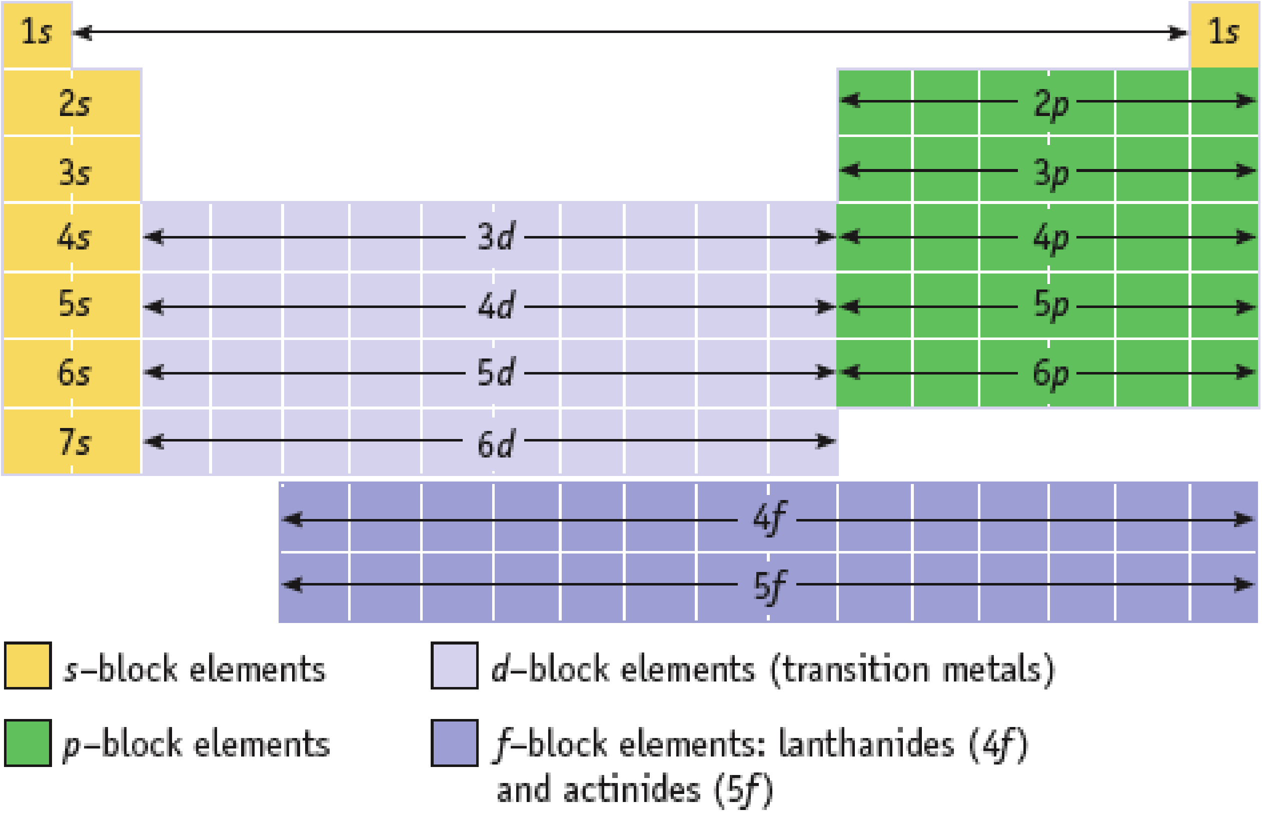
Chapter 8
ELECTRON CONFIGURATIONS and CHEMICAL PERIODICITY



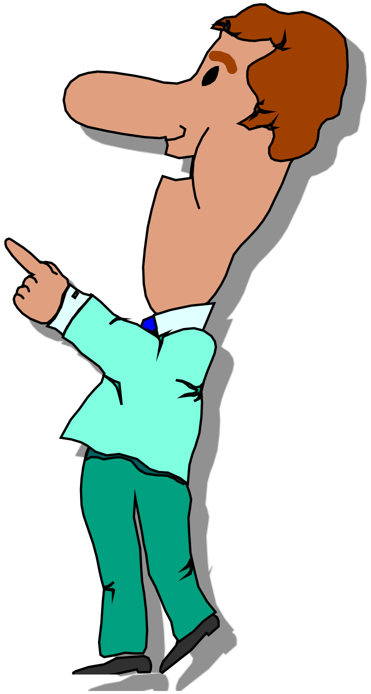
Quantum Numbers of an Electron

Electrons in atoms are arranged as

SHELLS (n)

SUBSHELLS (ℓ)

ORBITALS (m_ℓ)



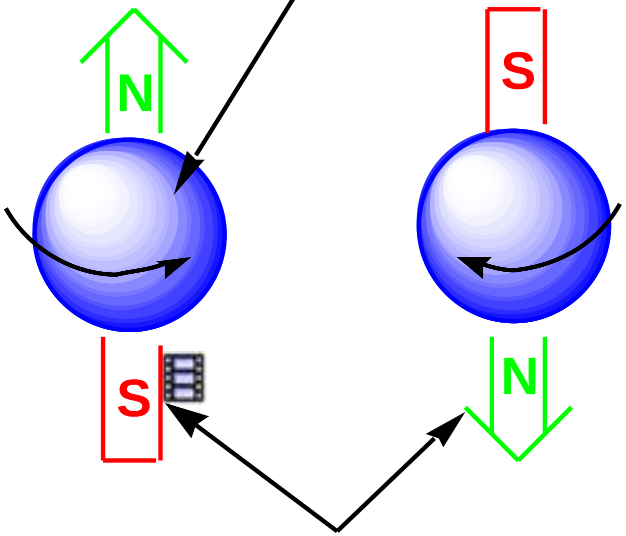
8.1 Characteristics of Many-Electron Atoms

The 3 additional features of a many-electron atom:

- **The fourth quantum number m_s**
- **The maximum number of electrons in an orbital**
- **The splitting of energy levels into sublevels**

Electron Spin Quantum Number, m_s = -1/2 or +1/2

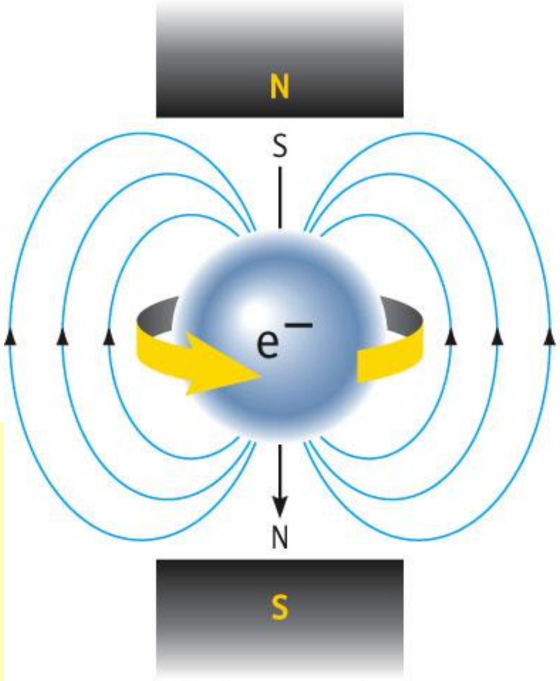
① A spinning electron generates a tiny magnetic field



② When their tiny magnetic fields are aligned N-S, the electron spins are paired

- Spin pairing means that electrons spin in opposite directions

Can be proved experimentally that electron has a spin. Two spin directions are given by quantum number m_s where m_s = +1/2 and -1/2.



(a) Electron spin

© Brooks/Cole, Cengage Learning

THE 4 QUANTUM NUMBERS

n (orbital energy or size)	1, 2, 3, 4, ...
ℓ (orbital shape)	0, 1, 2, ... n - 1
m_ℓ (orbital orientation)	$-\ell$ ... 0 ... $+\ell$
m_s (e^- spin direction)	+1/2 and -1/2

Pauli Exclusion Principle

No two electrons in the same atom can have the same set of 4 quantum numbers.

That is, each electron in any atom is described completely by a set of *four* quantum numbers, i.e., a unique ID or address.

Major consequence:

An atomic orbital can hold a maximum of TWO electrons, and they must have opposite spins

Numbers of Electrons Accommodated in Electron Shells and Subshells

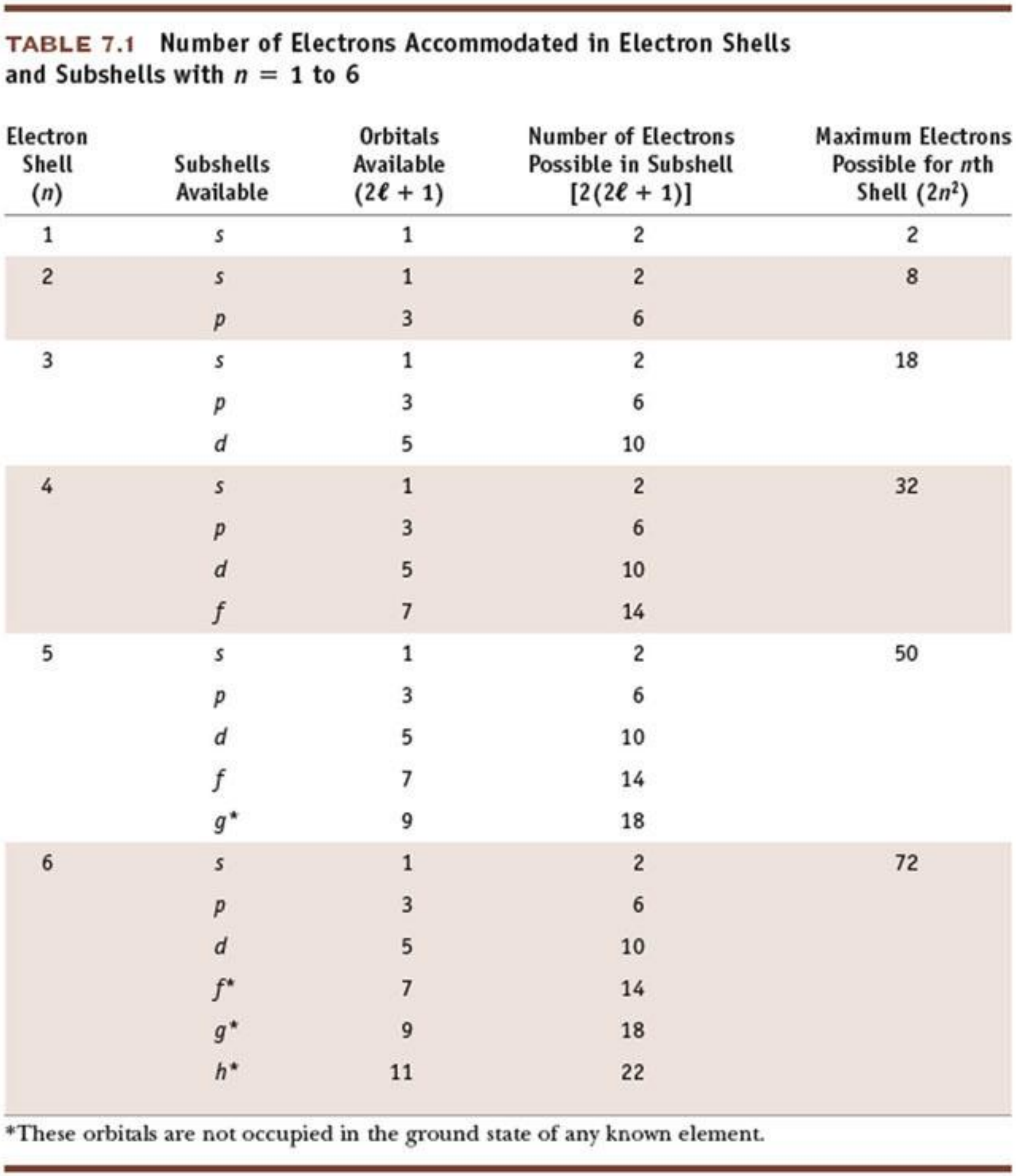


TABLE 7.1 Number of Electrons Accommodated in Electron Shells and Subshells with $n = 1$ to 6

Electron Shell (n)	Subshells Available	Orbitals Available ($2\ell + 1$)	Number of Electrons Possible in Subshell [$2(2\ell + 1)$]	Maximum Electrons Possible for nth Shell ($2n^2$)
1	*s*	1	2	2
2	*s*	1	2	8
	p	3	6	
3	*s*	1	2	18
	p	3	6	
	d	5	10	
4	*s*	1	2	32
	p	3	6	
	d	5	10	
	f	7	14	
5	*s*	1	2	50
	p	3	6	
	d	5	10	
	f	7	14	
	*g**	9	18	
6	*s*	1	2	72
	p	3	6	
	d	5	10	
	*f**	7	14	
	*g**	9	18	
	*h**	11	22	

*These orbitals are not occupied in the ground state of any known element.

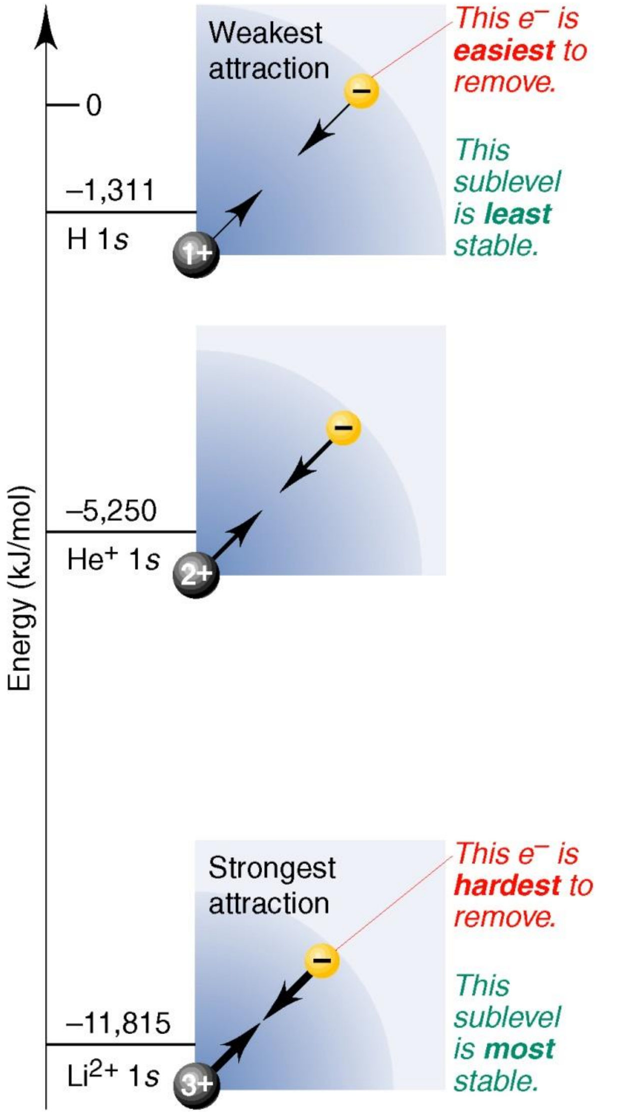
Electrostatic Effects & Energy-Level Splitting

Effect of Nuclear Charge Z, on Sublevel Energy

Greater nuclear charge

- increases nucleus-electron attractions

- lowers sublevel energy (stabilizing the atom)



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

Electrostatic Effects & Energy-Level Splitting

Shielding (Screening): Effect of Electron Repulsions on Sublevel Energy

- reduces the full nuclear charge to an EFFECTIVE NUCLEAR CHARGE, Z_{eff}, the nuclear charge that the electron actually experience.

- makes the electron easier to remove

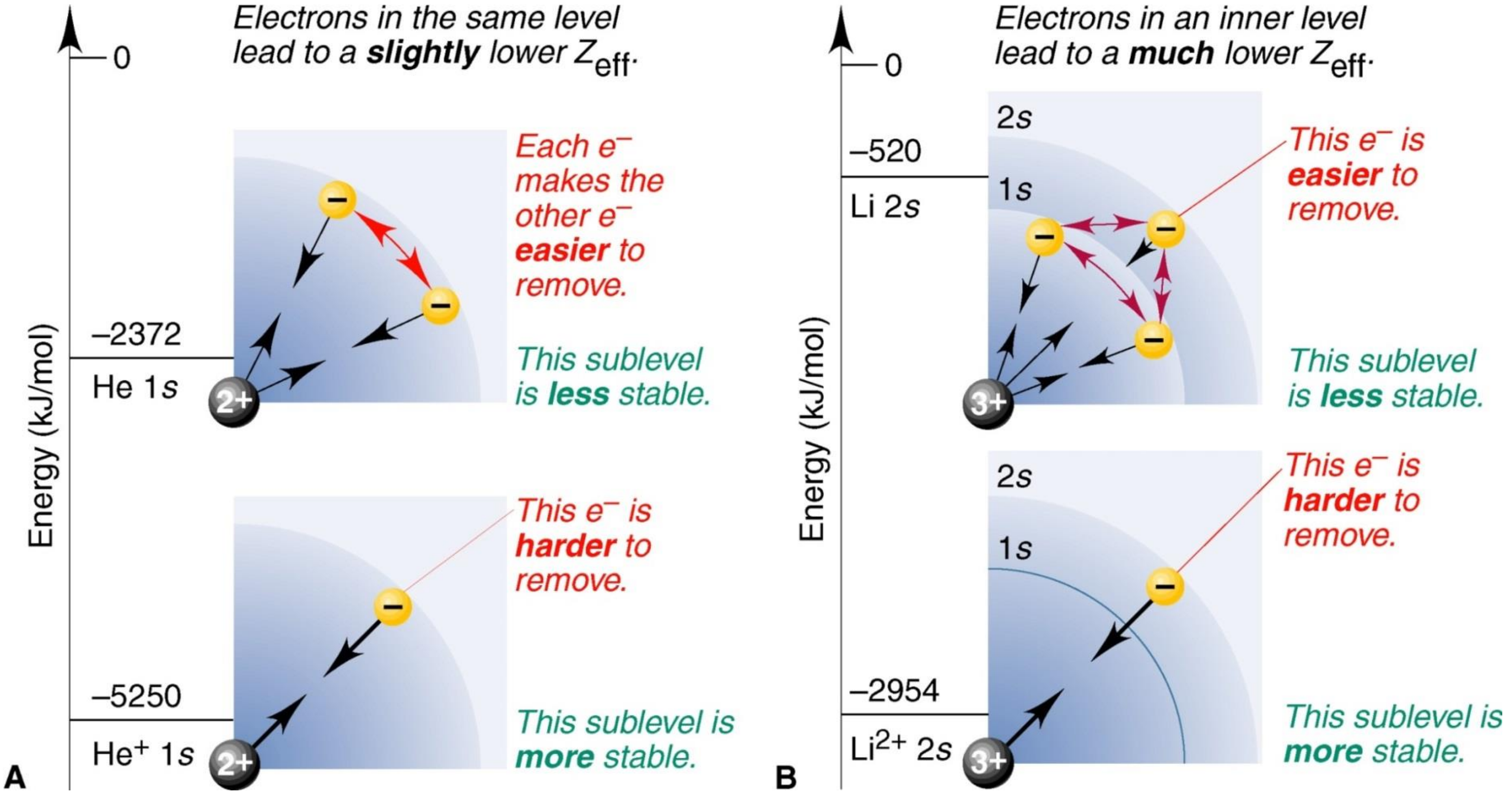
a) Shielding by other electrons in a given energy level

b) Shielding by electrons in inner energy levels: The inner electrons shield very effectively

The further from the nucleus an electron is, the lower the Z_{eff} for that particular electron.

Electrostatic Effects & Energy-Level Splitting

Shielding (Screening): Effect of Electron Repulsions on Sublevel Energy



Electrons in the same level lead to a **slightly** lower Z_{eff}.

Each e⁻ makes the other e⁻ ***easier*** *to remove.*

This sublevel is ***less*** *stable.*

This e⁻ is ***harder*** *to remove.*

This sublevel is ***more*** *stable.*

Electrons in an inner level lead to a **much** lower Z_{eff}.

This e⁻ is ***easier*** *to remove.*

This sublevel is ***less*** *stable.*

This e⁻ is ***harder*** *to remove.*

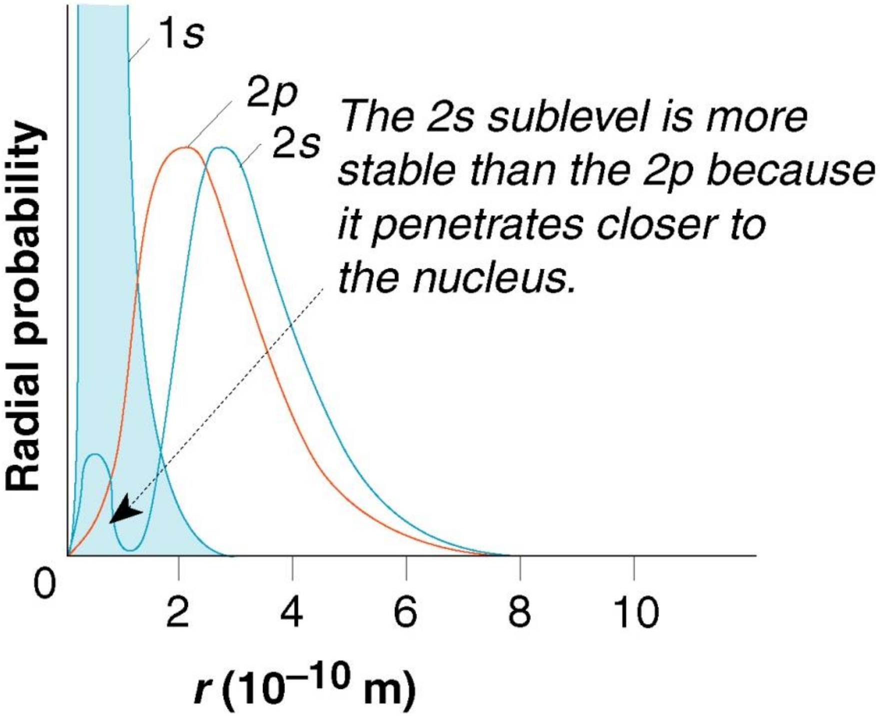
This sublevel is ***more*** *stable.*

Electrostatic Effects & Energy-Level Splitting

Penetration: Effect of Orbital Shape on Sublevel Energy

- increases the nuclear attraction for a 2s electron over that for a 2p electron

- decreases the shielding of a 2s electron by the 1s electron.



The 2s sublevel is more stable than the 2p because it penetrates closer to the nucleus.

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

Penetration increases nuclear attraction and decreases shielding.

Electrostatic Effects & Energy-Level Splitting

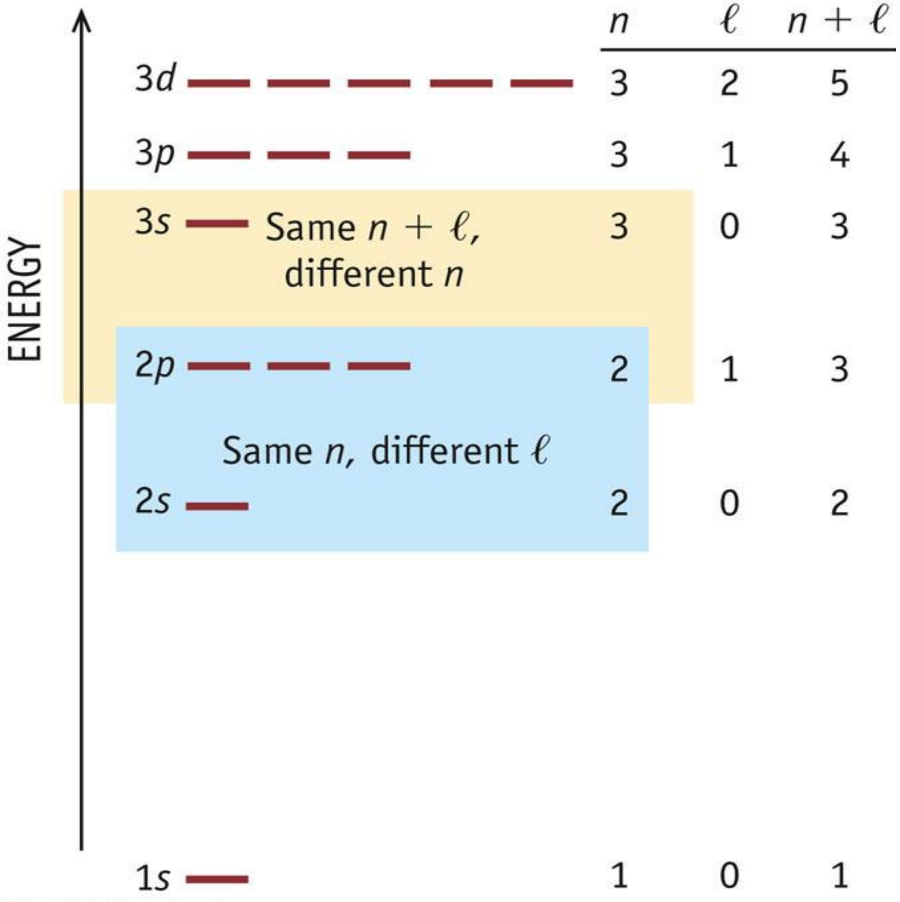
Splitting of Energy Levels into Sublevels of Different Energy

- **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 more stable sublevel.

Order of sublevel energies: *s* < *p* < *d* < *f*

Electrostatic Effects & Energy-Level Splitting

Splitting of Energy Levels into Sublevels of Different Energy



© Brooks/Cole, Cengage Learning

a) subshells increase in energy as value of n + ℓ increases.

b) subshells of same n + ℓ with lower n value is lower in energy.

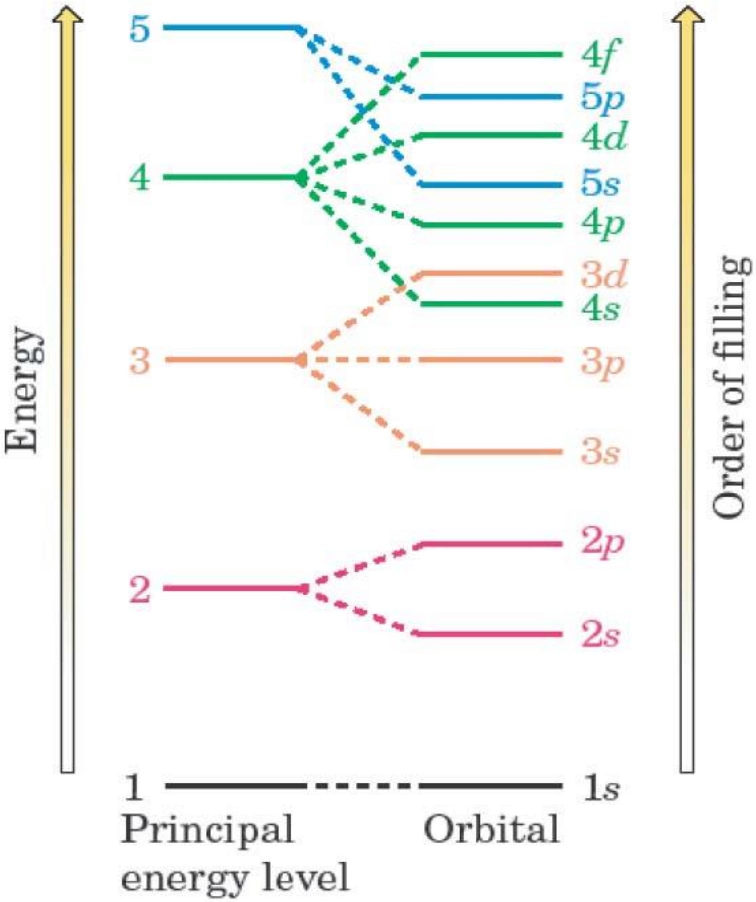
8.2 The Quantum Mechanical Model & The Periodic Table

- **In multi-electron atoms, sublevels within a principal energy level split such that**

s* < *p* < d < *f

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.



Orbital Diagram Rules

- determine how the electrons are filled in the principal energy levels and sublevels.
  - **Aufbau principle**

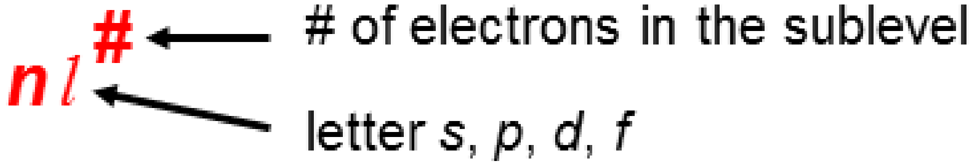
    Electrons fill orbitals starting with the lowest-energy orbitals.
  - **Pauli exclusion principle**

    A maximum of two electrons can occupy each orbital, and they must have opposite spins.
  - **Hund's rule**

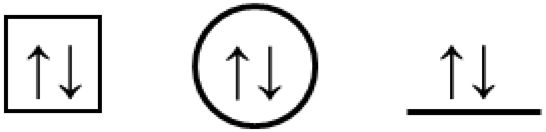
    Electrons are distributed into orbitals of identical energy (same sublevel) in such a way as to give the maximum number of unpaired electrons.
- Electrons are always filled in their ground state, or lowest energy state.

Electron Configurations & Orbital Diagrams

Electron configuration is indicated by a shorthand notation:



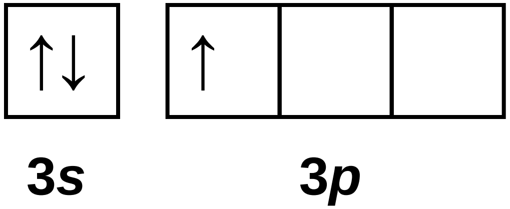
Orbital diagrams make use of a box, circle, or line for each orbital in the energy level. An arrow is used to represent an electron *and* its spin.



Partial Orbital Diagrams and Condensed Configurations

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

Al (Z = 13) $1s^22s^22p^63s^23p^1$



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

Al has the condensed configuration [Ne]$3s^23p^1$

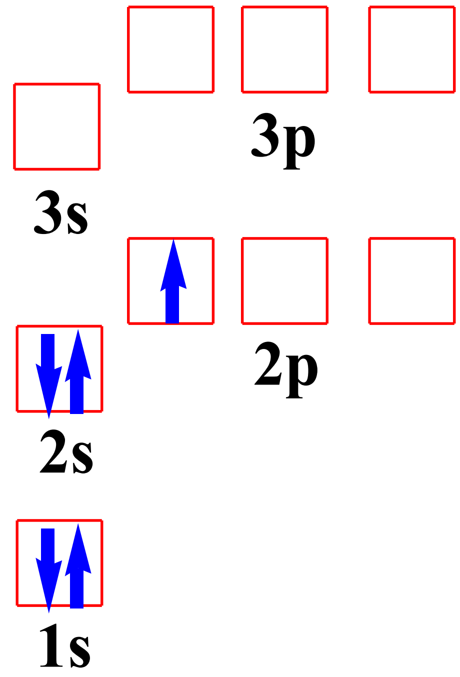
Orbital Diagram for Boron

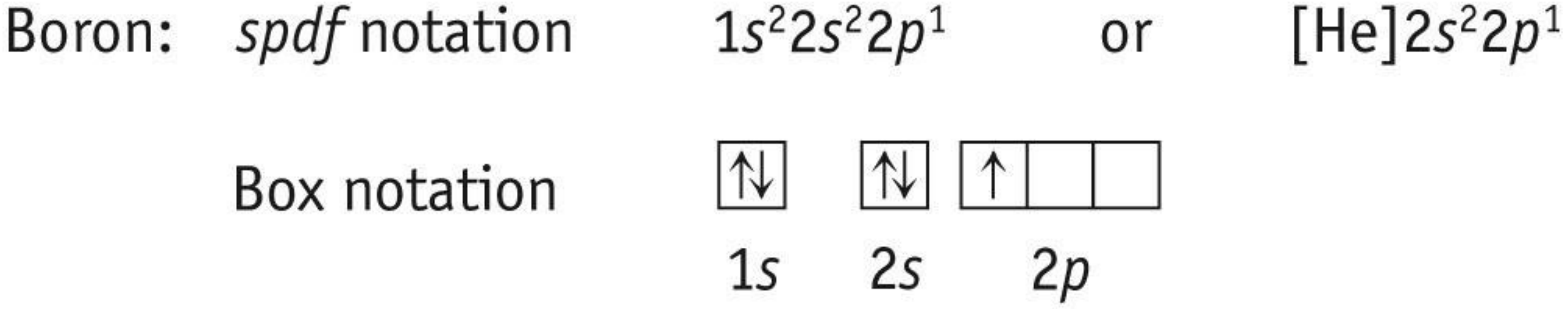
Full electron configuration or spdf notation

$1s^2 2s^2 2p^1$

Condensed, or abbreviated or noble gas notation.

$[He]\ 2s^2 2p^1$





Boron: *spdf* notation $1s^22s^22p^1$ or $[He]2s^22p^1$

Box notation

© Brooks/Cole, Cengage Learning

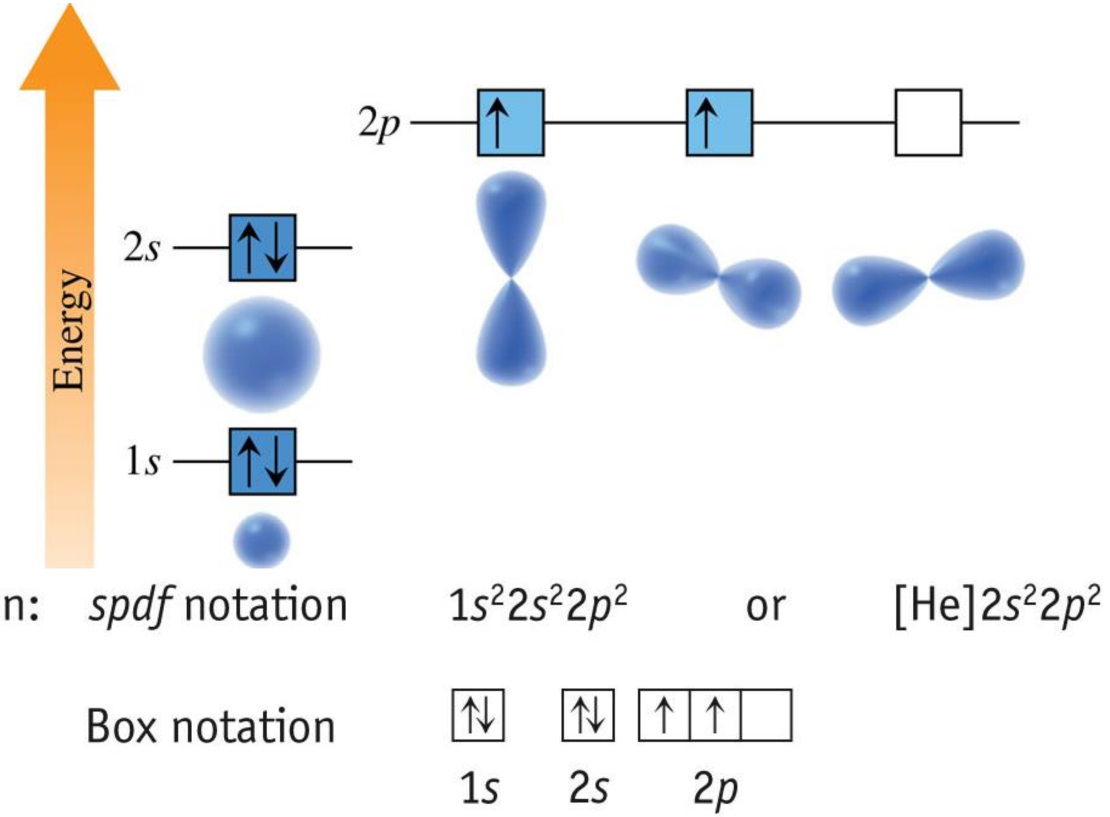
Orbital Diagram for Carbon

- Two electrons in the same orbital must have **opposite spins**, represented by the up and down arrows.

- Electrons occupy the lowest-energy orbitals first (**aufbau principle**).

- No more than two electrons occupy each orbital (**Pauli exclusion principle**).

- It takes a little bit of energy to pair up electrons, so single electrons occupy different orbitals of the same energy (**Hund's Rule**)



Carbon: *spdf* notation $1s^2 2s^2 2p^2$ or $[\mathrm{He}]2s^2 2p^2$

Problem

Write the full electron configuration and noble gas notation for bromine atom. Then write the set of quantum numbers for its sixth electron.

General Principles of Electron Configurations

Similar outer electron configurations within a group.

It *correlates with similar chemical behavior.*

In most cases, subshell ns is filled before subshell (n-1)d

Consider the alkali metals:

Li	$1s^2 2s^1$
Na	$1s^2 2s^2 2p^6 3s^1$
K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$

Consider the alkaline earth metals:

Be	$1s^2 2s^2$
Mg	$1s^2 2s^2 2p^6 3s^2$
K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$

General Principles of Electron Configurations

Similar outer electron configurations within a group.

It *correlates with similar chemical behavior.*

In most cases, subshell ns is filled before subshell (n-1)d

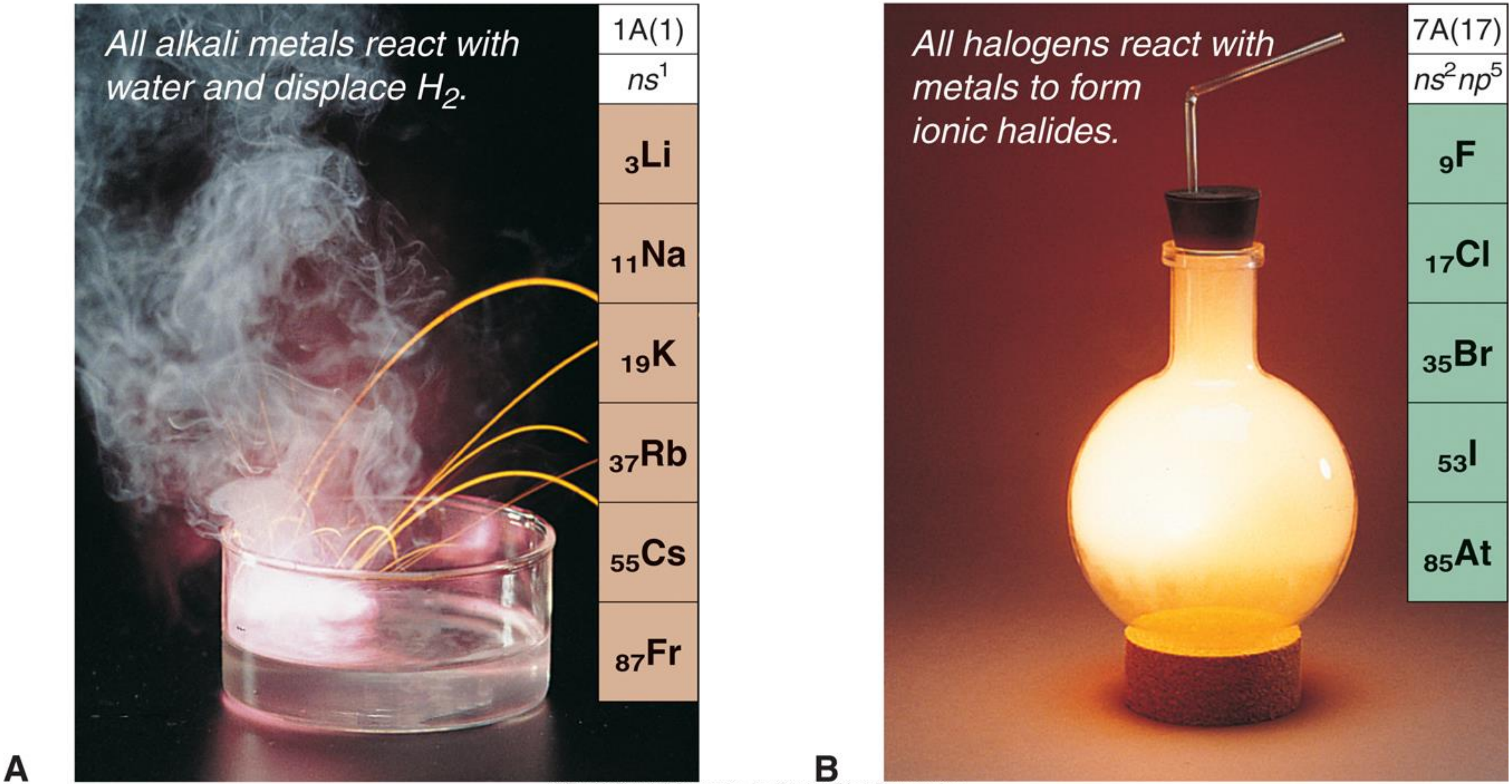
Consider a few of the halogens:

F	$1s^2 2s^2 2p^5$
Cl	$1s^2 2s^2 2p^6 3s^2 3p^5$
Br	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

Consider some of the noble gases:

Ne	$1s^2 2s^2 2p^6$
Ar	$1s^2 2s^2 2p^6 3s^2 3p^6$
Kr	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

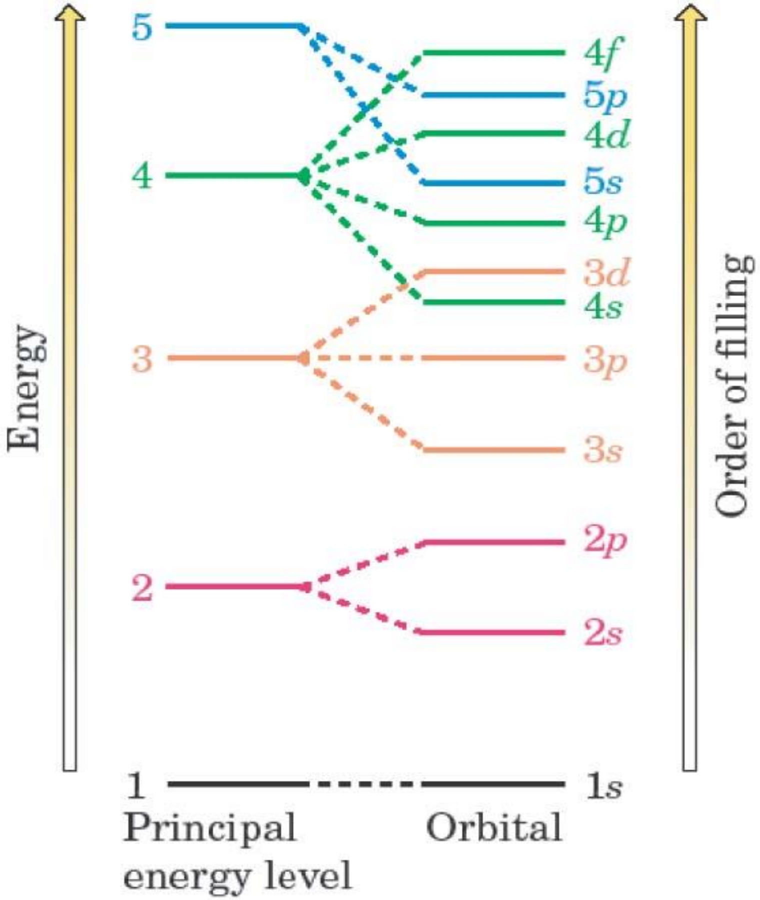
Similar outer electron configurations within a group *correlates with similar reactivities*

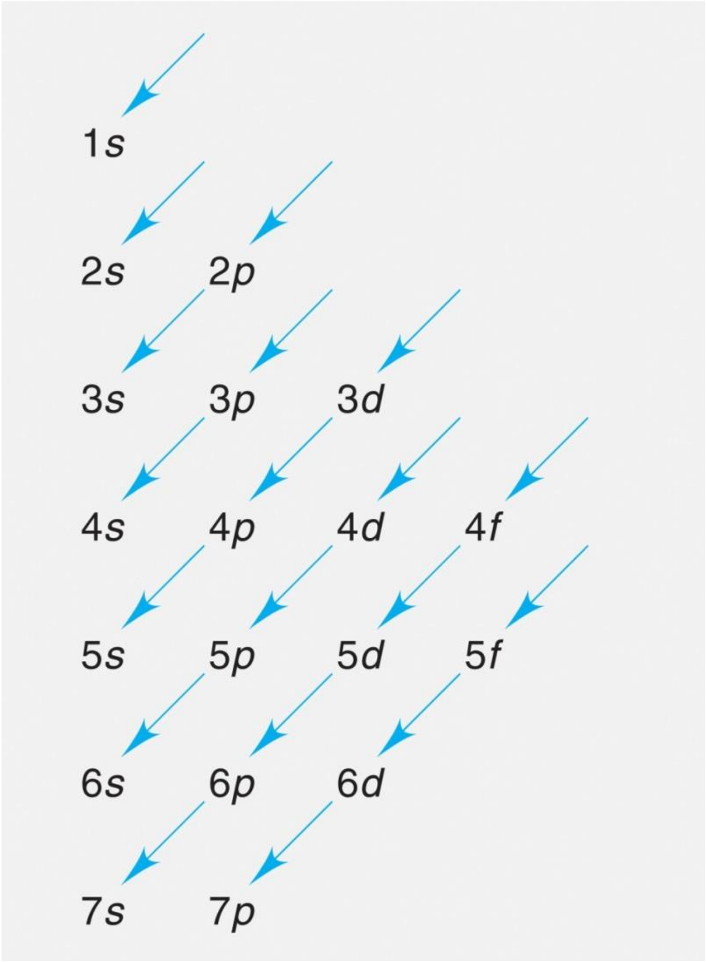


© The McGraw-Hill Companies, Inc./Stephen Frisch Photographer

Potassium reacting with water. **Chlorine reacting with potassium.**

Orbital Filling Order



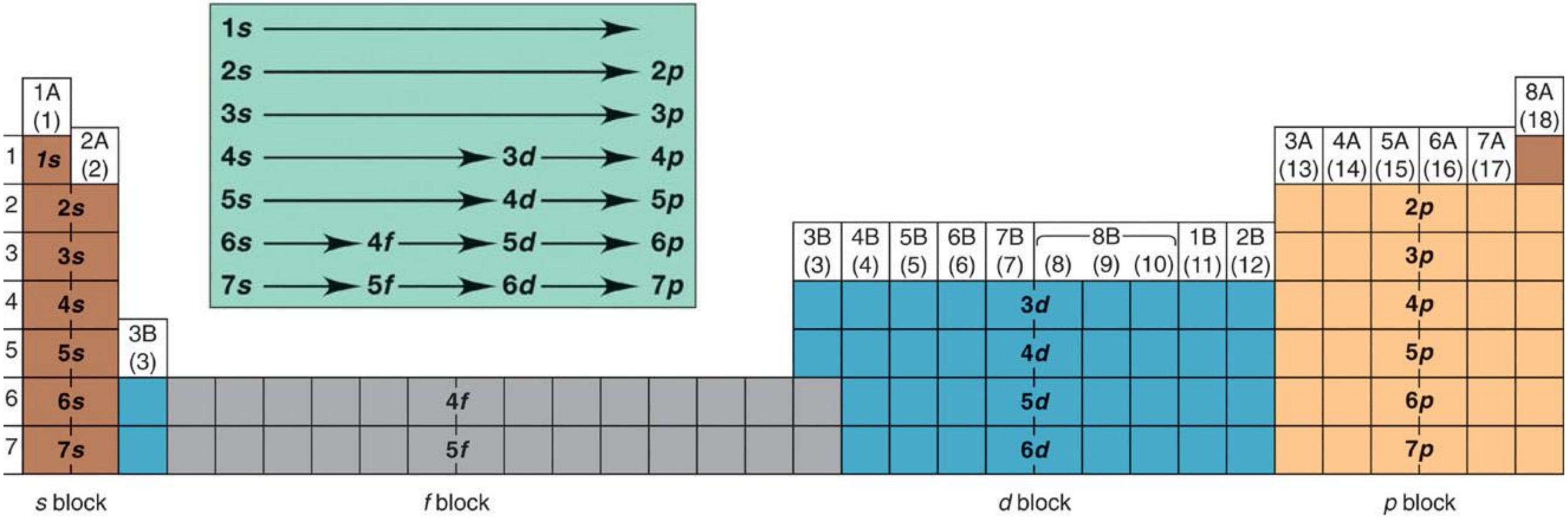


1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p, 8s

The ns orbital usually has a lower energy than the (n-1)*d* orbitals

Orbital Filling Order

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



Categories of Electrons

- **core (inner) electrons**: fill all the lower energy levels of an atom and any completed transition series.
- **Outer electrons**: farthest from the nucleus of an atom, in the highest energy level (highest n value).
- **valence electrons**: involve in forming compounds

for main-group elements:

valence electrons = the outer electrons

B $1s^2 2s^2 2p^1$ **Core = [He]** , **valence = $2s^2 2p^1$**

for transition elements:

valence electrons = the outer ns electrons

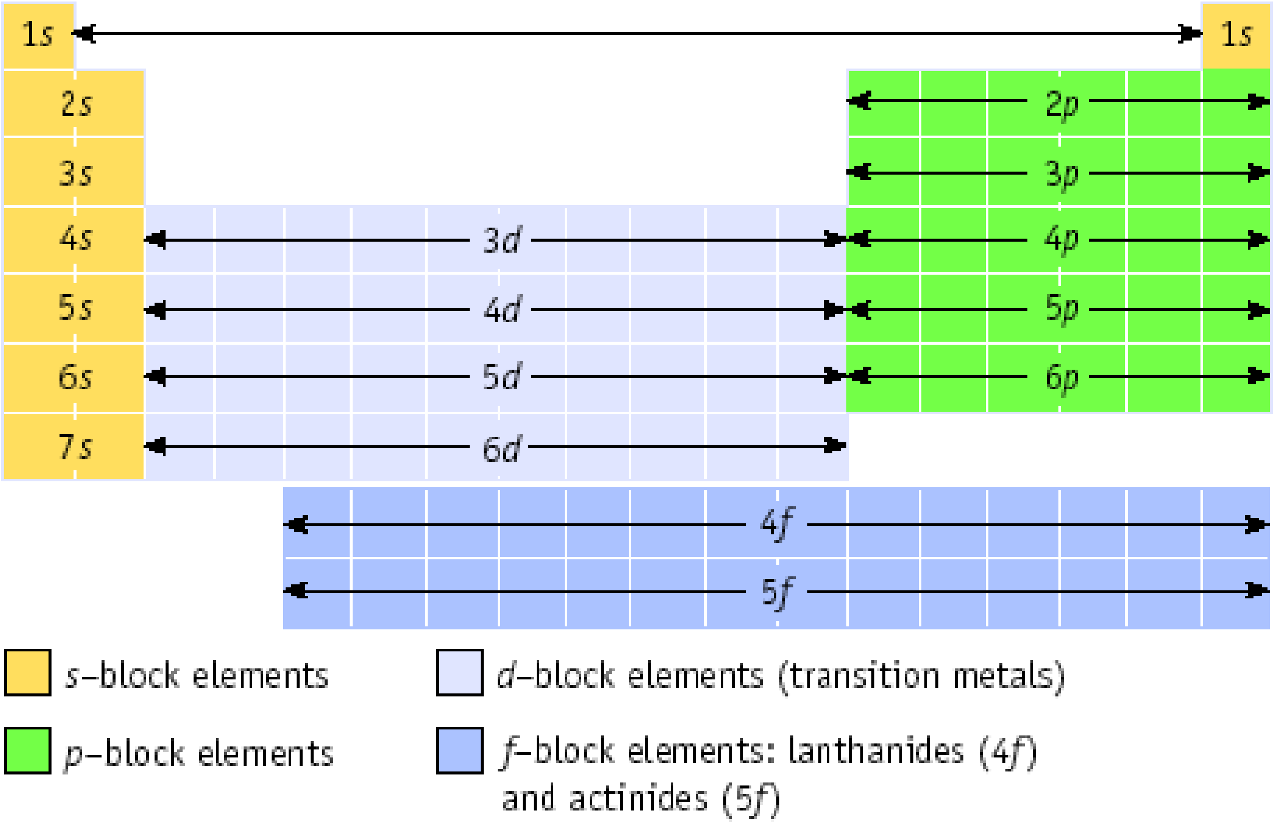
Fe [Ar] $4s^2 3d^6$ **Core = [Ar] $3d^6$**, **valence = $4s^2$**

(In some cases, (n-1)d electrons are also valence electrons)

Group and Period Numbers

For main-group elements:

- Group number = the number of outer electrons
- Period number = the n value of the highest energy level



Problems

A. The correct electron configuration for nitrogen is

1) $1s^2\ 2p^5$ 2) $1s^2\ 2s^2\ 2p^6$ 3) $1s^2\ 2s^2\ 2p^3$

B. The correct electron configuration for calcium is

1) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2$

2) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

3) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^8$

C. Write the full electron configuration, condensed notation, partial orbital diagrams showing valence electrons only, and number of inner electrons for each of the following elements:

Cl Fe Ba

Electron Configurations for Transition Elements

All 4th and 5th period elements (the **d-block elements**) have configuration

$[\text{core}]\ ns^2\ (n - 1)d^y$

Chromium

Iron

Copper

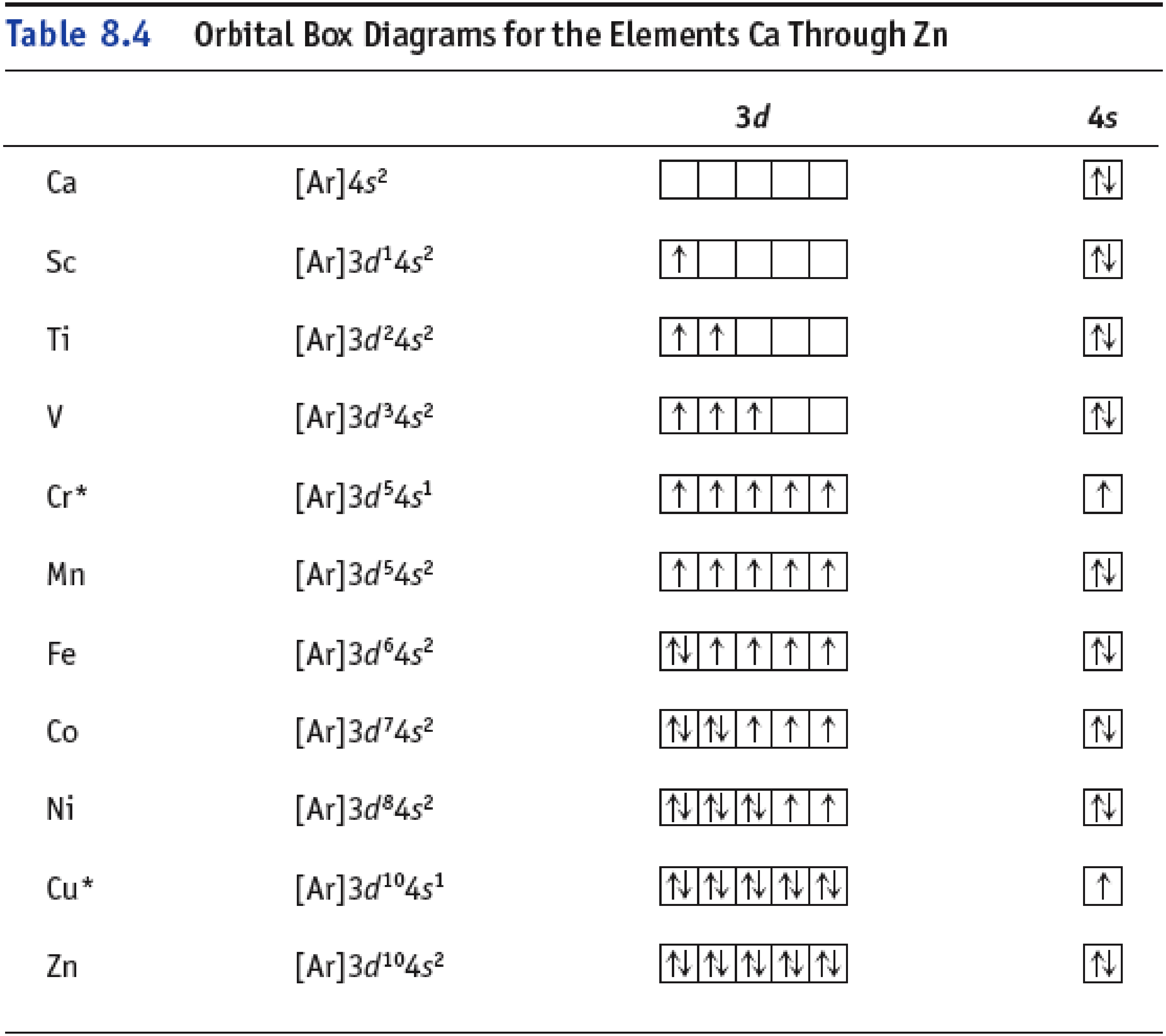


Table 8.4 Orbital Box Diagrams for the Elements Ca Through Zn

		3d					4s
Ca	$[\mathrm{Ar}]4s^2$						↑↓
Sc	$[\mathrm{Ar}]3d^1 4s^2$	↑					↑↓
Ti	$[\mathrm{Ar}]3d^2 4s^2$	↑	↑				↑↓
V	$[\mathrm{Ar}]3d^3 4s^2$	↑	↑	↑			↑↓
Cr*	$[\mathrm{Ar}]3d^5 4s^1$	↑	↑	↑	↑	↑	↑
Mn	$[\mathrm{Ar}]3d^5 4s^2$	↑	↑	↑	↑	↑	↑↓
Fe	$[\mathrm{Ar}]3d^6 4s^2$	↑↓	↑	↑	↑	↑	↑↓
Co	$[\mathrm{Ar}]3d^7 4s^2$	↑↓	↑↓	↑	↑	↑	↑↓
Ni	$[\mathrm{Ar}]3d^8 4s^2$	↑↓	↑↓	↑↓	↑	↑	↑↓
Cu*	$[\mathrm{Ar}]3d^{10} 4s^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑
Zn	$[\mathrm{Ar}]3d^{10} 4s^2$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓

*** Irregular filling pattern: The d-sublevels are particularly stable when electrons half-fill or completely fill them.**

Lanthanides and Actinides

Some **f-block elements** have configuration

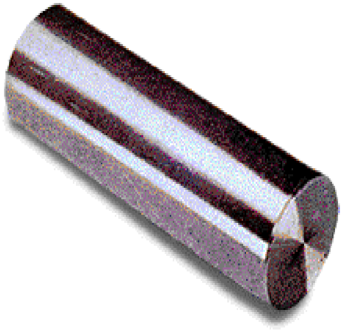
$[\text{core}]\ ns^2\ (n - 1)d^y\ (n - 2)f^z$

Cerium

$[\text{Xe}]\ 6s^2\ 5d^1\ 4f^1$

Uranium

$[\text{Rn}]\ 7s^2\ 6d^1\ 5f^3$



Others have configuration $[\text{core}]\ ns^2\ (n - 2)f^z$

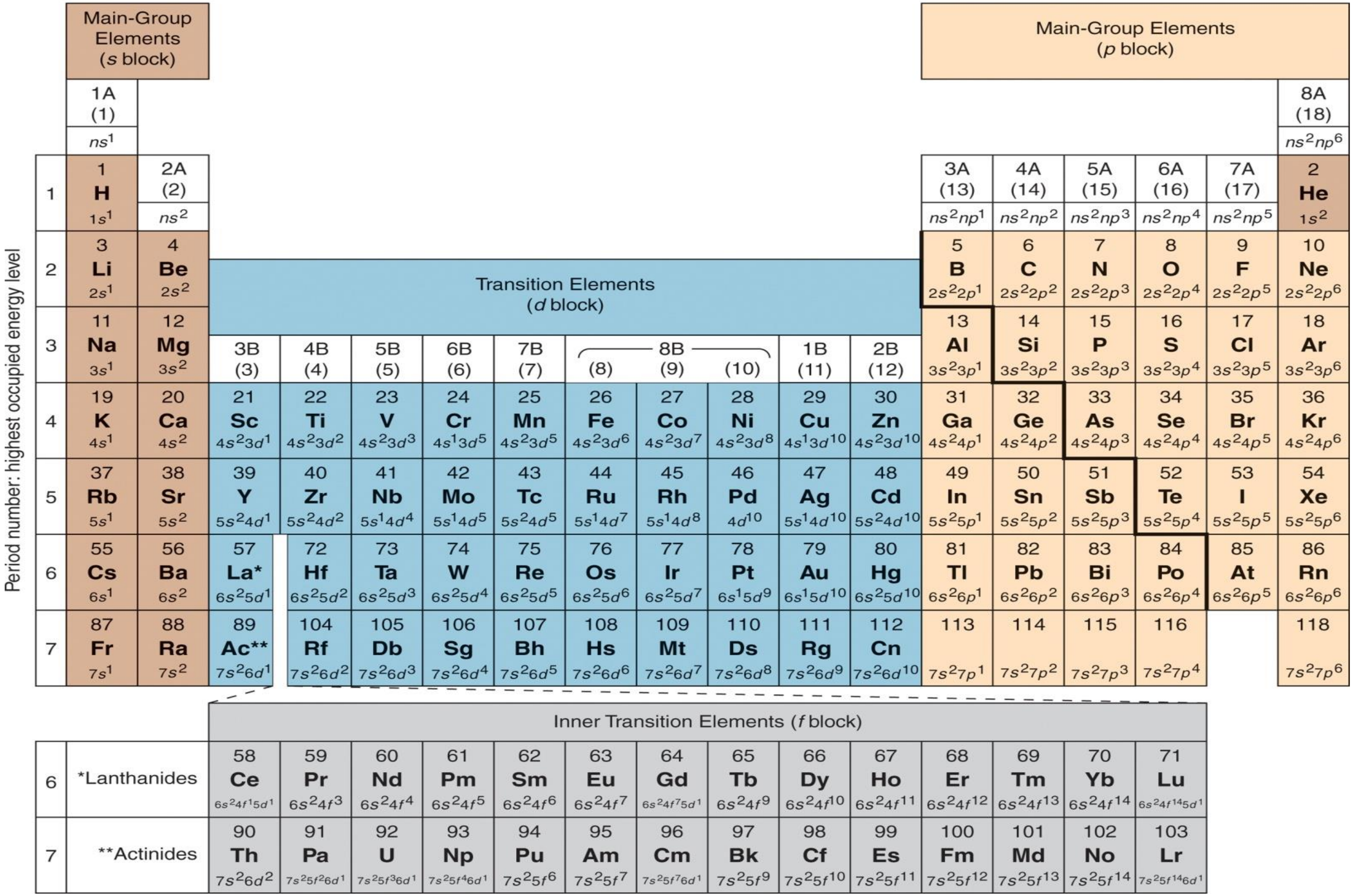
Neodymium

$[\text{Xe}]\ 6s^2\ 4f^4$

Californium

$[\text{Xe}]\ 7s^2\ 5f^{10}$

A periodic table of partial ground-state electron configurations



Main-Group Elements (*s* block)

Main-Group Elements (*p* block)

Transition Elements (*d* 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** $7s^25f^{14}6d^1$

Problems

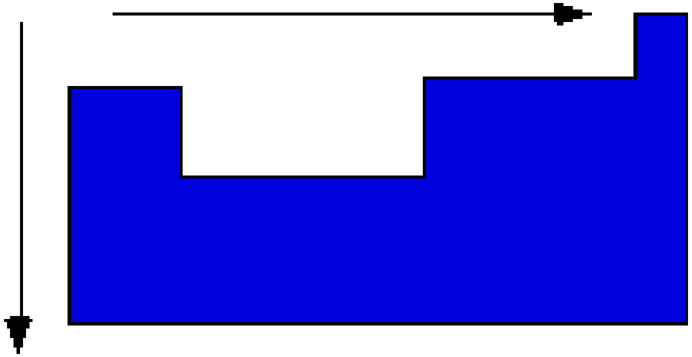
A. The partial electron configuration for Co and Sn

B. Using the periodic table, write the full electron configuration and condensed notation for each of the following elements:

Ag Pb Sr

8.3 Trends in Atomic Properties

- **Atomic size**
- **Ionization energy**
- **Electron affinity**



Higher effective nuclear charge Z_{eff}
Electrons held more tightly

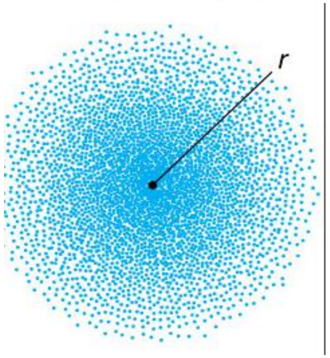
Larger orbitals.
Electrons held less tightly.

Atomic Size

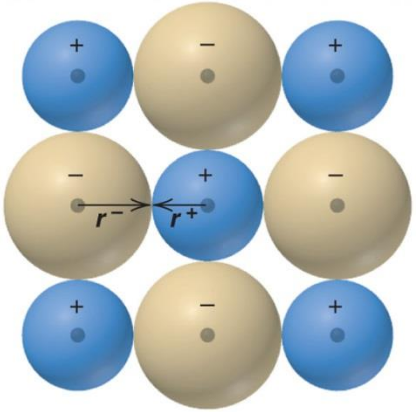
- is described by

the **atomic radius,**

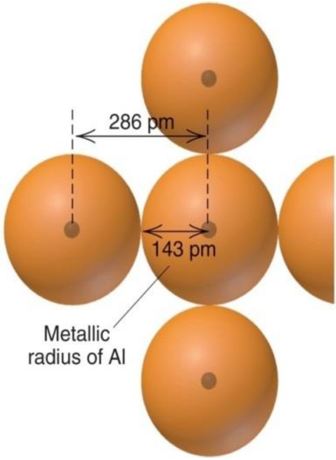
which is the distance from the nucleus to the valence electrons.



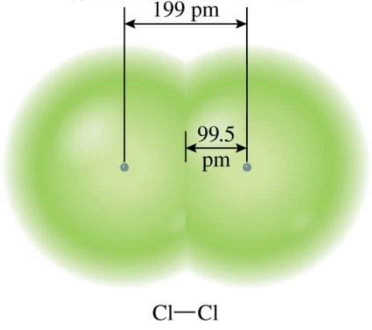
- **Metallic radius and/or ionic radius** is the shortest distance between nuclei of adjacent atoms in a crystal.

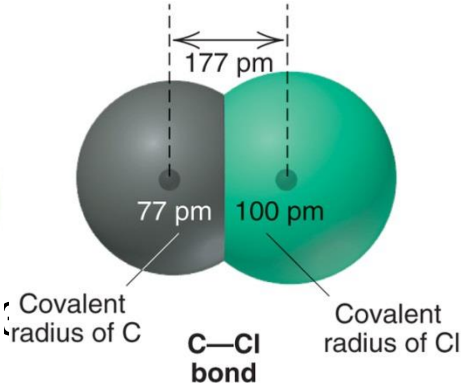


The metallic radius of aluminum

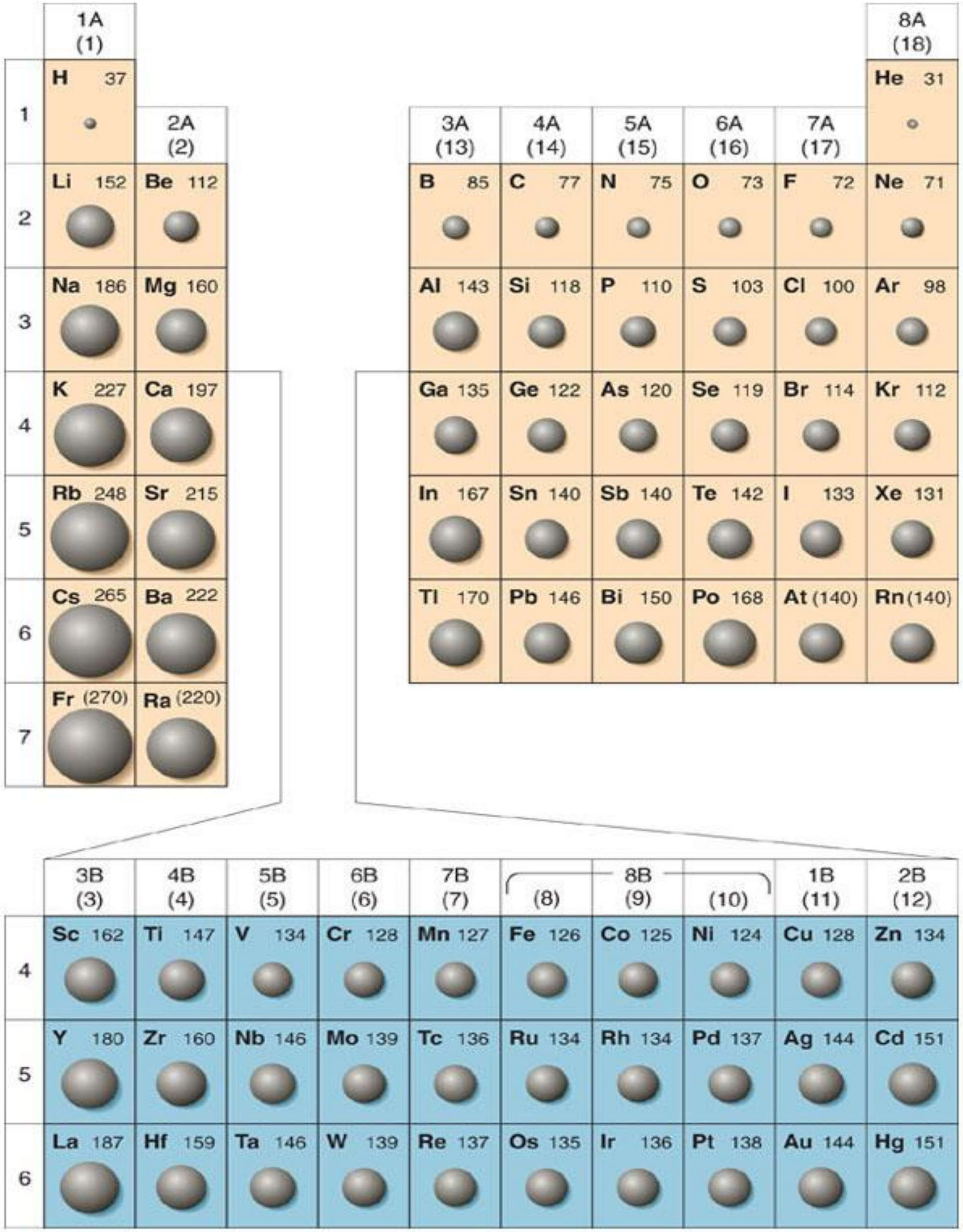


- **Covalent radius** is estimated as one-half the distance between the centers of two bonded identical atoms.





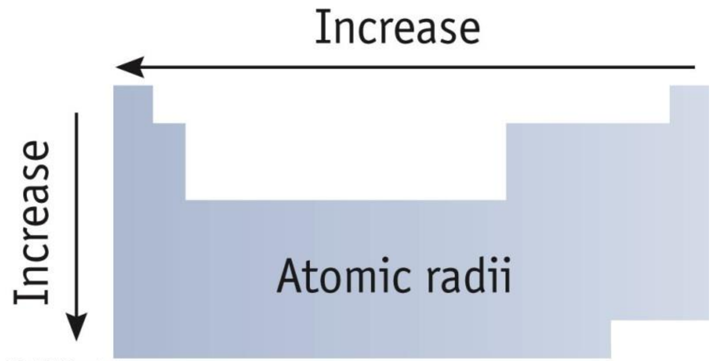
Atomic Radii for Main-Group Elements



	1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
1	H 37							He 31
2	Li 152	Be 112	B 85	C 77	N 75	O 73	F 72	Ne 71
3	Na 186	Mg 160	Al 143	Si 118	P 110	S 103	Cl 100	Ar 98
4	K 227	Ca 197	Ga 135	Ge 122	As 120	Se 119	Br 114	Kr 112
5	Rb 248	Sr 215	In 167	Sn 140	Sb 140	Te 142	I 133	Xe 131
6	Cs 265	Ba 222	Tl 170	Pb 146	Bi 150	Po 168	At (140)	Rn (140)
7	Fr (270)	Ra (220)						

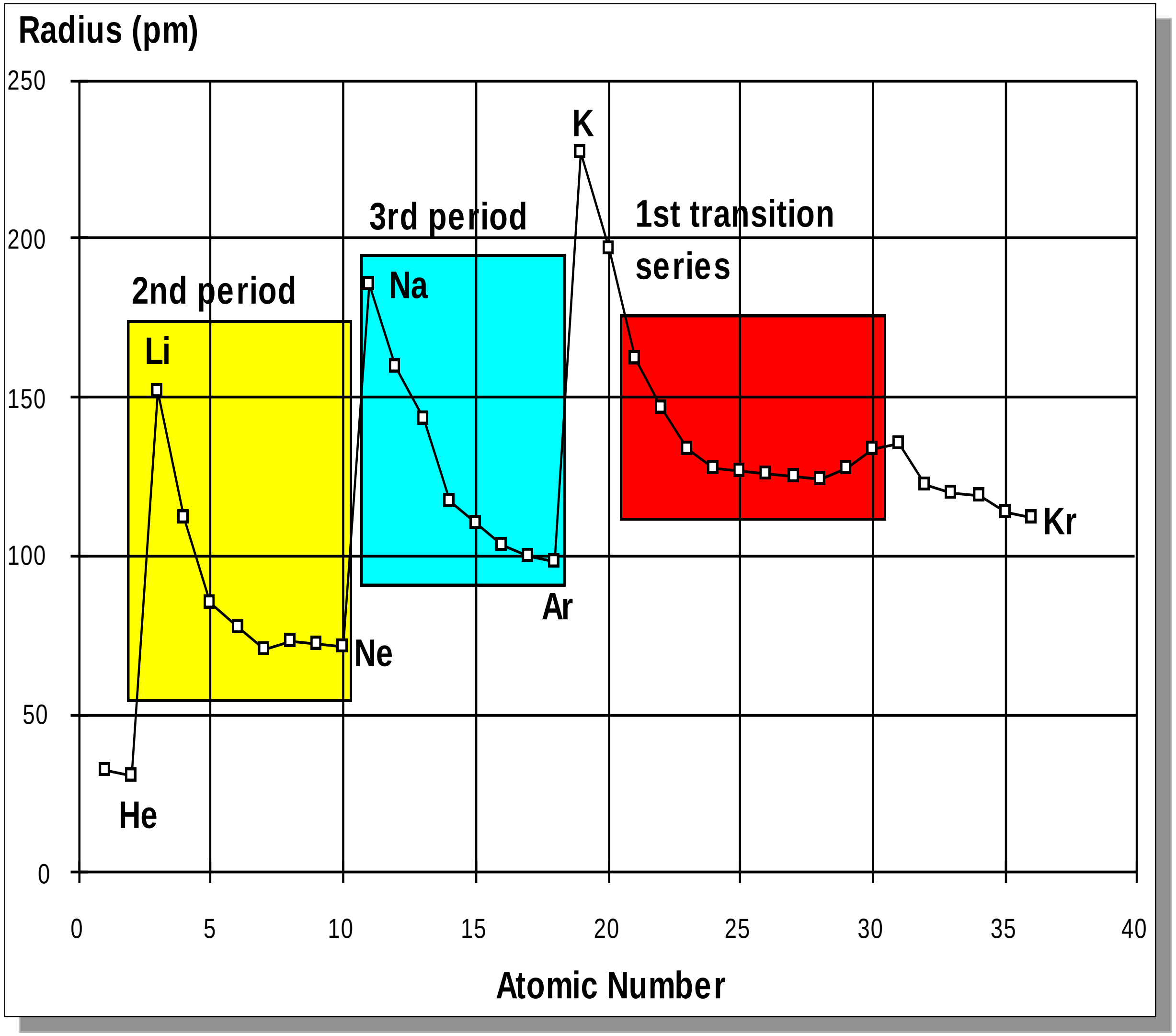
	3B	4B	5B	6B	7B	8B			1B	2B
	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)
4	Sc 162	Ti 147	V 134	Cr 128	Mn 127	Fe 126	Co 125	Ni 124	Cu 128	Zn 134
5	Y 180	Zr 160	Nb 146	Mo 139	Tc 136	Ru 134	Rh 134	Pd 137	Ag 144	Cd 151
6	La 187	Hf 159	Ta 146	W 139	Re 137	Os 135	Ir 136	Pt 138	Au 144	Hg 151

- **Size DECREASES** from left to right of a period due to increasing Z_{eff} across a period
- **Size INCREASES** from top to bottom of a group, because outer electrons are farther and feel less attraction from the nucleus



© Brooks/Cole, Cengage Learning

Trends in Atomic Size of Transition Elements



3d subshell is inside the 4s subshell.

4s electrons feel a more or less constant Z_{eff}.

Sizes stay about the same and chemistries are similar!

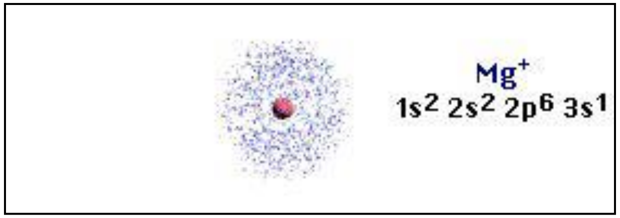
Ionization Energy (IE)

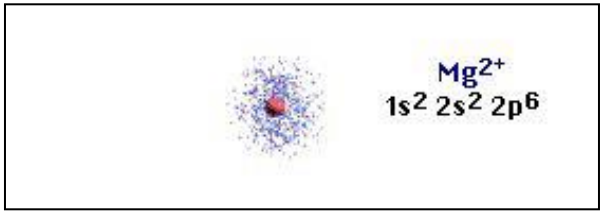
IE = energy required to remove 1 mol of electrons from 1 mol of gaseous atoms or ions. It indicates how tight the atom holds electrons.

Mg (g) + 738 kJ ---> Mg^+ (g) + e- IE_1

Mg^+ (g) + 1451 kJ ---> Mg^{2+} (g) + e- $IE_2 > IE_1$

Mg^{2+} (g) + 7733 kJ ---> Mg^{3+} (g) + e- $IE_3 >>> IE_3$





Energy cost very high to dip into a shell of lower n.

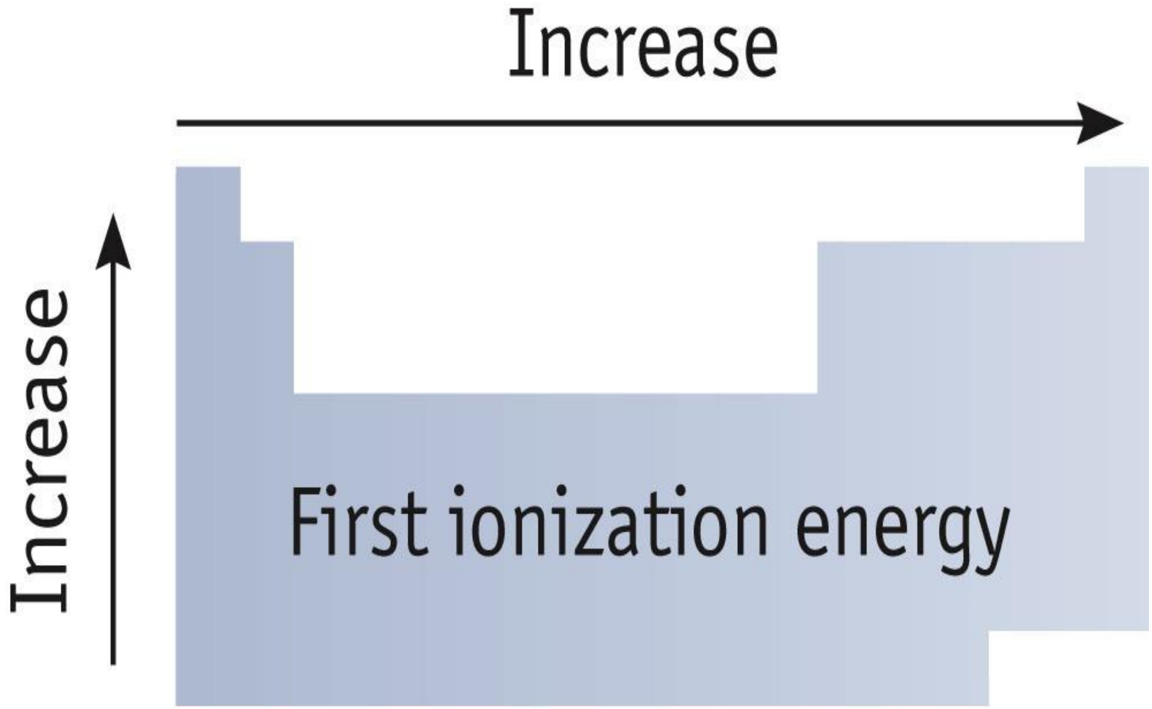
This is why ion charge # = Group #

Trends in First Ionization Energy (IE_1)

- IE increases across a period because Z_{eff} increases.
- IE decreases down a group because value of n and size increases.

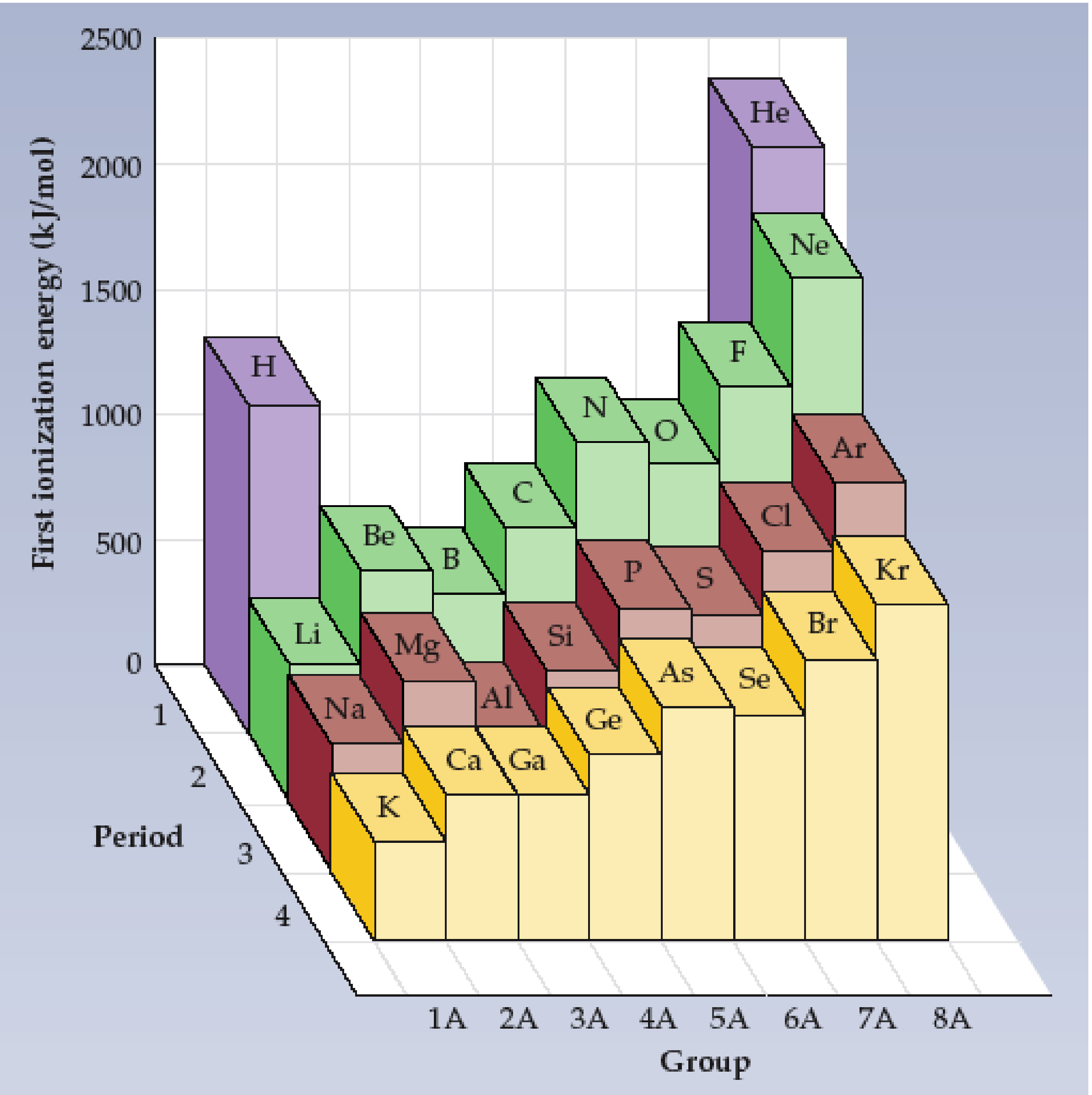
Atoms with a *low IE* tend to form *cations*.

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



© Brooks/Cole, Cengage Learning

Trends in the First Ionization Energy



Two exceptions:

In each period, 1) IE$_1$ of elements in group 3A is smaller than IE$_1$ of elements in group 2A

2) IE$_1$ of elements in group 6A smaller than those in group 5A. (Hints: Z$_{eff}$ and/or electron pair repulsion)

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).

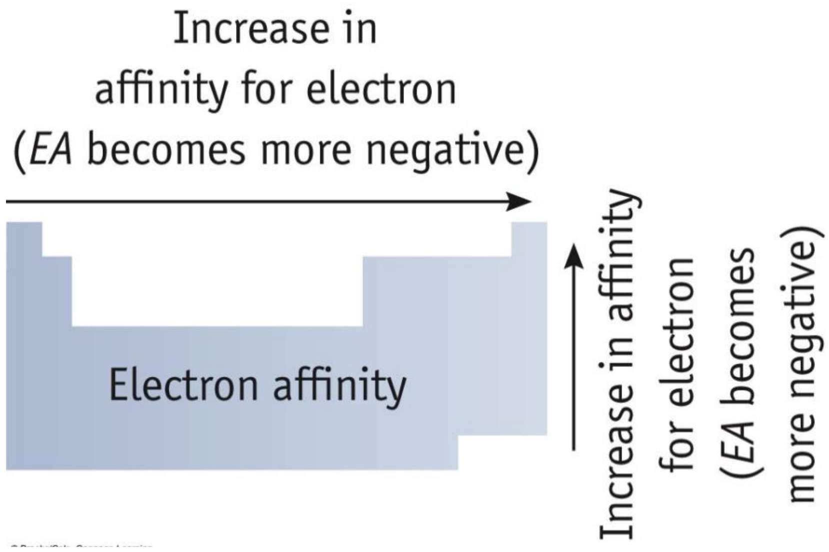
The complete electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^3$.

Electron Affinity (EA)

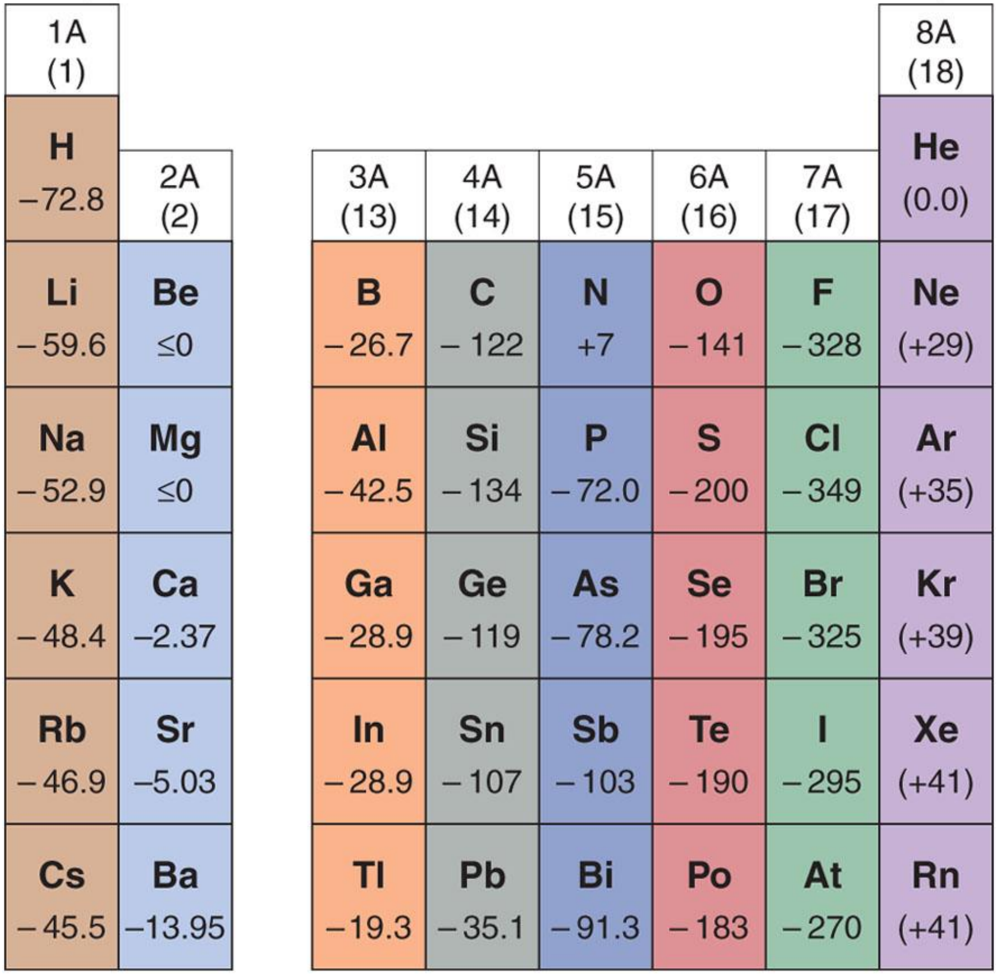
Electron affinity is the energy change when one mole of electrons is added to one mole of gaseous atoms. It indicates the ability of an atom to attract electron.

$A\,(g) + e^- \longrightarrow A^-\,(g) \quad EA = \Delta E$ (negative)

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

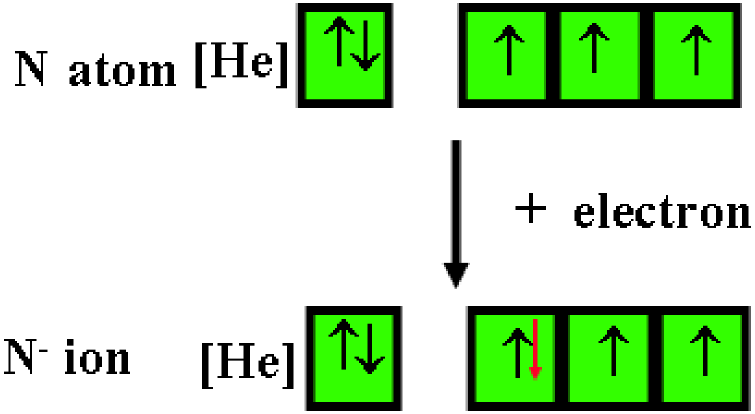


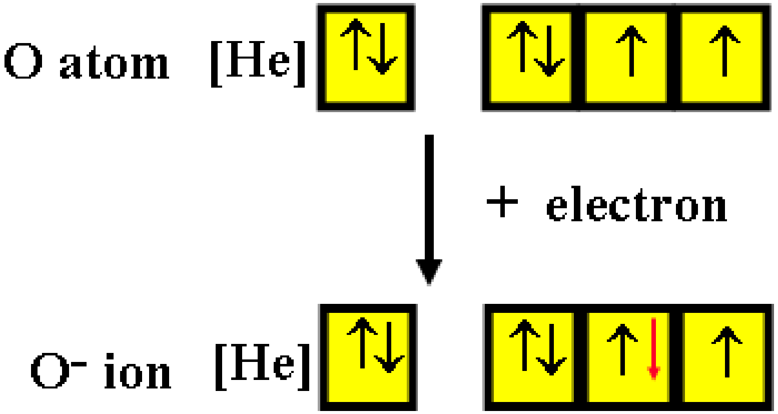
Periodicity in Electron Affinity (EA)



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)

Electron Affinity





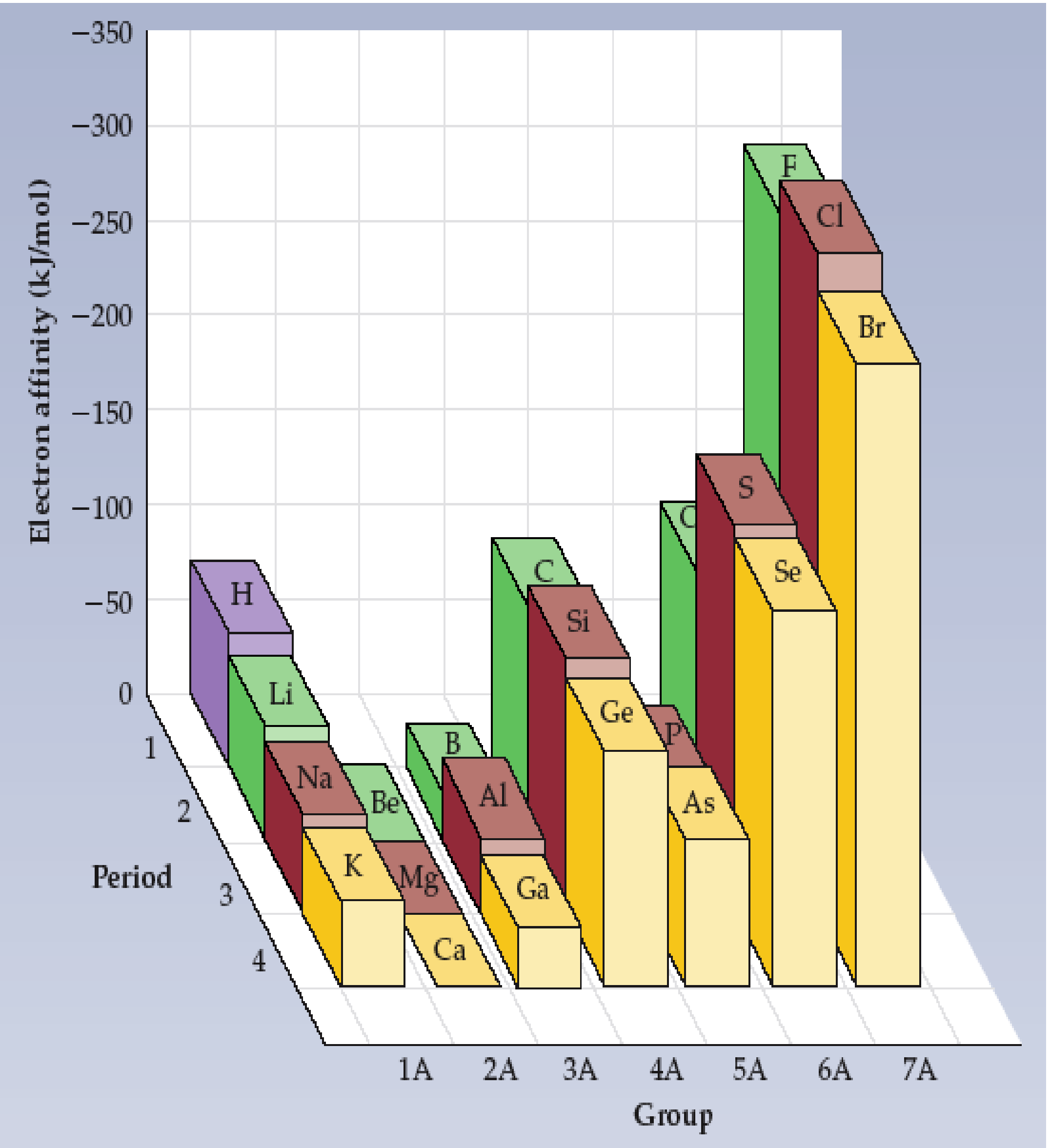
EA = +7 kJ

EA = -141 kJ

ΔE is ENDOthermic for N due to electron-electron repulsions.

ΔE is EXOthermic because O has an affinity for an e-.

Trends in Electron Affinity (EA)



Two exceptions:

(1) in the same period, elements in group 2A have lower EA than those in group 3A; similarly, those in group 5A lower than those in group 6A. (Hints: Z* and/or electron pair repulsion)

(2) Except groups 1A and 2A, EA of the first element is lower than that of the second element in the same group. (Hint: atomic/ionic sizes)

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.

Problems

Select the element in each pair with the larger atomic radius.

A. Li or K

B. K or Br

C. P or Cl

Select the element in each pair with the higher Ionization energy.

A. Li or K

B. K or Br

C. P or Cl

8.4 Atomic Properties & Chemical Reactivity

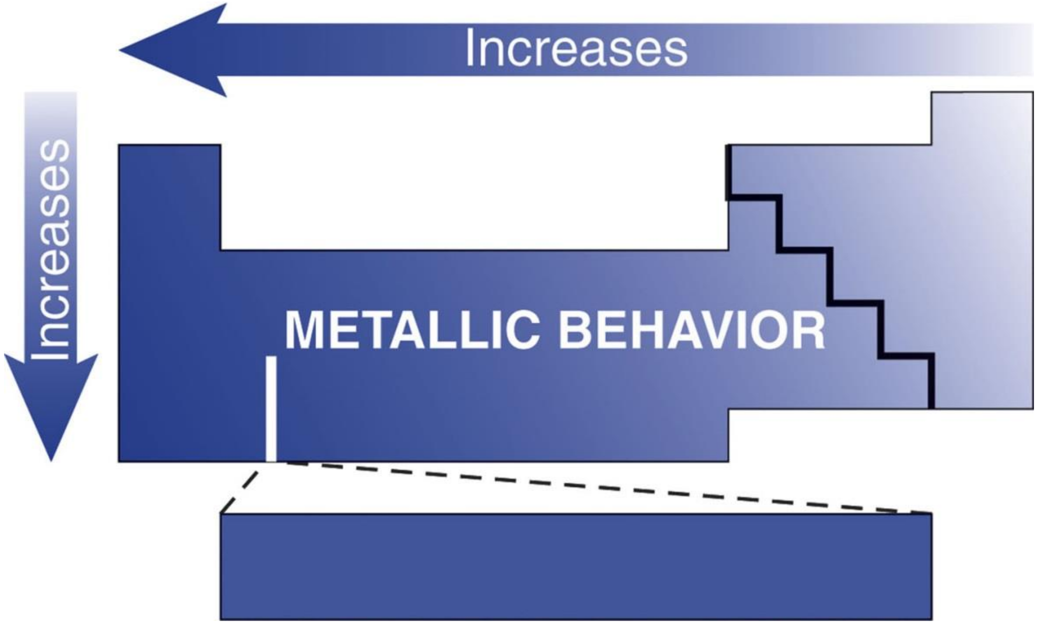
Metallic Behavior

- **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.**

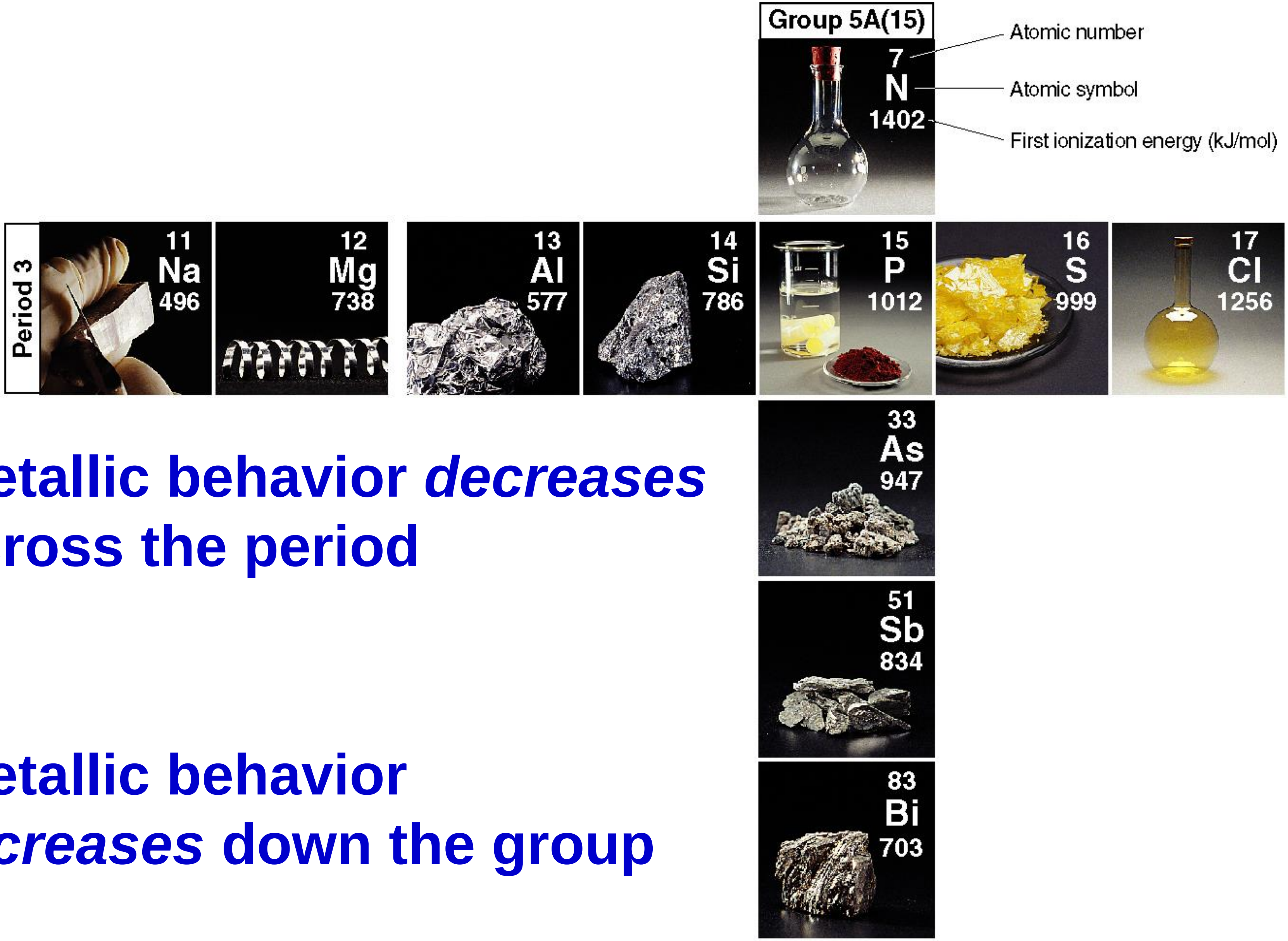
Metallic behavior

Metallic behavior of an element indicates its tendency to lose electrons.

- **Metals** tend to lose electrons to nonmetals during reactions. The more active the metal, the more easily it loses electrons to form ions.
- **Nonmetals** tend to gain electrons.



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



Metallic behavior *decreases* across the period

Metallic behavior *increases* down the group

The McGraw-Hill Companies, Inc./Stephen Frisch Photographer

Metallic Behavior

Redox Behavior of Main-Group Elements:

- **Metals** of the Main Groups are reducing agents with the **highest oxidation number = Group number**

e.g., Group 1A(1) = 1+

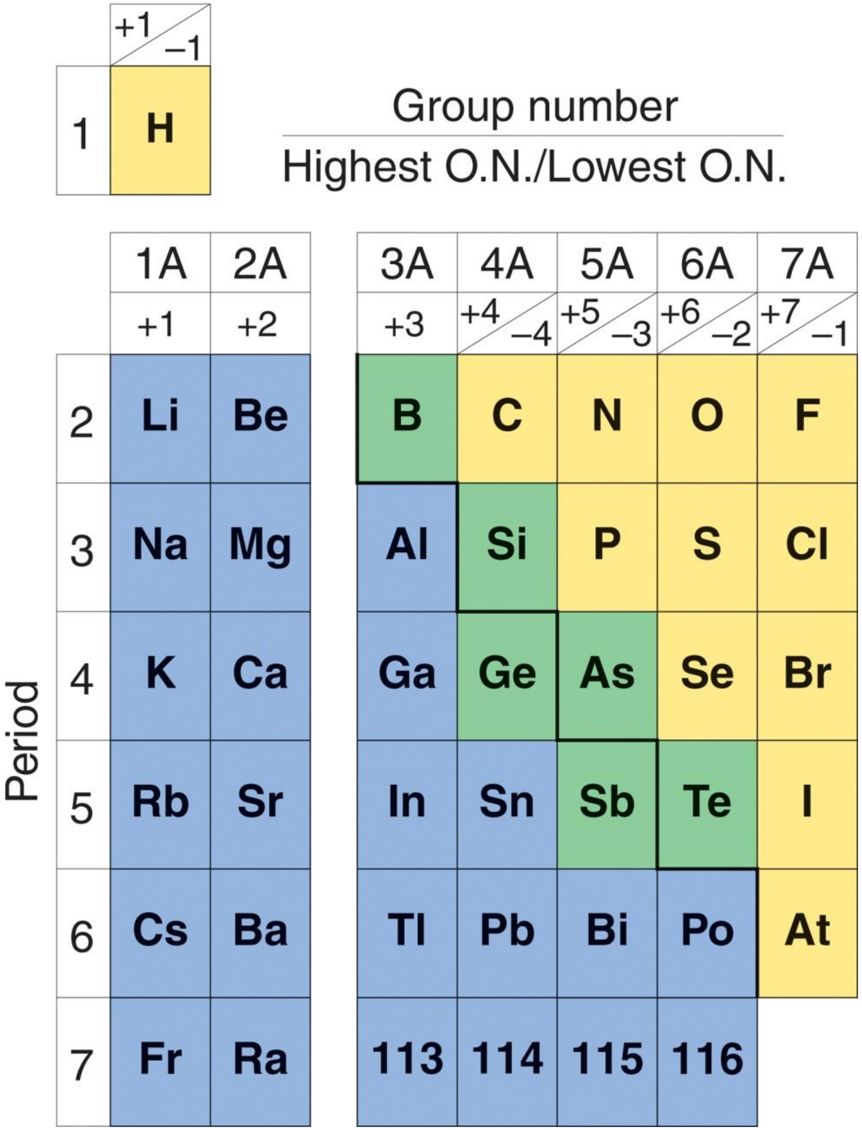
Group 2A(2) = 2+

- **Nonmetals and metalloids** are oxidizing agents with **highest oxidation number = group number – 8 or 18**

e.g., Group 6A(16) = 6 - 8 = 2 -

or 16 - 18 = 2 -

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



	+1/−1
1	H

Group number
Highest O.N./Lowest O.N.

Period	1A	2A	3A	4A	5A	6A	7A
	+1	+2	+3	+4/−4	+5/−3	+6/−2	+7/−1
2	Li	Be	B	C	N	O	F
3	Na	Mg	Al	Si	P	S	Cl
4	K	Ca	Ga	Ge	As	Se	Br
5	Rb	Sr	In	Sn	Sb	Te	I
6	Cs	Ba	Tl	Pb	Bi	Po	At
7	Fr	Ra	113	114	115	116	

Metallic Behavior

Acid-Base Behavior of Oxides

- **Metal oxides** are **ionic**. In water they act as **bases**, producing OH^- ions and reacting with acids

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

  $CaO\ (s) + H_2O\ (l) \rightarrow Ca(OH)_2\ (s)$

- **Nonmetal oxides** are **covalent**. In water they act as **acids**, producing H^+ ions and reacting with bases.

  $N_2O_5\ (s) + H_2O\ (l) \rightarrow 2\ HNO_3\ (aq)$

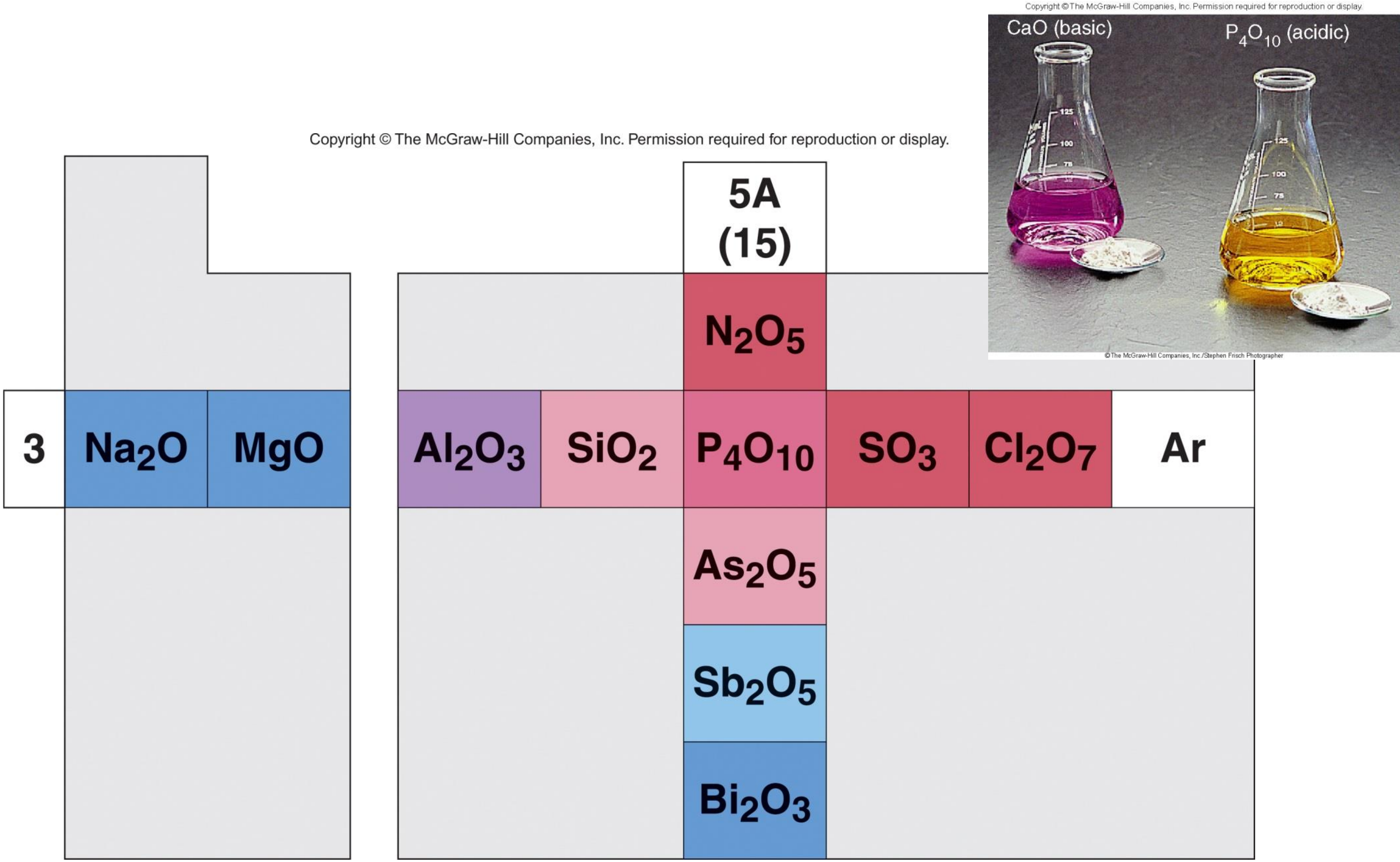
  $P_4O_{10}\ (s) + 6\ H_2O\ (l) \rightarrow 4\ H_3PO_4\ (aq)$

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

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

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

Acid-base behavior of some element oxides



©The McGraw-Hill Companies, Inc./Stephen Frisch Photographer

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

8.4 Atomic Properties & Chemical Reactivity

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.

To form cations from metals, remove 1 or more e- from subshell of highest n

$$\mathrm{Na}(1s^22s^22p^63s^1) \rightarrow e^- + \mathrm{Na}^+([\mathrm{He}]2s^22p^6)$$

[isoelectronic with Ne]

To form anions from nonmetals, add 1 or more e- to subshell of highest n

$$\mathrm{Br}([\mathrm{Ar}]4s^23d^{10}4p^5) + e^- \rightarrow \mathrm{Br}^-([\mathrm{Ar}]4s^23d^{10}4p^6)$$

[isoelectronic with Kr]

Electron configurations of Monatomic Ions

Ions without a noble gas configuration

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 remains with a filled (*n*-1)*d* sublevel.

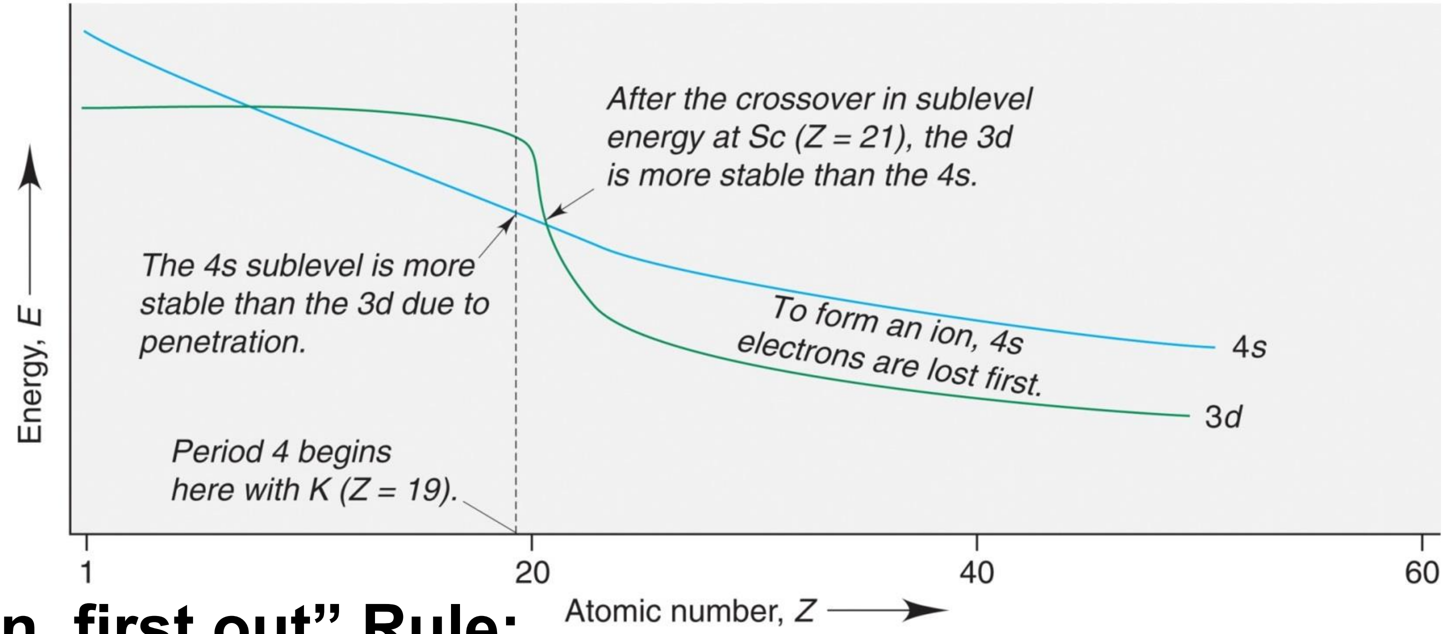
$$\mathrm{Sn}([\mathrm{Kr}]5s^2 4d^{10} 5p^2) \rightarrow 4e^- + \mathrm{Sn}^{4+}\ ([\mathrm{Kr}]4d^{10})$$

An *inert pair configuration* is attained when a metal loses only the *np* electrons.
The ion attains the stability of a filled *ns* and (*n*-1)*d* sublevels.

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

Properties of Monoatomic Ions

Electron configurations of transition metal ions



"First in, first out" Rule:

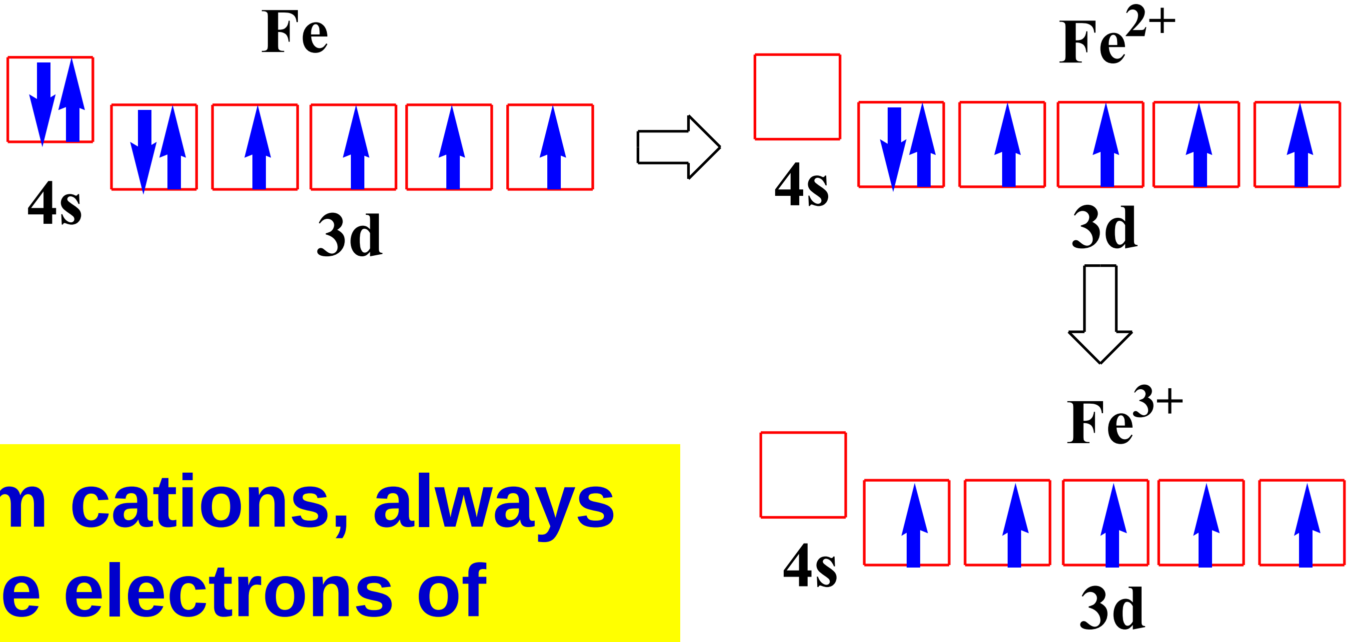
To fill the electrons in energy levels, the ns subshell is filled before the (n-1) d subshell.
To form a transition metal ion, remove electrons of highest n subshells first!

Properties of Monoatomic Ions

Electron configurations of transition metals

Remove **ns electrons first, then (n - 1)d electrons**.

Fe [Ar] **$4s^2$** $3d^6$ ---> Fe^{2+} [Ar] **$4s^0$** $3d^6$ + $2e^-$

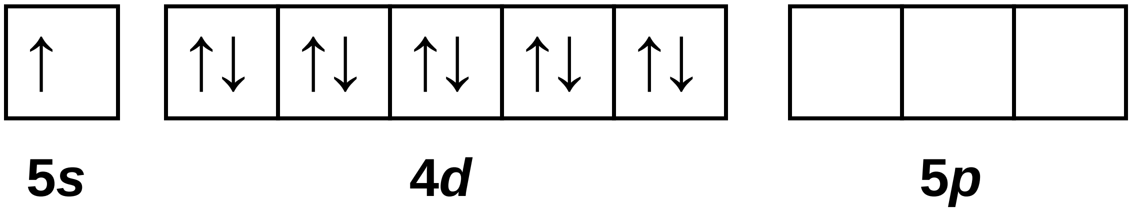


To form cations, always remove electrons of highest n value first!

Magnetic Properties of Transition Metal ions

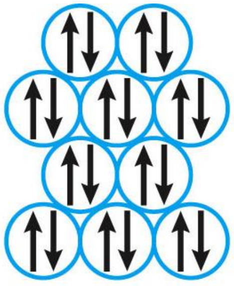
A species with one or more unpaired electrons exhibits *paramagnetism* – it is attracted by a magnetic field.

Ag (Z = 47)

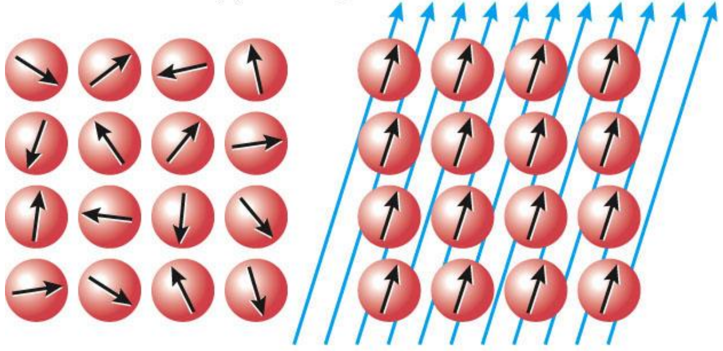


A species with all its electrons paired exhibits *diamagnetism* – it is not attracted (and is slightly repelled) by a magnetic field.

(b) Diamagnetism



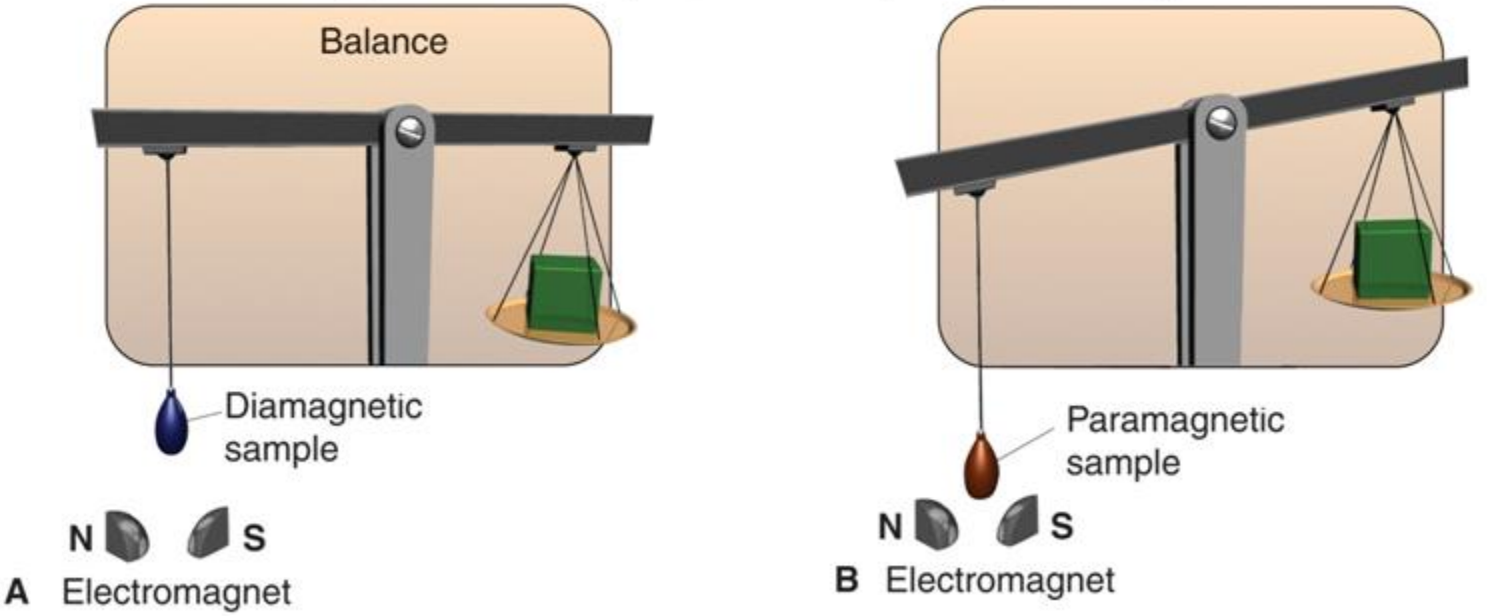
(a) Paramagnetism



© Brooks/Cole, Cengage Learning

Magnetic Properties of Transition Metal ions

Measuring the magnetic behavior of a sample

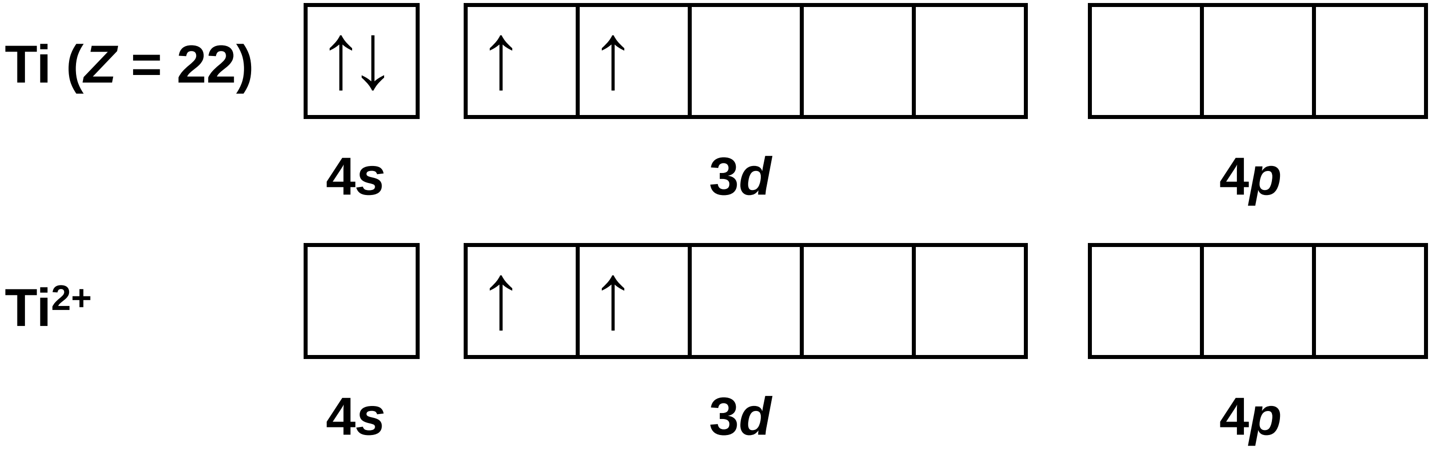


The apparent mass of a diamagnetic substance is unaffected by the magnetic field.

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.



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

Problems

A) Write the electron configuration in full and condensed notation for the following ions. Indicate whether they are (a) isoelectric to noble gas configuration, (b) pseudo-noble gas configuration, or (c) inert pair configuration

Br^-	N^{3-}	K^+
Sr^{2+}	S^{2-}	Ni^{2+}

B) Give the symbol of the element that has

1. [Ar]$4s^2\ 3d^5$
2. Four 3*p* electrons
3. Six electrons in the 4*d* sublevel
4. The element that has the electron configuration

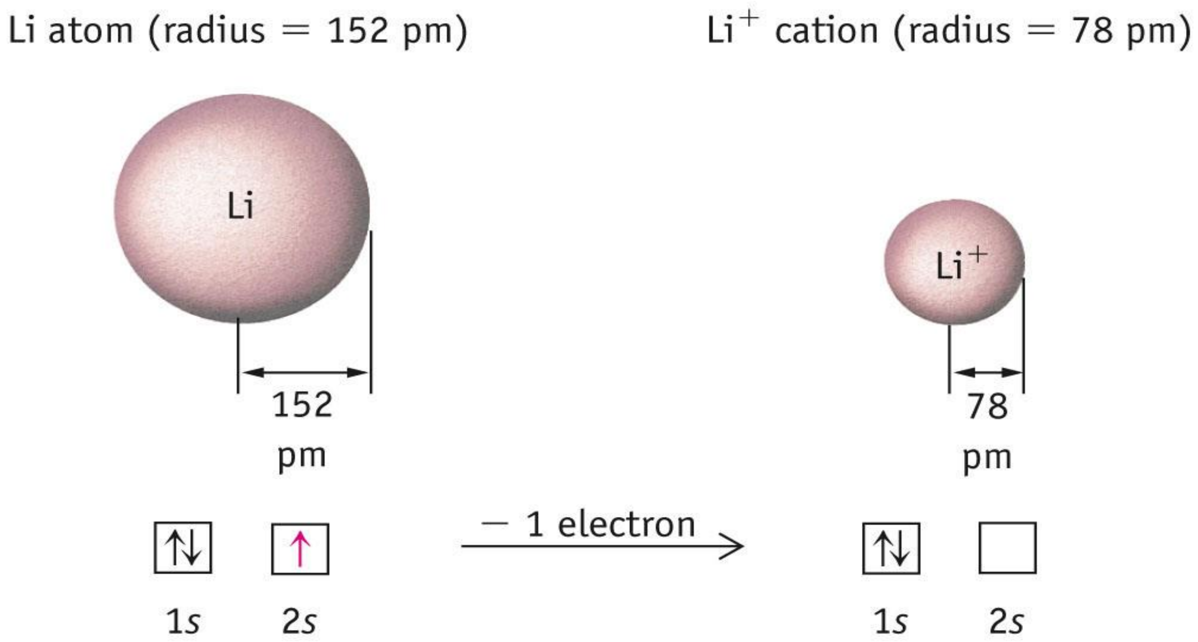
$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$$

Which one(s) is/are paramagnetic?

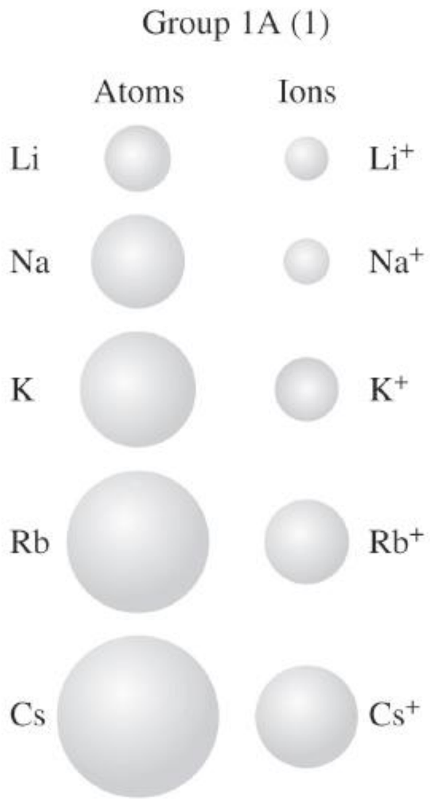
C) Are K, S, I paramagnetic or diamagnetic?

Ionic Size & Atomic Size

Forming a cation.



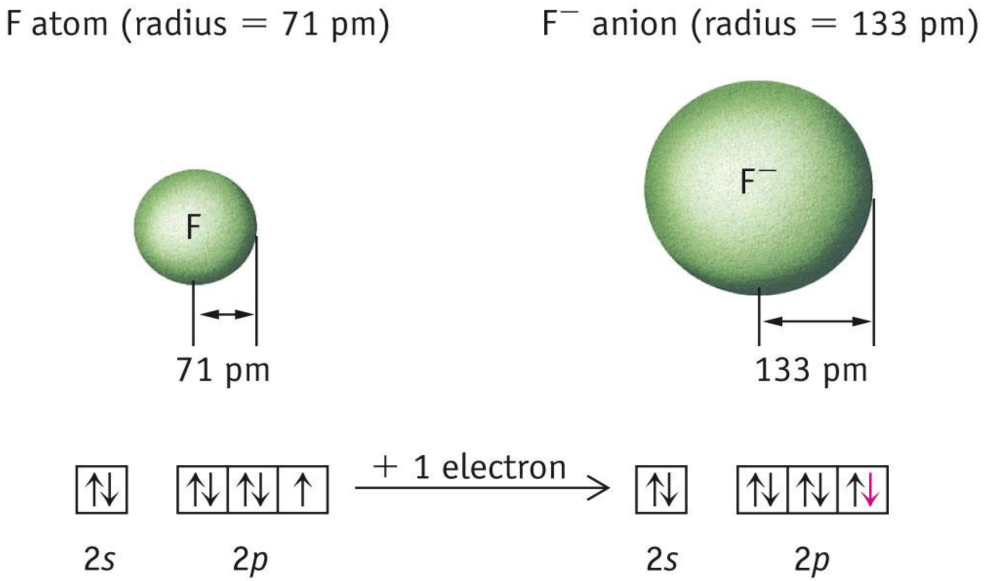
© Brooks/Cole, Cengage Learning



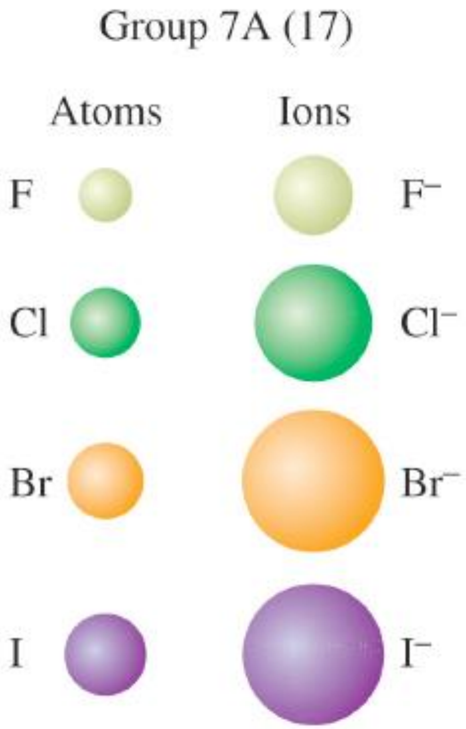
- **CATIONS are SMALLER than their parent atoms.**
- **Due to Z_{eff}, the electron/proton attraction has gone UP and so size DECREASES.**

Ionic Size & Atomic Size

Forming an anion.



© Brooks/Cole, Cengage Learning

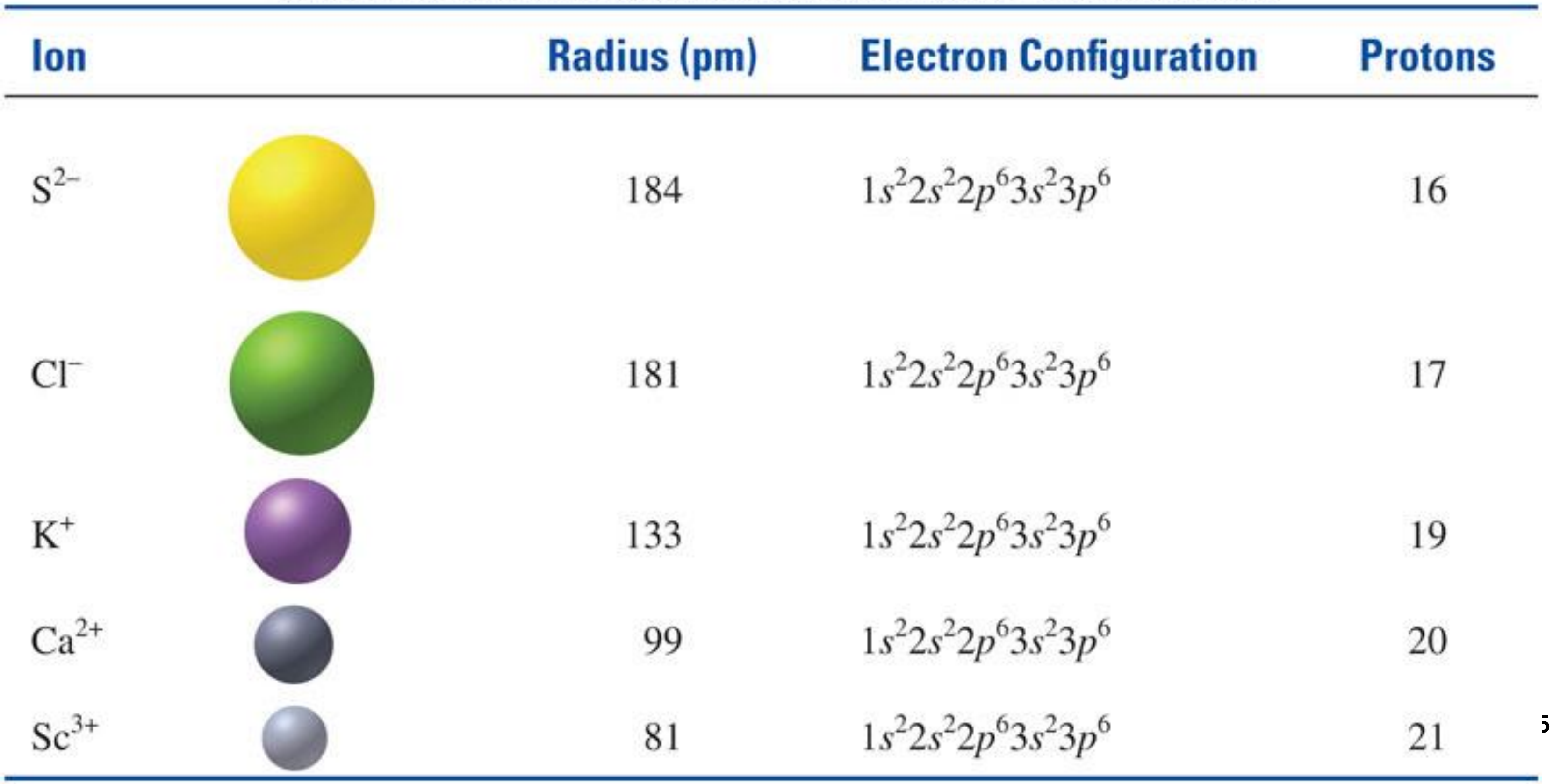


- **ANIONS are LARGER than their parent atoms.**
- **Due to the increasing number of core electrons, the electron/proton attraction has gone DOWN and so size INCREASES.**

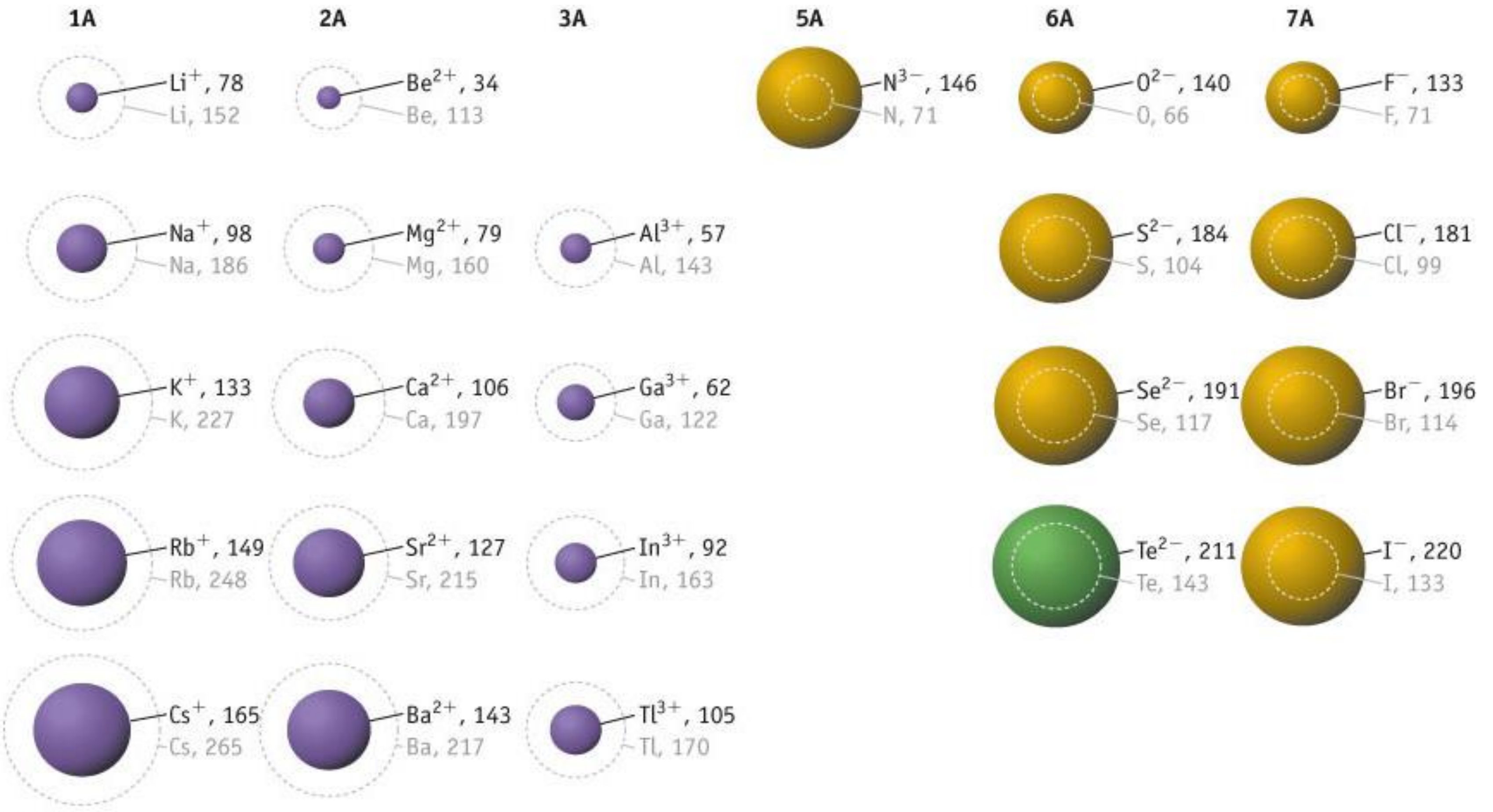
Ionic Size & Atomic Size

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+



Ion	Radius (pm)	Electron Configuration	Protons
S^{2-}	184	$1s^22s^22p^63s^23p^6$	16
Cl^-	181	$1s^22s^22p^63s^23p^6$	17
K^+	133	$1s^22s^22p^63s^23p^6$	19
Ca^{2+}	99	$1s^22s^22p^63s^23p^6$	20
Sc^{3+}	81	$1s^22s^22p^63s^23p^6$	21



Explain why K^+ ion is smaller than that of Cl^- ion; Rb^+ ion size < Br^- ion size; Cs^+ ion size smaller than I^- size.

Select the larger atom or ion in the following:

1. Mg or Mg^{2+} 2. S or S^{2-} 3. Na^+ or F^-