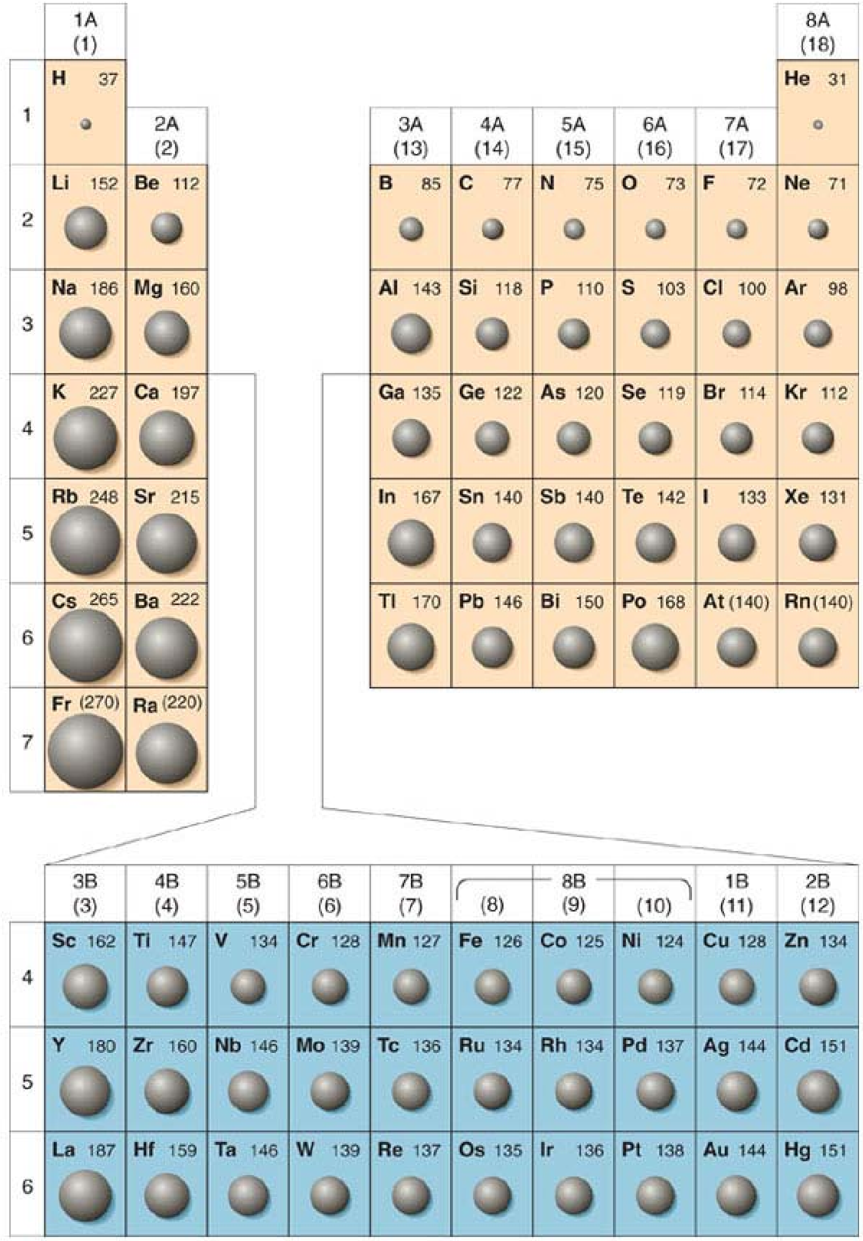
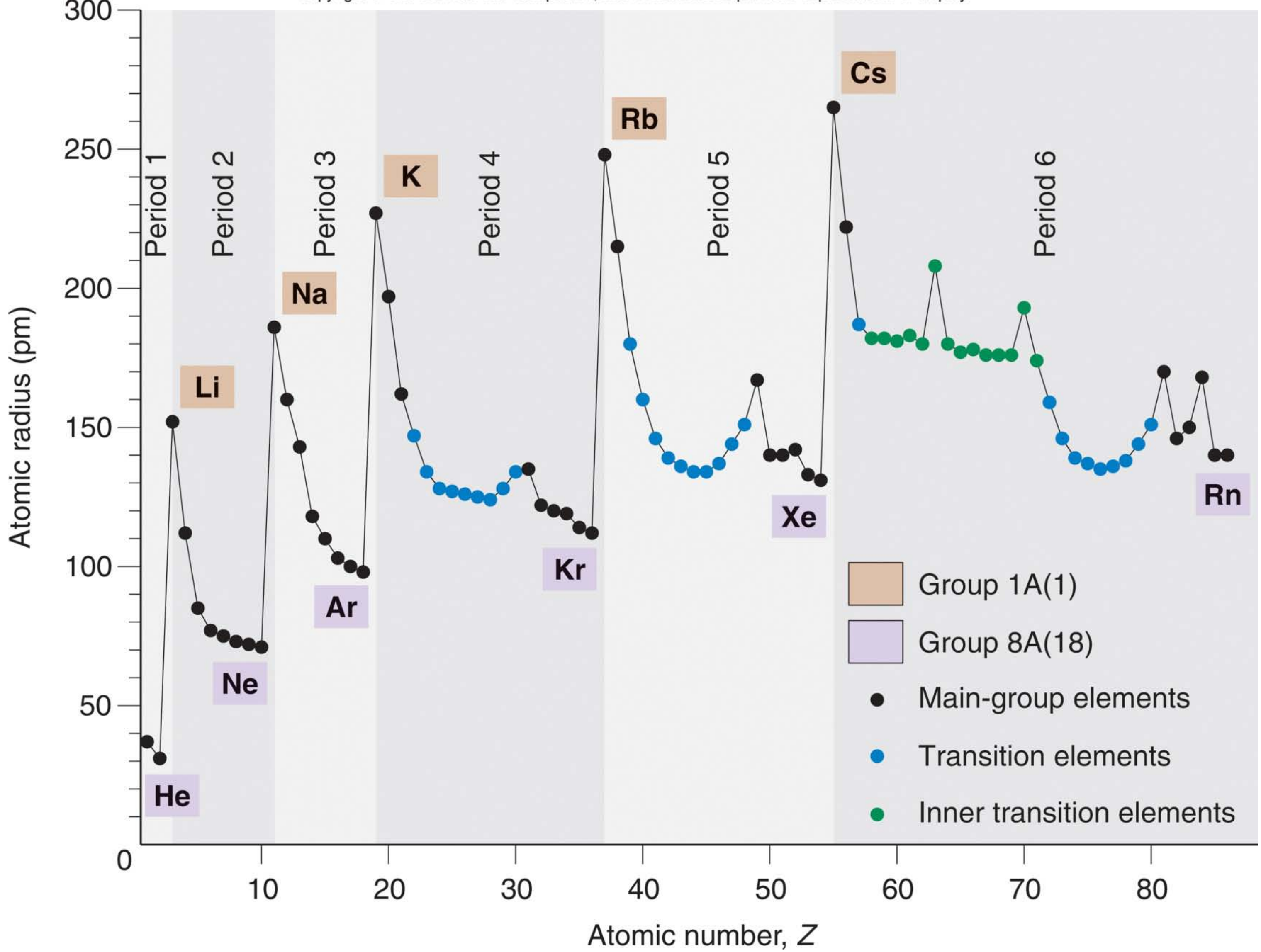


Figure 8.14 Periodicity of atomic radius.



Sample Problem 8.3 **Ranking Elements by Atomic Size**

PROBLEM: Using only the periodic table (not Figure 8.15), rank each set of main-group elements in order of *decreasing* atomic size:

(a) Ca, Mg, Sr **(b)** K, Ga, Ca

(c) Br, Rb, Kr **(d)** Sr, Ca, Rb

PLAN: Locate each element on the periodic table. Main-group elements increase in size down a group and decrease in size across the period.

Sample Problem 8.3

SOLUTION:

(a) Sr > Ca > Mg
Ca, Mg, and Sr are in Group 2A. Size increases down the group.

(b) K > Ca > Ga
K, Ga, and Ca are all in Period 4. Size decreases across the period.

(c) Rb > Br > Kr
Rb is the largest because it has one more energy level than the other elements. Kr is smaller than Br because Kr is further to the right in the same period.

(d) Rb > Sr > Ca
Ca is the smallest because it has one fewer energy level. Sr is smaller than Rb because it is smaller to the right in the same period.

Trends in Ionization Energy

Ionization energy (IE) is the energy required for the ***complete removal*** of 1 mol of electrons from 1 mol of gaseous atoms or ions.

Atoms with a ***low IE*** tend to form ***cations***.
Atoms with a ***high IE*** tend to form ***anions*** (except the noble gases).

Ionization energy tends to ***decrease*** down a group and ***increase*** across a period.

Figure 8.15 Periodicity of first ionization energy (IE_1).

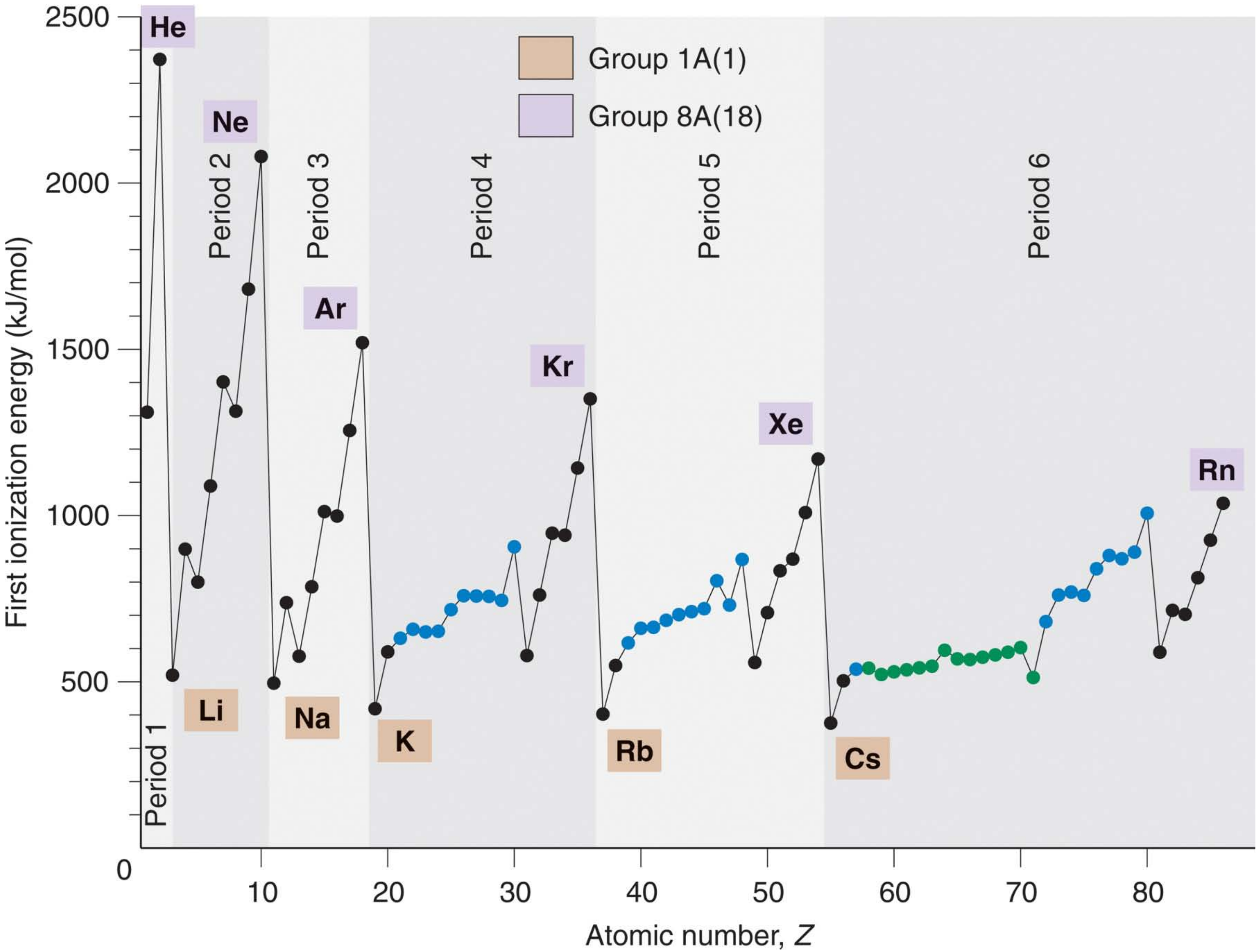
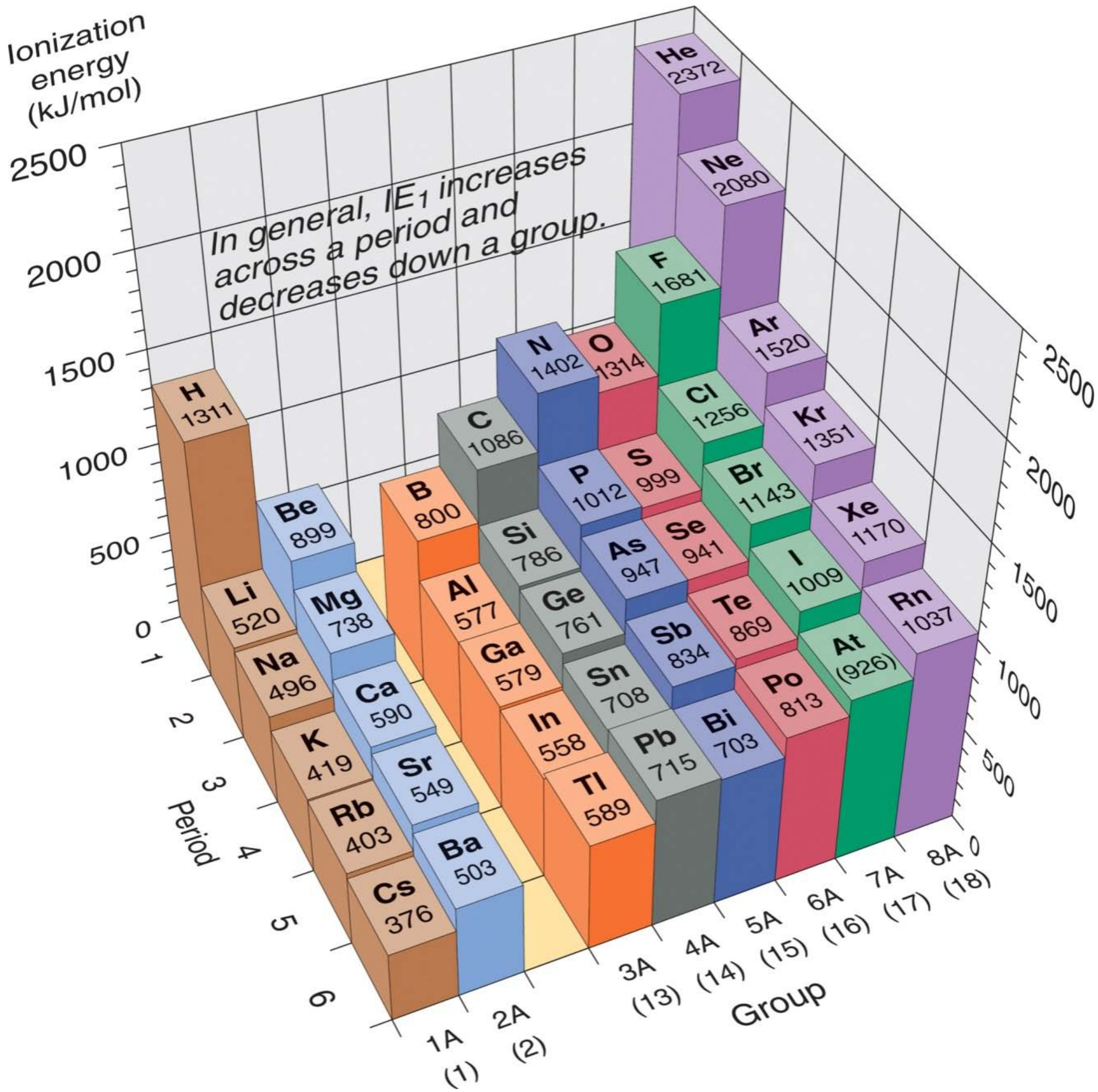


Figure 8.16 First ionization energies of the main-group elements.



Sample Problem 8.4 **Ranking Elements by First Ionization Energy**

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

(a) Kr, He, Ar **(b)** Sb, Te, Sn

(c) K, Ca, Rb **(d)** I, Xe, Cs

PLAN: Find each element on the periodic table. IE_1 generally decreases down a group and increases across a period.

SOLUTION:

(a) He > Ar > Kr

Kr, He, and Ar are in Group 8A. IE_1 decreases down the group.

Sample Problem 8.4

SOLUTION:

(b) Te > Sb > Sn

Sb, Te, and Sn are in Period 5. IE_1 increases across a period.

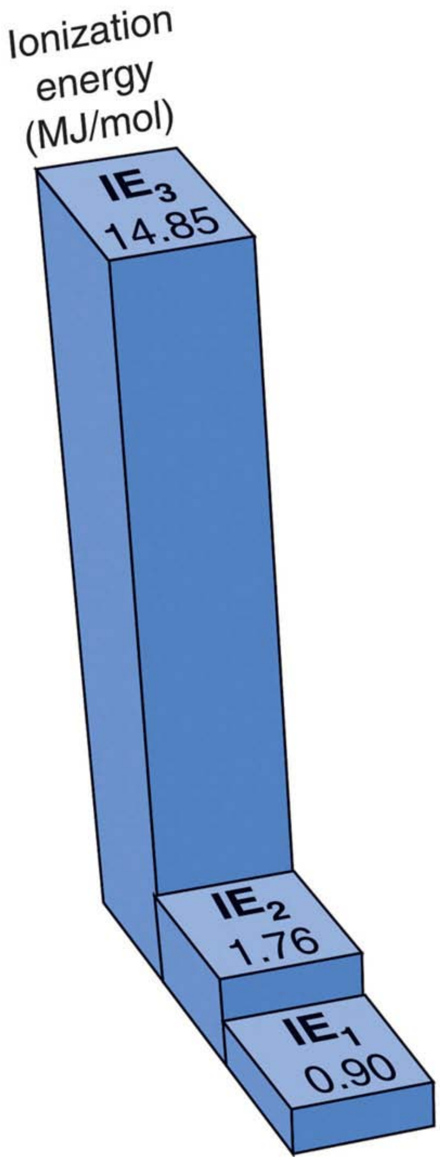
(c) Ca > K > Rb

K has a higher IE_1 than Rb because K is higher up in Group 1A. Ca has a higher IE_1 than K because Ca is further to the right in Period 4.

(d) Xe > I > Cs

Xe has a higher IE_1 than I because Xe is further to the right in the same period. Cs has a lower IE_1 than I because it is further to the left in a higher period.

Figure 8.17 The first three ionization energies of beryllium.



Beryllium has 2 valence electrons, so IE_3 is much larger than IE_2.

Table 8.4 Successive Ionization Energies of the Elements Lithium Through Sodium

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

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

Sample Problem 8.5 **Identifying an Element from Its Ionization Energies**

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

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

PLAN: Look for a large increase in IE, which occurs after all valence electrons have been removed.

SOLUTION:

The largest increase occurs after IE_5, that is, after the 5th valence electron has been removed. The Period 3 element with 5 valence electrons is **phosphorus (P; *Z* = 15**).

The complete electron configuration is $\mathbf{1s^2 2s^2 2p^6 3s^2 3p^3}$.

Trends in Electron Affinity

Electron Affinity (EA) is the energy change that occurs when 1 mol of electrons is ***added*** to 1 mol of gaseous atoms or ions.

Atoms with a ***low EA*** tend to form ***cations***.
Atoms with a ***high EA*** tend to form ***anions***.

The trends in electron affinity are not as regular as those for atomic size or IE.

Figure 8.18 Electron affinities of the main-group elements (in kJ/mol).

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)

Behavior Patterns for IE and EA

Reactive nonmetals have high IEs and highly negative EAs.

These elements attract electrons strongly and tend to form negative ions in ionic compounds.

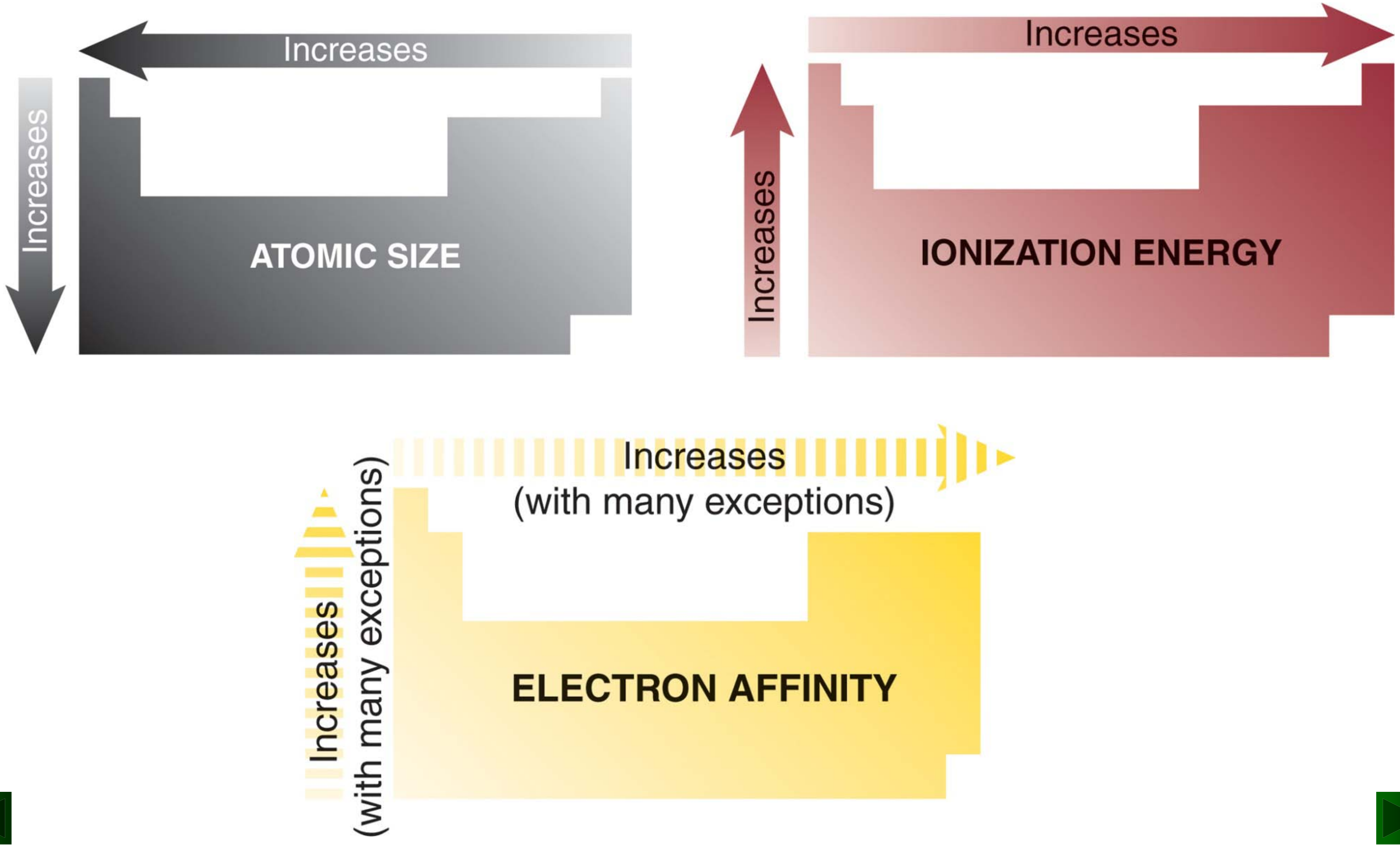
Reactive metals have low IEs and slightly negative EAs.

These elements lose electrons easily and tend to form positive ions in ionic compounds.

Noble gases have very high IEs and slightly positive EAs.

These elements tend to neither lose nor gain electtons.

Figure 8.19 **Trends in three atomic properties.**



Metallic Behavior

- Metals are typically shiny solids with moderate to high melting points.
- Metals are good conductors of heat and electricity, and can easily be shaped.
- Metals tend to lose electrons and form cations, i.e., they are easily ***oxidized***.
- Metals are generally ***strong reducing agents***.
- Most metals form ionic oxides, which are ***basic*** in aqueous solution.

Figure 8.20 Trends in metallic behavior.

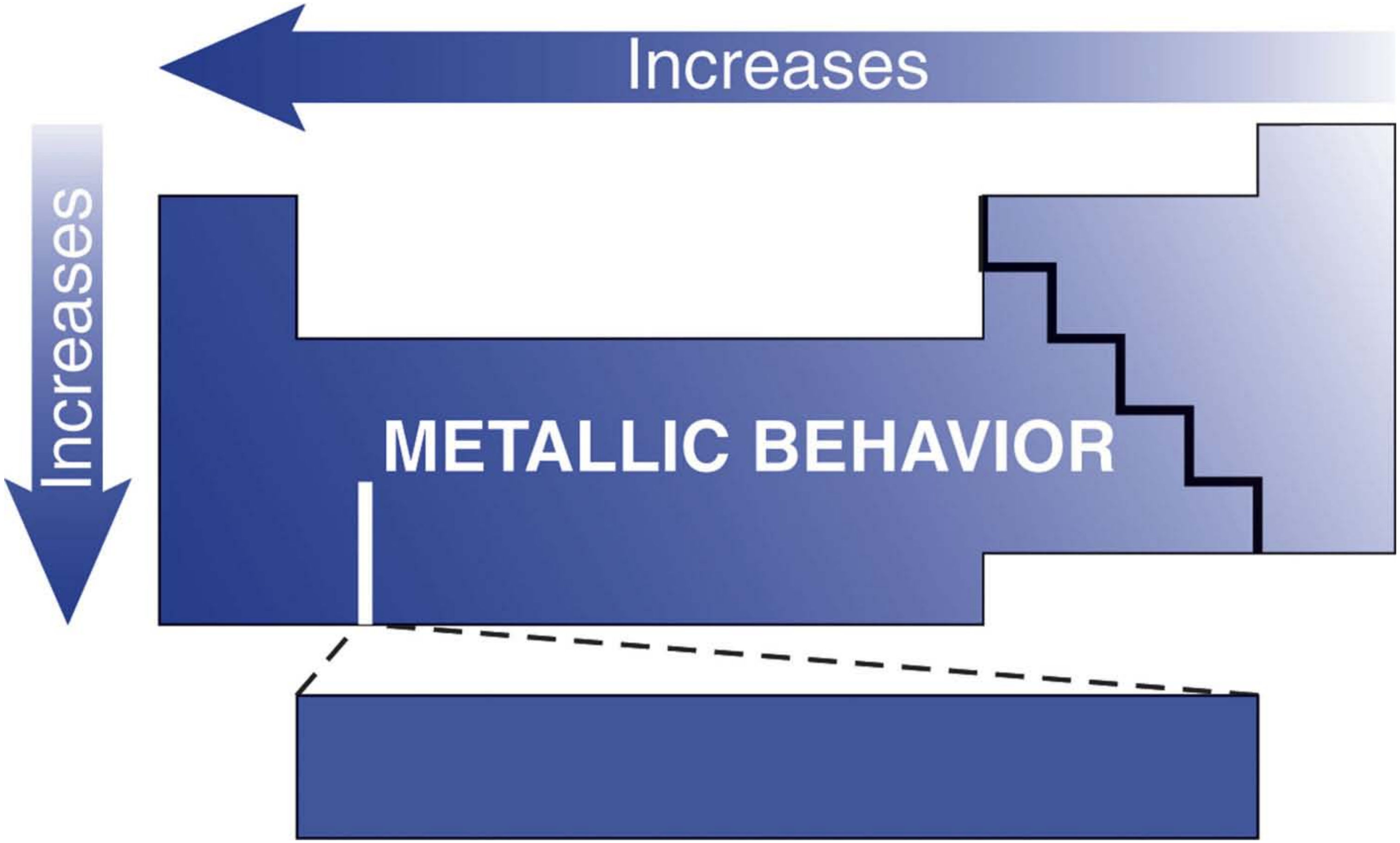
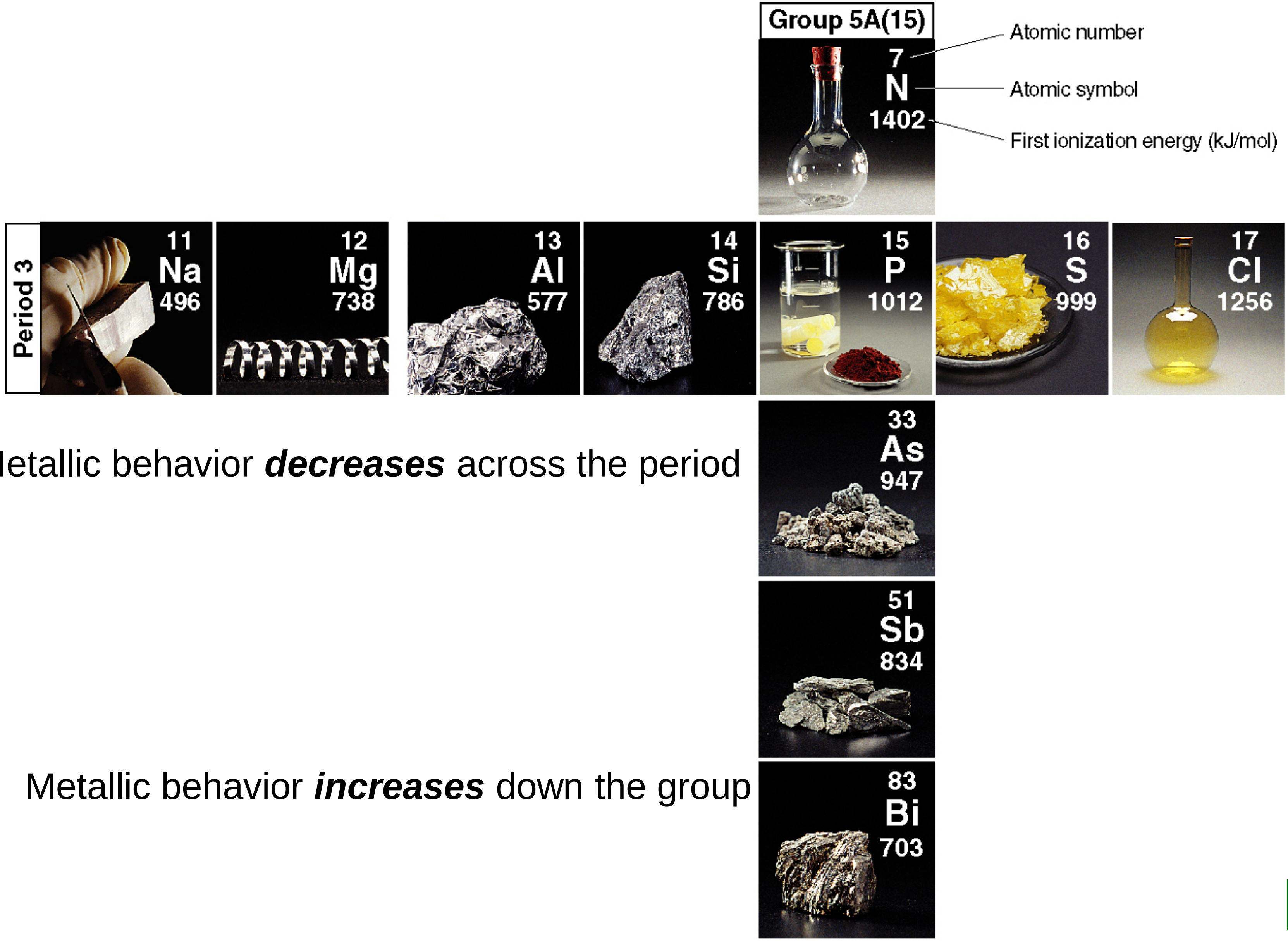


Figure 8.21 Metallic behavior in Group 5A(15) and Period 3.



Metallic behavior ***decreases*** across the period

Metallic behavior ***increases*** down the group

Figure 8.22 **Highest and lowest O.N.s of reactive main-group elements.**

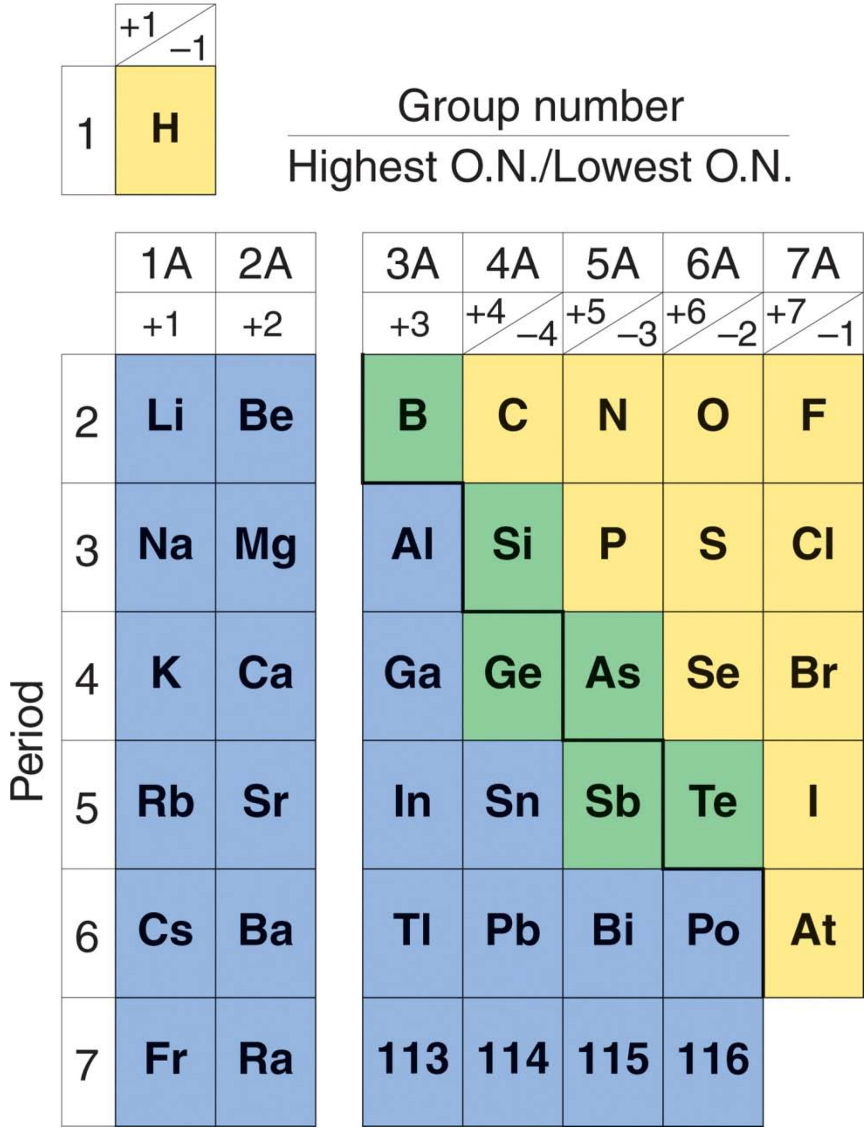
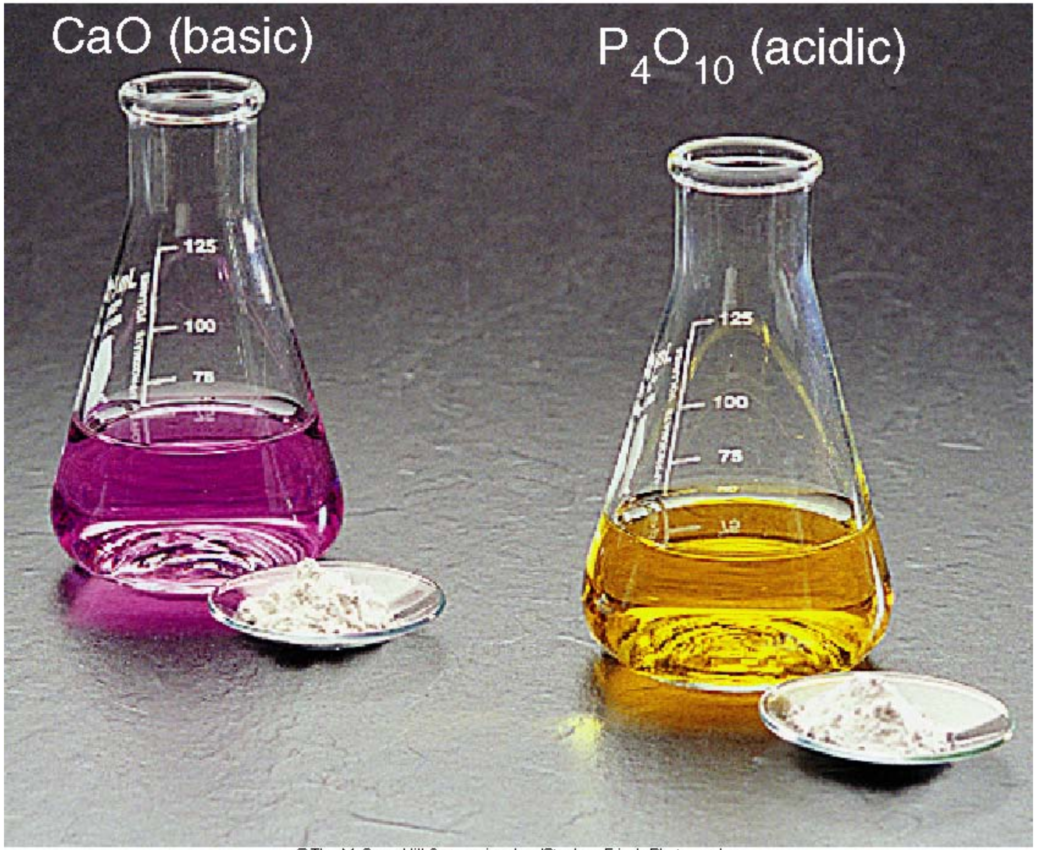


Figure 8.23 Oxide acidity.



CaO, the oxide of a main-group ***metal***, is strongly ***basic.***
P_4O_{10}, the oxide of a main-group ***nonmetal***, is ***acidic***.

Acid-Base Behavior of Oxides

Main-group metals form ***ionic oxides***, which are ***basic*** in aqueous solution.

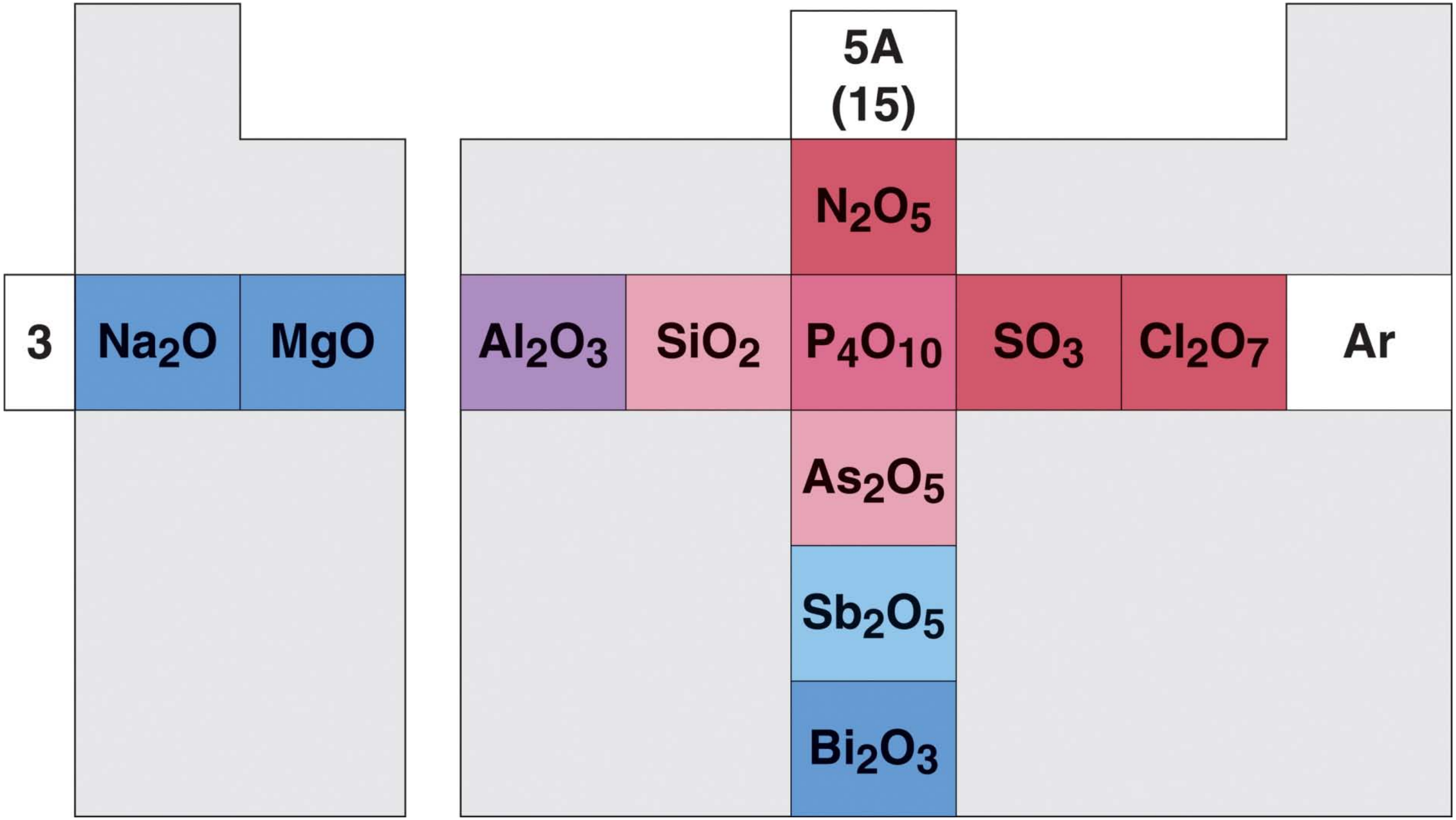
Main-group nonmetals form ***covalent oxides***, which are ***acidic*** in aqueous solution.

Some metals and metalloids from ***amphoteric oxides***, which can act as acids or bases in water:

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

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

Figure 8.24 Acid-base behavior of some element oxides.



Oxides become more basic down a group and more acidic across a period.

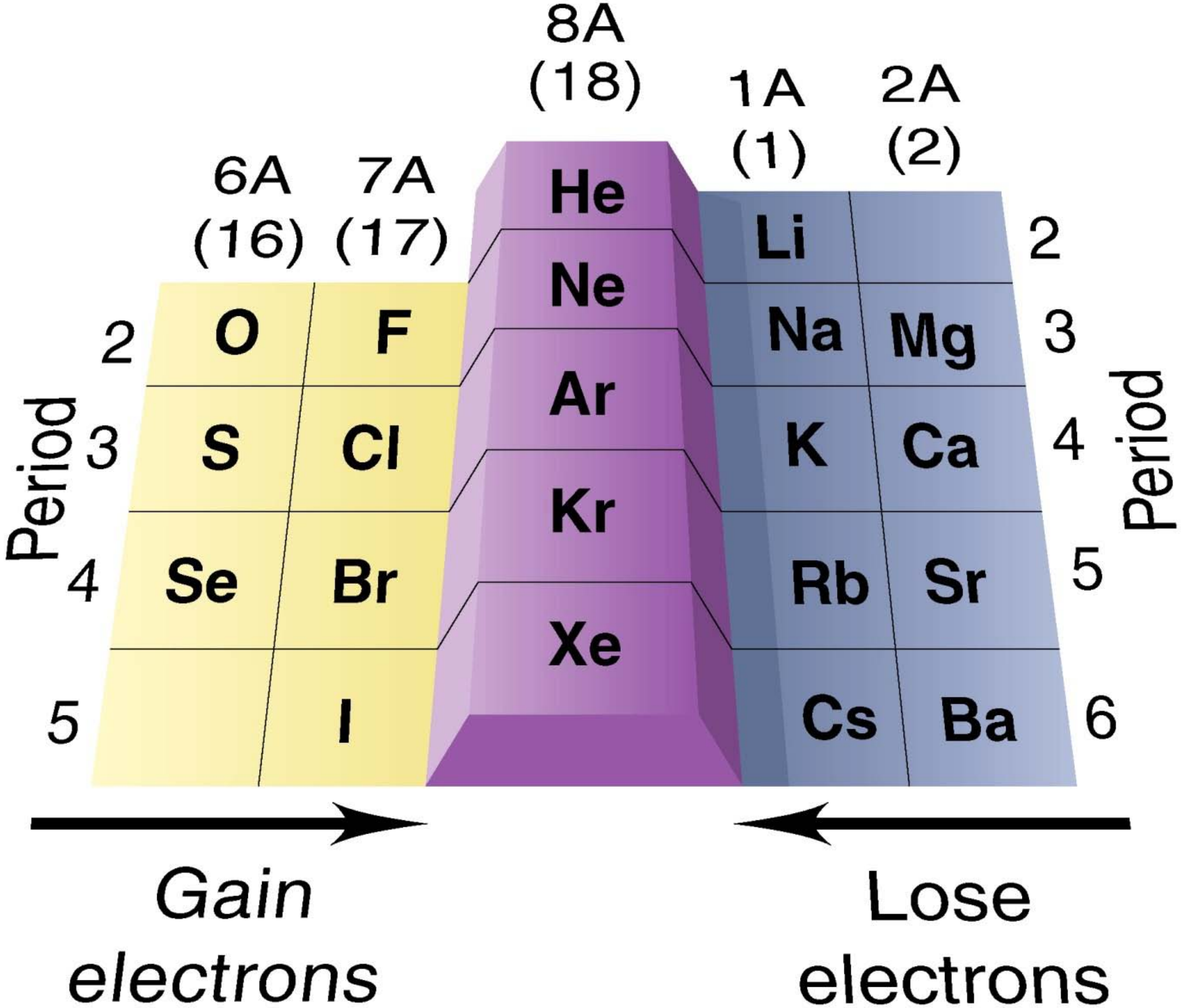
Electron configurations of Monatomic Ions

Elements at either end of a period gain or lose electrons to attain a filled outer level. The resulting ion will have a ***noble gas electron configuration*** and is said to be ***isoelectronic*** with that noble gas.

$Na(1s^2 2s^2 2p^6 \mathbf{3s^1}) \rightarrow \mathbf{e^-} + Na^+([He]2s^2 2p^6)$
[isoelectronic with Ne]

$Br([Ar]4s^2 3d^{10} \mathbf{4p^5}) + \mathbf{e^-} \rightarrow Br^- ([Ar]4s^2 3d^{10} 4p^6)$
[isoelectronic with Kr]

Figure 8.25 **Main-group elements whose ions have noble gas electron configurations.**



Electron configurations of Monatomic Ions

A ***pseudo-noble gas configuration*** is attained when a metal atom empties its highest energy level.

The ion attains the stability of empty *ns* and *np* sublevels and a filled (*n*-1)*d* sublevel.

$$Sn([Kr]\mathbf{5s^2}4d^{10}\mathbf{5p^2}) \rightarrow \mathbf{4e^-} + Sn^{4+}\ ([Kr]4d^{10})$$

A metal may lose only the *np* electrons to attain an ***inert pair configuration***.

The ion attains the stability of a filled *ns* and (*n*-1)*d* sublevels.

$$Sn([Kr]5s^24d^{10}\mathbf{5p^2}) \rightarrow \mathbf{2e^-} + Sn^{2+}\ ([Kr]5s^24d^{10})$$

Sample Problem 8.6 **Writing Electron Configurations of Main-Group Ions**

PROBLEM: Using condensed electron configurations, write reactions for the formation of the common ions of the following elements:

(a) Iodine (Z = 53) **(b)** Potassium (Z = 19) **(c)** Indium (Z = 49)

PLAN: Identify the position of each element on the periodic table and recall that:

- Ions of elements in Groups 1A(1), 2A(2), 6A(16), and 7A(17) are usually isoelectronic with the nearest noble gas.
- Metals in Groups 3A(13) to 5A(15) can lose the *ns* and *np* electrons or just the *np* electrons.

Sample Problem 8.6

SOLUTION:

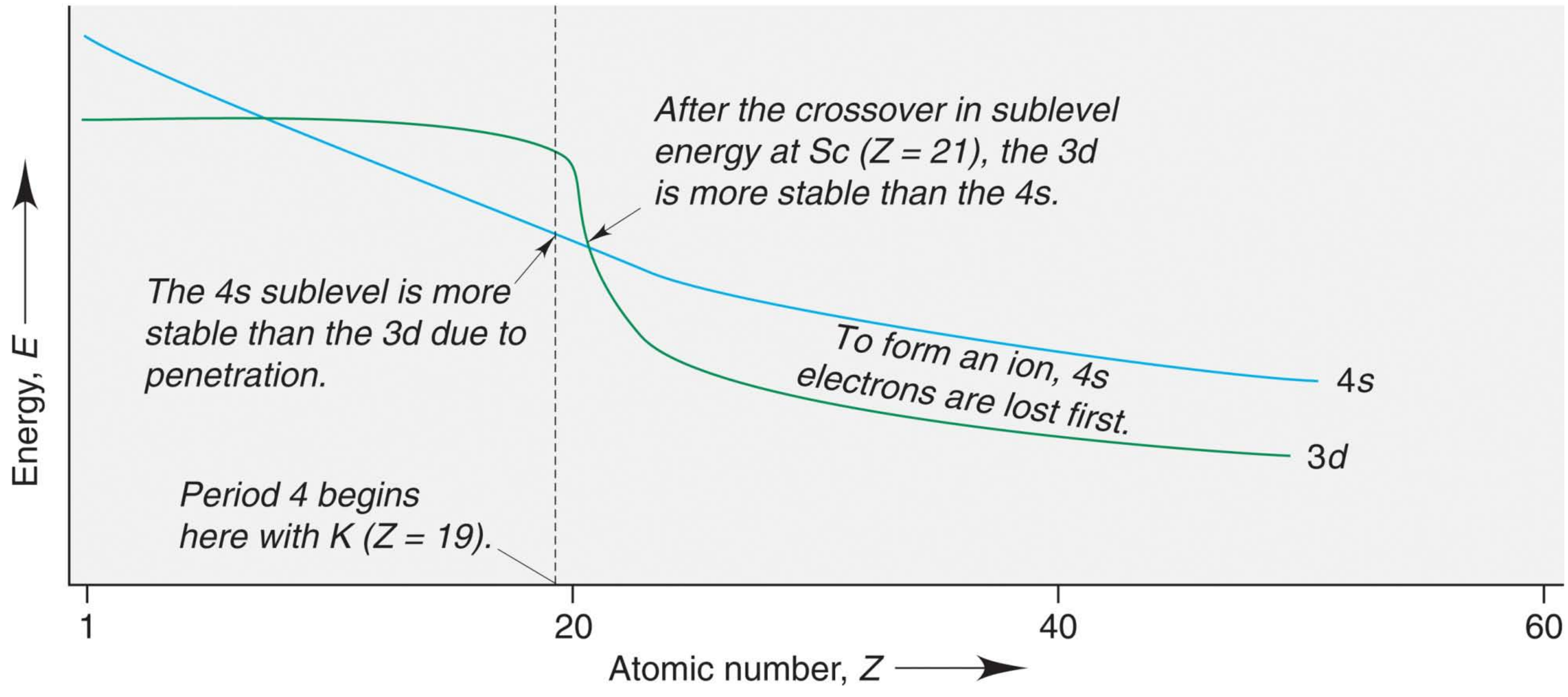
(a) Iodine (Z = 53) is in Group 7A(17) and will gain one electron to be isoelectronic with Xe: I ([Kr] $5s^2 4d^{10} 5p^5$) + e^- $\longrightarrow$ I^- ([Kr] $5s^2 4d^{10} 5p^6$)

(b) Potassium (Z = 19) is in Group 1A(1) and will lose one electron to be isoelectronic with Ar: K ([Ar] $4s^1$) $\longrightarrow$ K^+ ([Ar]) + e^-

(c) Indium (Z = 49) is in Group 3A(13) and can lose either one electron or three electrons: In ([Kr] $5s^2 4d^{10} 5p^1$) $\longrightarrow$ In^+ ([Kr] $5s^2 4d^{10}$) + e^-

In ([Kr] $5s^2 4d^{10} 5p^1$) $\longrightarrow$ In^{3+}([Kr] $4d^{10}$) + $3e^-$

Figure 8.26 The crossover of sublevel energies in Period 4.



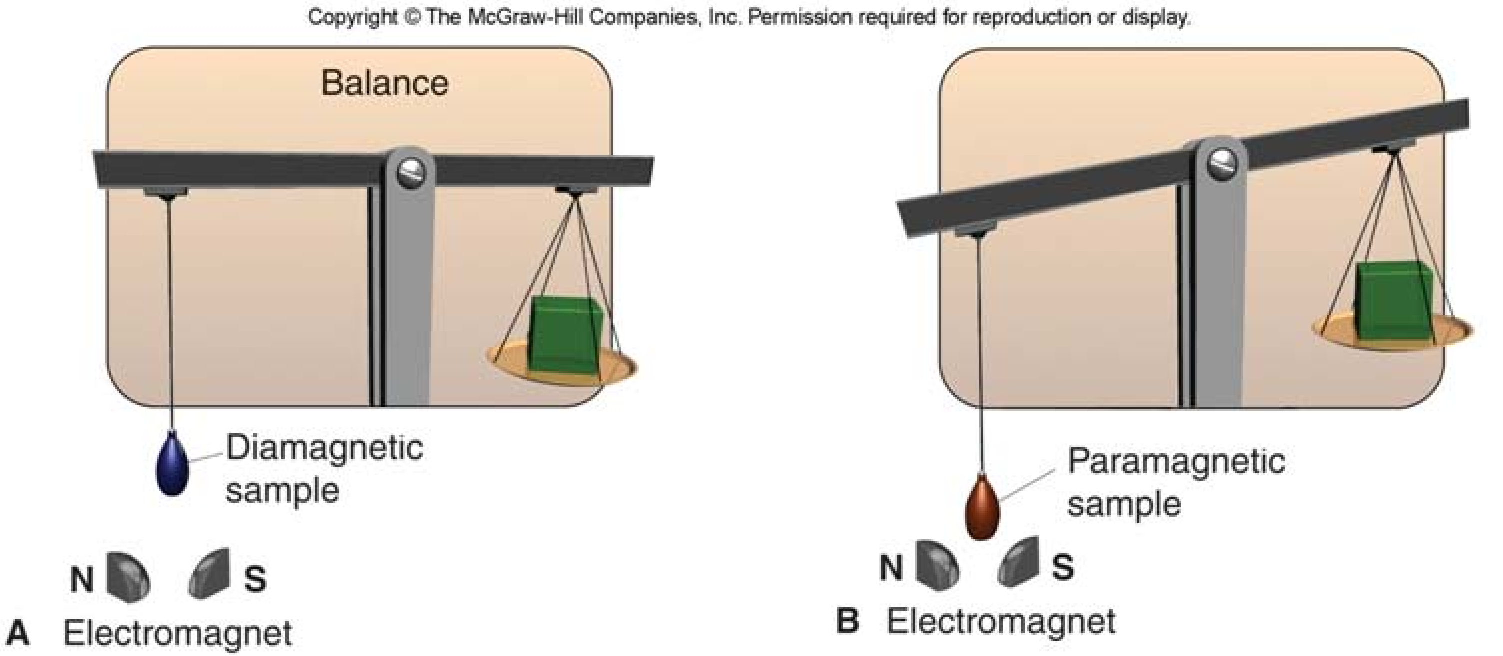
Magnetic Properties of Transition Metal ions

A species with one or more unpaired electrons exhibits ***paramagnetism*** – it is attracted by a magnetic field.

	5*s*	4*d*					5*p*		
Ag (Z = 47)	↑	↑↓	↑↓	↑↓	↑↓	↑↓			

A species with all its electrons paired exhibits ***diamagnetism*** – it is not attracted (and is slightly repelled) by a magnetic field.

Figure 8.27 Measuring the magnetic behavior of a sample.



The apparent mass of a diamagnetic substance is unaffected by the magnetic field.

The apparent mass of a paramagnetic substance increases as it is attracted by the magnetic field.

Magnetic Properties of Transition Metal ions

Magnetic behavior can provide evidence for the electron configuration of a given ion.

	4*s*	3*d*					4*p*		
Ti (Z = 22)	↑↓	↑	↑						
Ti^{2+}		↑	↑						

Ti^{2+} has 2 unpaired electrons and is paramagnetic, providing evidence that the 4*s* electrons are lost before the 3*d* electrons.

Sample Problem 8.7

Writing Electron Configurations and Predicting Magnetic Behavior of Transition Metal Ions

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

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

PLAN: Write the condensed electron configuration for each atom, recalling the irregularity for Cr. Remove electrons, beginning with the *ns* electrons, and determine if there are any unpaired electrons.

Sample Problem 8.7

SOLUTION:

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

Since there are 5 *d* electrons they are all unpaired. Mn^{2+} is **paramagnetic**.

(b) Cr^{3+}(Z = 24)   Cr ([Ar] $4s^1 3d^5$) $\longrightarrow$ Cr^{3+} ([Ar] $3d^3$) + $3e^-$

Since there are 3 *d* electrons they are all unpaired. Cr^{3+} is **paramagnetic.**

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

The 4*f* and the 5*s* sublevels are filled, so there are no unpaired electrons. Hg^{2+} is **diamagnetic**.

Ionic Size vs. Atomic Size

Cations are ***smaller*** than their parent atoms while anions are ***larger***.

Ionic radius ***increases*** down a group as *n* increases.

Cation size ***decreases*** as charge ***increases***.

An ***isoelectronic series*** is a series of ions that have the same electron configuration. Within the series, ion size ***decreases*** with increasing nuclear charge.

3- > 2- > 1- > 1+ > 2+ > 3+

Figure 8.28 **Ionic radius.**

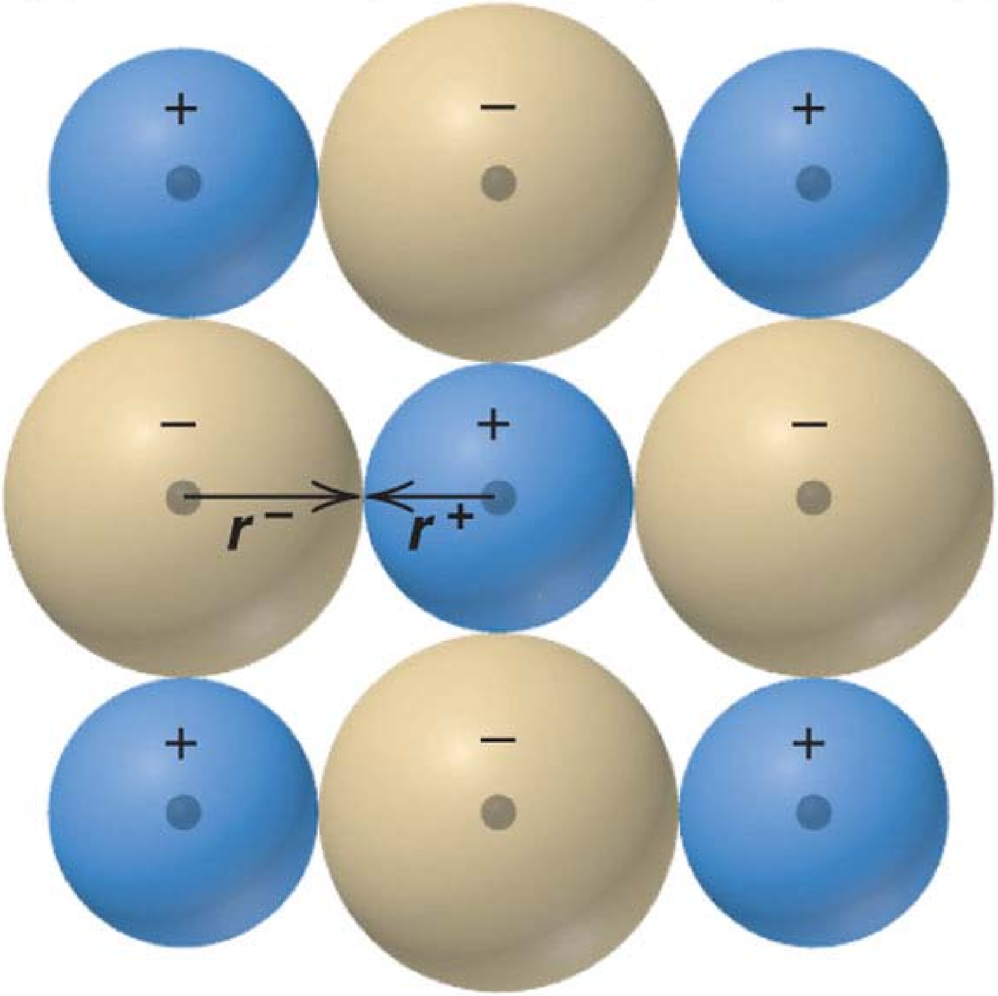
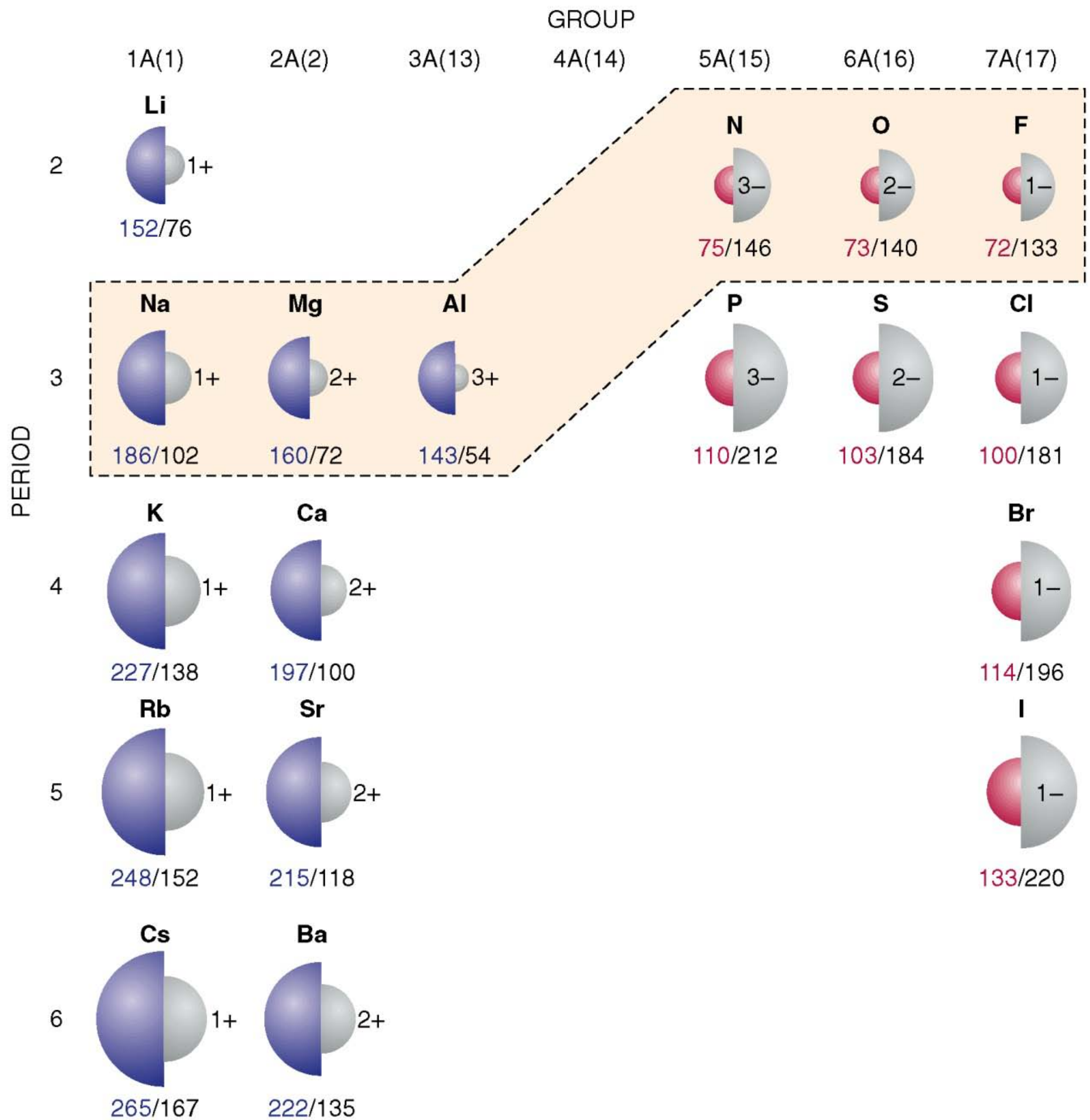


Figure 8.29 Ionic vs. atomic radii.



Sample Problem 8.8 **Ranking Ions by Size**

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

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

PLAN: Find the position of each element on the periodic table and apply the trends for ionic size.

SOLUTION:

(a) $\mathbf{Sr^{2+} > Ca^{2+} > Mg^{2+}}$

All these ions are from Group 2A, so size increases down the group.

Sample Problem 8.8

SOLUTION:

(b) $S^{2-} > Cl^{-} > K^{+}$

These ions are isoelectronic, so size decreases as nuclear charge increases.

(c) $Au^{+} > Au^{3+}$

Cation size decreases as charge increases.