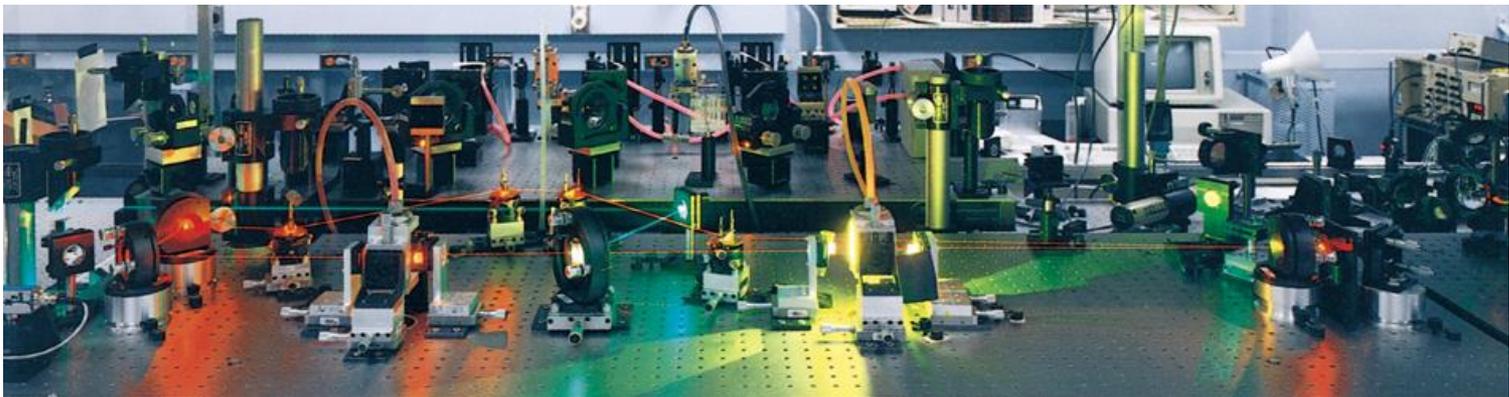


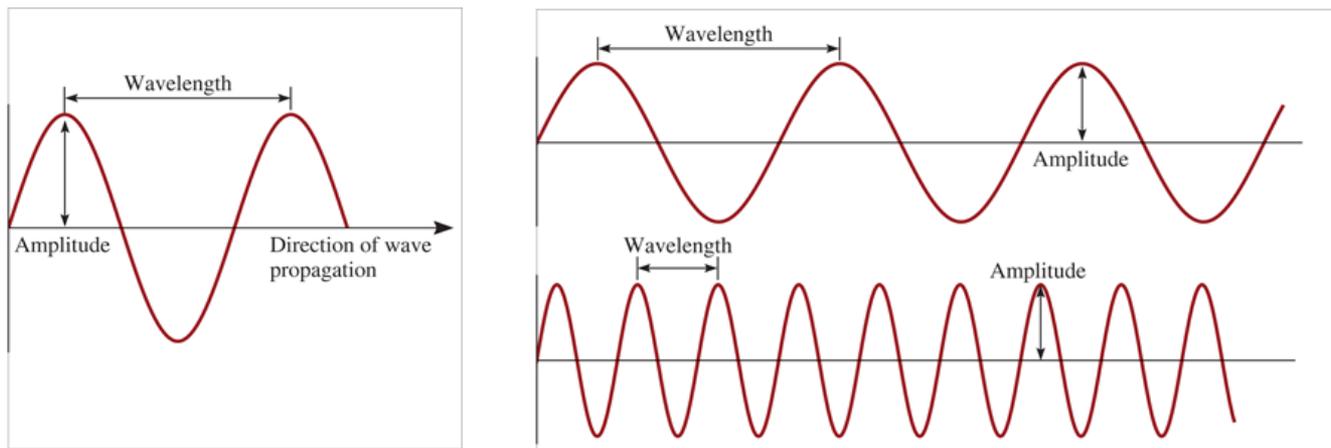
Chapter 7

Quantum Theory and the Electronic Structure of Atoms



Properties of Waves

Quantum theory enables us to predict and understand the critical role that electrons play in chemistry.



Wavelength (λ) is the distance between identical points on successive waves.

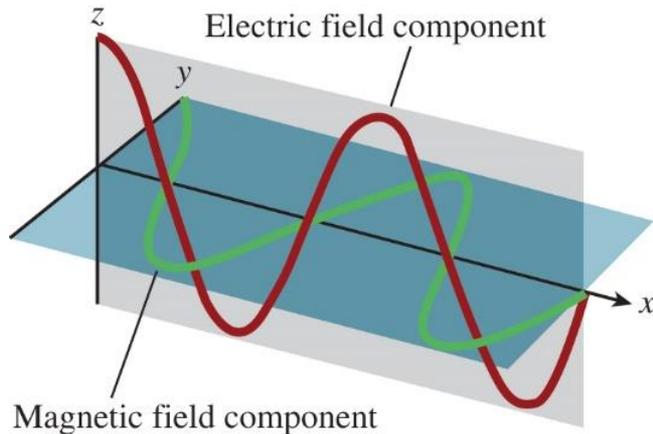
Amplitude is the vertical distance from the midline of a wave to the peak or trough.

Frequency (ν) is the number of waves that pass through a particular point in 1 second ($\text{Hz} = 1 \text{ cycle/s}$).

$$\text{The speed } (u) \text{ of the wave} = \lambda \times \nu$$

Light as a Wave

Maxwell (1873), proposed that **visible light consists of electromagnetic waves.**



Electromagnetic radiation is the emission and transmission of energy in the form of electromagnetic waves.

Speed of light (c) in vacuum = 3.00×10^8 m/s

All electromagnetic radiation

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

EXAMPLE 7.1

The wavelength of the green light from a traffic signal is centered at 522 nm. What is the frequency of this radiation?

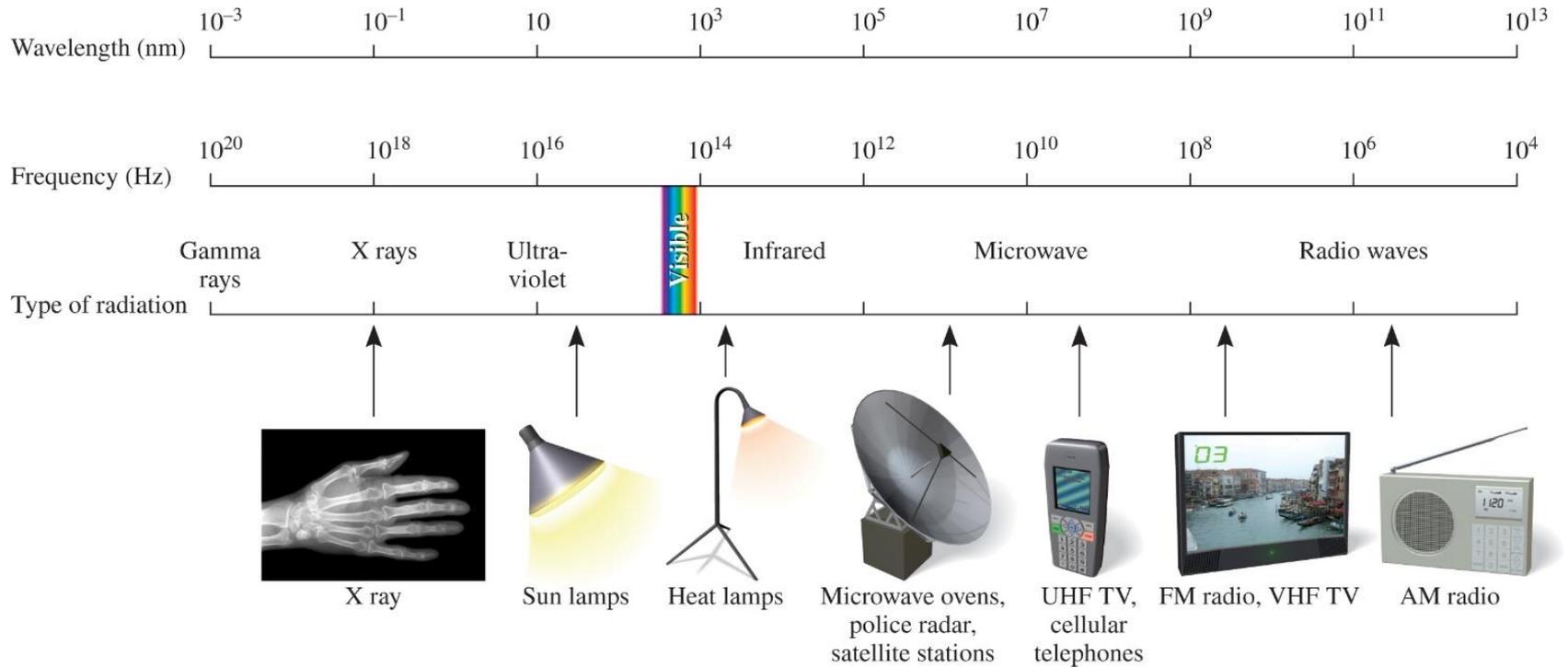
Solution Because the speed of light is given in meters per second, it is convenient to first convert wavelength to meters. Recall that $1 \text{ nm} = 1 \times 10^{-9} \text{ m}$ (see Table 1.3). We write

$$\begin{aligned}\lambda &= 522 \text{ nm} \times \frac{1 \times 10^{-9} \text{ m}}{1 \text{ nm}} = 522 \times 10^{-9} \text{ m} \\ &= 5.22 \times 10^{-7} \text{ m}\end{aligned}$$

Substituting in the wavelength and the speed of light ($3.00 \times 10^8 \text{ m/s}$), the frequency is

$$\begin{aligned}\nu &= \frac{3.00 \times 10^8 \text{ m/s}}{5.22 \times 10^{-7} \text{ m}} \\ &= 5.75 \times 10^{14} \text{ /s, or } 5.75 \times 10^{14} \text{ Hz}\end{aligned}$$

Electromagnetic Spectrum



(a)



(b)

Mystery #1, “Heated Solids Problem” Solved by Planck in 1900

When solids are heated, they emit electromagnetic radiation over a wide range of wavelengths.

Radiant energy emitted by an object at a certain temperature depends on its wavelength.

Planck: Energy (light) is emitted or absorbed in discrete units (quantum).

Quantum to the smallest quantity of energy that can be emitted (or absorbed) in the form of electromagnetic radiation

$$E = h \times \nu \quad \longrightarrow \quad E = h \frac{c}{\lambda}$$

Planck’s constant (h) = 6.63×10^{-34} J.s

Mystery #2, “Photoelectric Effect” Solved by Einstein in 1905

The **photoelectric effect**, a phenomenon in which electrons are ejected from the surface of certain metals exposed to light of at least a certain minimum frequency, called the threshold frequency

Light has both:

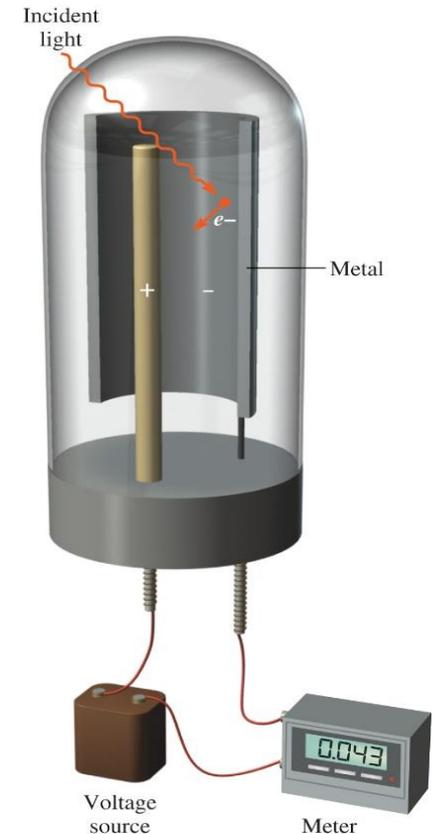
- 1) wave nature
- 2) particle nature

Photon is a “particle” of light

$$h\nu = KE + W$$

$$KE = h\nu - W$$

where **W** is the work function and depends how strongly electrons are held in the metal.



EXAMPLE 7.2

Calculate the energy (in joules) of (a) a photon with a wavelength of 5.00×10^4 nm (infrared region) and (b) a photon with a wavelength of 5.00×10^{-2} nm (X ray region).

Strategy In both (a) and (b) we are given the wavelength of a photon and asked to calculate its energy. We need to use Equation (7.3) to calculate the energy. Planck's constant is given in the text and also on the back inside cover.

Solution (a) From Equation (7.3),

$$\begin{aligned} E &= h \frac{c}{\lambda} \\ &= \frac{(6.63 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m/s})}{(5.00 \times 10^4 \text{ nm}) \frac{1 \times 10^{-9} \text{ m}}{1 \text{ nm}}} \\ &= 3.98 \times 10^{-21} \text{ J} \end{aligned}$$

This is the energy of a single photon with a 5.00×10^4 nm wavelength.

(b) Following the same procedure as in (a), we can show that the energy of the photon that has a wavelength of 5.00×10^{-2} nm is 3.98×10^{-15} J.

EXAMPLE 7.3

The work function of cesium metal is 3.42×10^{-19} J. (a) Calculate the minimum frequency of light required to release electrons from the metal. (b) Calculate the kinetic energy of the ejected electron if light of frequency $1.00 \times 10^{15} \text{ s}^{-1}$ is used for irradiating the metal.

Solution (a) Setting $\text{KE} = 0$ in Equation (7.4), we write

$$h\nu = W$$

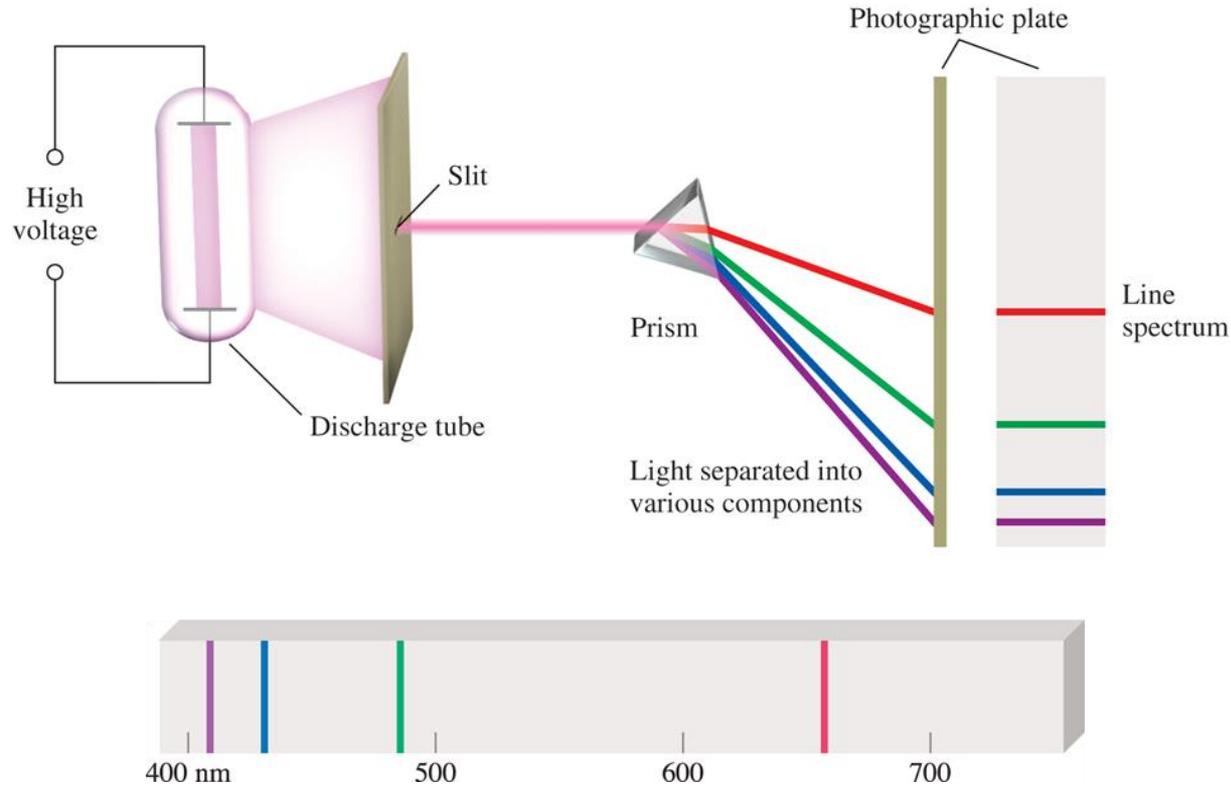
Thus,

$$\begin{aligned}\nu &= \frac{W}{h} = \frac{3.42 \times 10^{-19} \text{ J}}{6.63 \times 10^{-34} \text{ J} \cdot \text{s}} \\ &= 5.16 \times 10^{14} \text{ s}^{-1}\end{aligned}$$

(b) Rearranging Equation (7.4) gives

$$\begin{aligned}\text{KE} &= h\nu - W \\ &= (6.63 \times 10^{-34} \text{ J} \cdot \text{s})(1.00 \times 10^{15} \text{ s}^{-1}) - 3.42 \times 10^{-19} \text{ J} \\ &= 3.21 \times 10^{-19} \text{ J}\end{aligned}$$

Line Emission Spectrum of Hydrogen Atoms

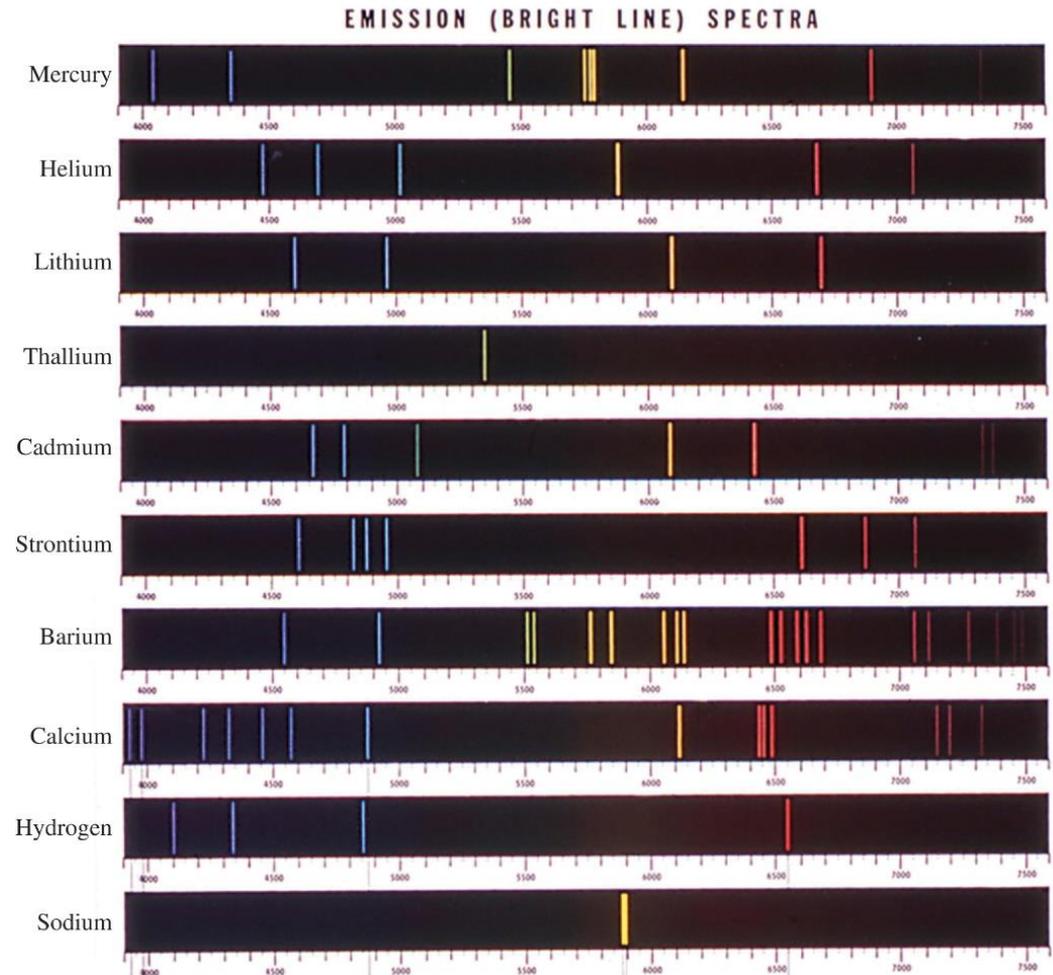


- The four lines, of the hydrogen atom, that appear are all in the visible region, because the photographic plate only detects visible light.
- The line at 410 nm appears violet, and the line at 434 nm appears blue. The emission line at 486 nm appears blue-green, and the line at 656 nm appears red.

Emission Spectra of Some Elements

hydrogen (4 lines) sodium (1 line), barium (20 lines) and calcium (14 lines) etc..

- Every element has a unique emission spectrum.
- The characteristic lines in atomic spectra can be used in chemical analysis to identify unknown atoms, much as fingerprints are used to identify people.
- When the lines of the emission spectrum of a known element exactly match the lines of the emission spectrum of an unknown sample, the identity of the sample is established.



Bohr's Model of the Atom (1913)

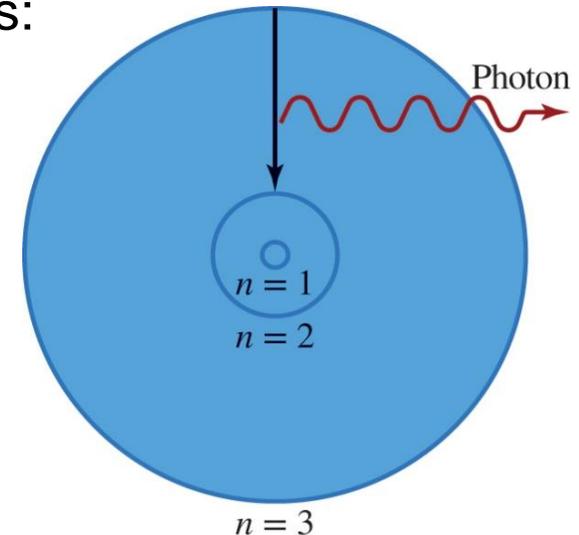
The energy of the photon is equal to the energy difference between the two n levels in the atom.

1. e^- can only have specific (quantized) energy values.
2. light is emitted as e^- moves from one energy level to a lower energy level
3. The radius of each circular orbit depends on n^2
4. e^- is allowed to occupy orbits with energies as:

$$E_n = -R_H \left(\frac{1}{n^2} \right)$$

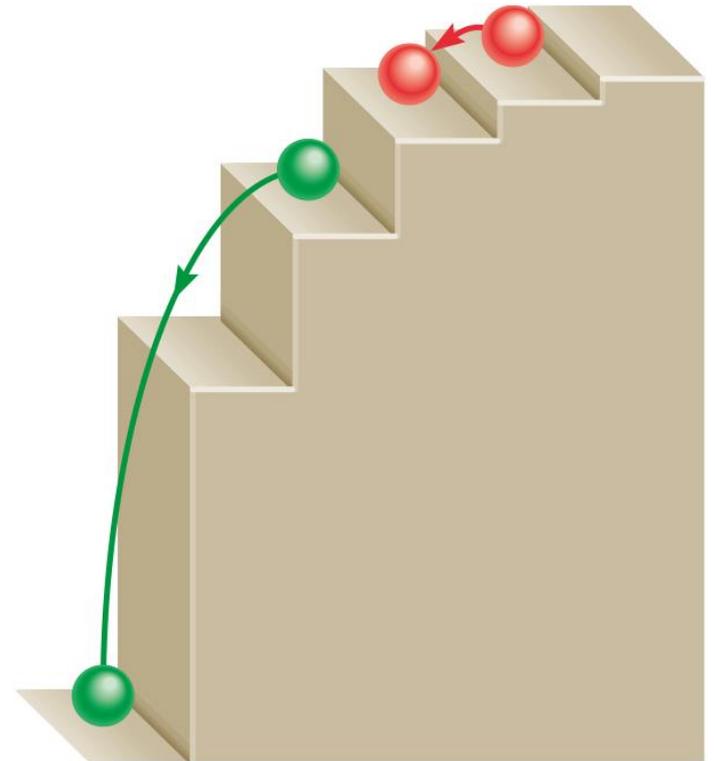
n (principle quantum number) = 1, 2, 3,

R_H (Rydberg constant) = 2.18×10^{-18} J

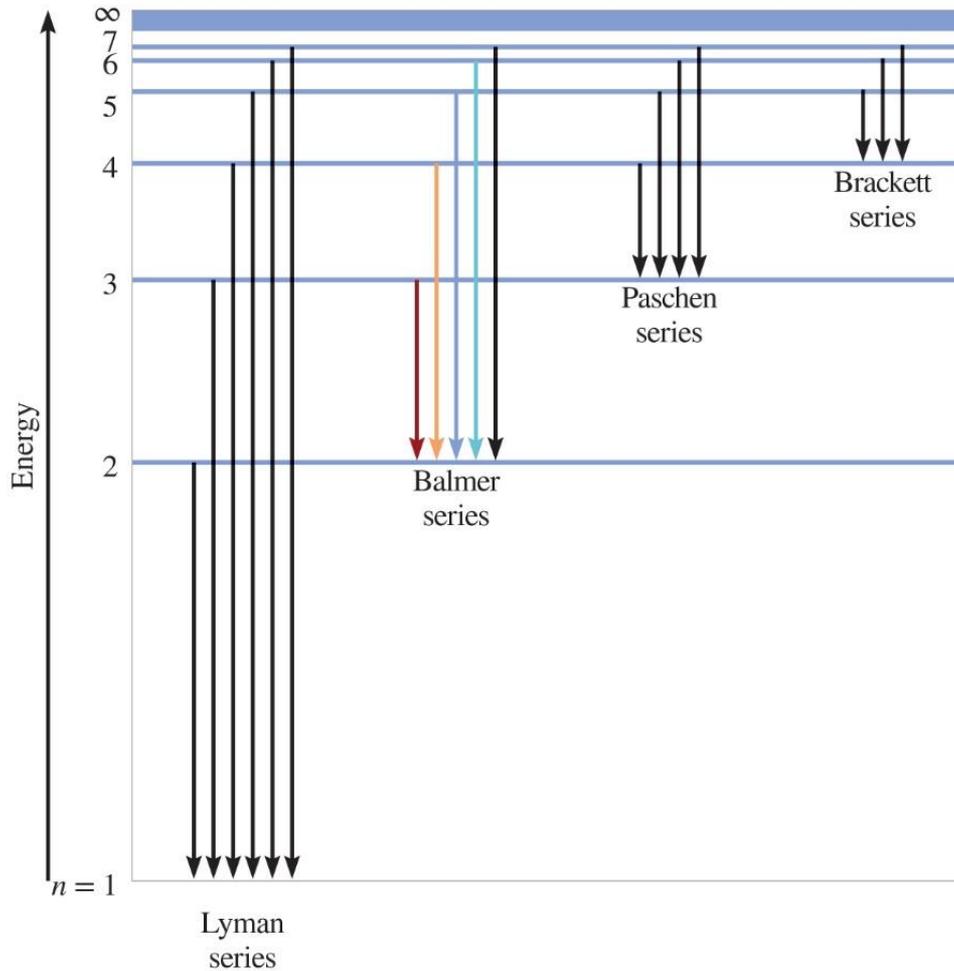


Quantized Energy

- When an electron undergoes a transition, it can only exist in an energy level, not between energy levels, because energy levels of the atom are quantized.
- A way to think about this is a ball on a staircase. The ball can only rest on steps, not between steps, much like how an electron can only exist in an energy level, not between energy levels.



Energy Transitions of the Hydrogen Atom



$$E_{\text{photon}} = \Delta E = E_f - E_i$$

$$E_f = -R_H \left(\frac{1}{n_f^2} \right)$$

$$E_i = -R_H \left(\frac{1}{n_i^2} \right)$$

$$\Delta E = R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$$

n_i = initial orbit and n_f = final orbit

Hydrogen Atom Emission Series

TABLE 7.1 The Various Series in Atomic Hydrogen Emission Spectrum

Series	n_f	n_i	Spectrum Region
Lyman	1	2, 3, 4, . . .	Ultraviolet
Balmer	2	3, 4, 5, . . .	Visible and ultraviolet
Paschen	3	4, 5, 6, . . .	Infrared
Brackett	4	5, 6, 7, . . .	Infrared

EXAMPLE 7.4

What is the wavelength of a photon (in nanometers) emitted during a transition from the $n_i = 5$ state to the $n_f = 2$ state in the hydrogen atom?

Strategy We are given the initial and final states in the emission process. We can calculate the energy of the emitted photon using Equation (7.6). Then from Equations (7.2) and (7.1) we can solve for the wavelength of the photon. The value of Rydberg's constant is given in the text.

Solution From Equation (7.6) we write

$$\begin{aligned}\Delta E &= R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) \\ &= 2.18 \times 10^{-18} \text{ J} \left(\frac{1}{5^2} - \frac{1}{2^2} \right) \\ &= -4.58 \times 10^{-19} \text{ J}\end{aligned}$$

(Continue..)

The negative sign indicates that this is energy associated with an emission process. To calculate the wavelength, we will omit the minus sign for ΔE because the wavelength of the photon must be positive. Because $\Delta E = h\nu$ or $\nu = \Delta E/h$, we can calculate the wavelength of the photon by writing

$$\begin{aligned}\lambda &= \frac{c}{\nu} \\ &= \frac{ch}{\Delta E} \\ &= \frac{(3.00 \times 10^8 \text{ m/s})(6.63 \times 10^{-34} \text{ J} \cdot \text{s})}{4.58 \times 10^{-19} \text{ J}} \\ &= 4.34 \times 10^{-7} \text{ m} \\ &= 4.34 \times 10^{-7} \text{ m} \times \left(\frac{1 \text{ nm}}{1 \times 10^{-9} \text{ m}} \right) = 434 \text{ nm}\end{aligned}$$

Quantization of Electron Energy

Why is e- energy quantized?

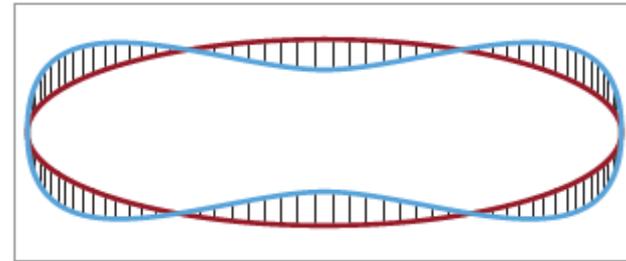
De Broglie (1924) reasoned that e- is **both particle and wave**.

$$2\pi r = n\lambda \quad \lambda = \frac{h}{mu}$$

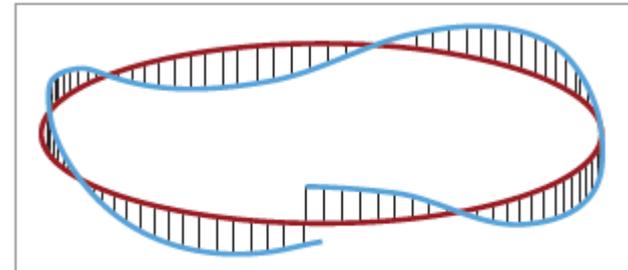
u = velocity of e^-

m = mass of e^-

$2\pi r$ = circumference of the allowed orbit



(a)



(b)

(a) The circumference of the orbit is equal to an integral number of wavelengths. This is an allowed orbit.

(b) The circumference of the orbit is not equal to an integral number of wavelengths. As a result, the electron wave does not close in on itself. This is a nonallowed orbit.

EXAMPLE 7.5

Calculate the wavelength of the “particle” in the following two cases: (a) The fastest serve in tennis is about 150 miles per hour, or 68 m/s. Calculate the wavelength associated with a 6.0×10^{-2} -kg tennis ball traveling at this speed. (b) Calculate the wavelength associated with an electron (9.1094×10^{-31} kg) moving at 68 m/s.

Strategy We are given the mass and the speed of the particle in (a) and (b) and asked to calculate the wavelength so we need Equation (7.8). Note that because the units of Planck’s constants are $\text{J} \cdot \text{s}$, m and u must be in kg and m/s ($1 \text{ J} = 1 \text{ kg m}^2/\text{s}^2$), respectively.

Solution (a) Using Equation (7.8) we write

$$\begin{aligned}\lambda &= \frac{h}{mu} \\ &= \frac{6.63 \times 10^{-34} \text{ J} \cdot \text{s}}{(6.0 \times 10^{-2} \text{ kg}) \times 68 \text{ m/s}} \\ &= 1.6 \times 10^{-34} \text{ m}\end{aligned}$$

Comment This is an exceedingly small wavelength considering that the size of an atom itself is on the order of 1×10^{-10} m. For this reason, the wave properties of a tennis ball cannot be detected by any existing measuring device.

(Continued)

(b) In this case,

$$\begin{aligned}\lambda &= \frac{h}{mu} \\ &= \frac{6.63 \times 10^{-34} \text{ J} \cdot \text{s}}{(9.1094 \times 10^{-31} \text{ kg}) \times 68 \text{ m/s}} \\ &= 1.1 \times 10^{-5} \text{ m}\end{aligned}$$

Comment This wavelength ($1.1 \times 10^{-5} \text{ m}$ or $1.1 \times 10^4 \text{ nm}$) is in the infrared region. This calculation shows that only electrons (and other submicroscopic particles) have measurable wavelengths.

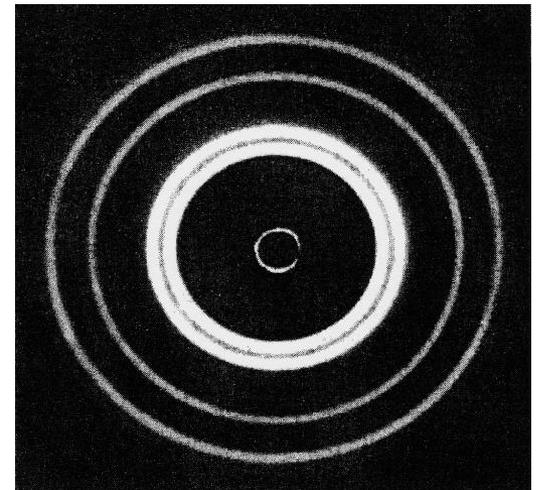
Practice Exercise Calculate the wavelength (in nanometers) of a H atom (mass = $1.674 \times 10^{-27} \text{ kg}$) moving at $7.00 \times 10^2 \text{ cm/s}$.

Schrodinger Wave Equation ¹

In 1926 Schrodinger wrote an equation that described both the **particle and wave nature of the e-**
Wave function (ψ) describes:

1. Energy of e- with a with a given ψ
2. probability of finding e- in a volume of space

Schrodinger's equation can only be solved exactly for the hydrogen atom Must approximate its solution for multi electron systems.



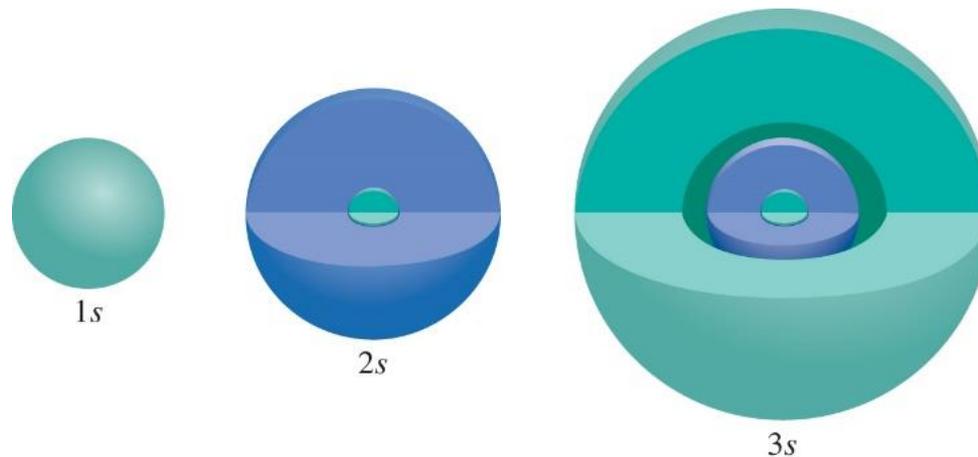
Schrodinger Wave Equation ₂

ψ is a function of four numbers called
quantum numbers (n, l, m_l, m_s)

Principal Quantum Number n

$$n = 1, 2, 3, 4, \dots$$

distance of e^- from the nucleus



Schrodinger Wave Equation ³

quantum numbers : (n, l, m_l, m_s)

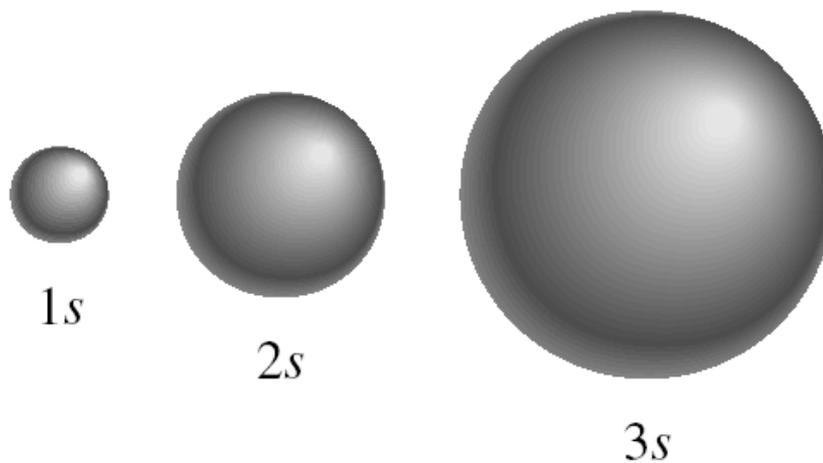
Angular Momentum Quantum Number l

The shape of the “volume” of space that the e^- occupies

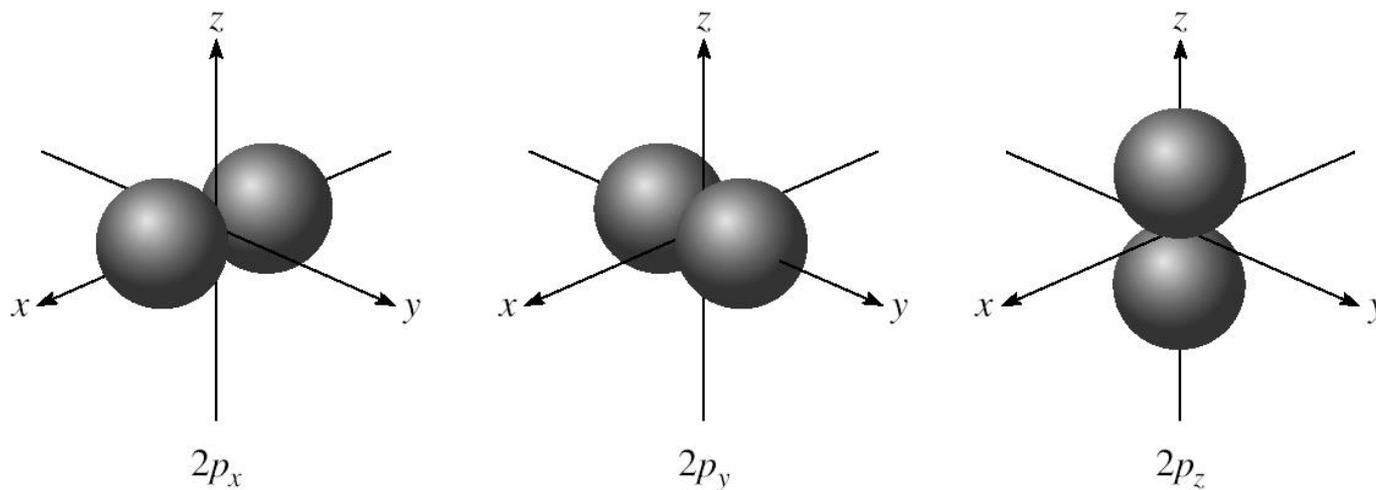
for a given value of n , $l = 0, 1, 2, 3, \dots n-1$

$n = 1, l = 0$	$l = 0$	s orbital
$n = 2, l = 0$ or 1	$l = 1$	p orbital
$n = 3, l = 0, 1,$ or 2	$l = 2$	d orbital
$n = 4, l = 0, 1, 2,$ or 3	$l = 3$	f orbital

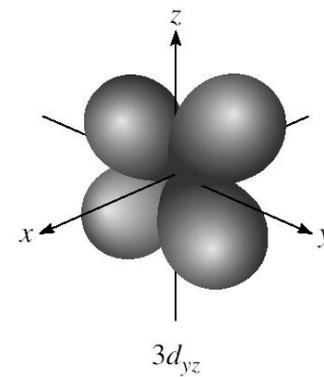
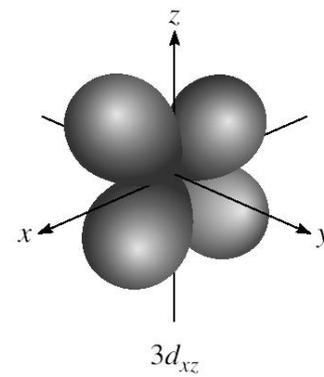
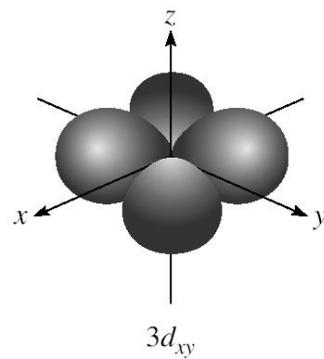
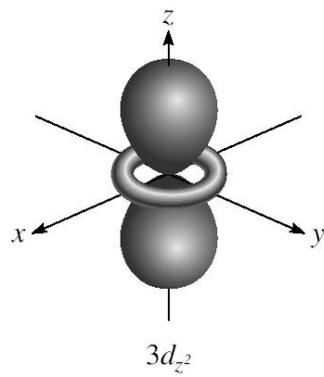
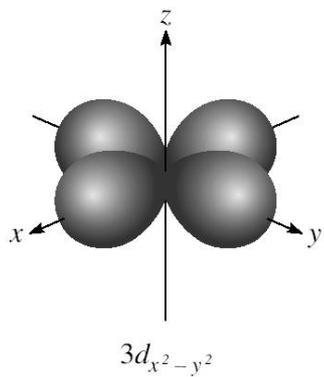
$l = 0$ (s orbitals)



$l = 1$ (p orbitals)



$l = 2$ (d orbitals)



Schrodinger Wave Equation 4

quantum numbers : (n, l, m_l, m_s)

Magnetic Quantum Number m_l

for a given value of l

$$m = (2l+1)$$

$$m_l = -l, \dots, 0, \dots, +l$$

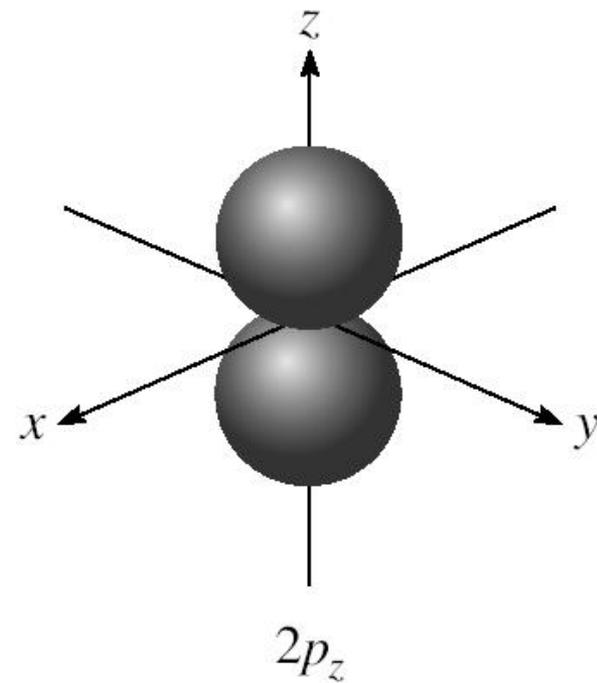
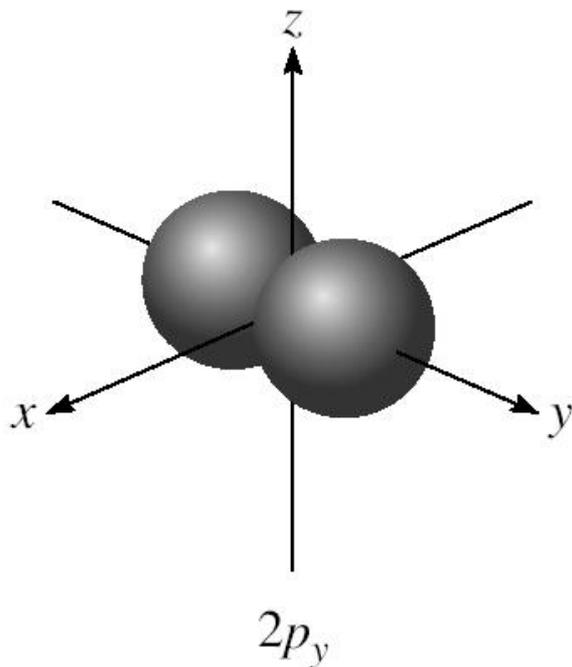
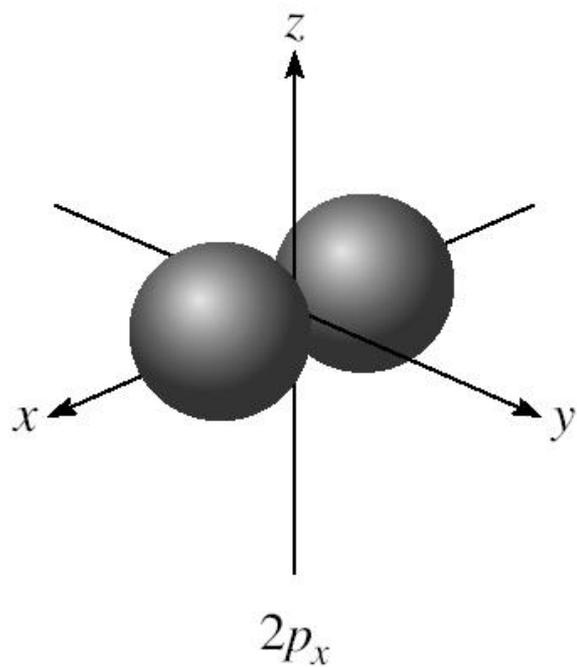
If $l = 1$ (p orbital), $m_l = -1, 0, \text{ or } +1$

If $l = 2$ (d orbital), $m_l = -2, -1, 0, +1, +2$

orientation of the orbital in space

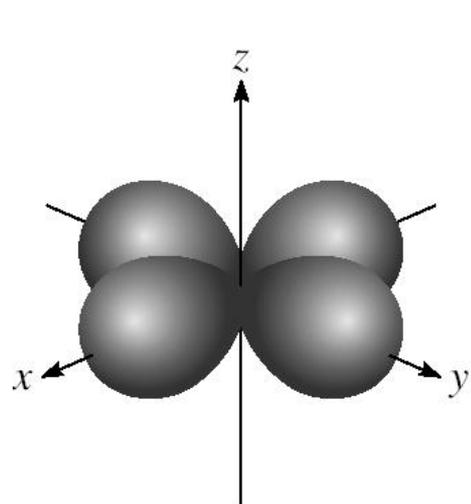
$$m_l = -1, 0, \text{ or } 1$$

3 orientations in space

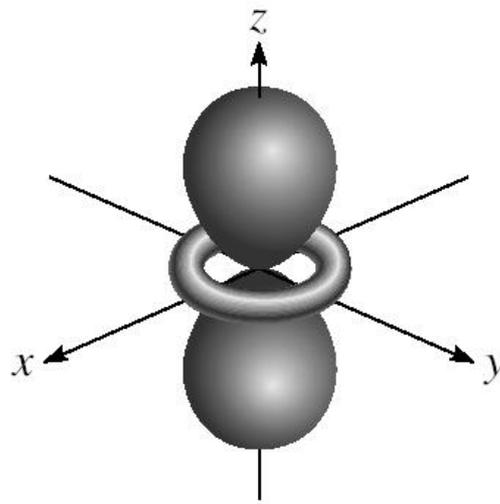


$m_l = -2, -1, 0, 1, \text{ or } 2$

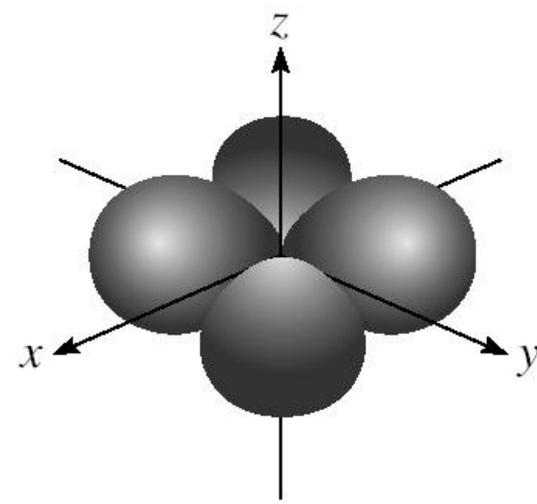
5 orientations in space



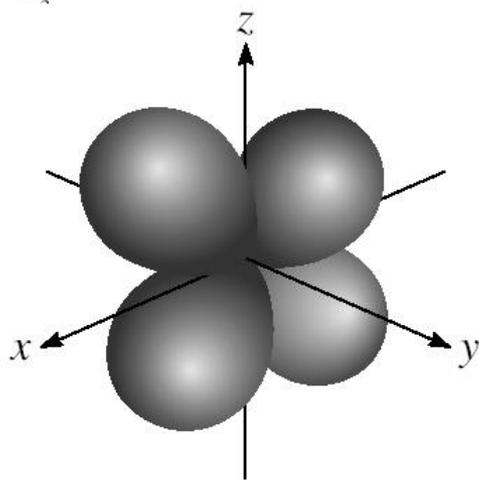
$3d_{x^2-y^2}$



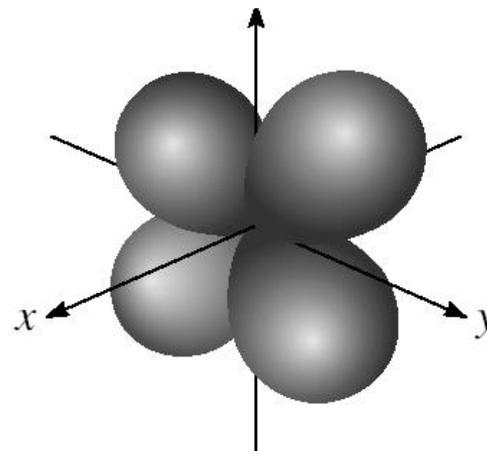
$3d_{z^2}$



$3d_{xy}$



$3d_{xz}$



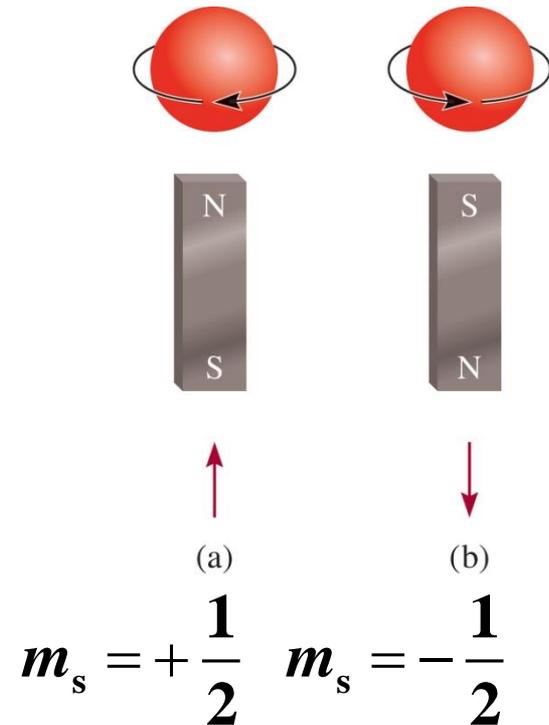
$3d_{yz}$

Schrodinger Wave Equation 5

$$(n, l, m_l, m_s)$$

Spin Quantum Number m_s

$$m_s = +\frac{1}{2} \text{ or } -\frac{1}{2}$$

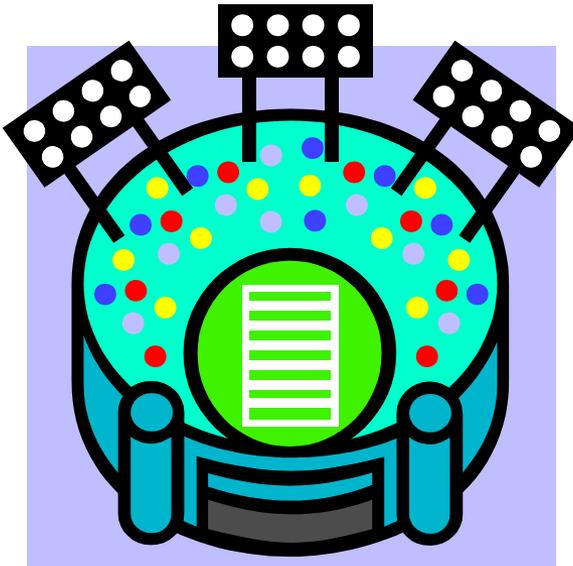


An electron that spins counterclockwise is given a spin value of $-1/2$. These electrons, known as "spin down", are indicated by an arrow pointing downward.

Quantum Numbers: (n, l, m_l, m_s)

Existence (and energy) of electron in atom is described by its *unique* wave function ψ .

Pauli exclusion principle - no two electrons in an atom can have the same four quantum numbers.



Each seat is uniquely identified (E, R12, S8)
Each seat can hold only one individual at a time

TABLE 7.2 Relation Between Quantum Numbers and Atomic Orbitals

n	ℓ	m_ℓ	Number of Orbitals	Atomic Orbital Designations
1	0	0	1	1s
2	0	0	1	2s
	1	-1, 0, 1	3	$2p_x, 2p_y, 2p_z$
3	0	0	1	3s
	1	-1, 0, 1	3	$3p_x, 3p_y, 3p_z$
	2	-2, -1, 0, 1, 2	5	$3d_{xy}, 3d_{yz}, 3d_{xz},$ $3d_{x^2-y^2}, 3d_{z^2}$
⋮	⋮	⋮	⋮	⋮

Quantum Numbers: (n, l, m_l, m_s)

Shell – electrons with the same value of n

Subshell – electrons with the same values of n **and** l

Orbital – electrons with the same values of $n, l, \mathbf{and} m_l$

How many electrons can an orbital hold?

If $n, l,$ and m_l are fixed, then $m_s = 1/2$ or $-1/2$

$\psi = (n, l, m_l, 1/2)$ **or** $\psi = (n, l, m_l, -1/2)$

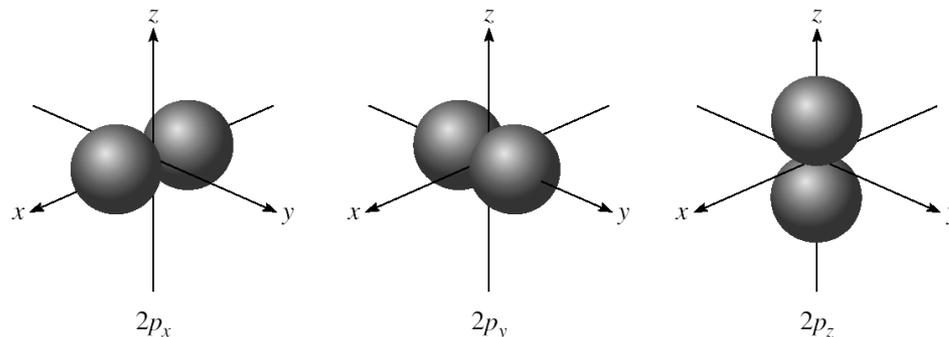
An orbital can hold **2 electrons**

How many 2p orbitals are there in an atom?

$$\begin{array}{c} n = 2 \\ \downarrow \\ 2p \\ \uparrow \\ l = 1 \end{array}$$

If $l = 1$, then $m_l = -1, 0, \text{ or } +1$

3 orbitals



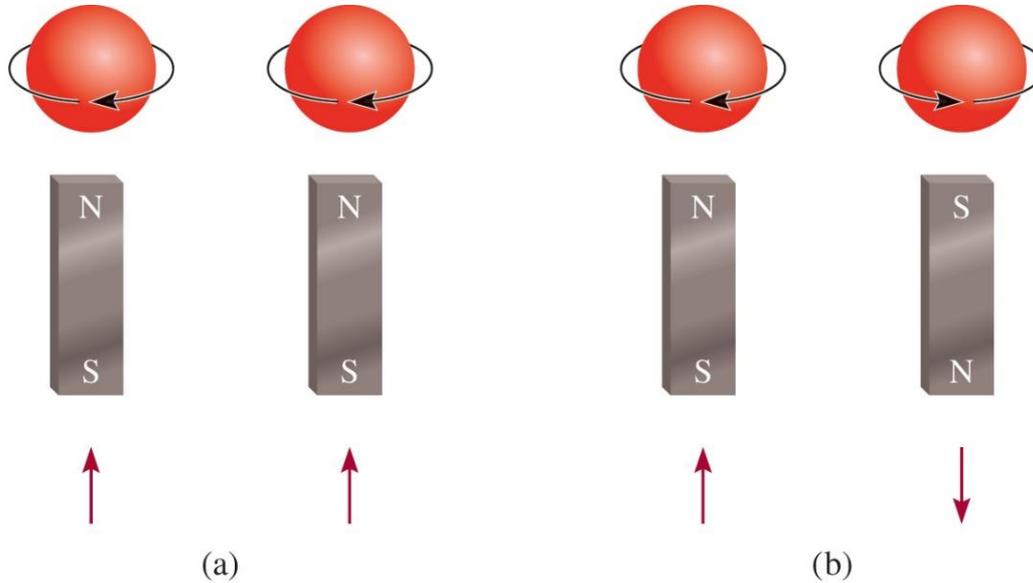
How many electrons can be placed in the 3d subshell?

$$\begin{array}{c} n = 3 \\ \downarrow \\ 3d \\ \uparrow \\ l = 2 \end{array}$$

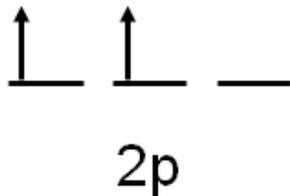
If $l = 2$, then $m_l = -2, -1, 0, +1, \text{ or } +2$

5 orbitals which can hold a total of **10 e⁻**

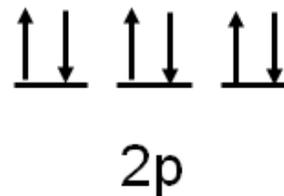
Paramagnetism and Diamagnetism



Paramagnetic
unpaired electrons



Diamagnetic
all electrons paired



EXAMPLE 7.8

Write the four quantum numbers for an electron in a $3p$ orbital.

Strategy What do the “3” and “ p ” designate in $3p$? How many orbitals (values of m_ℓ) are there in a $3p$ subshell? What are the possible values of electron spin quantum number?

Solution To start with, we know that the principal quantum number n is 3 and the angular momentum quantum number ℓ must be 1 (because we are dealing with a p orbital).

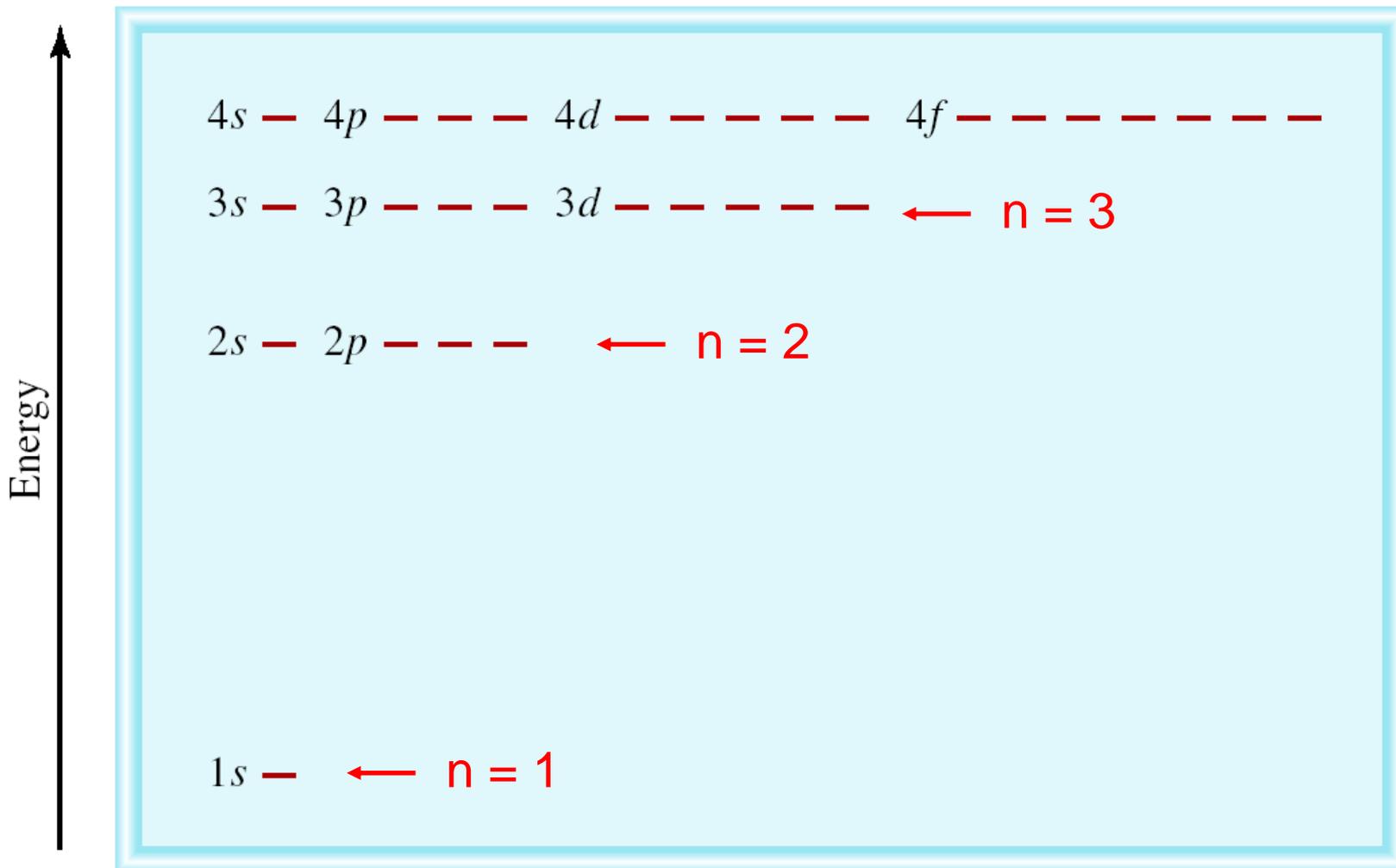
For $\ell = 1$, there are three values of m_ℓ given by -1 , 0 , and 1 . Because the electron spin quantum number m_s can be either $+\frac{1}{2}$ or $-\frac{1}{2}$, we conclude that there are six possible ways to designate the electron using the (n, ℓ, m_ℓ, m_s) notation:

$$\begin{array}{ll} (3, 1, -1, +\frac{1}{2}) & (3, 1, -1, -\frac{1}{2}) \\ (3, 1, 0, +\frac{1}{2}) & (3, 1, 0, -\frac{1}{2}) \\ (3, 1, 1, +\frac{1}{2}) & (3, 1, 1, -\frac{1}{2}) \end{array}$$

The Energies of Orbitals

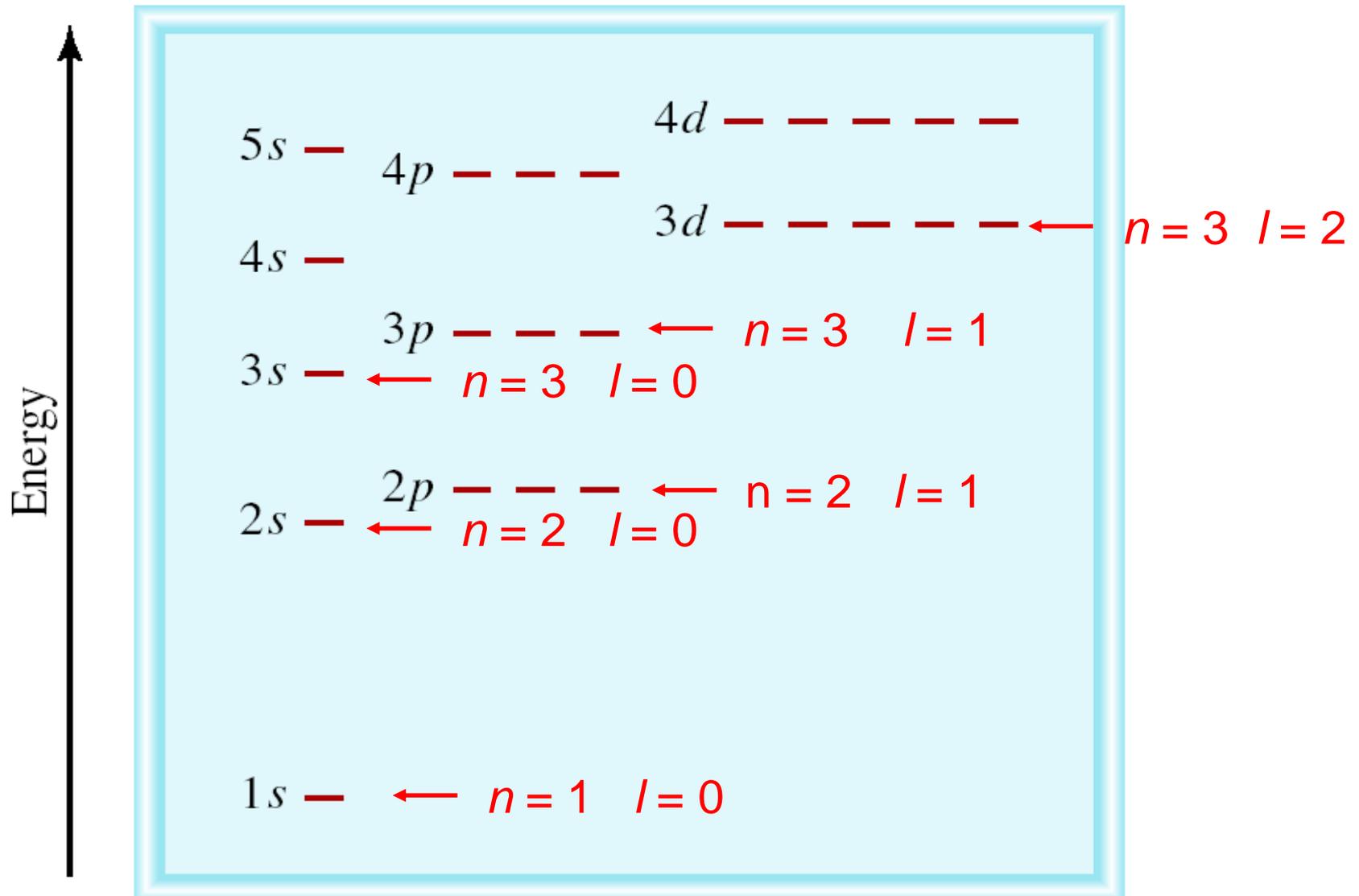
Energy of orbitals in a *single* electron atom

Energy only depends on principal quantum number *n*



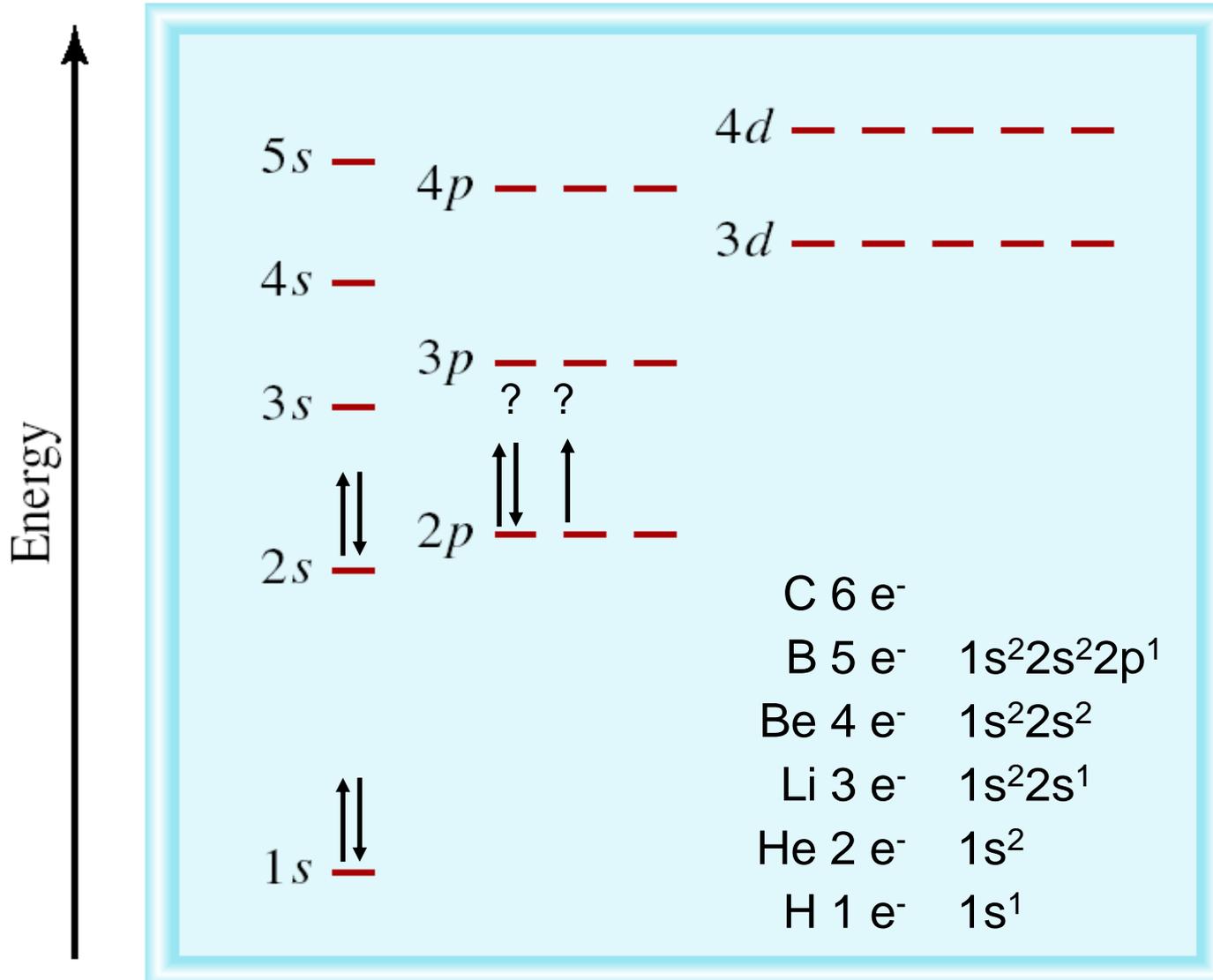
Energy of orbitals in a multi-electron atom

Energy depends on n and l



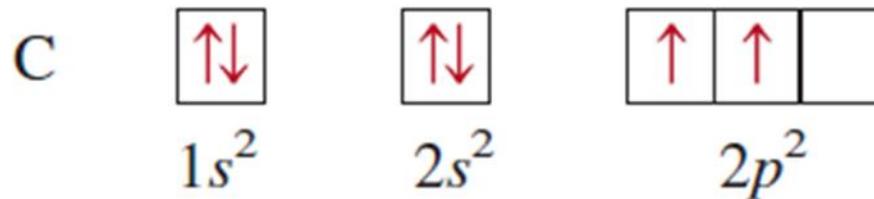
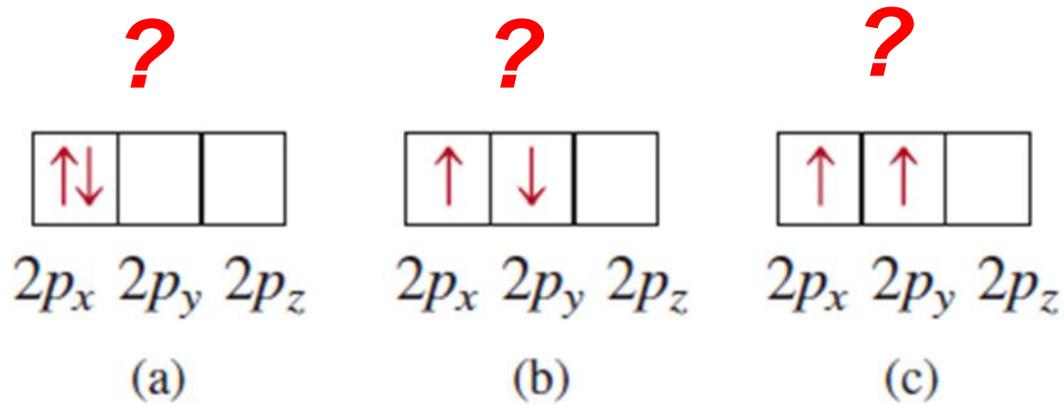
Aufbau Principle

“Fill up” electrons in lowest energy orbitals (*Aufbau principle*)

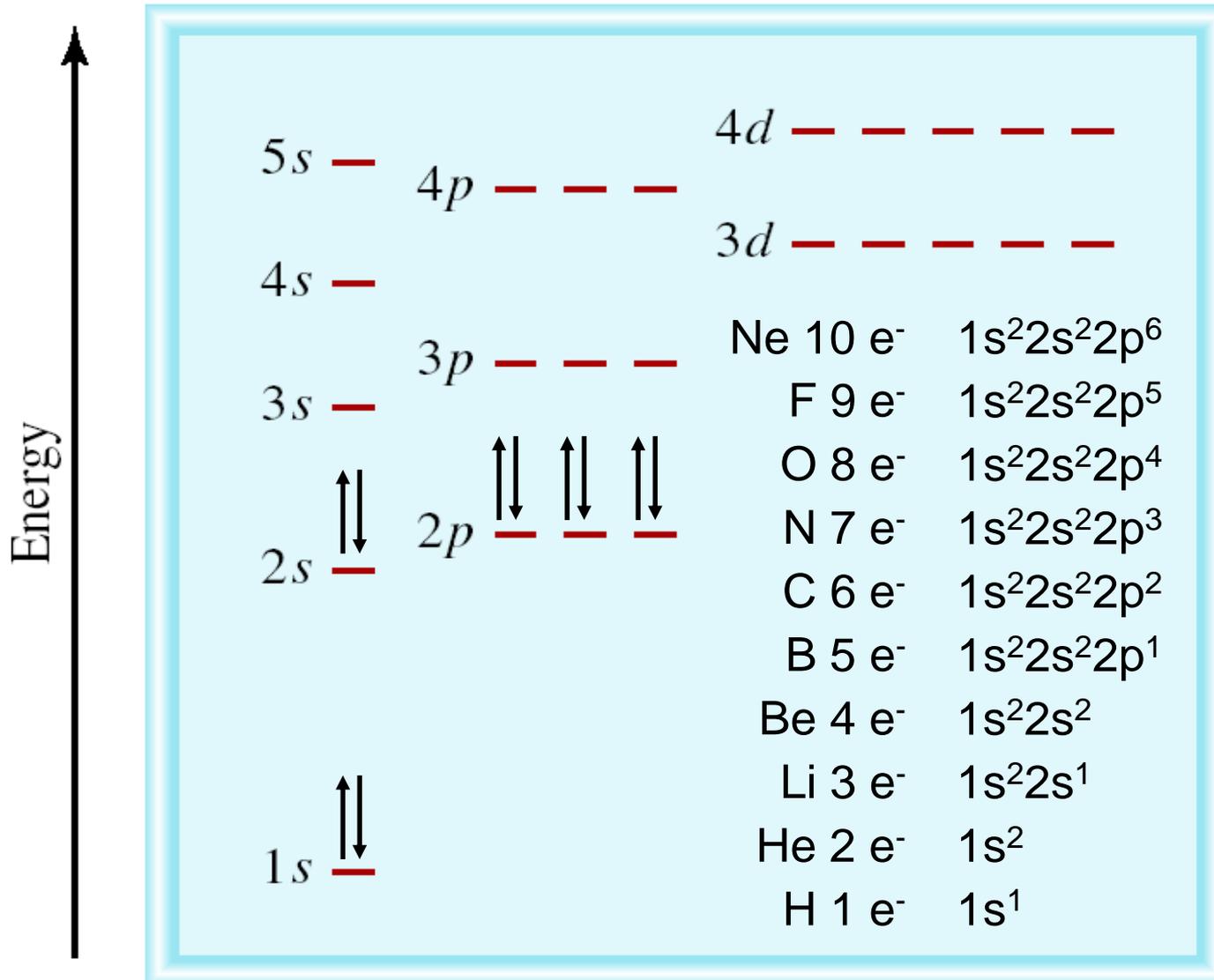


Hund's Rule

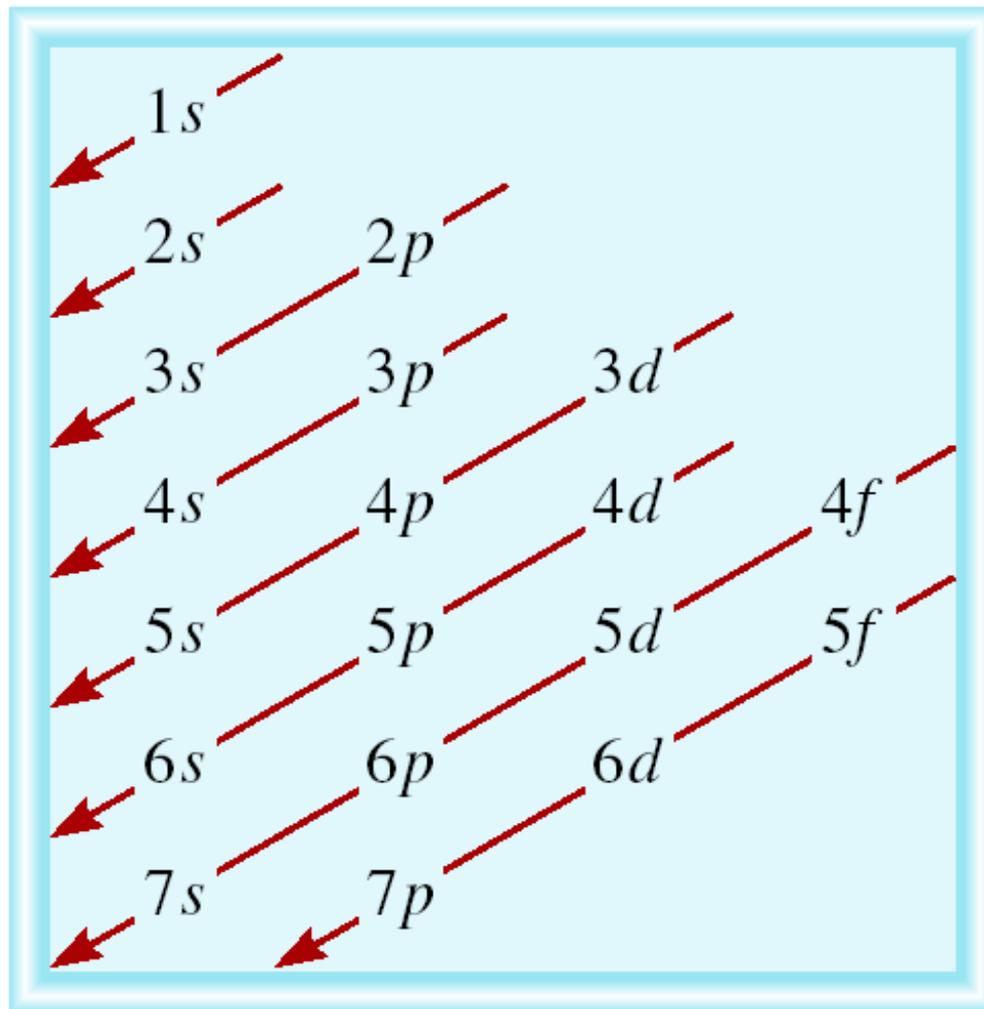
The most stable arrangement of electrons in subshells is the one with the greatest number of parallel spins (**Hund's rule**).



Aufbau Principle + Hund's Rule

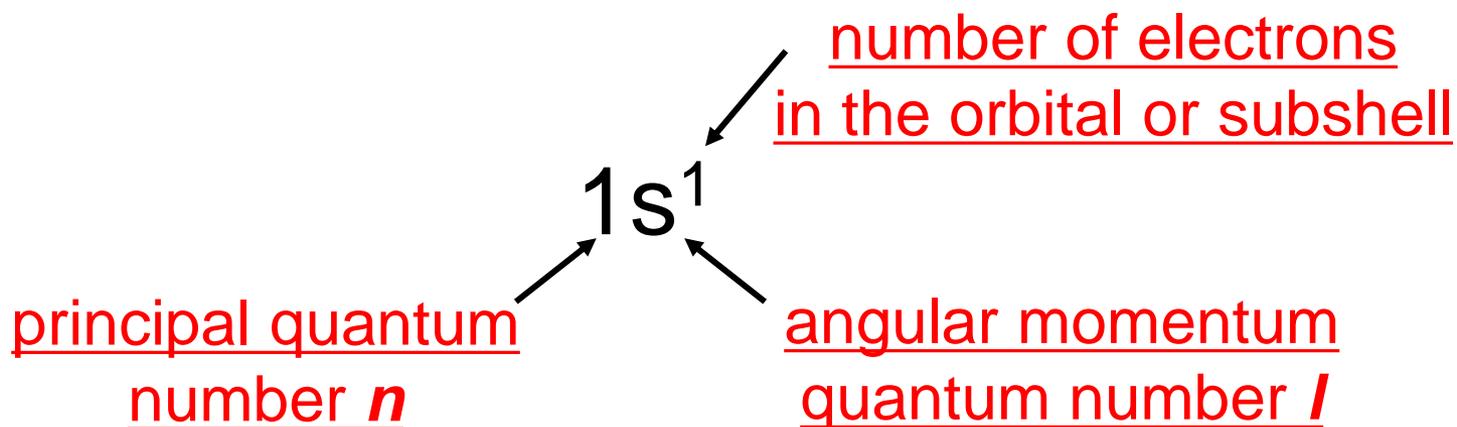


Order of Orbitals (filling) in Multi-electron Atom



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

Electron configuration is how the electrons are distributed among the various atomic orbitals in an atom.



Orbital diagram



Outermost subshell being filled with electrons

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

What is the electron configuration of Mg?

Mg 12 electrons

$1s < 2s < 2p < 3s < 3p < 4s$

$1s^2 2s^2 2p^6 3s^2$ $2 + 2 + 6 + 2 = 12$ electrons

Abbreviated as $[\text{Ne}]3s^2$ $[\text{Ne}] 1s^2 2s^2 2p^6$

What are the possible quantum numbers for the last (outermost) electron in Cl?

Cl 17 electrons $1s < 2s < 2p < 3s < 3p < 4s$

$1s^2 2s^2 2p^6 3s^2 3p^5$ $2 + 2 + 6 + 2 + 5 = 17$ electrons

Last electron added to **3p** orbital

$n = 3$ $l = 1$ $m_l = -1, 0, \text{ or } +1$ $m_s = 1/2 \text{ or } -1/2$

Example 7.12 ₁

An oxygen atom has a total of eight electrons. Write the four quantum numbers for each of the eight electrons in the ground state.

Example 7.12 ₂

Strategy

We start with $n = 1$ and proceed to fill orbitals in the order shown in Figure 7.24.

For each value of n we determine the possible values of l .

For each value of l , we assign the possible values of m_l .

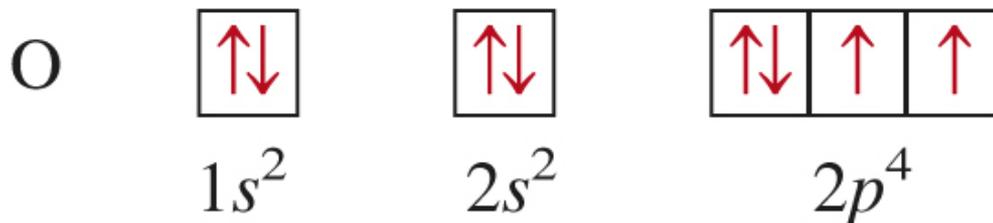
We can place electrons in the orbitals according to the Pauli exclusion principle and Hund's rule.

Example 7.12 ₃

Solution

We start with $n = 1$, so $l = 0$, a subshell corresponding to the $1s$ orbital. This orbital can accommodate a total of two electrons.

Next, $n = 2$, and l may be either 0 or 1. The $l = 0$ subshell contains one $2s$ orbital, which can accommodate two electrons. The remaining four electrons are placed in the $l = 1$ subshell, which contains three $2p$ orbitals. The orbital diagram is



Example 7.12 ₄

The results are summarized in the following table:

Electron	n	l	m_l	m_s	Orbital
1	1	0	0	$+\frac{1}{2}$	$1s$
2	1	0	0	$-\frac{1}{2}$	$1s$
3	2	0	0	$+\frac{1}{2}$	$2s$
4	2	0	0	$-\frac{1}{2}$	$2s$
5	2	1	-1	$+\frac{1}{2}$	$2p_x, 2p_y, 2p_z$
6	2	1	0	$+\frac{1}{2}$	$2p_x, 2p_y, 2p_z$
7	2	1	1	$+\frac{1}{2}$	$2p_x, 2p_y, 2p_z$
8	2	1	1	$-\frac{1}{2}$	$2p_x, 2p_y, 2p_z$

Of course, the placement of the eighth electron in the orbital labeled $m_l = 1$ is completely arbitrary. It would be equally correct to assign it to $m_l = 0$ or $m_l = -1$.

Ground State Electron Configurations

Table 7.3 The Ground-State Electron Configurations of the Elements*

Atomic Number	Symbol	Electron Configuration	Atomic Number	Symbol	Electron Configuration	Atomic Number	Symbol	Electron Configuration
1	H	1s ¹	41	Nb	[Kr]5s ¹ 4d ⁴	81	Tl	[Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ¹
2	He	1s ²	42	Mo	[Kr]5s ¹ 4d ⁵	82	Pb	[Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ²
3	Li	[He]2s ¹	43	Tc	[Kr]5s ² 4d ⁵	83	Bi	[Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ³
4	Be	[He]2s ²	44	Ru	[Kr]5s ¹ 4d ⁷	84	Po	[Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ⁴
5	B	[He]2s ² 2p ¹	45	Rh	[Kr]5s ¹ 4d ⁸	85	At	[Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ⁵
6	C	[He]2s ² 2p ²	46	Pd	[Kr]4d ¹⁰	86	Rn	[Xe]6s ² 4f ¹⁴ 5d ¹⁰ 6p ⁶
7	N	[He]2s ² 2p ³	47	Ag	[Kr]5s ¹ 4d ¹⁰	87	Fr	[Rn]7s ¹
8	O	[He]2s ² 2p ⁴	48	Cd	[Kr]5s ² 4d ¹⁰	88	Ra	[Rn]7s ²
9	F	[He]2s ² 2p ⁵	49	In	[Kr]5s ² 4d ¹⁰ 5p ¹	89	Ac	[Rn]7s ² 6d ¹
10	Ne	[He]2s ² 2p ⁶	50	Sn	[Kr]5s ² 4d ¹⁰ 5p ²	90	Th	[Rn]7s ² 6d ²
11	Na	[Ne]3s ¹	51	Sb	[Kr]5s ² 4d ¹⁰ 5p ³	91	Pa	[Rn]7s ² 5f ² 6d ¹
12	Mg	[Ne]3s ²	52	Te	[Kr]5s ² 4d ¹⁰ 5p ⁴	92	U	[Rn]7s ² 5f ³ 6d ¹
13	Al	[Ne]3s ² 3p ¹	53	I	[Kr]5s ² 4d ¹⁰ 5p ⁵	93	Np	[Rn]7s ² 5f ⁴ 6d ¹
14	Si	[Ne]3s ² 3p ²	54	Xe	[Kr]5s ² 4d ¹⁰ 5p ⁶	94	Pu	[Rn]7s ² 5f ⁶
15	P	[Ne]3s ² 3p ³	55	Cs	[Xe]6s ¹	95	Am	[Rn]7s ² 5f ⁷
16	S	[Ne]3s ² 3p ⁴	56	Ba	[Xe]6s ²	96	Cm	[Rn]7s ² 5f ⁶ 6d ¹
17	Cl	[Ne]3s ² 3p ⁵	57	La	[Xe]6s ² 5d ¹	97	Bk	[Rn]7s ² 5f ⁹
18	Ar	[Ne]3s ² 3p ⁶	58	Ce	[Xe]6s ² 4f ¹ 5d ¹	98	Cf	[Rn]7s ² 5f ¹⁰
19	K	[Ar]4s ¹	59	Pr	[Xe]6s ² 4f ³	99	Es	[Rn]7s ² 5f ¹¹
20	Ca	[Ar]4s ²	60	Nd	[Xe]6s ² 4f ⁴	100	Fm	[Rn]7s ² 5f ¹²
21	Sc	[Ar]4s ² 3d ¹	61	Pm	[Xe]6s ² 4f ⁵	101	Md	[Rn]7s ² 5f ¹³
22	Ti	[Ar]4s ² 3d ²	62	Sm	[Xe]6s ² 4f ⁶	102	No	[Rn]7s ² 5f ¹⁴
23	V	[Ar]4s ² 3d ³	63	Eu	[Xe]6s ² 4f ⁷	103	Lr	[Rn]7s ² 5f ¹⁴ 6d ¹
24	Cr	[Ar]4s ¹ 3d ⁵	64	Gd	[Xe]6s ² 4f ⁷ 5d ¹	104	Rf	[Rn]7s ² 5f ¹⁴ 6d ²
25	Mn	[Ar]4s ² 3d ⁵	65	Tb	[Xe]6s ² 4f ⁹	105	Db	[Rn]7s ² 5f ¹⁴ 6d ³
26	Fe	[Ar]4s ² 3d ⁶	66	Dy	[Xe]6s ² 4f ¹⁰	106	Sg	[Rn]7s ² 5f ¹⁴ 6d ⁴
27	Co	[Ar]4s ² 3d ⁷	67	Ho	[Xe]6s ² 4f ¹¹	107	Bh	[Rn]7s ² 5f ¹⁴ 6d ⁵
28	Ni	[Ar]4s ² 3d ⁸	68	Er	[Xe]6s ² 4f ¹²	108	Hs	[Rn]7s ² 5f ¹⁴ 6d ⁶
29	Cu	[Ar]4s ¹ 3d ¹⁰	69	Tm	[Xe]6s ² 4f ¹³	109	Mt	[Rn]7s ² 5f ¹⁴ 6d ⁷
30	Zn	[Ar]4s ² 3d ¹⁰	70	Yb	[Xe]6s ² 4f ¹⁴	110	Ds	[Rn]7s ² 5f ¹⁴ 6d ⁸
31	Ga	[Ar]4s ² 3d ¹⁰ 4p ¹	71	Lu	[Xe]6s ² 4f ¹⁴ 5d ¹	111	Rg	[Rn]7s ² 5f ¹⁴ 6d ⁹
32	Ge	[Ar]4s ² 3d ¹⁰ 4p ²	72	Hf	[Xe]6s ² 4f ¹⁴ 5d ²	112	Cn	[Rn]7s ² 5f ¹⁴ 6d ¹⁰
33	As	[Ar]4s ² 3d ¹⁰ 4p ³	73	Ta	[Xe]6s ² 4f ¹⁴ 5d ³	113	Nh	[Rn]7s ² 5f ¹⁴ 6d ¹⁰ 7p ¹
34	Se	[Ar]4s ² 3d ¹⁰ 4p ⁴	74	W	[Xe]6s ² 4f ¹⁴ 5d ⁴	114	Fl	[Rn]7s ² 5f ¹⁴ 6d ¹⁰ 7p ²
35	Br	[Ar]4s ² 3d ¹⁰ 4p ⁵	75	Re	[Xe]6s ² 4f ¹⁴ 5d ⁵	115	Mc	[Rn]7s ² 5f ¹⁴ 6d ¹⁰ 7p ³
36	Kr	[Ar]4s ² 3d ¹⁰ 4p ⁶	76	Os	[Xe]6s ² 4f ¹⁴ 5d ⁶	116	Lv	[Rn]7s ² 5f ¹⁴ 6d ¹⁰ 7p ⁴
37	Rb	[Kr]5s ¹	77	Ir	[Xe]6s ² 4f ¹⁴ 5d ⁷	117	Ts	[Rn]7s ² 5f ¹⁴ 6d ¹⁰ 7p ⁵
38	Sr	[Kr]5s ²	78	Pt	[Xe]6s ¹ 4f ¹⁴ 5d ⁹	118	Og	[Rn]7s ² 5f ¹⁴ 6d ¹⁰ 7p ⁶
39	Y	[Kr]5s ² 4d ¹	79	Au	[Xe]6s ¹ 4f ¹⁴ 5d ¹⁰			
40	Zr	[Kr]5s ² 4d ²	80	Hg	[Xe]6s ² 4f ¹⁴ 5d ¹⁰			

*The symbol [He] is called the helium core and represents 1s². [Ne] is called the neon core and represents 1s²2s²2p⁶. [Ar] is called the argon core and represents [Ne]3s²3p⁶. [Kr] is called the krypton core and represents [Ar]4s²3d¹⁰4p⁶. [Xe] is called the xenon core and represents [Kr]5s²4d¹⁰5p⁶. [Rn] is called the radon core and represents [Xe]6s²4f¹⁴5d¹⁰6p⁶.

Example 7.13 ₁

Write the ground-state electron configurations for

(a) sulfur (S)

(b) palladium (Pd), which is diamagnetic

	1																	18
		2											13	14	15	16	17	
			3	4	5	6	7	8	9	10	11	12				S		
										Pd								

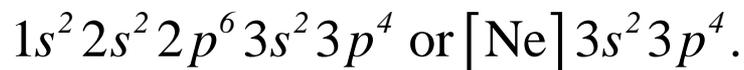
Example 7.13 ₂

(a) *Strategy*

How many electrons are in the S ($Z = 16$) atom? We start with $n = 1$ and proceed to fill orbitals in the order shown in Figure 7.24. For each value of ℓ , we assign the possible values of m_ℓ . We can place electrons in the orbitals according to the Pauli exclusion principle and Hund's rule and then write the electron configuration. The task is simplified if we use the noble-gas core preceding S for the inner electrons.

Solution

Sulfur has 16 electrons. The noble gas core in this case is [Ne]. (Ne is the noble gas in the period preceding sulfur.) [Ne] represents $1s^2 2s^2 2p^6$. This leaves us 6 electrons to fill the 3s subshell and partially fill the 3p subshell. Thus, the electron configuration of S is



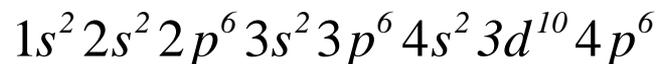
Example 7.13 ₃

(b) Strategy

We use the same approach as that in (a). What does it mean to say that Pd is a diamagnetic element?

Solution

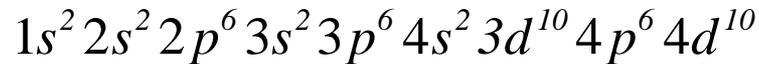
Palladium has 46 electrons. The noble-gas core in this case is [Kr]. (Kr is the noble gas in the period preceding palladium.) [Kr] represents



The remaining 10 electrons are distributed among the 4*d* and 5*s* orbitals. The three choices are (1) 4*d*¹⁰, (2) 4*d*⁹5*s*¹, and (3) 4*d*⁸5*s*².

Example 7.13 ₄

Because palladium is diamagnetic, all the electrons are paired and its electron configuration must be



or simply $[\text{Kr}]4d^{10}$. The configurations in (2) and (3) both represent paramagnetic elements.

Check

To confirm the answer, write the orbital diagrams for (1), (2), and (3).