Chapter 4-5. Quantum Theory and Atomic Structure

6.1 The Nature of Light

6.2 Atomic Spectra

6.3 The Wave-Particle Duality of Matter and Energy

6.4 The Quantum-Mechanical Model of the Atom

6.1 The Wave Nature of Light

Visible light is a type of ***electromagnetic radiation***.

The wave properties of electromagnetic radiation are described by three variables:

- **frequency (ν)**, cycles per second
- **wavelength (λ)**, the distance a wave travels in one cycle
- **amplitude**, the height of a wave crest or depth of a trough.

The ***speed of light*** is a constant:

$$c = \nu \times \lambda$$

$$= 3.00 \times 10^8 \text{ m/s in a vacuum}$$

Figure 6.1 The reciprocal relationship of frequency and wavelength.

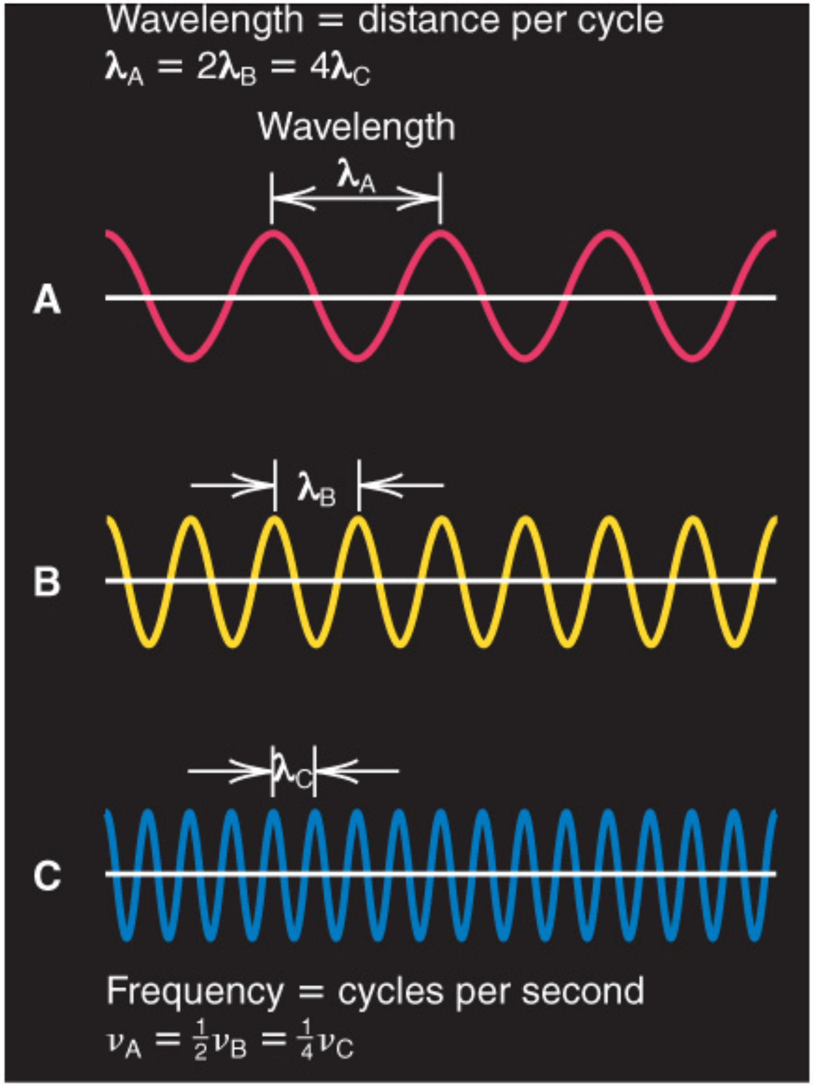


Figure 6.4 Different behaviors of waves and particles.

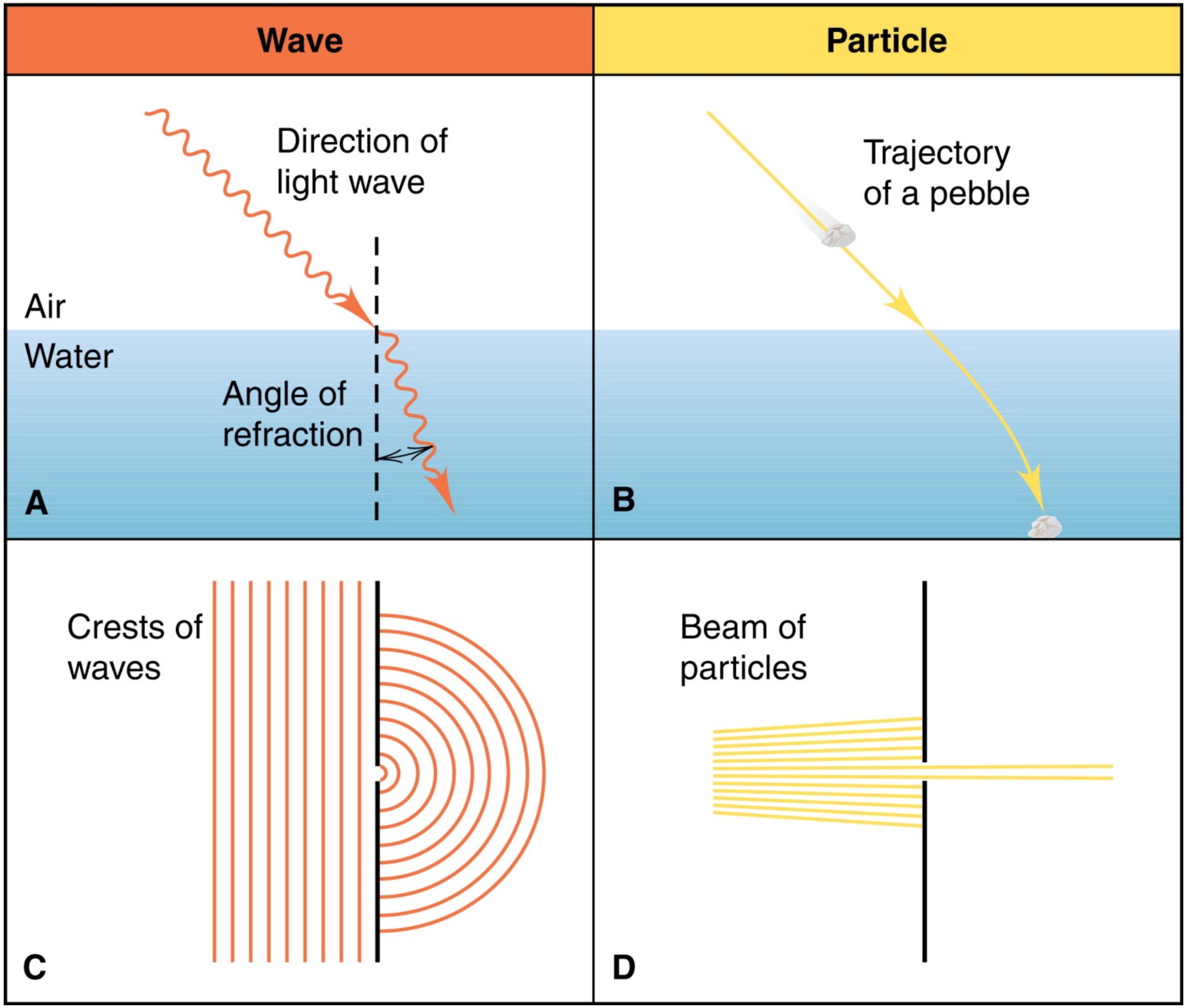
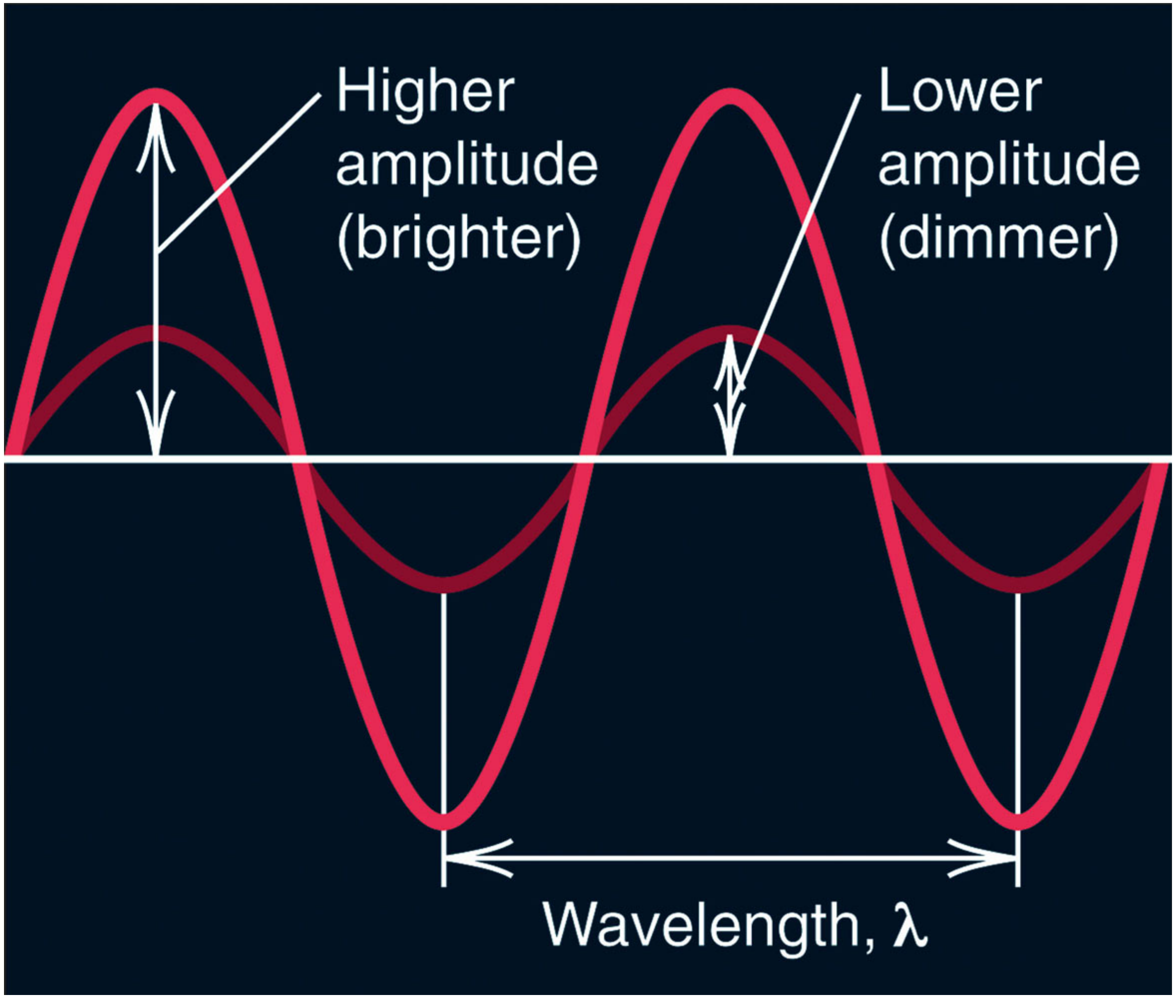


Figure 6.2 Differing amplitude (brightness, or intensity) of a wave.



Energy and frequency

A solid object emits visible light when it is heated to about 1000 K. This is called ***blackbody radiation***.

The *color* (and the intensity) of the light changes as the temperature changes. Color is related to ***wavelength*** and ***frequency***, while temperature is related to ***energy***.

Energy is therefore related to frequency and wavelength:

$$\boxed{E = n\,h\nu}$$

E = energy
n is a positive integer
h is Planck's constant
ν is frequency

The Quantum Theory of Energy

Any object (including atoms) can emit or absorb only ***certain quantities*** of energy.

Energy is ***quantized***; it occurs in fixed quantities, rather than being continuous. Each fixed quantity of energy is called a ***quantum***.

An atom changes its energy state by emitting or absorbing one or more ***quanta*** of energy.

$\Delta E = nh\nu$ where n can only be a whole number.

Energy of one photon (quantum) , **$\Delta E = h\nu$**

Figure 6.8 The photoelectric effect.

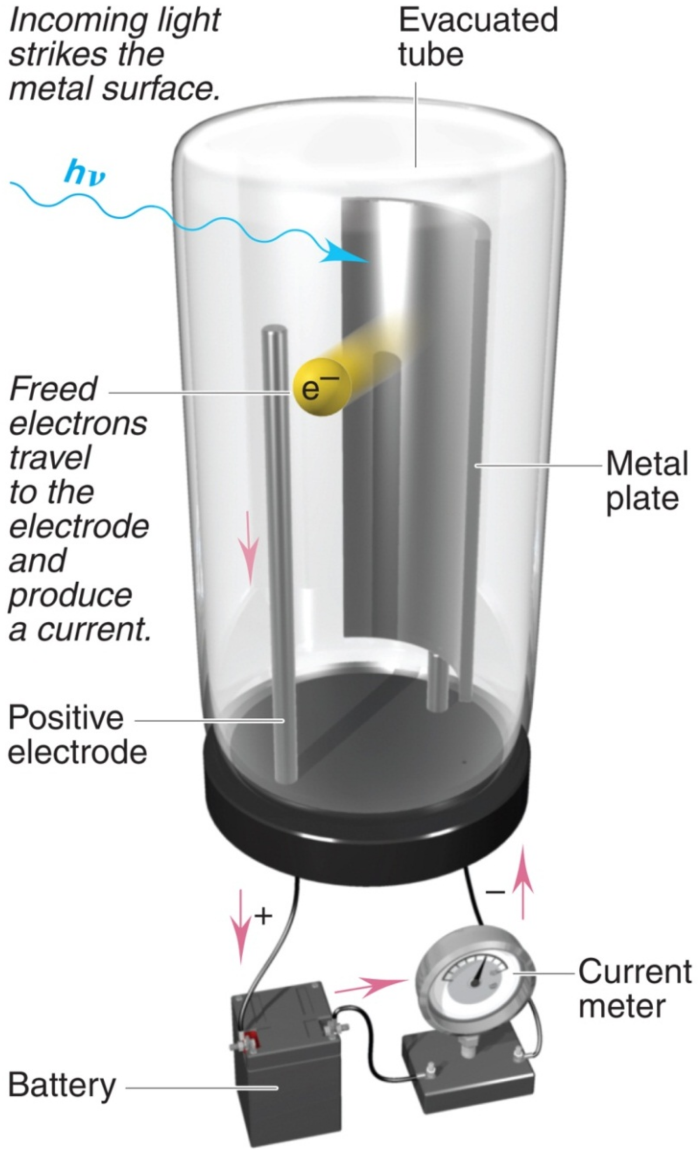
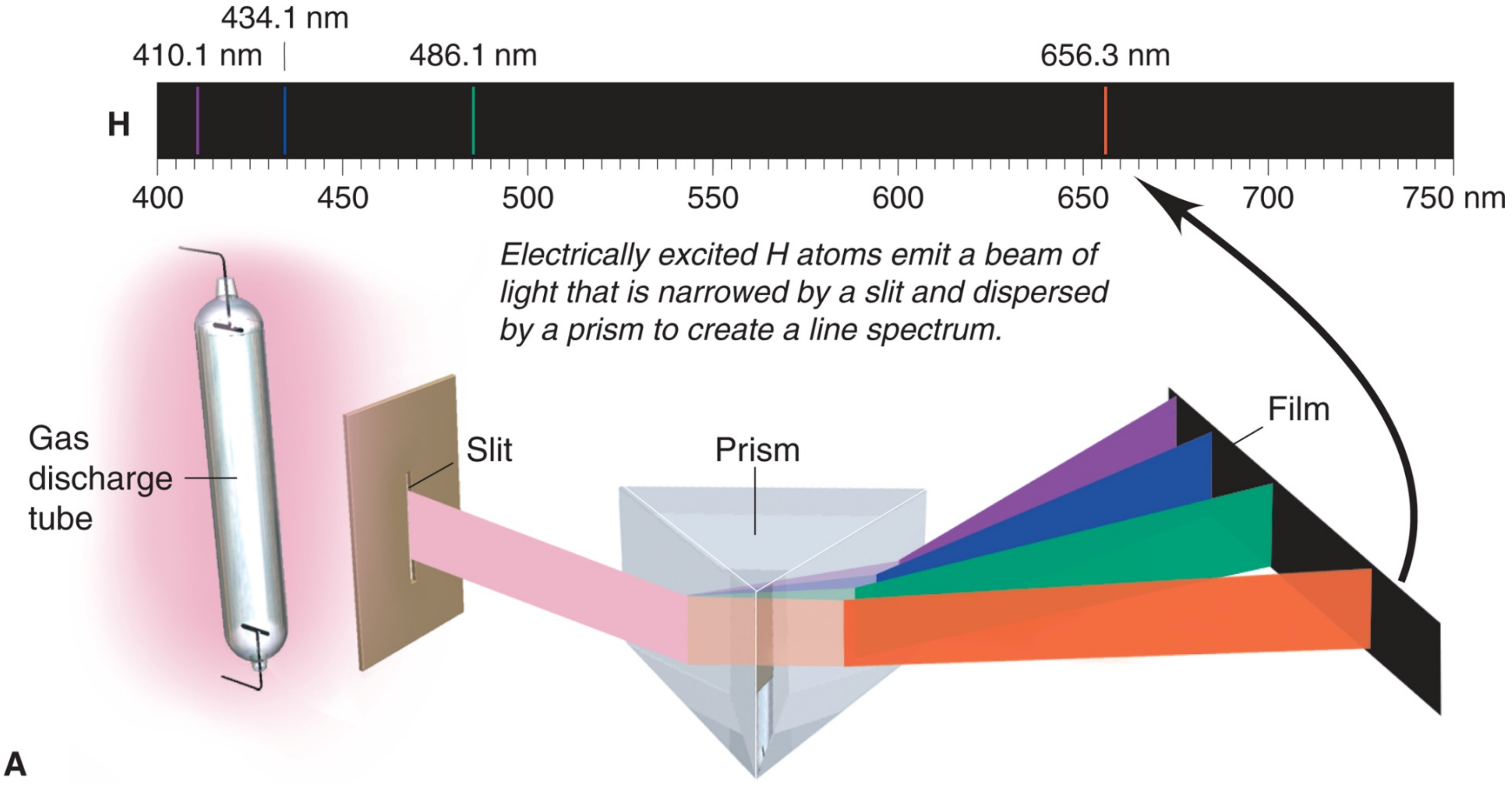


Figure 6.9A The line spectrum of atomic hydrogen.



Electrically excited H atoms emit a beam of light that is narrowed by a slit and dispersed by a prism to create a line spectrum.

Figure 6.9B The line spectra of Hg and Sr.

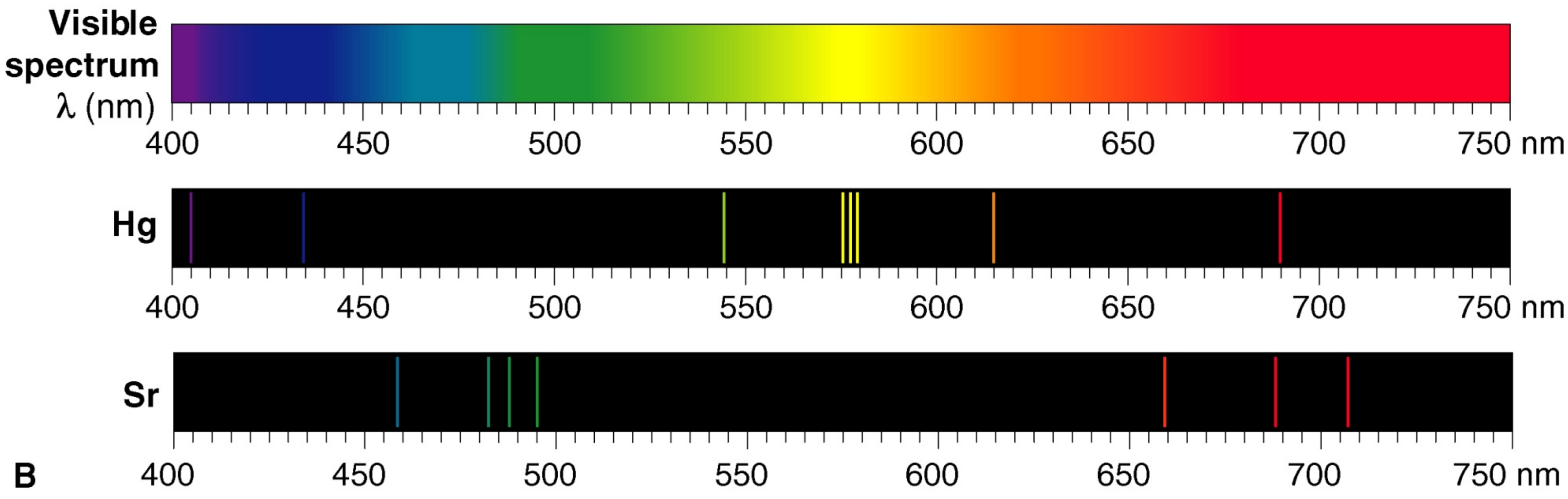
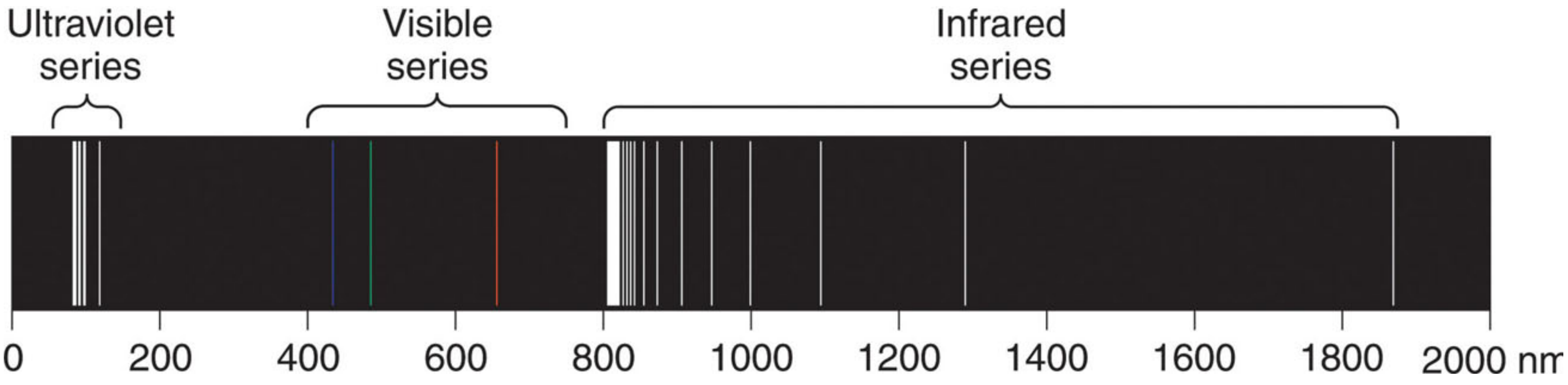


Figure 6.10 Three series of spectral lines of atomic hydrogen.



Rydberg equation

$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

R is the Rydberg constant = 1.096776x10^{7} m^{-1}

for the visible series, n_1 = 2 and n_2 = 3, 4, 5, ...

The Bohr Model of the Hydrogen Atom

Bohr's atomic model postulated the following:

- The H atom has only certain energy levels, which Bohr called ***stationary states***.
  - Each state is associated with a fixed circular orbit of the electron around the nucleus.
  - The higher the energy level, the farther the orbit is from the nucleus.
  - When the H electron is in the first orbit, the atom is in its lowest energy state, called the ***ground state***.

- The atom does not radiate energy while in one of its stationary states.
- The atom changes to another stationary state only by absorbing or emitting a photon.
  - The energy of the photon ($h\nu$) equals the difference between the energies of the two energy states.
  - When the E electron is in any orbit higher than $n = 1$, the atom is in an ***excited state***.

Figure 6.12 The Bohr explanation of three series of spectral lines emitted by the H atom.

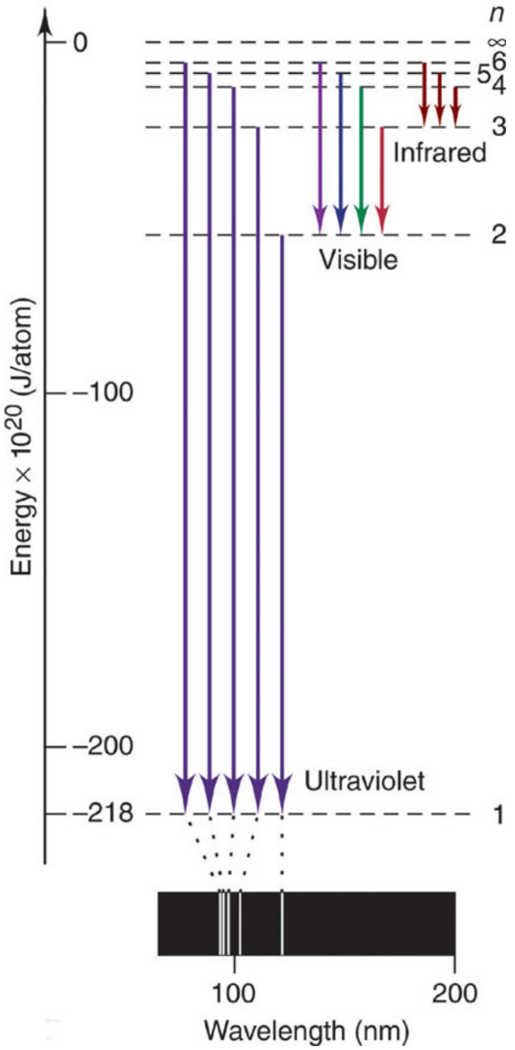
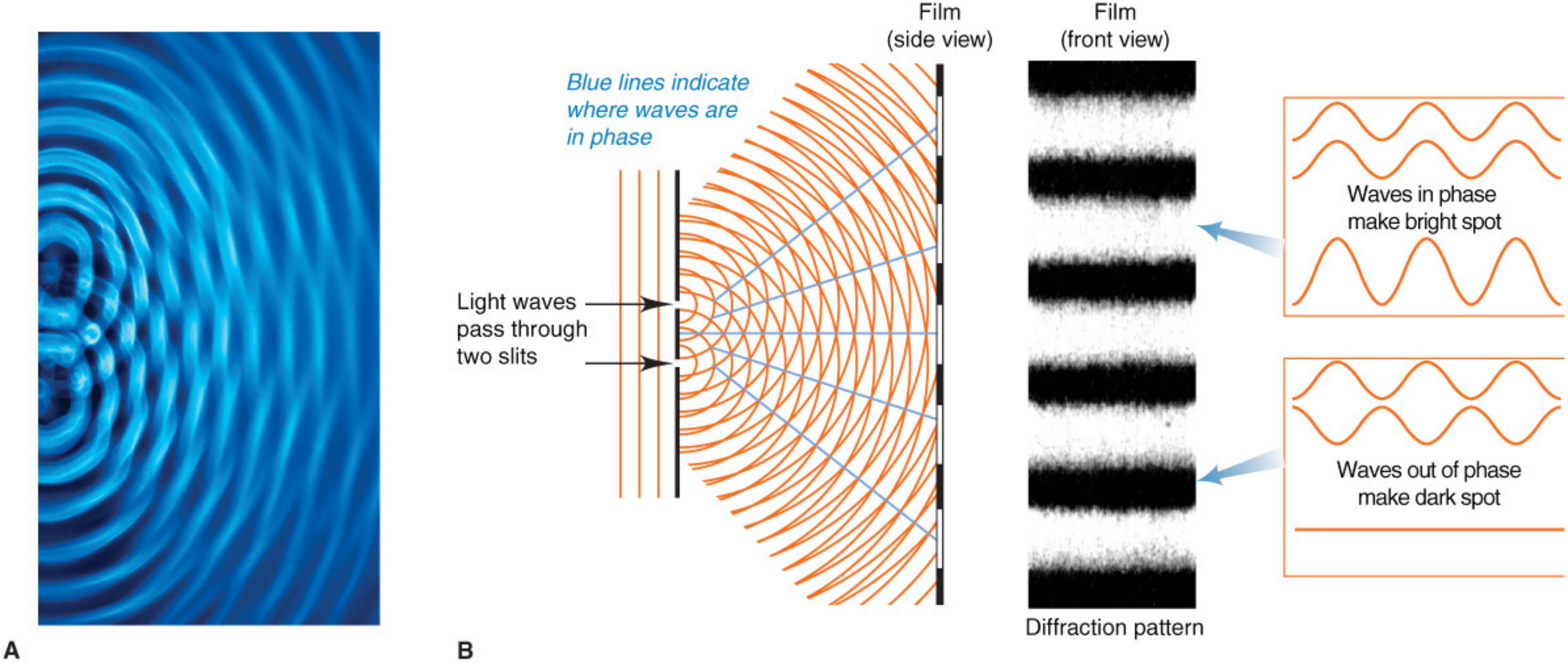


Figure 6.5 Formation of a diffraction pattern.



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The Wave-Particle Duality of Matter and Energy

Matter and Energy are alternate forms of the same entity.

$$\boxed{E = mc^2}$$

All matter exhibits properties of ***both particles and waves***. Electrons have wave-like motion and therefore have only certain allowable frequencies and energies.

Matter behaves as though it moves in a wave, and the ***de Broglie wavelength*** for any particle is given by:

$$\lambda = \frac{h}{mu}$$

m = mass

u = speed in m/s

Table 6.1 The de Broglie Wavelengths of Several Objects

Substance	Mass (kg)	Speed (m/s)	λ (m)
slow electron	$9x10^{-31}$	1.0	$7x10^{-4}$
fast electron	$9x10^{-31}$	$5.9x10^{6}$	$1x10^{-10}$
alpha particle	$6.6x10^{-27}$	$1.5x10^{7}$	$7x10^{-15}$
one-gram mass	$1.09x10^{-3}$	0.01	$7x10^{-29}$
baseball	0.142	25.0	$2x10^{-34}$
Earth	$6.0x10^{24}$	$3.0x10^{4}$	$4x10^{-63}$

Sample Problem 6.5 Calculating the de Broglie Wavelength of an Electron

PROBLEM: Find the deBroglie wavelength of an electron with a speed of 1.00×10^{6} m/s (electron mass = 9.11×10^{-31} kg; $h = 6.626 \times 10^{-34}$ kg·m²/s).

PLAN: We know the speed and mass of the electron, so we substitute these into Equation 7.5 to find λ.

SOLUTION:

$$\lambda = \frac{h}{mu}$$

$$\lambda = \frac{6.626 \times 10^{-34}\ \text{kg}\cdot\text{m}^2/\text{s}}{9.109\ \times 10^{-31}\ \text{kg} \times 1.00 \times 10^{6}\ \text{m/s}} \quad \boxed{= \mathbf{7.27 \times 10^{-10}\ m}}$$

Heisenberg's Uncertainty Principle

Heisenberg's Uncertainty Principle states that it is not possible to know both the position *and* momentum of a moving particle at the same time.

$$\Delta x \cdot \Delta p \geq \frac{h}{4\pi}$$

$$\Delta x \cdot m \Delta u \geq \frac{h}{4\pi}$$

x = position

u = speed

p = momentum

The more accurately we know the speed, the less accurately we know the position, and vice versa.

Sample Problem 6.6

Applying the Uncertainty Principle

PROBLEM: An electron moving near an atomic nucleus has a speed $6\text{x}10^6$ m/s ± 1%. What is the uncertainty in its position (Δx)?

PLAN: The uncertainty in the speed (Δu) is given as ±1% (0.01) of $6\text{x}10^6$ m/s. We multiply u by 0.01 and substitute this value into Equation 7.6 to solve for Δx.

SOLUTION: $\Delta u = (0.01)(6\text{x}10^6\ \text{m/s}) = 6\text{x}10^4\ \text{m/s}$

$$\Delta x \cdot m\Delta u \geq \frac{h}{4\pi}$$

$$\Delta x \geq \frac{h}{4\pi m\Delta u} \geq \frac{6.626\text{x}10^{-34}\ \text{kg}\cdot\text{m}^2/\text{s}}{4\pi\ (9.11\text{x}10^{-31}\ \text{kg})(6\text{x}10^4\ \text{m/s})} \quad \boxed{\geq 1\text{x}10^{-9}\ \text{m}}$$

Orbital Shape

$$\underbrace{E = nh\nu \qquad E_{tot} = E_{kin} + E_{pot} + E_{magn} \qquad \Delta(mv)\Delta(x) = h \qquad \lambda = h/mv}$$

$$E\psi = H\psi \longrightarrow E =- 2.178 \times 10^{-28}(1/n^2)$$

$$\downarrow$$

$$\Psi = \overset{\text{Position}}{R} \times \overset{\text{momentum}}{\Theta} \times \overset{\text{magnetism}}{\Phi}$$

$n = 1, l = 0, m = 0$ $\qquad 2\exp^{-r/a}/a^{3/2} \qquad$ const $\qquad$ const

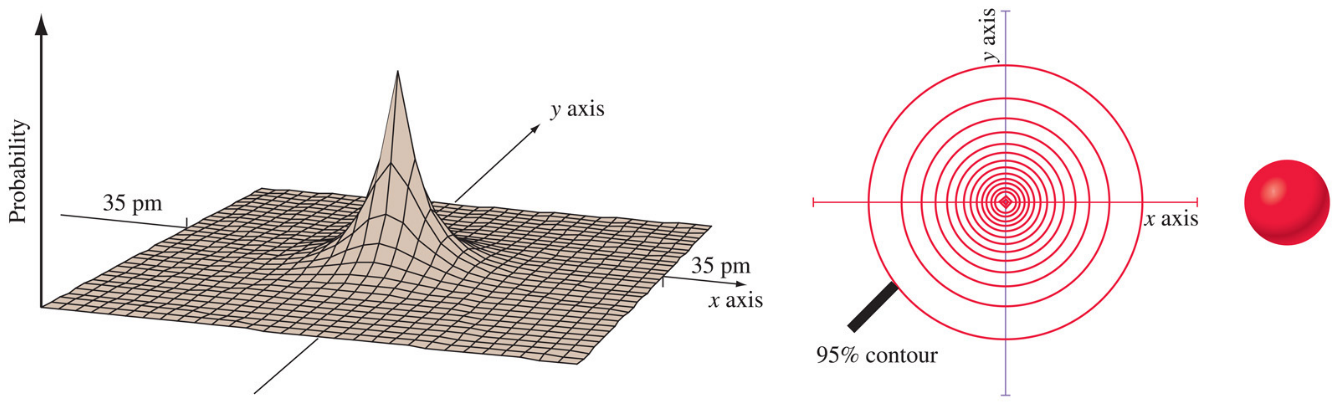
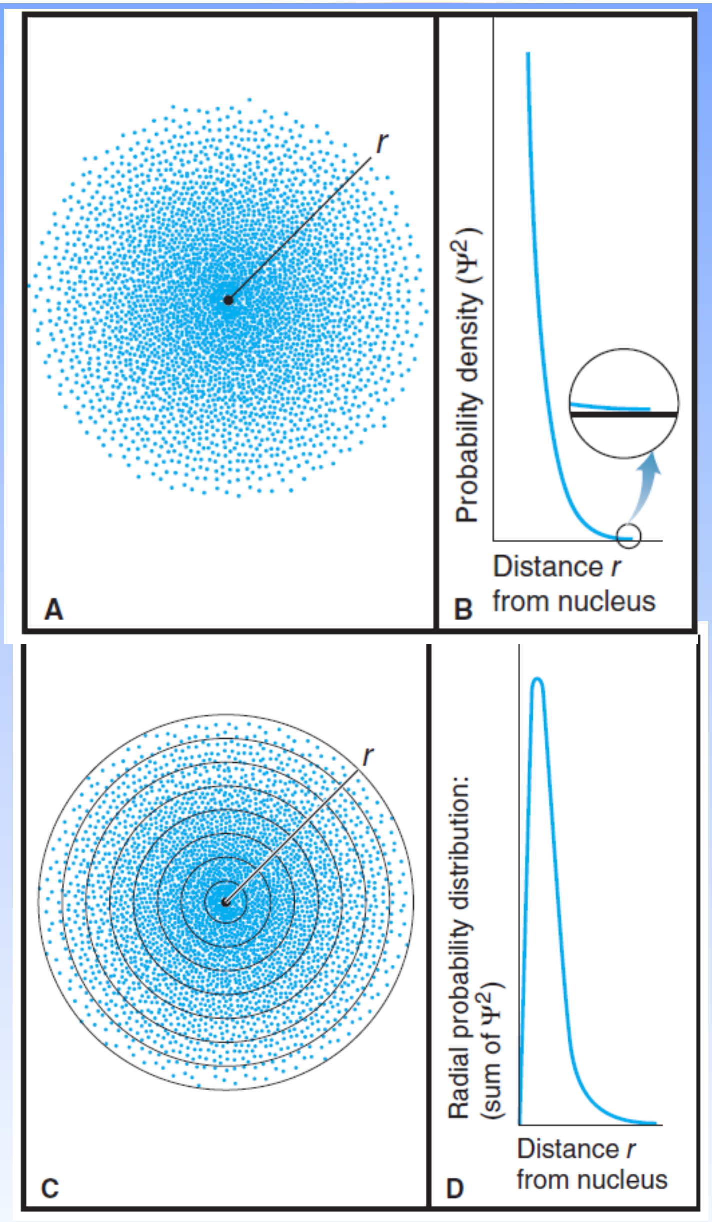
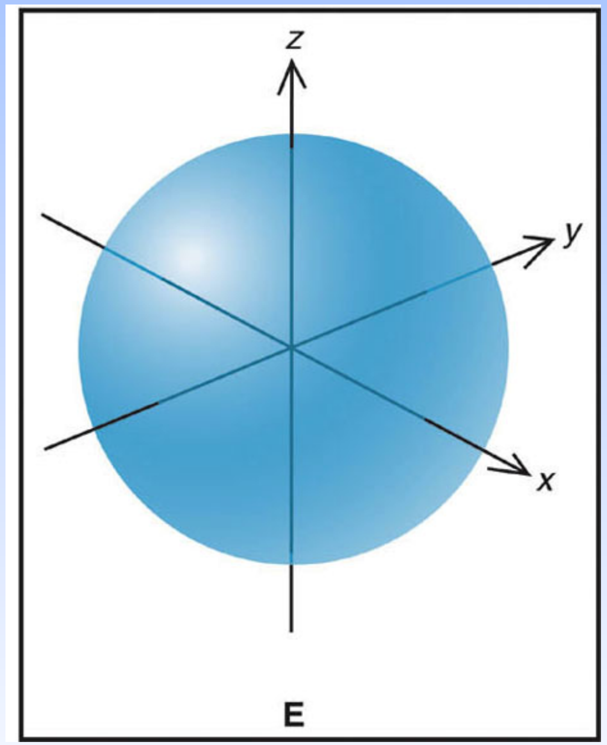


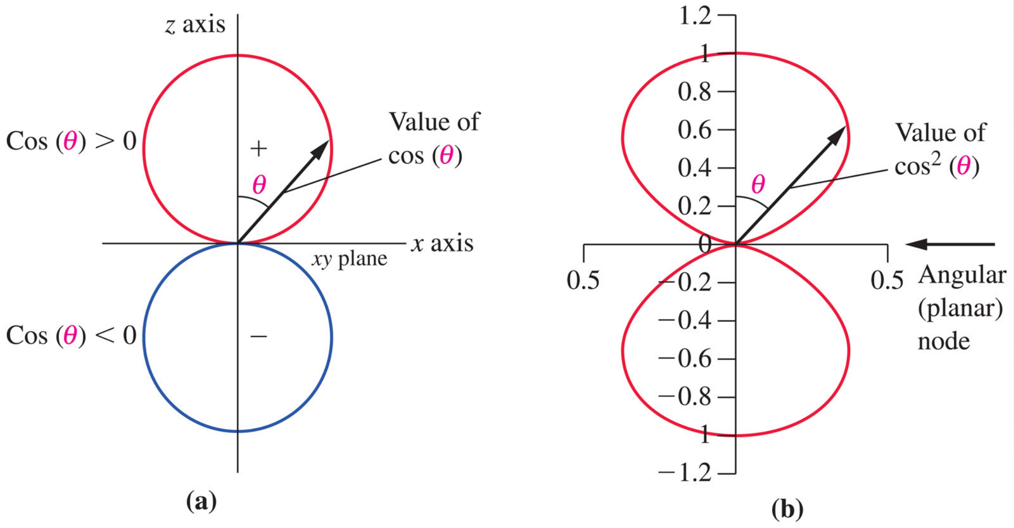
Figure 6.17

Electron probability density in the ground-state H atom.





Ψ =	R	x Θ	x Φ
n = 2, l = 0, m = 0	= 2exp-r/a/a3/2	const	const
l = 1, m = 0,1,-1	= const$(1/a)^{3/2}$ (r/a)exp$^{-r/2a}$	cos θ	const (x,y,z)



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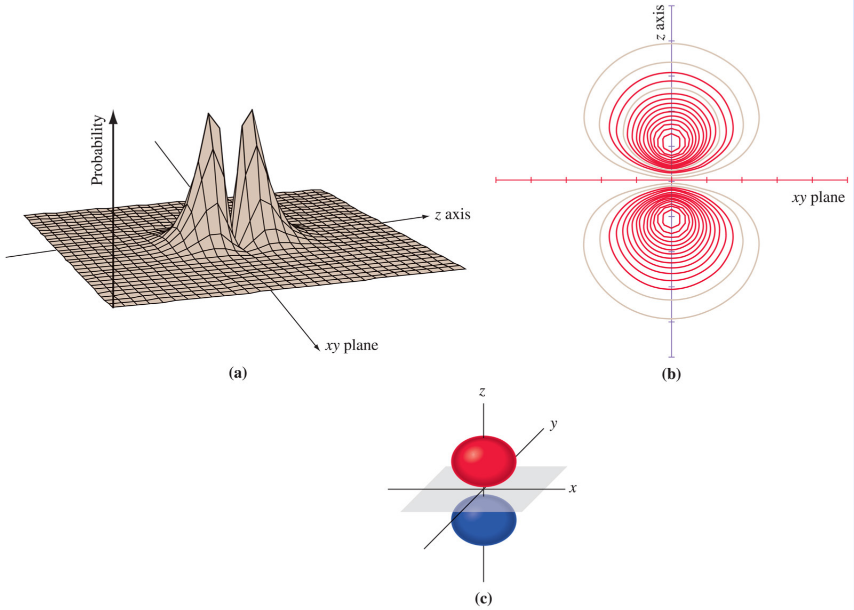

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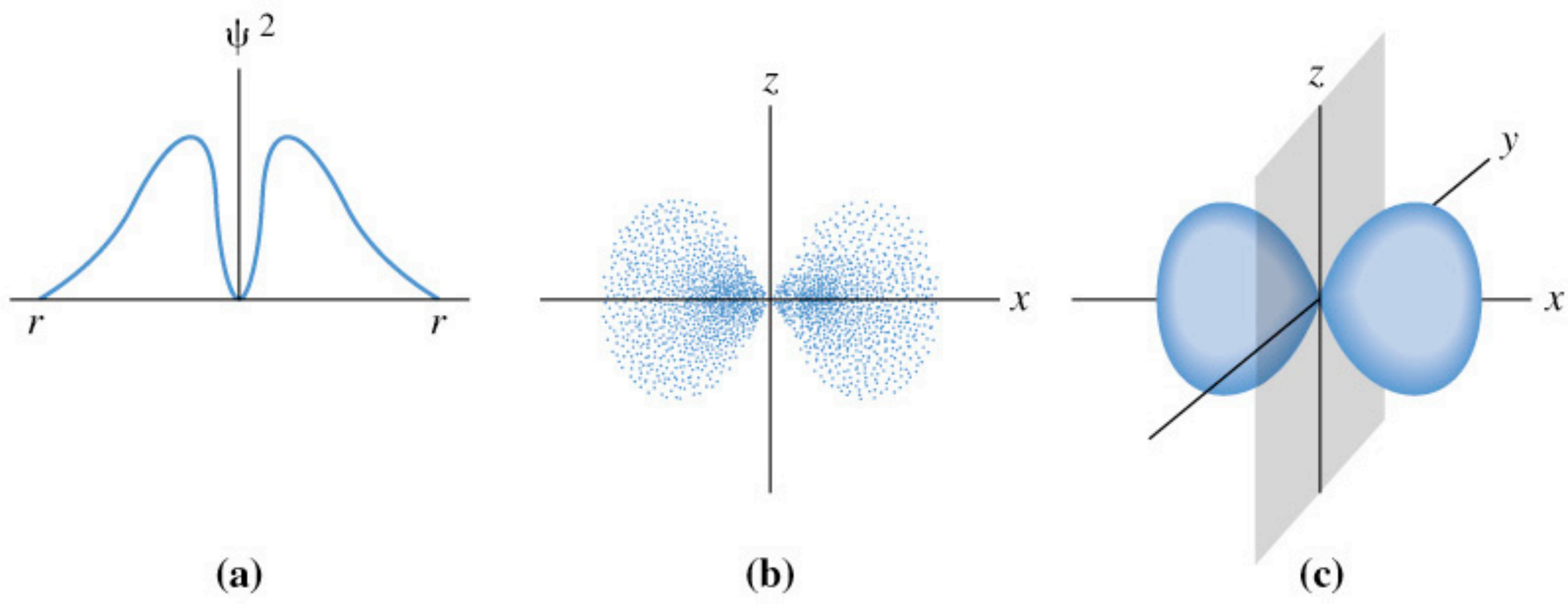
p Orbitals



p Orbitals
Three orientations

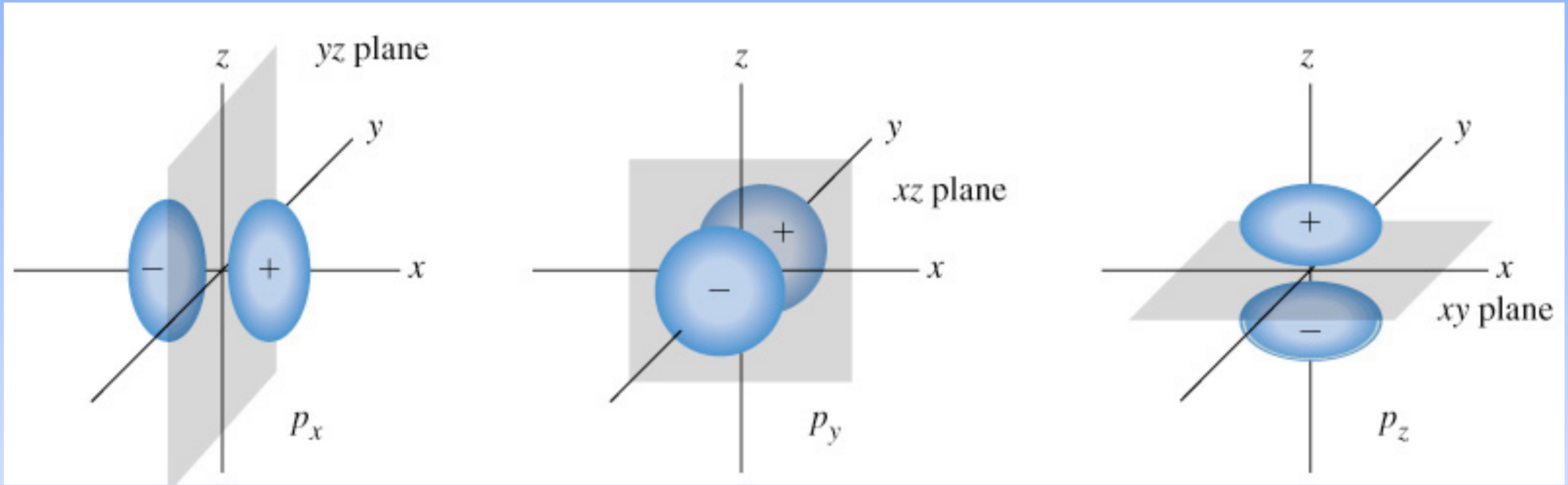
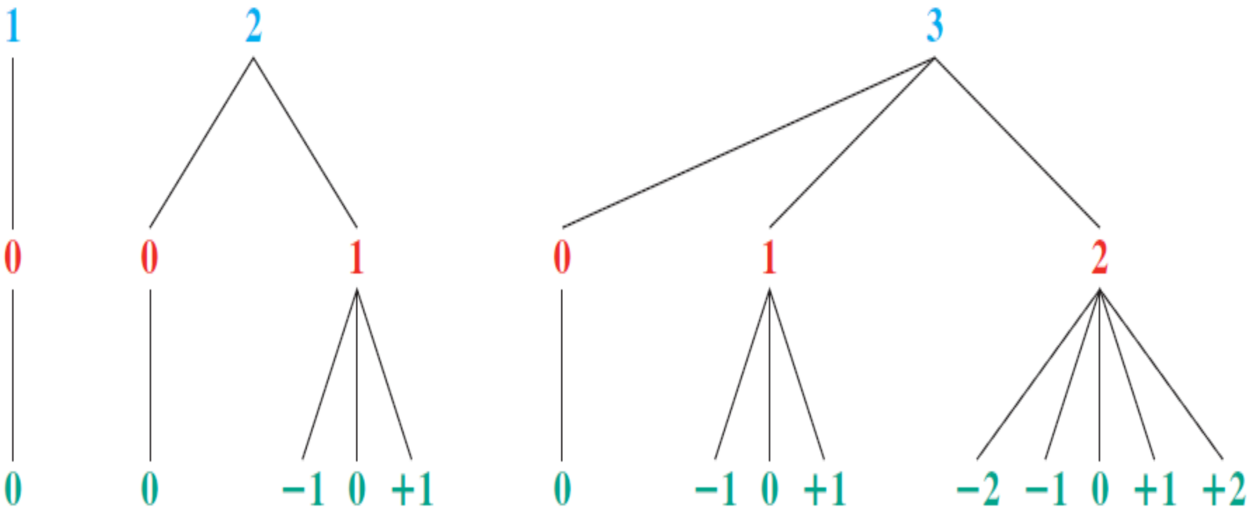


Table 6.2 The Hierarchy of Quantum Numbers for Atomic Orbitals

Name, Symbol (Property)	Allowed Values	Quantum Numbers
Principal, n (size, energy)	Positive integer (1, 2, 3, . . .)	1 2 3
Angular momentum, l (shape)	0 to $n - 1$	0 0 1 0 1 2
Magnetic, m_l (orientation)	$-l$, . . ., 0, . . ., $+l$	0 0 −1 0 +1 0 −1 0 +1 −2 −1 0 +1 +2



Sample Problem 6.8

Determining Sublevel Names and Orbital Quantum Numbers

PROBLEM: Give the name, magnetic quantum numbers, and number of orbitals for each sublevel with the following quantum numbers:

(a) $n = 3, l = 2$ **(b)** $n = 2, l = 0$ **(c)** $n = 5, l = 1$ **(d)** $n = 4, l = 3$

PLAN: Combine the n value and l designation to name the sublevel. Knowing l, we can find m_l and the number of orbitals.

SOLUTION:

	n	l	sublevel name	possible m_l values	# of orbitals
(a)	3	2	3*d*	-2, -1, 0, 1, 2	5
(b)	2	0	2*s*	0	1
(c)	5	1	5*p*	-1, 0, 1	3
(d)	4	3	4*f*	-3, -2, -1, 0, 1, 2, 3	7

Sample Problem 6.9 Identifying Incorrect Quantum Numbers

PROBLEM: What is wrong with each of the following quantum numbers designations and/or sublevel names?

	n	l	m_l	Name
(a)	1	1	0	1*p*
(b)	4	3	+1	4*d*
(c)	3	1	-2	3*p*

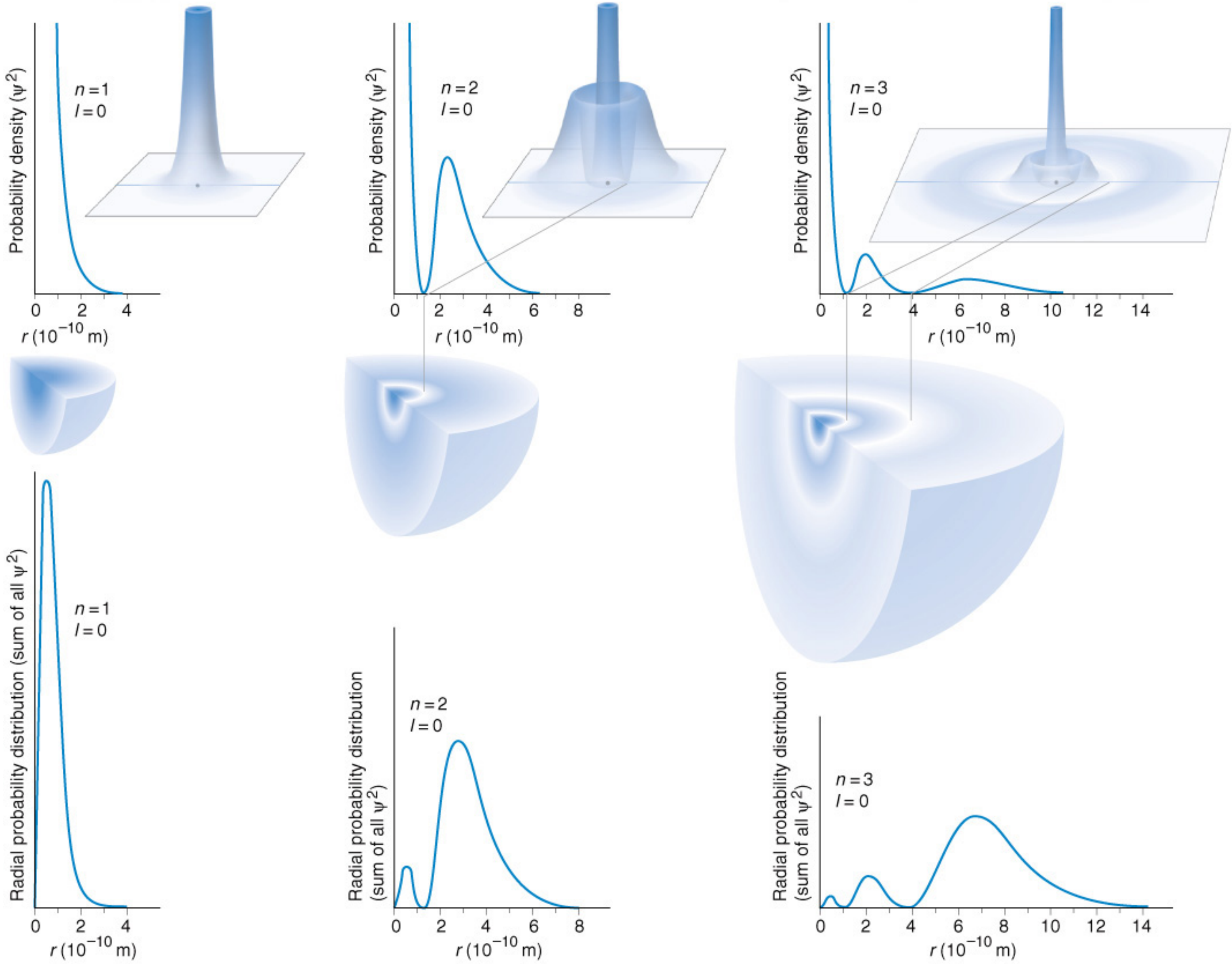
SOLUTION:

(a) A sublevel with $n = 1$ can only have $l = 0$, not $l = 1$. The only possible sublevel name is 1*s*.

(b) A sublevel with $l = 3$ is an *f* sublevel, to a *d* sublevel. The name should be 4*f*.

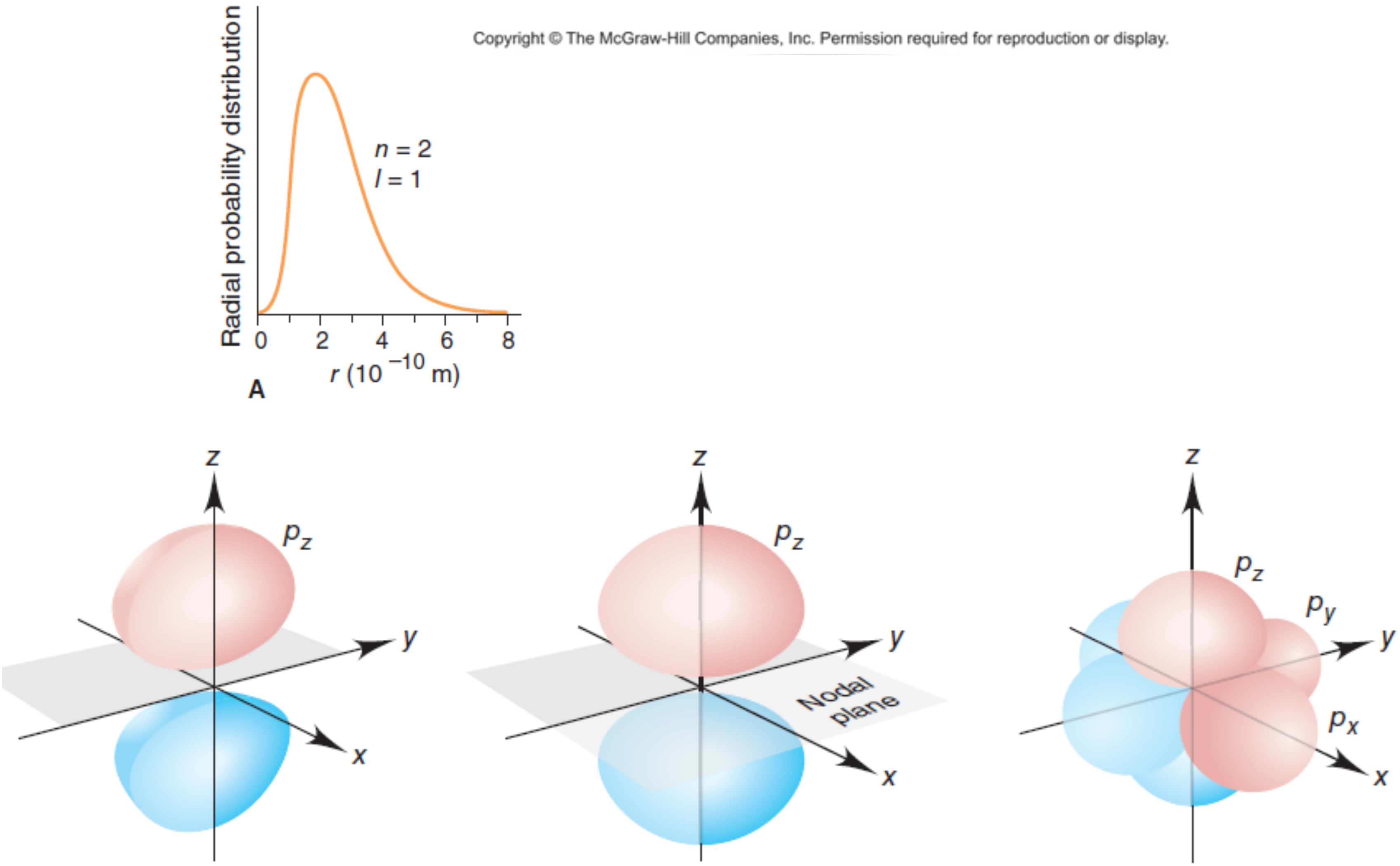
(c) A sublevel with $l = 1$ can only have m_l values of -1, 0, or +1, not -2.

Figure 6.18 Representations of the 1*s*, 2*s*, and 3*s* orbitals.



A 1*s* orbital **B** 2*s* orbital **C** 3*s* orbital

Figure 6.19 The 2*p* orbitals.



B Cross section of electron cloud depiction

C Accurate probability contour

D The three *p* orbitals

Figure 6.20 The 3*d* orbitals.

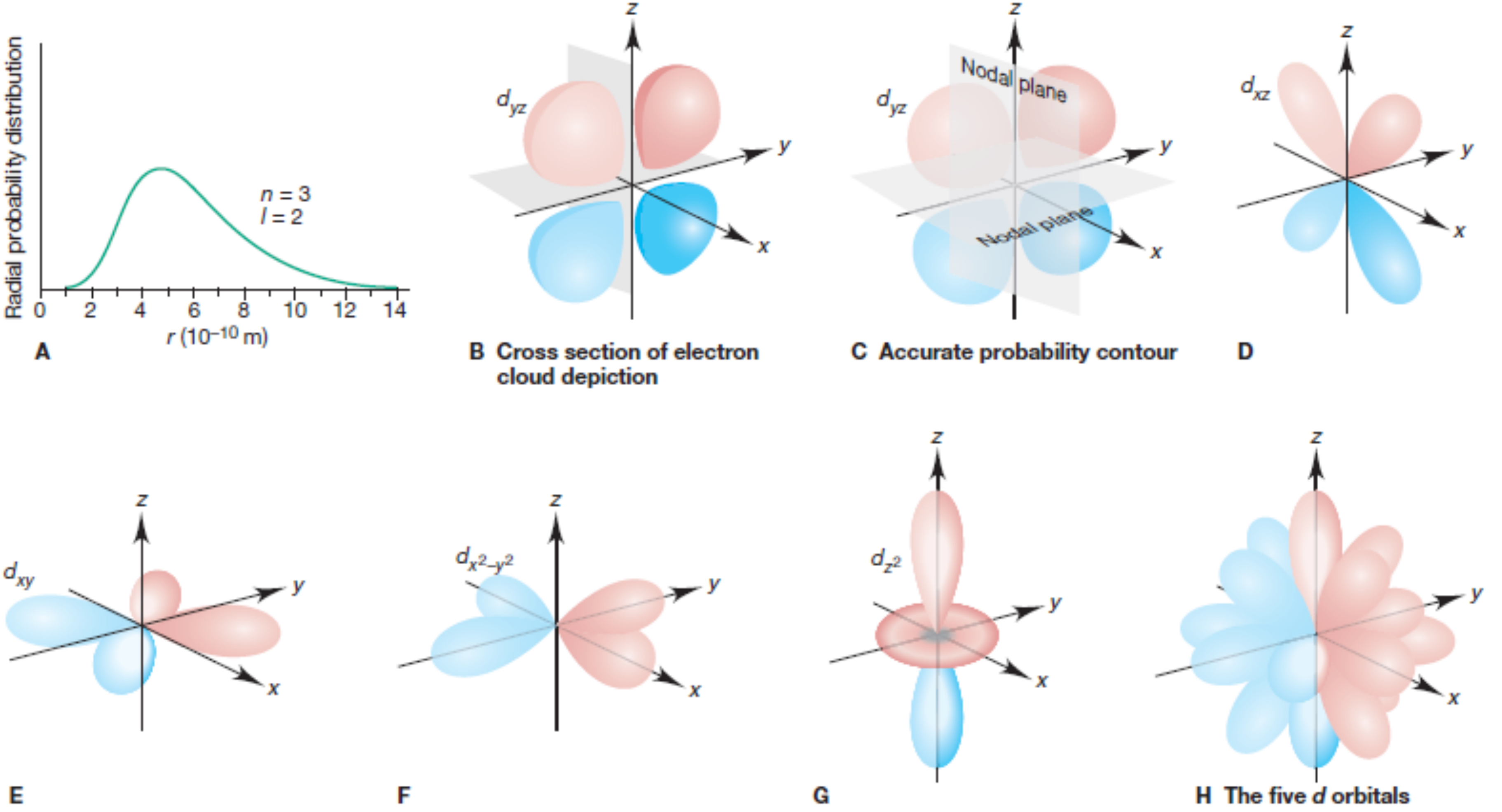
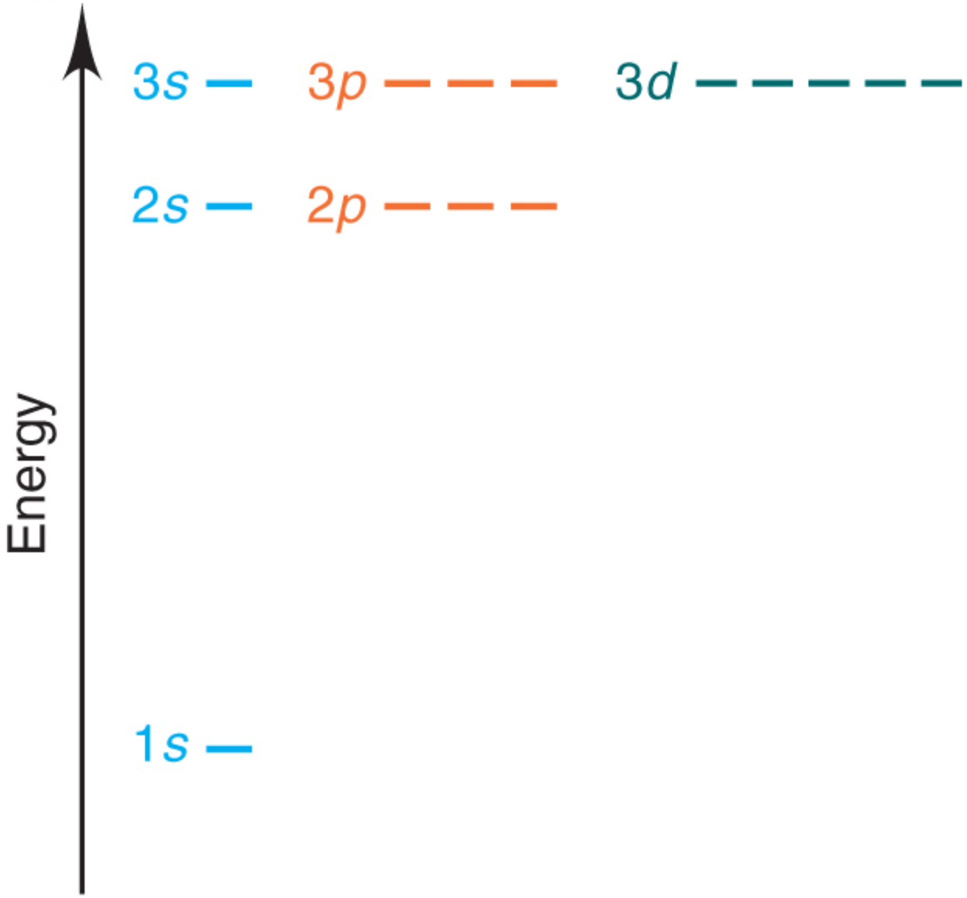


Figure 6.22 Energy levels of the H atom.



Chapter 7

Electron Configuration and Chemical Periodicity

7.1

Figure 7.1 The effect of electron spin.

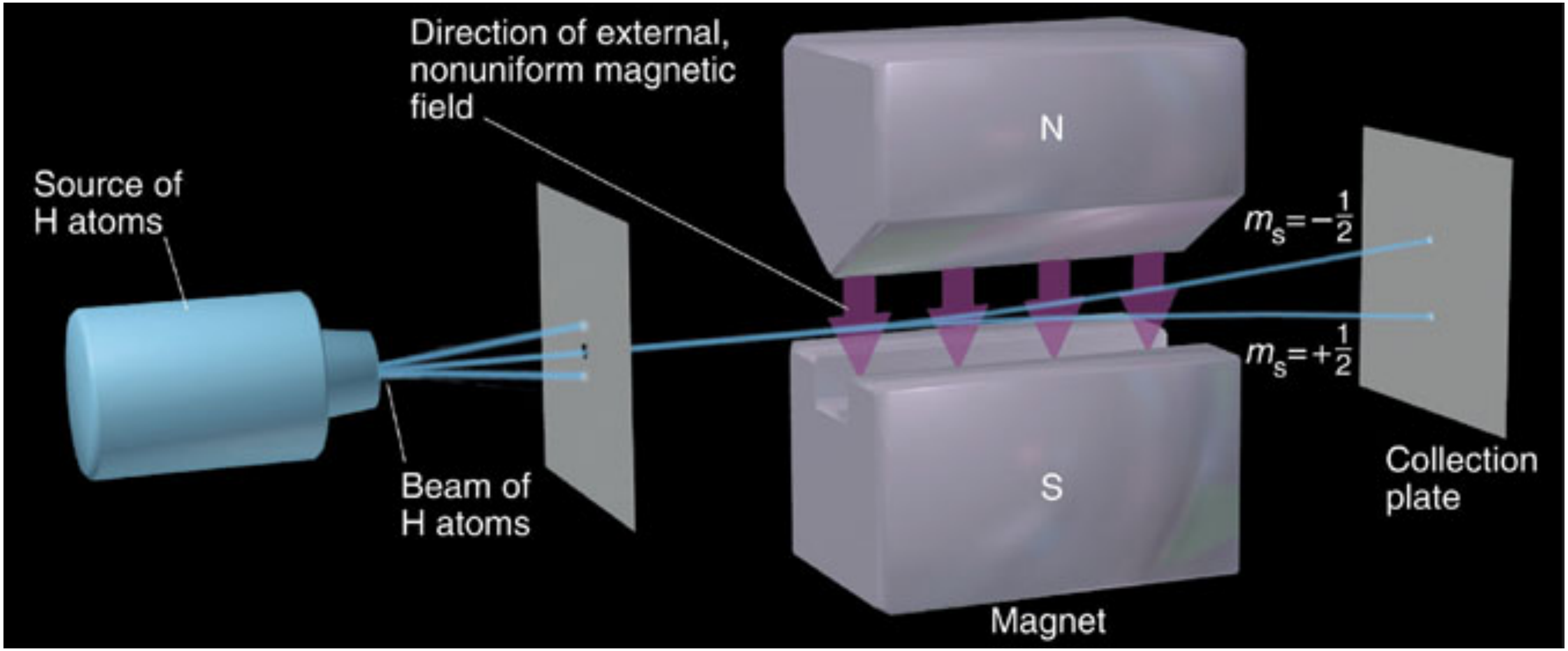


Table 7.1 Summary of Quantum Numbers of Electrons in Atoms

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

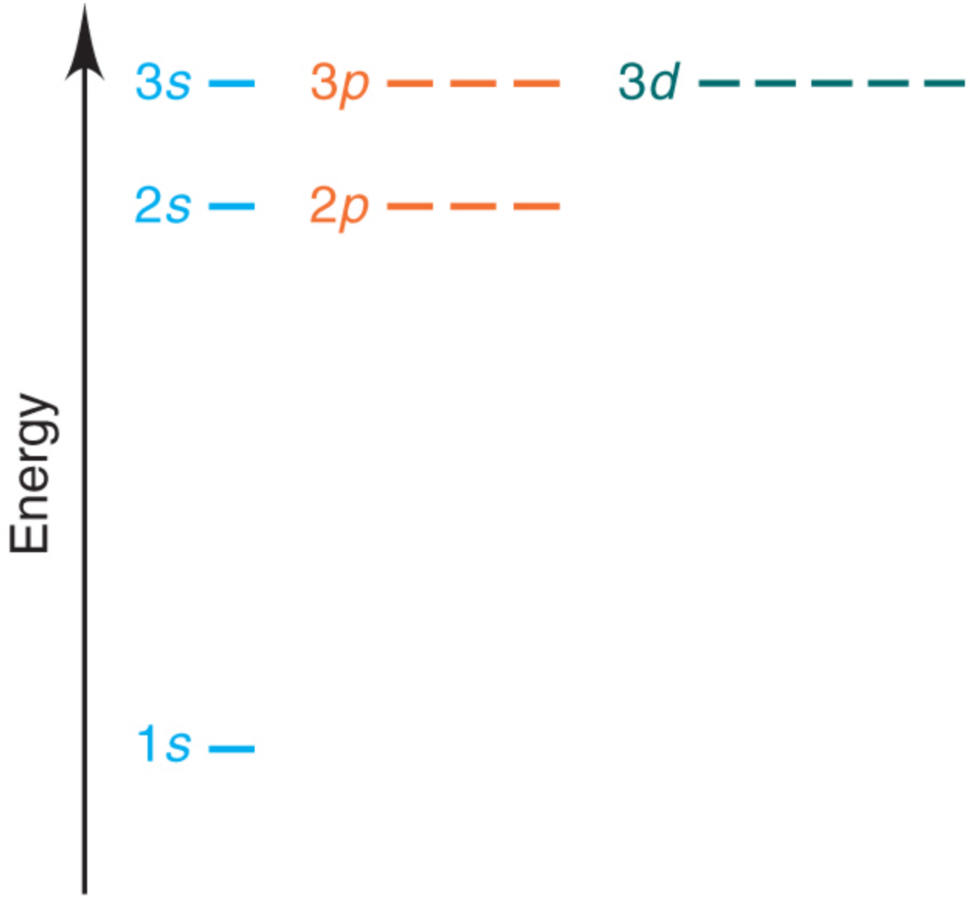
Each electron in any atom is described completely by a set of ***four*** quantum numbers.

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***.

Figure 6.22 Energy levels of the H atom.



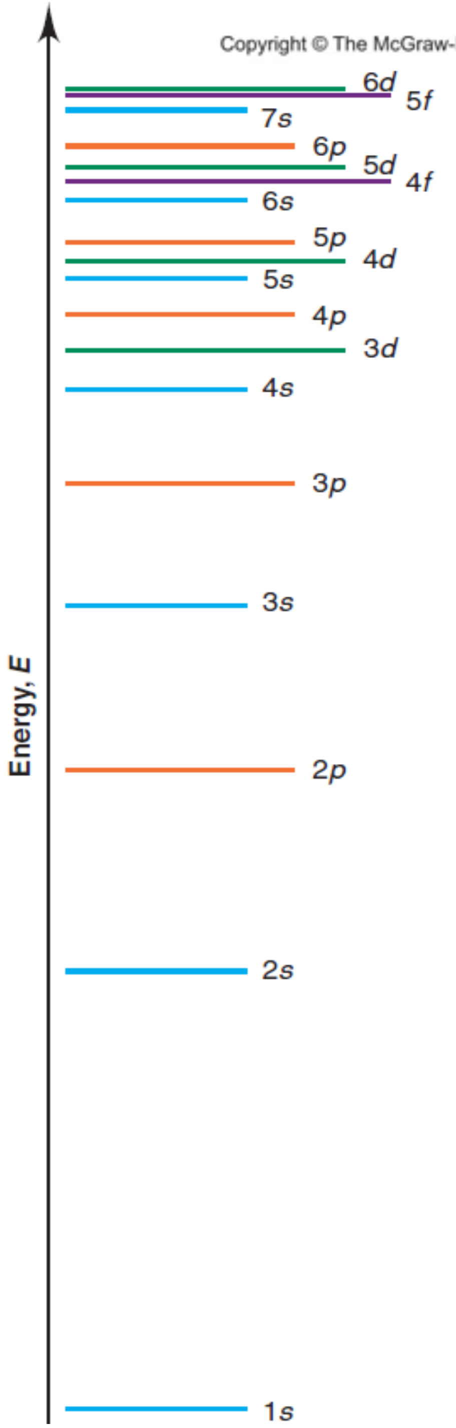


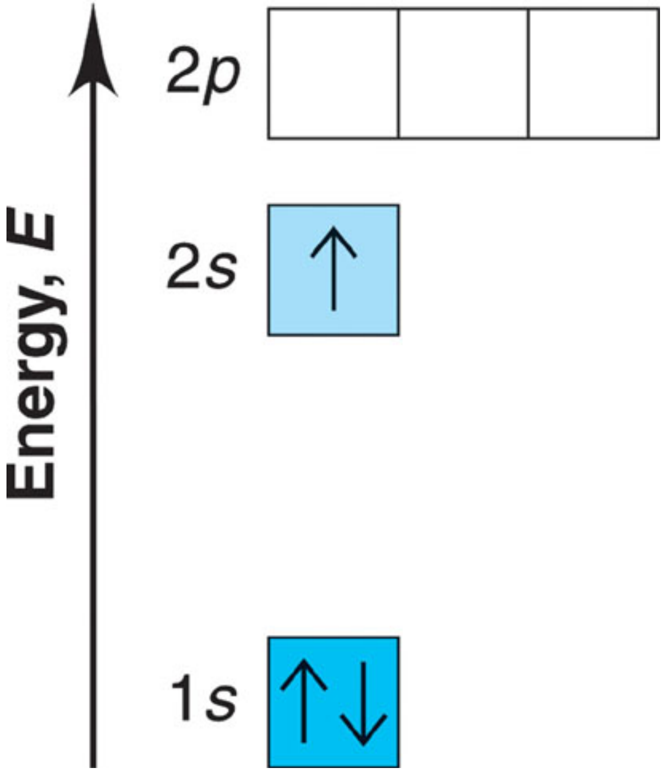
Figure 7.5

The order for filling shells with electrons.

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

As *n* increases, some sublevels overlap.

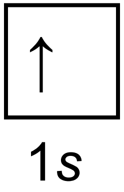
Figure 7.6 A vertical orbital diagram for the Li ground state.



Building Orbital Diagrams

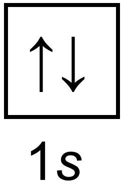
The **Aufbau principle** is applied – electrons are always placed in the lowest energy sublevel available.

H (Z = 1) 1s^1



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

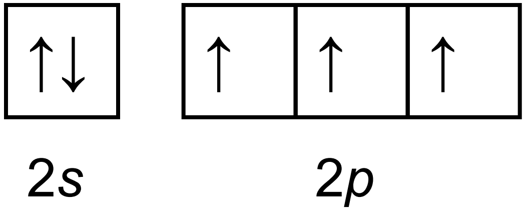
He (Z = 2) 1s^2



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.

$\mathrm{N}\ (Z = 7)\ 1s^2 2s^2 2\mathrm{p}^3$



Sample Problem 7.1

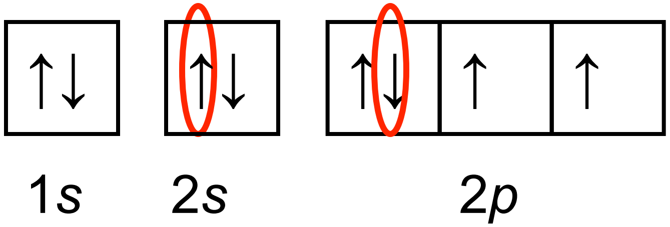
Determining Quantum Numbers from Orbital Diagrams

PROBLEM: Use the orbital diagrams,write a set of quantum numbers for(a) the third electron and (b)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) $1s^22s^22p^5$



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 7.7 Depicting orbital occupancy for the first 10 elements.

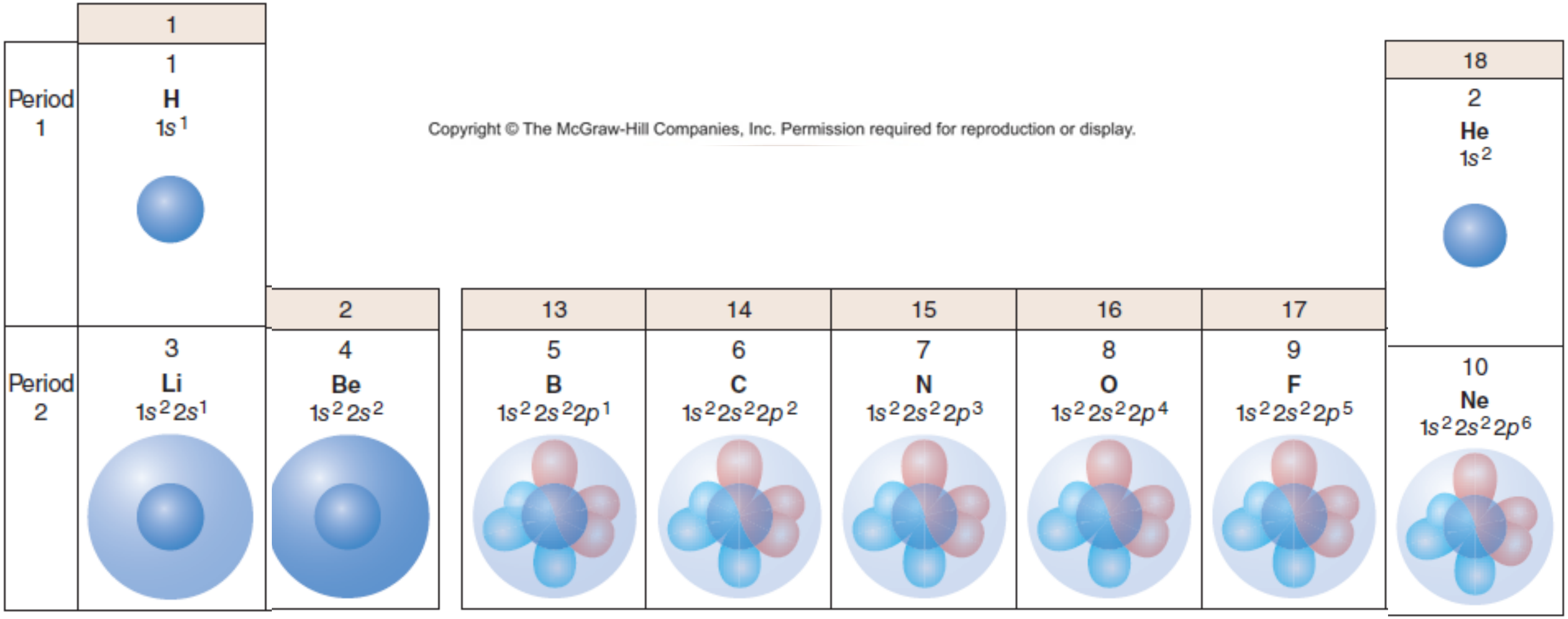


Table 7.2 Partial Orbital Diagrams and Electron Configurations* for the Elements in Period 3.

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	↑				$[1s^22s^22p^6]\ 3s^1$	[Ne] $3s^1$
12	Mg	↑↓				$[1s^22s^22p^6]\ 3s^2$	[Ne] $3s^2$
13	Al	↑↓	↑			$[1s^22s^22p^6]\ 3s^23p^1$	[Ne] $3s^23p^1$
14	Si	↑↓	↑	↑		$[1s^22s^22p^6]\ 3s^23p^2$	[Ne] $3s^23p^2$
15	P	↑↓	↑	↑	↑	$[1s^22s^22p^6]\ 3s^23p^3$	[Ne] $3s^23p^3$
16	S	↑↓	↑↓	↑	↑	$[1s^22s^22p^6]\ 3s^23p^4$	[Ne] $3s^23p^4$
17	Cl	↑↓	↑↓	↑↓	↑	$[1s^22s^22p^6]\ 3s^23p^5$	[Ne] $3s^23p^5$
18	Ar	↑↓	↑↓	↑↓	↑↓	$[1s^22s^22p^6]\ 3s^23p^6$	[Ne] $3s^23p^6$

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

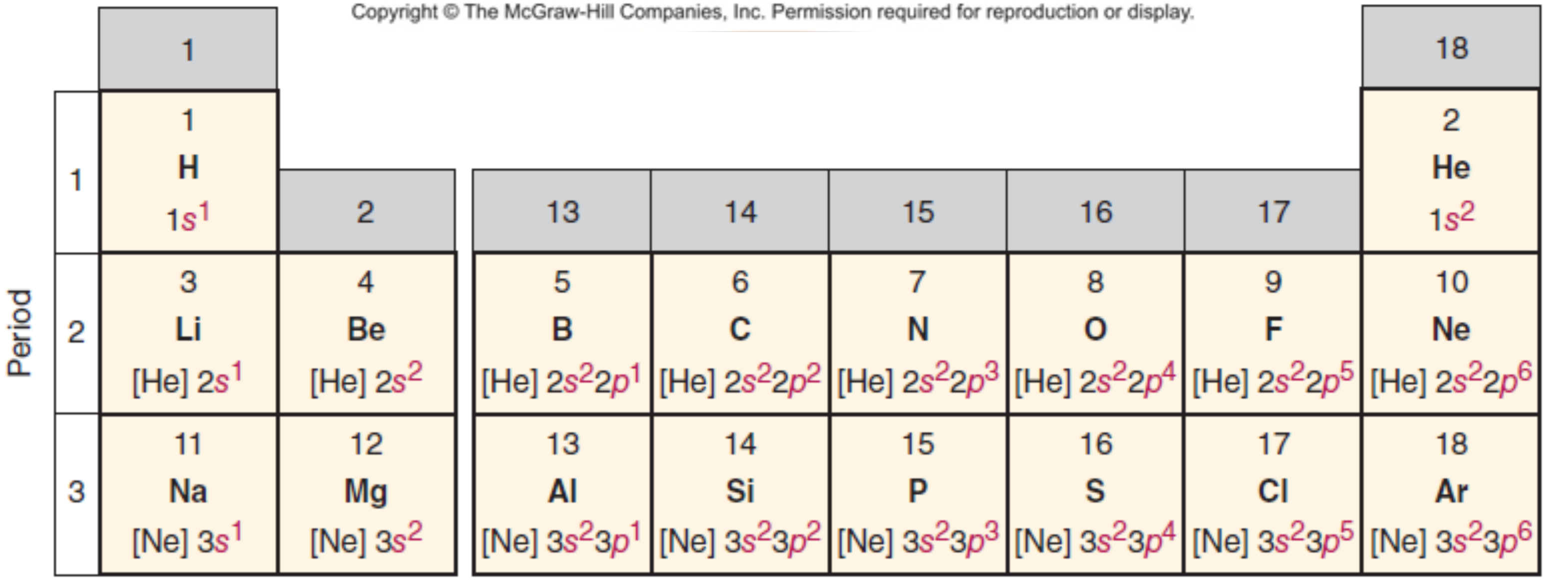
Electron Configuration and Group

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.

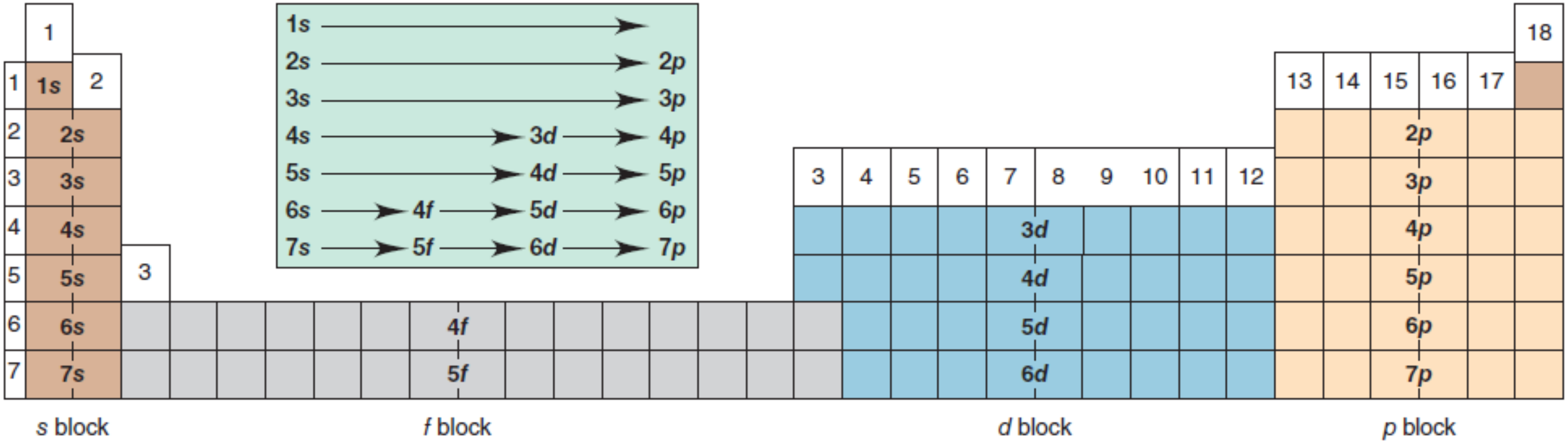
Similar outer electron configurations correlate with similar chemical behavior.

Figure 7.8 Condensed electron configurations in the first three periods.



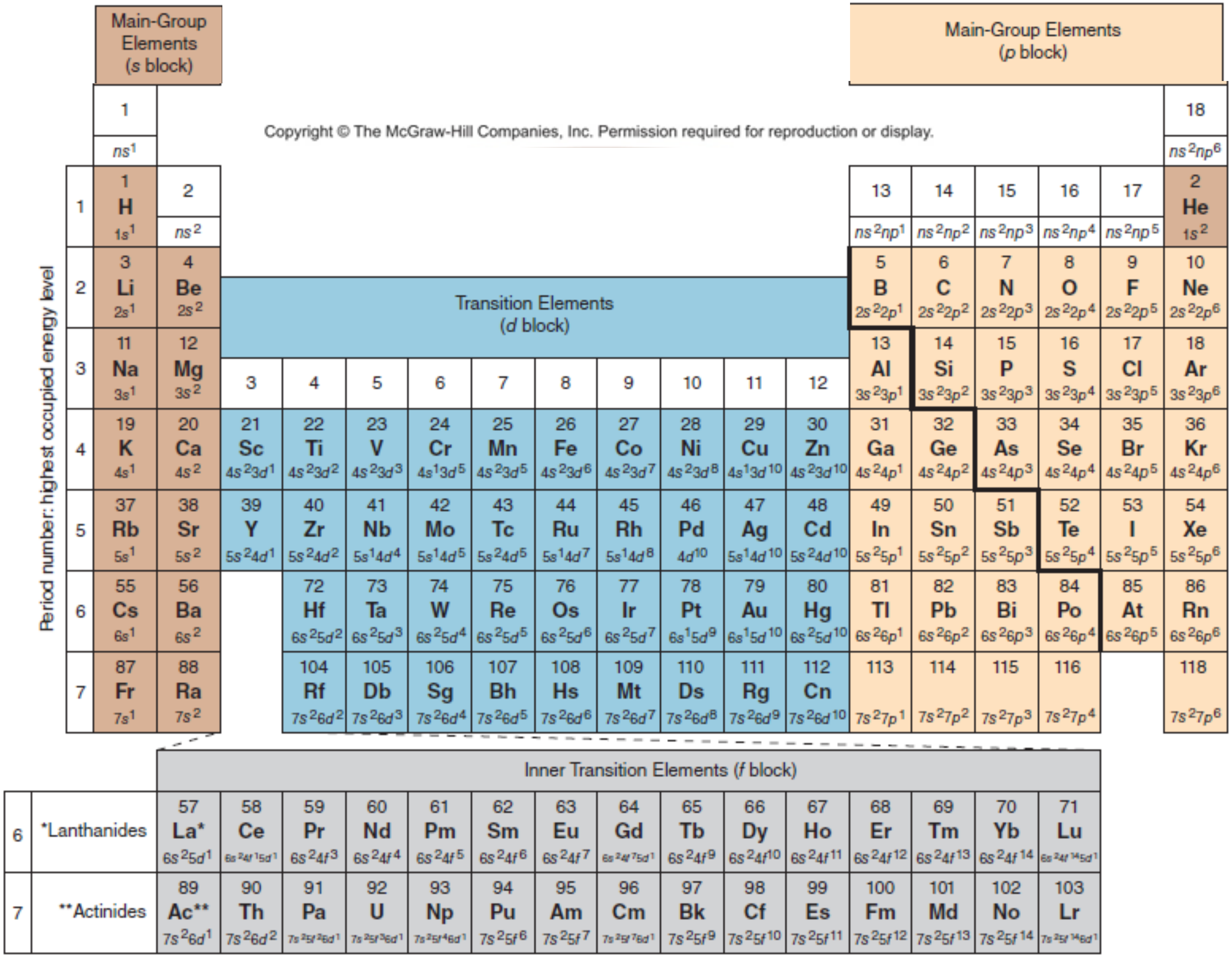
Period	1	2	13	14	15	16	17	18
1	1 **H** $1s^1$							2 **He** $1s^2$
2	3 **Li** [He] $2s^1$	4 **Be** [He] $2s^2$	5 **B** [He] $2s^2 2p^1$	6 **C** [He] $2s^2 2p^2$	7 **N** [He] $2s^2 2p^3$	8 **O** [He] $2s^2 2p^4$	9 **F** [He] $2s^2 2p^5$	10 **Ne** [He] $2s^2 2p^6$
3	11 **Na** [Ne] $3s^1$	12 **Mg** [Ne] $3s^2$	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 7.11 Orbital filling and the periodic table.



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

Figure 7.10 A periodic table of partial ground-state electron configurations.



Main-Group Elements (*s* block)

Main-Group Elements (*p* block)

Transition Elements (*d* block)

	1 ns^1	2 ns^2	3	4	5	6	7	8	9	10	11	12	13 ns^2np^1	14 ns^2np^2	15 ns^2np^3	16 ns^2np^4	17 ns^2np^5	18 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** $6s^2$		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** $7s^2$		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)

6	*Lanthanides	57 **La*** $6s^25d^1$	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	89 **Ac**** $7s^26d^1$	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** $7s^25f^{14}6d^1$

Sample Problem 7.2 Determining Electron Configurations

PROBLEM: Using the periodic table on the inside cover of the text (not Figure 7.10 or Table 7.3), 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 7.2

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 core electrons.

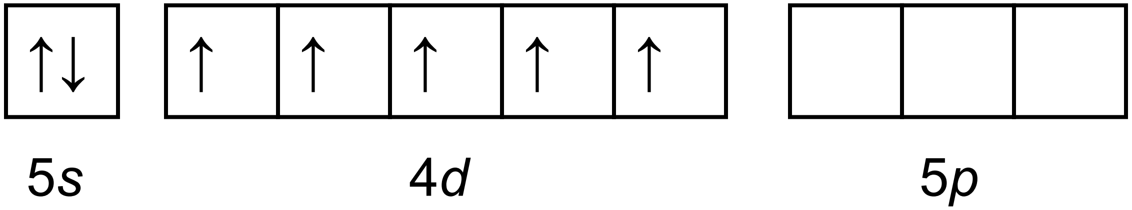
Sample Problem 7.2

SOLUTION:

(b) For Tc (Z = 43)

full configuration	$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^5$
condensed configuration	$[Kr]5s^24d^5$

partial orbital diagram



There are 36 core electrons.

Sample Problem 7.2

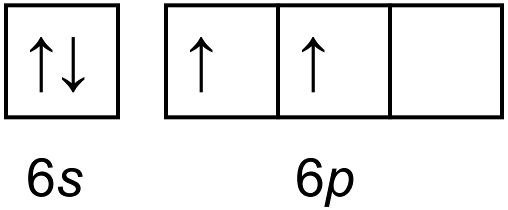
SOLUTION:

(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



There are 78 core electrons.

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 7.15 Periodicity of first ionization energy (IE_1).

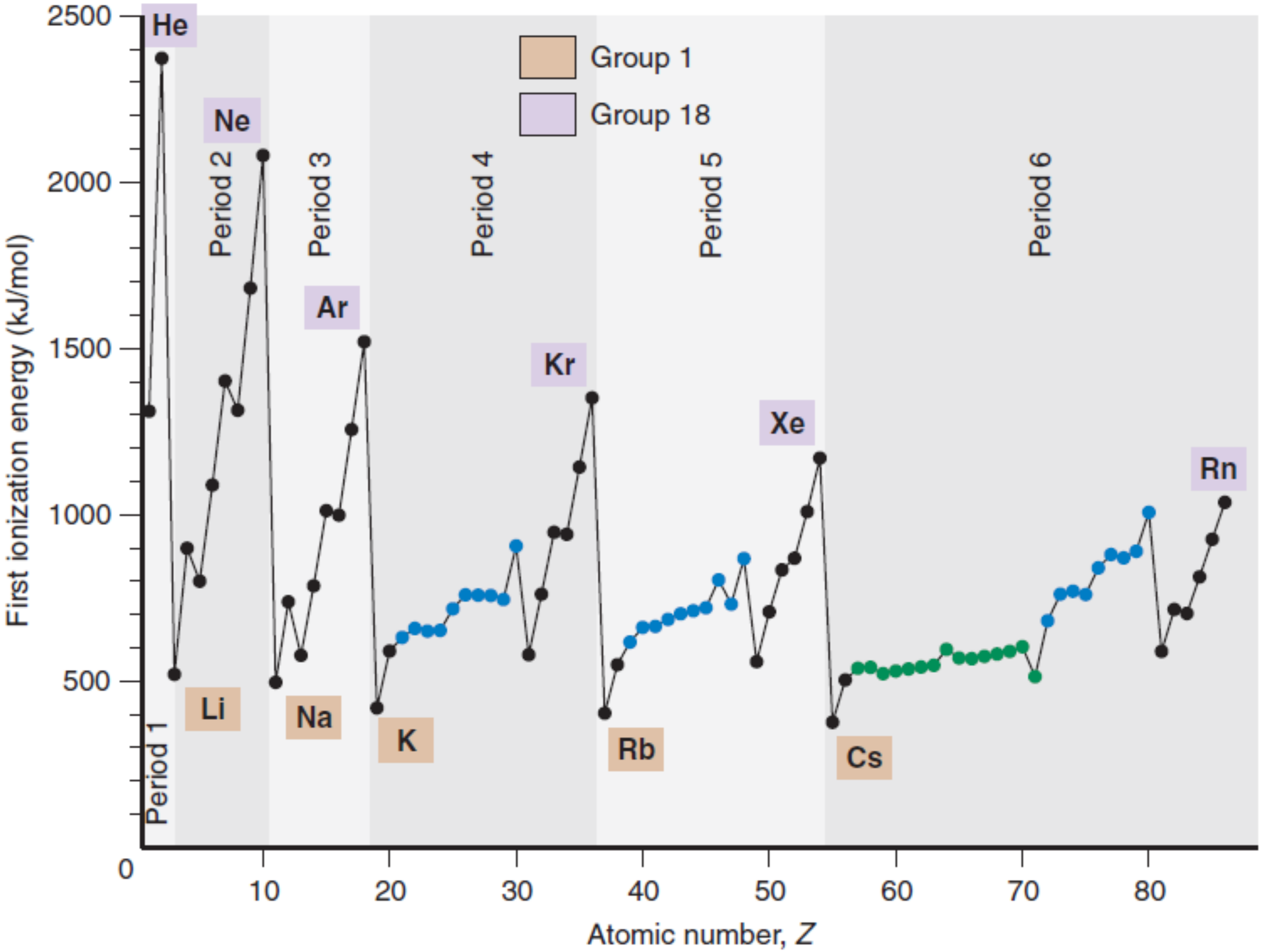
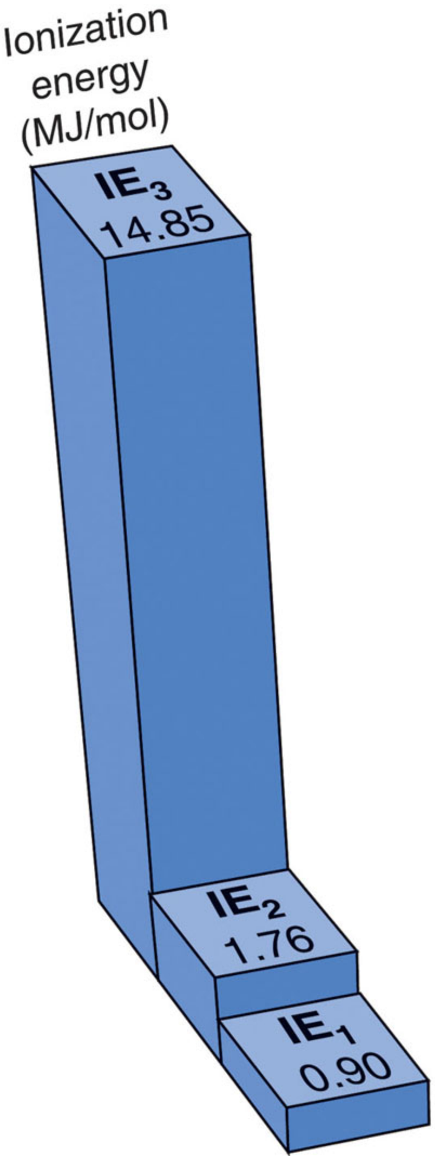


Figure 7.17 The first three ionization energies of beryllium.



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

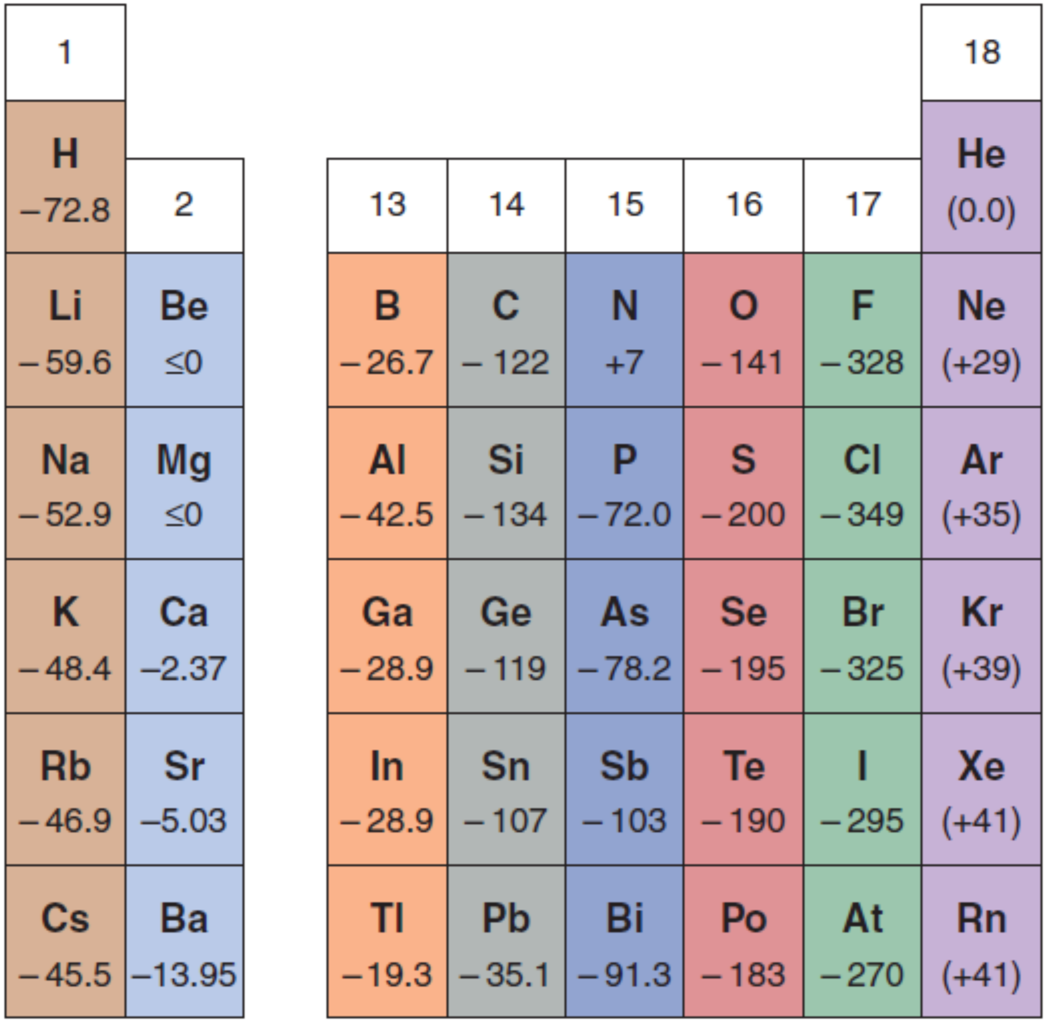
Trends in Electron Affinity

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***.

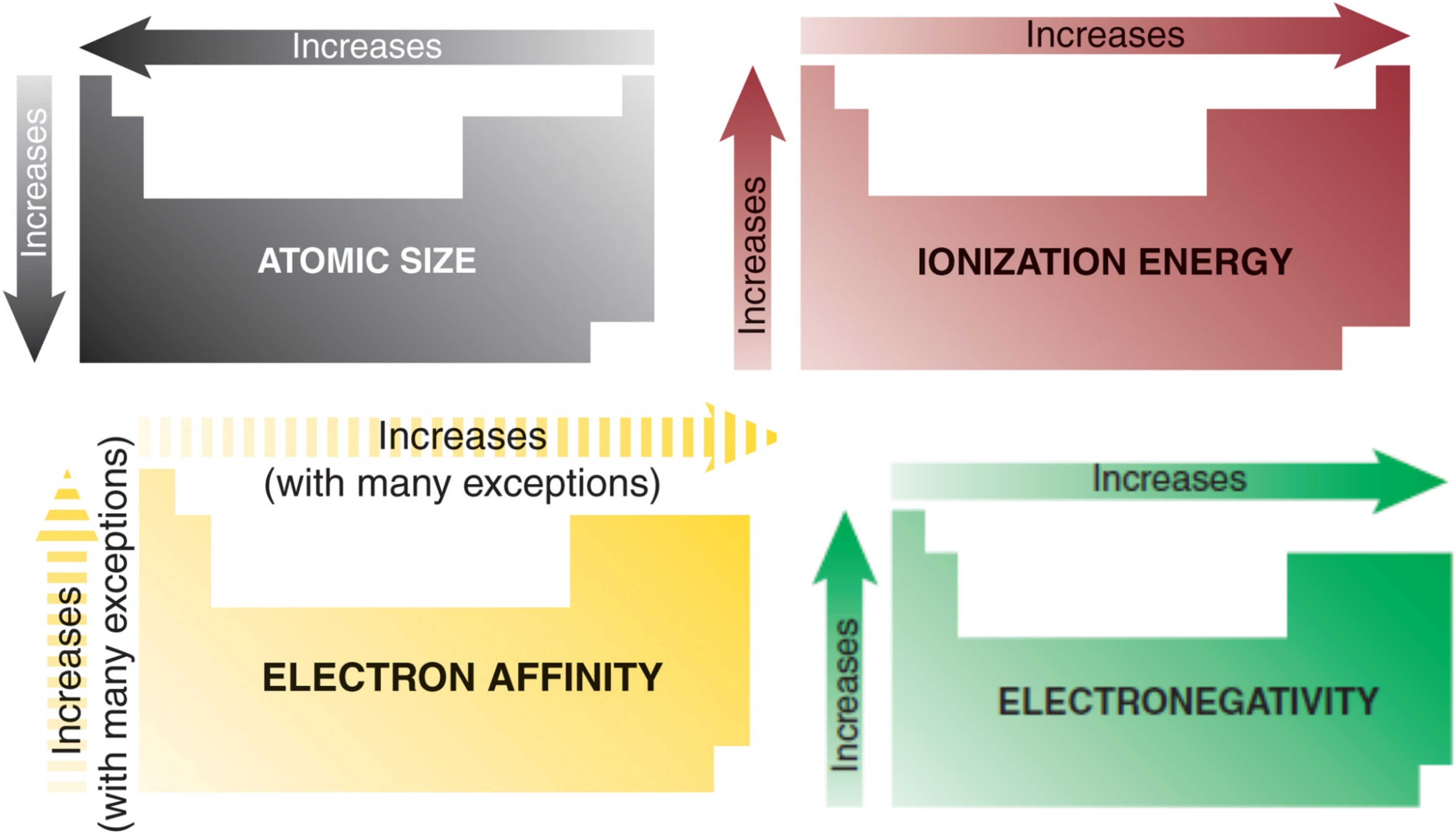
The trends in electron affinity are not as regular as those for atomic size or IE.

Figure 7.18 Electron affinities of the main-group elements (in kJ/mol).



1	2	13	14	15	16	17	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)

Figure 7.19 Trends in three atomic properties.



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.

$\mathrm{Na}(1s^22s^22p^6\mathbf{3s^1}) \rightarrow \mathbf{e^-} + \mathrm{Na}^+([\mathrm{He}]2s^22p^6)$

[isoelectronic with Ne]

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

[isoelectronic with Kr]

Figure 7.22 Main-group elements whose ions have noble gas electron configurations.