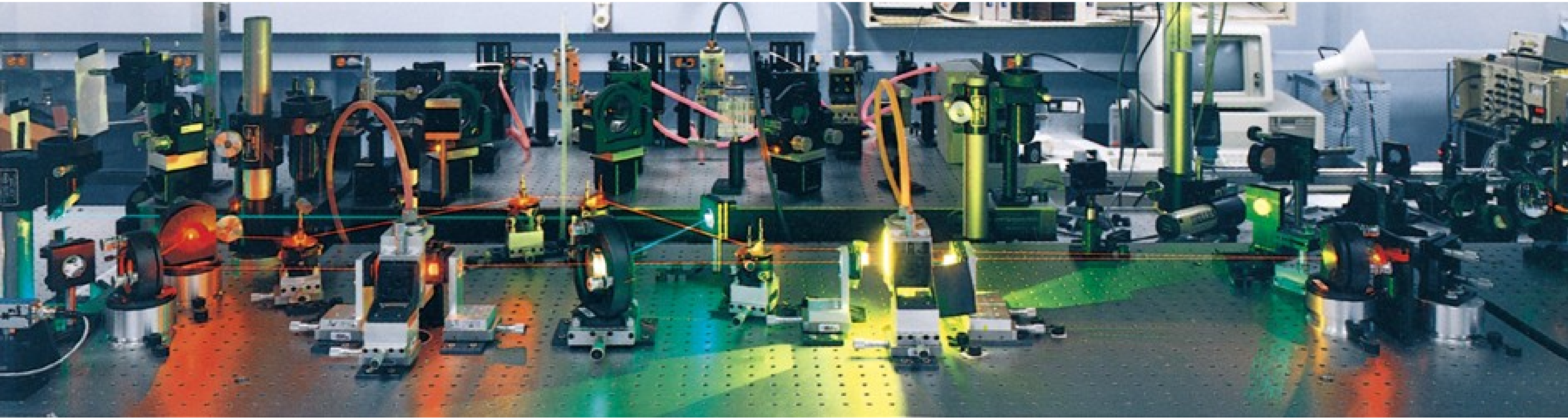
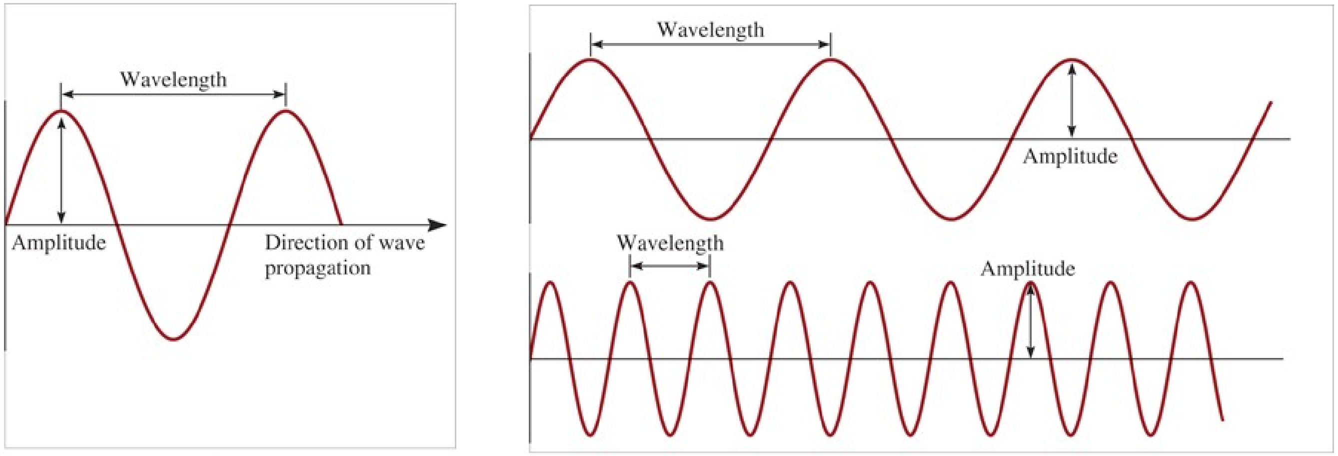


Quantum Theory and the Electronic Structure of Atoms

Chapter 7

Properties of Waves



(a) (b)

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

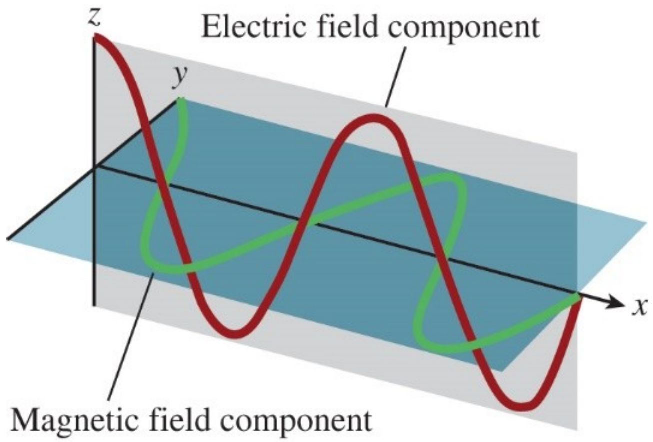
Amplitude is the vertical distance from the midline of a wave to the peak or trough.

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

The speed (u) 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 =

All electromagnetic radiation

$\lambda \times \nu = c$

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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 nm = 1×10^{-9} m (see Table 1.3). We write

$$\lambda = 522 \cancel{\text{nm}} \times \frac{1 \times 10^{-9}\text{ m}}{1 \cancel{\text{nm}}} = 522 \times 10^{-9}\text{ m}$$

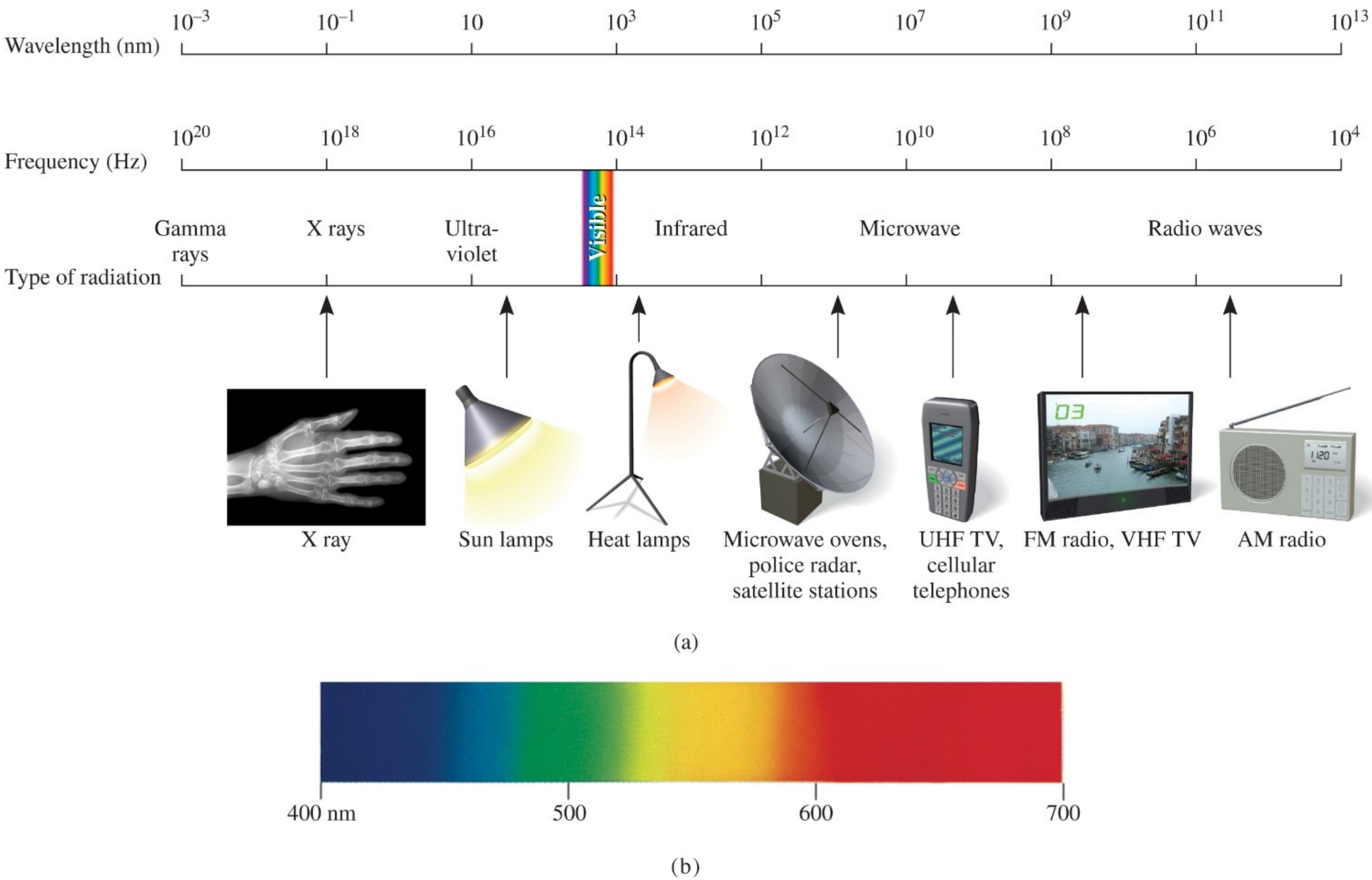
$$= 5.22 \times 10^{-7}\text{ m}$$

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

$$\nu = \frac{3.00 \times 10^{8}\text{ m/s}}{5.22 \times 10^{-7}\text{ m}}$$

$$= 5.75 \times 10^{14}\text{/s, or } 5.75 \times 10^{14}\text{ Hz}$$

Electromagnetic Spectrum



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a: (X ray): Ted Kinsman/Science Source

"Photoelectric Effect" Solved by Einstein in 1905

Light has both:

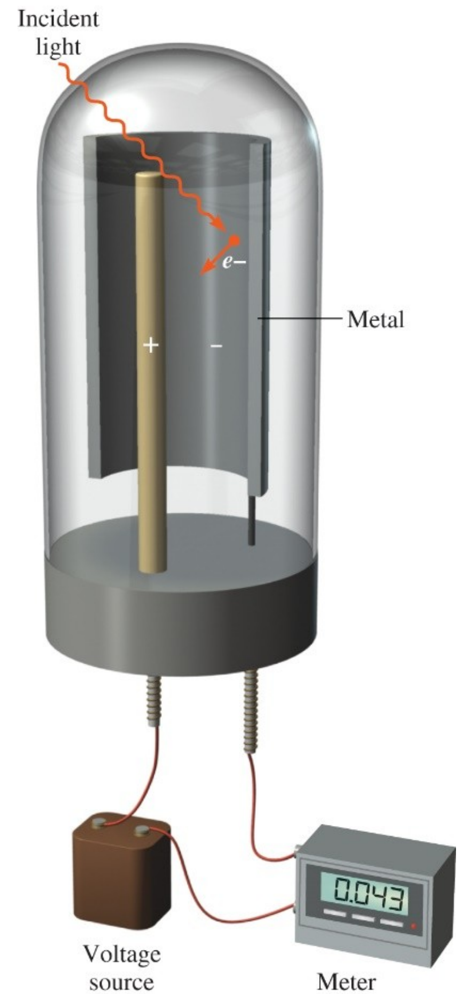
1. **wave nature**
2. **particle nature**

Photon is a "particle" of light

$$h\text{v} = \text{KE} + \text{W}$$

$$\text{KE} = h\text{v} - \text{W}$$

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



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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),

$$E = h\frac{c}{\lambda}$$

$$= \frac{(6.63 \times 10^{-34}\ \text{J} \cdot \text{s})(3.00 \times 10^8\ \text{m/s})}{(5.00 \times 10^4\ \text{nm})\dfrac{1 \times 10^{-9}\ \text{m}}{1\ \text{nm}}}$$

$$= 3.98 \times 10^{-21}\ \text{J}$$

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 $1.00 \times 10^{15}\ s^{-1}$ is used for irradiating the metal.

Solution (a) Setting KE = 0 in Equation (7.4), we write

$$h\nu = W$$

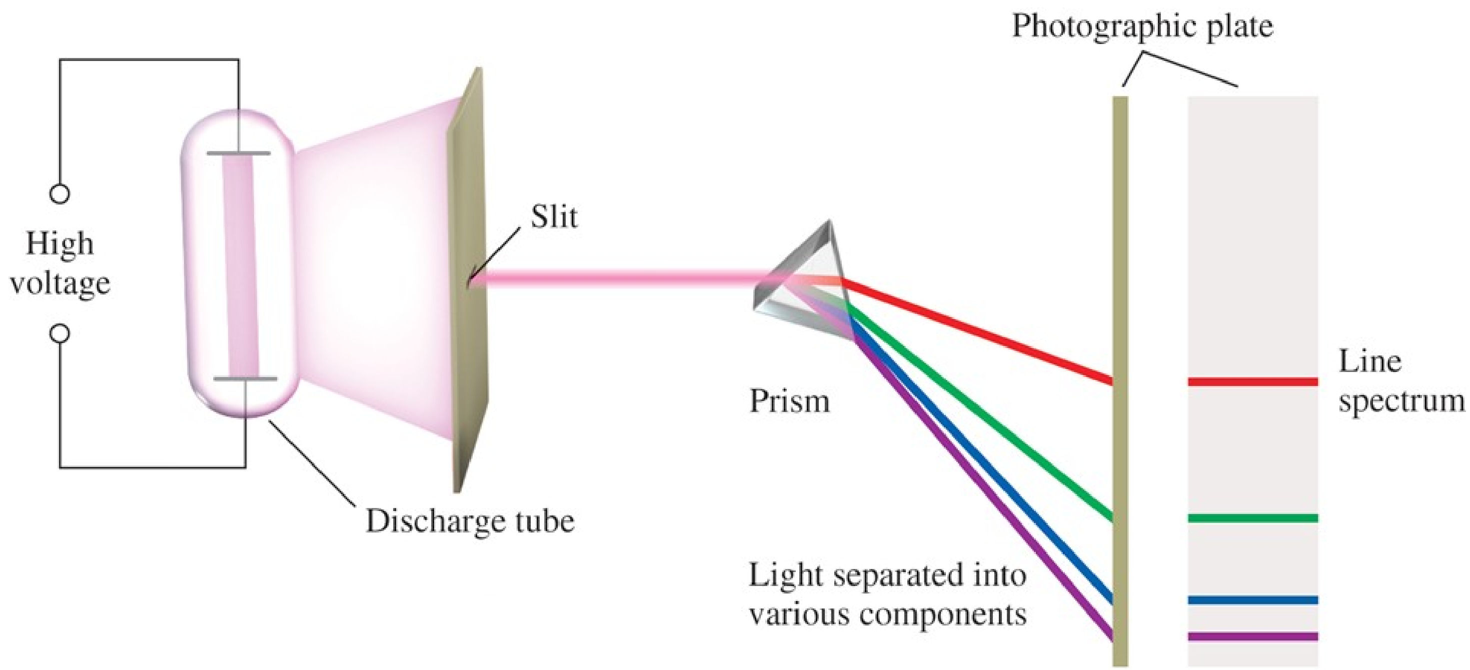
Thus,

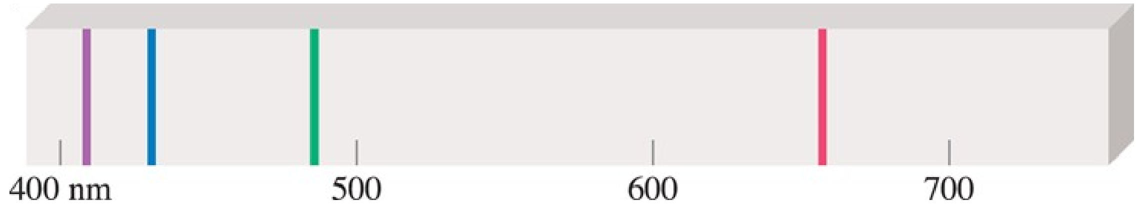
$$\nu = \frac{W}{h} = \frac{3.42 \times 10^{-19}\ \text{J}}{6.63 \times 10^{-34}\ \text{J} \cdot \text{s}}$$
$$= 5.16 \times 10^{14}\ \text{s}^{-1}$$

(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





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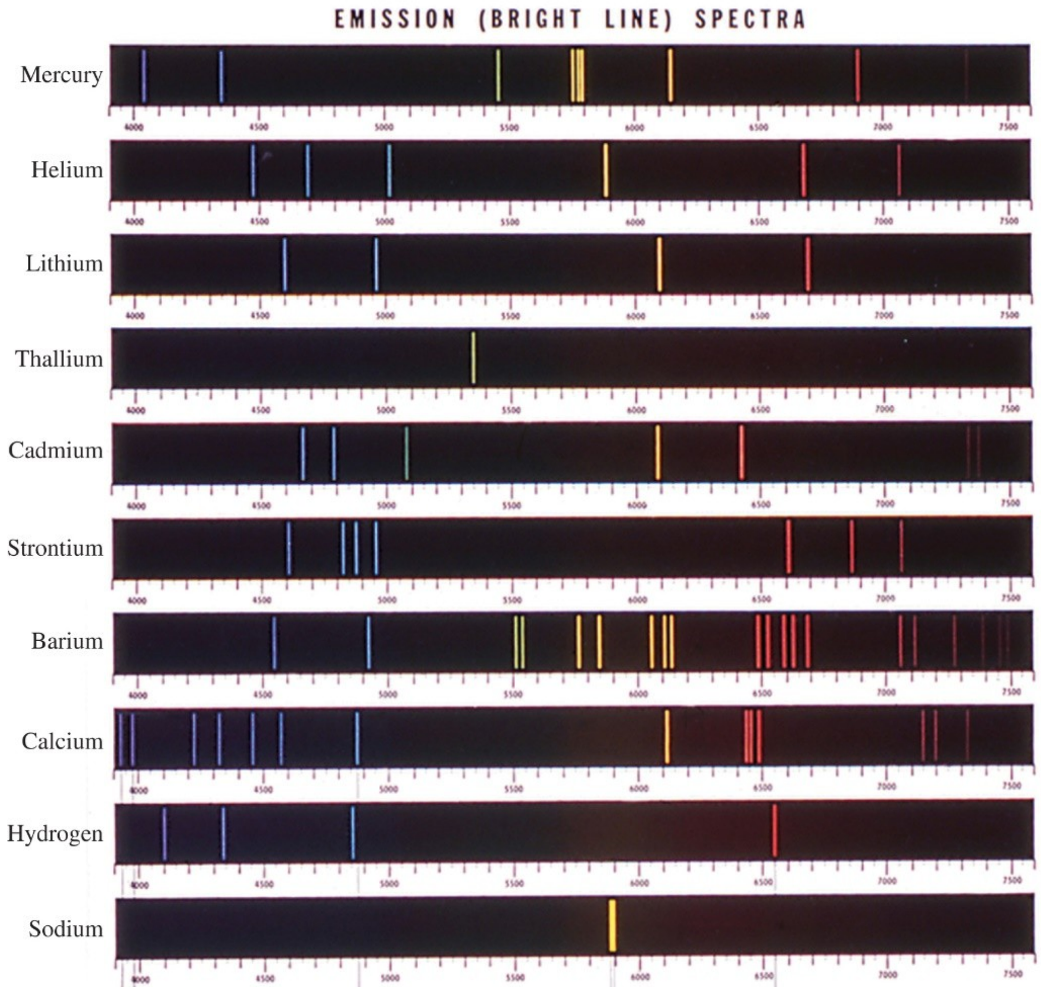
Emission Spectra of Some Elements



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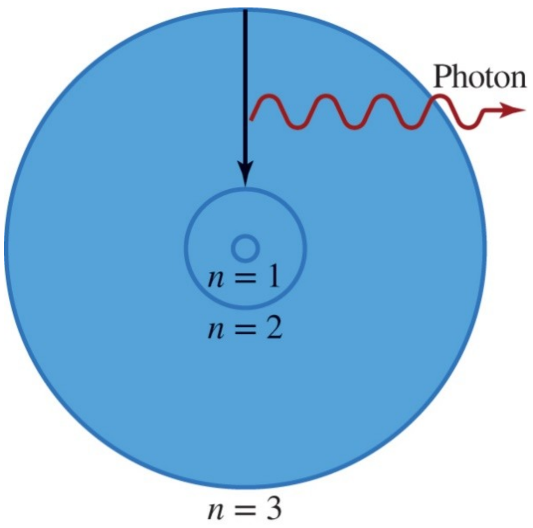
Courtesy of Wabash Instrument Corporation

Bohr's Model of the Atom (1913)

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



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

n (principle quantum number) =

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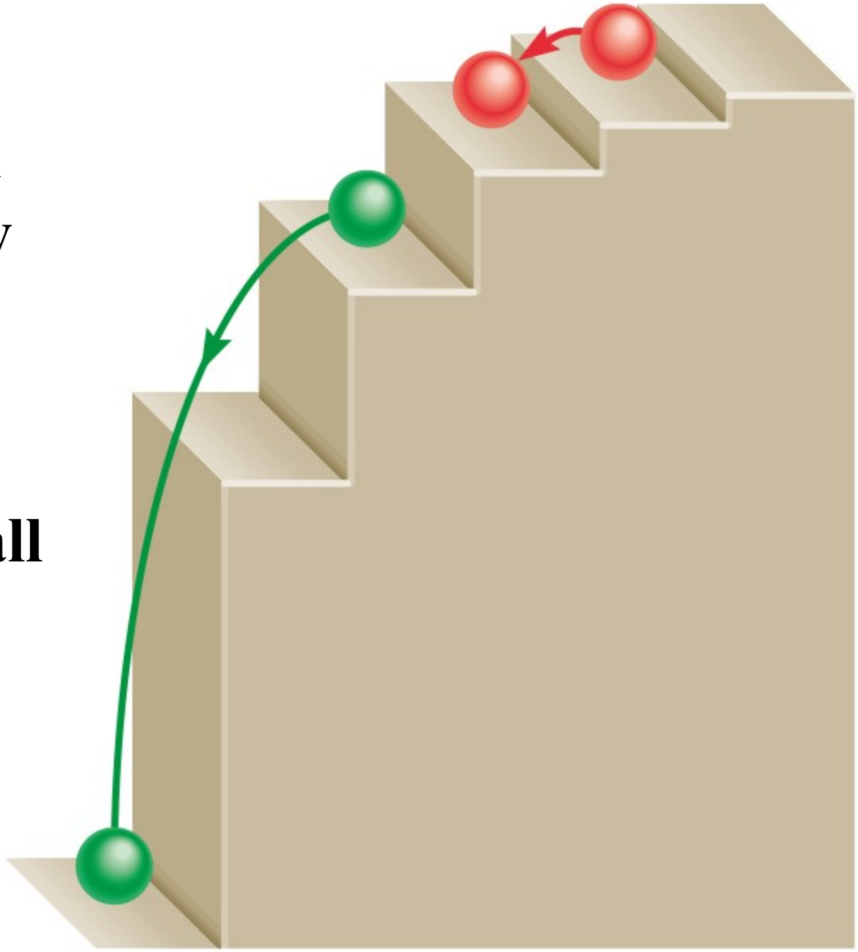
Quantized Energy

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.



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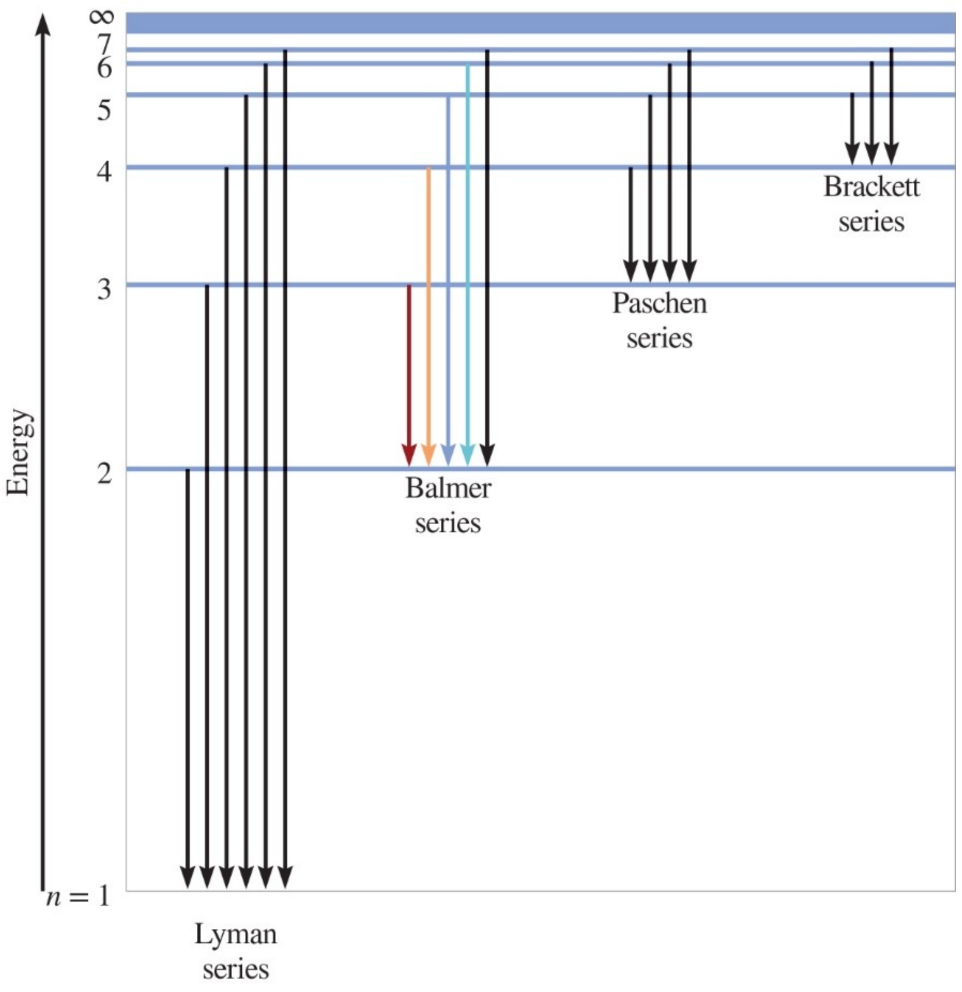
Energy Transitions of the Hydrogen Atom



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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

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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..)

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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

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

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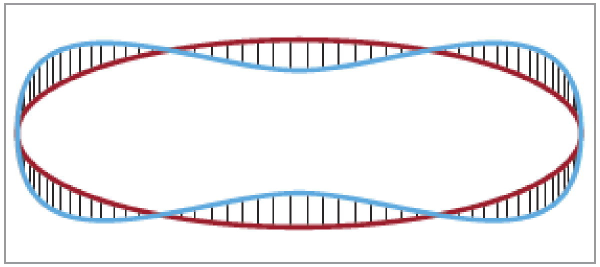
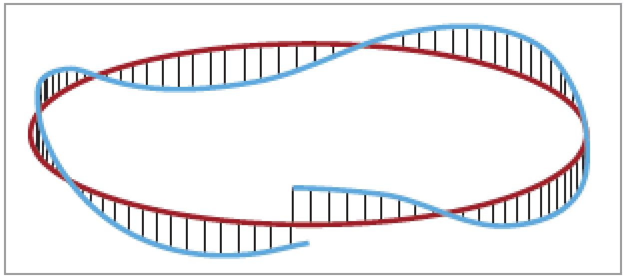
Quantization of Electron Energy

- Why is e^- • energy quantized?



(a)



(b)

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

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 J · 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,

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

Comment This wavelength (1.1×10^{-5} m or 1.1×10^{4} 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 = 1.674×10^{-27} kg) moving at 7.00×10^{2} cm/s.

Schrodinger Wave Equation

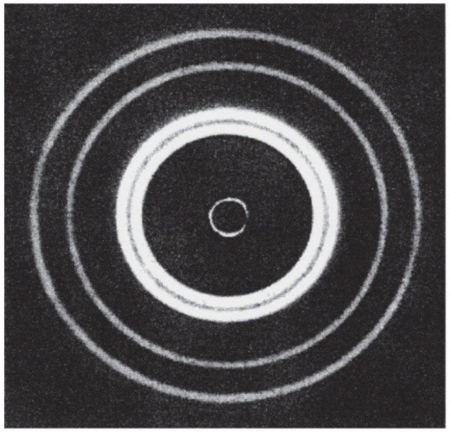
In 1926 Schrodinger wrote an **equation that described both the particle and wave nature of the** e^- Wave function (ψ) describes:

1. energy of e^- with a 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.



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Schrodinger Wave Equation

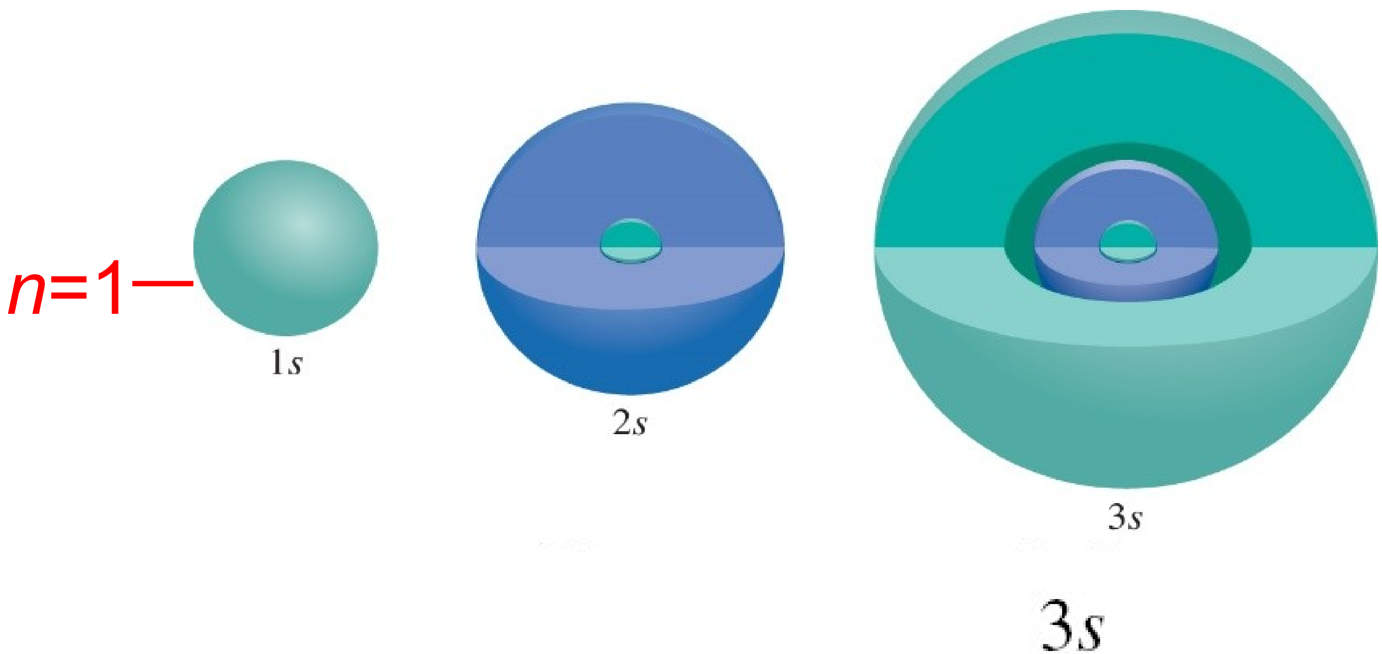
ψ is a function of four numbers called

quantum numbers (n, l, m_l, m_s)

principal quantum number n

n = 1, 2, 3, 4, ….

distance of e^- from the nucleus

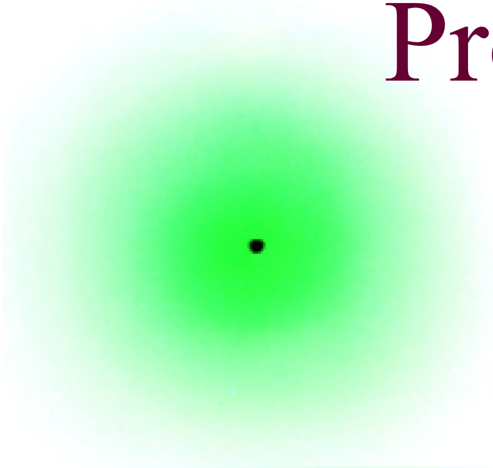


Principal Quantum Number (*n*)

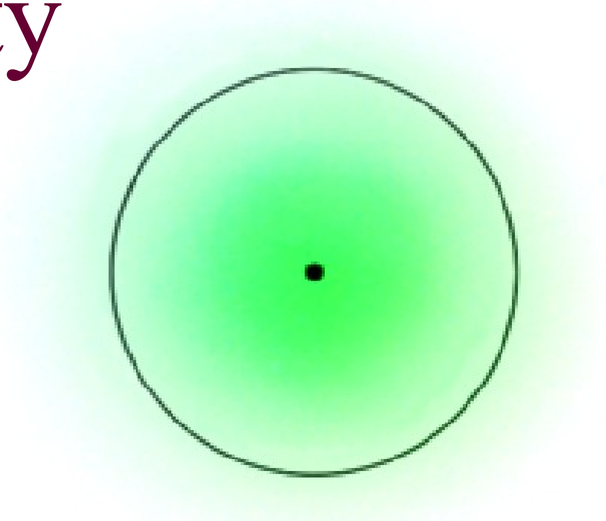
- The principal quantum number, *n*, describes the energy level on which the orbital resides.
- The values of *n* are integers ≥ 1.
  - Allowed values: n = 1, 2, 3, 4, etc
  - describes the energy of the electron
    - as n increases, the energy of the electron increases
    - as n increases, the electron is more loosely bound to the atom
  - indicates the <u>average</u> distance of the electron from the nucleus

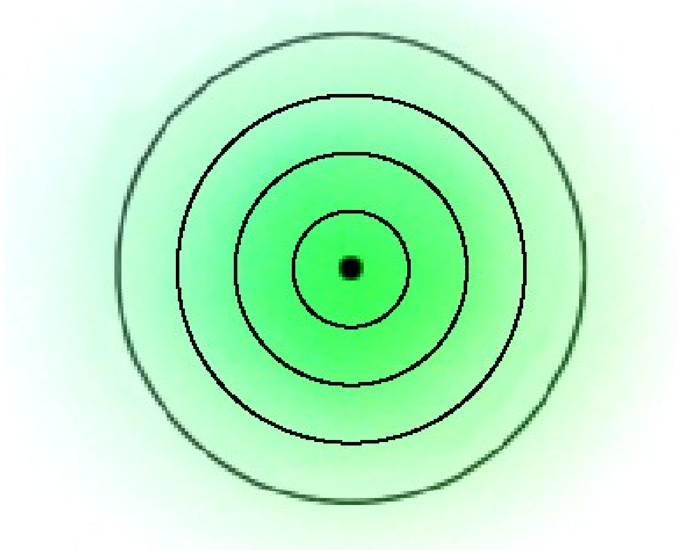
    as n increases, the average distance from the nucleus increases

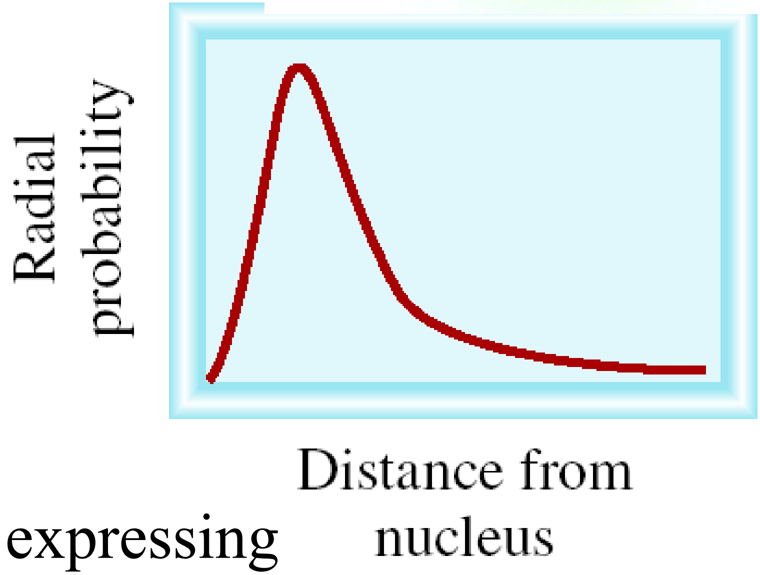
Probability Density



Where 90% of the e^- density is found – for the 1s orbital







Electron density is another way of expressing probability. A region of high electron density is one where there is a high probability of finding an electron.

Schrodinger Wave Equation

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

angular momentum quantum number l

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

$n = 1$, $l = 0$
$n = 2$, $l = 0$ **or** 1
$n = 3$, $l = 0, 1,$ **or** 2

$l = 0$ *s* orbital
$l = 1$ *p* orbital
$l = 2$ *d* orbital
$l = 3$ *f* orbital

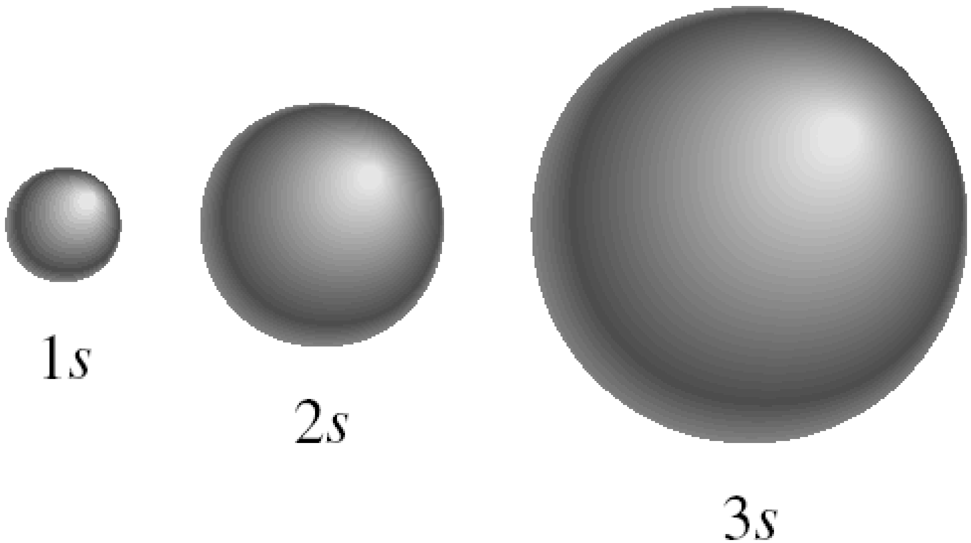
Shape of the "volume" of space that the e^- occupies

Angular Momentum Quantum Number (*l*)

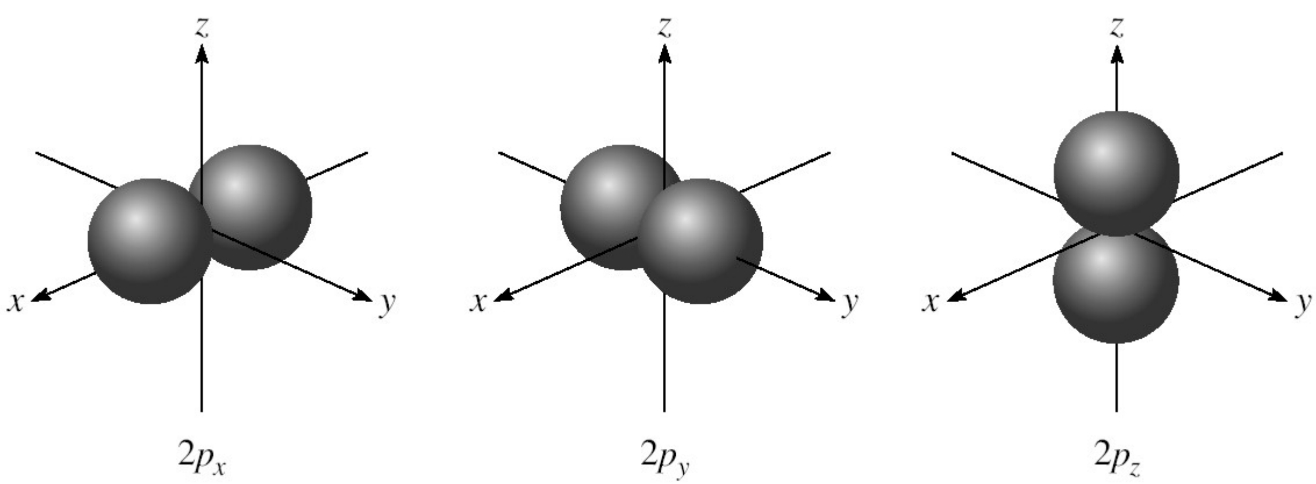
- This quantum number defines the shape of the orbital.
- Allowed values of *l* are integers ranging from 0 to $n - 1$.
- We use letter designations to communicate the different values of *l* and, therefore, the shapes and types of orbitals.
  - Allowed values: l = 0, 1, 2,….(n-1)
    - Example: If n = 2, then l = 0 or 1
  - defines the shape of the orbital

Value of *l*	0	1	2	3
Type of orbital	*s*	*p*	*d*	*f*

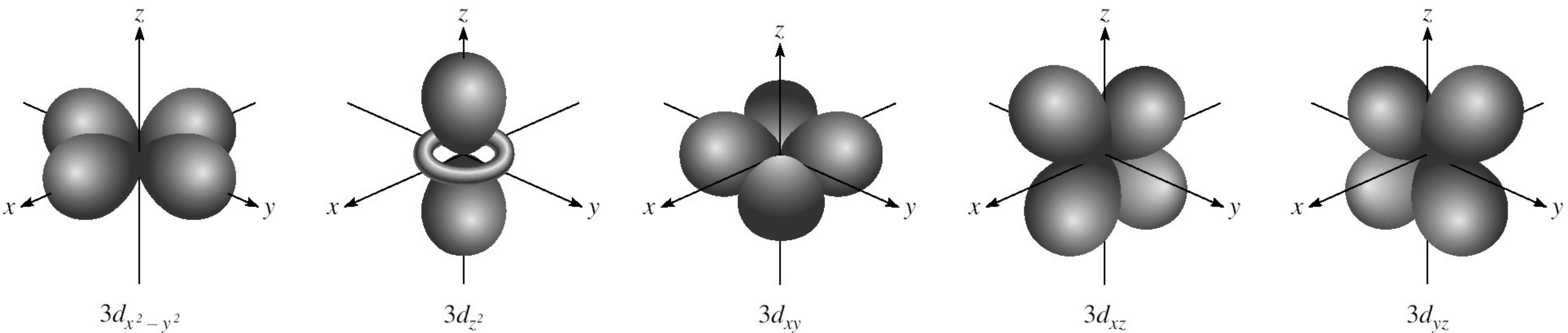
l = 0 (s orbitals)



l = 1 (p orbitals)



l = 2 (*d* orbitals)



Schrodinger Wave Equation

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

magnetic quantum number m_l

for a given value of l

$m_l = -l, \ldots., 0, \ldots. +l$

if $l = 1$ (p orbital), $m_l = -1, 0,$ **or** 1

if $l = 2$ (d orbital), $m_l = -2, -1, 0, 1,$ **or** 2

orientation of the orbital in space

Magnetic Quantum Number (m_l)

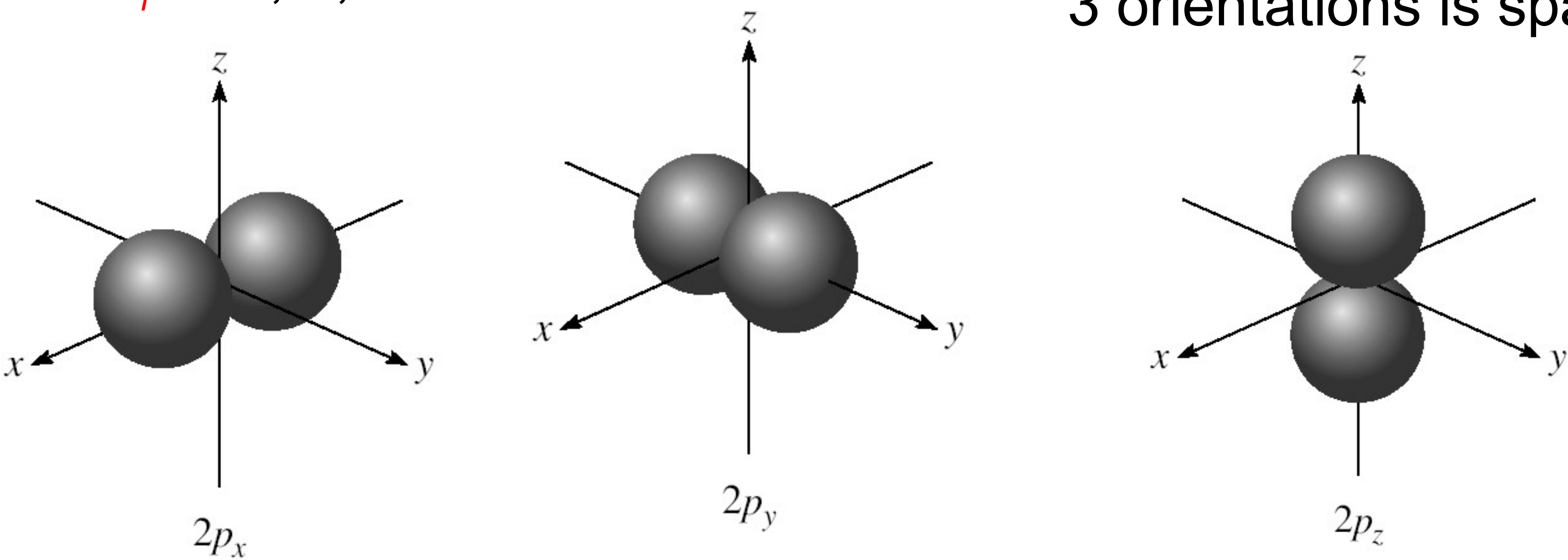
- The magnetic quantum number describes the three-dimensional orientation of the orbital.
- Allowed values of m_l are integers ranging from -l to l :

$$-l \leq m_l \leq l.$$

- Therefore, on any given energy level, there can be up to 1 *s* orbital, 3 *p* orbitals, 5 *d* orbitals, 7 *f* orbitals, etc.
  - Allowed values: integers from l to –l
    - If l = 1, then m_l = 1, 0, -1
  - describes the orientation in space of the orbital

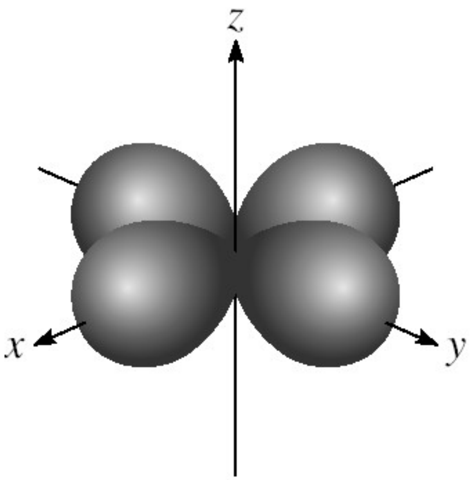
m_l = -1, 0, **or** 1

3 orientations is space

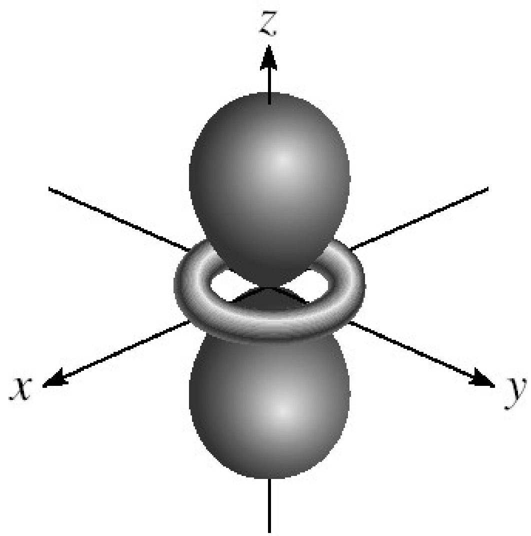


n	Possible Values of l	Subshell Designation	Possible Values of m_l	Number of Orbitals in Subshell	Total Number of Orbitals in Shell
1	0	1*s*	0	1	1
2	0	2*s*	0	1	
	1	2*p*	1, 0, −1	3	4
3	0	3*s*	0	1	
	1	3*p*	1, 0, −1	3	
	2	3*d*	2, 1, 0, −1, −2	5	9
4	0	4*s*	0	1	
	1	4*p*	1, 0, −1	3	
	2	4*d*	2, 1, 0, −1, −2	5	
	3	4*f*	3, 2, 1, 0, −1, −2, −3	7	16

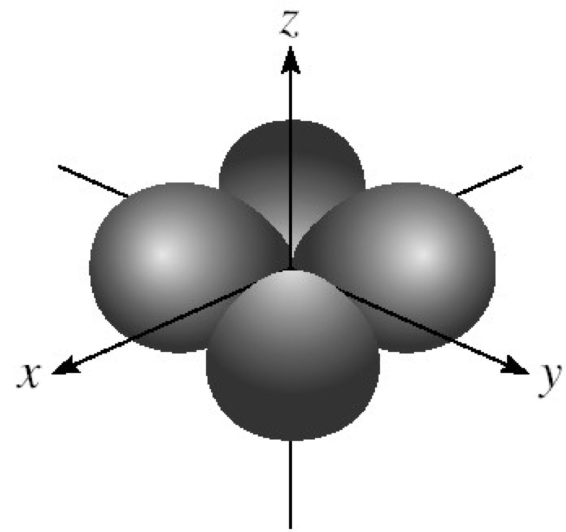
m_l = -2, -1, 0, 1, **or** 2 5 orientations is space



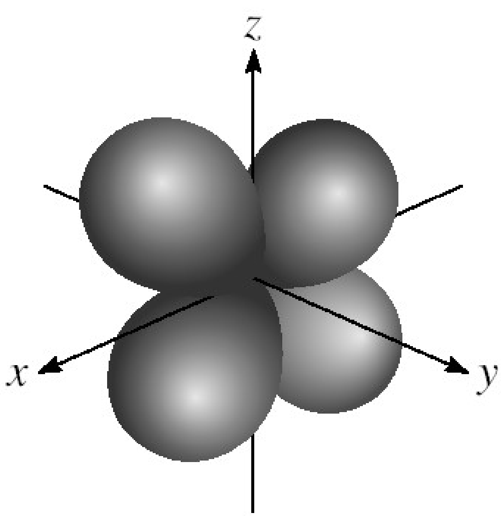
$3d_{x^2-y^2}$



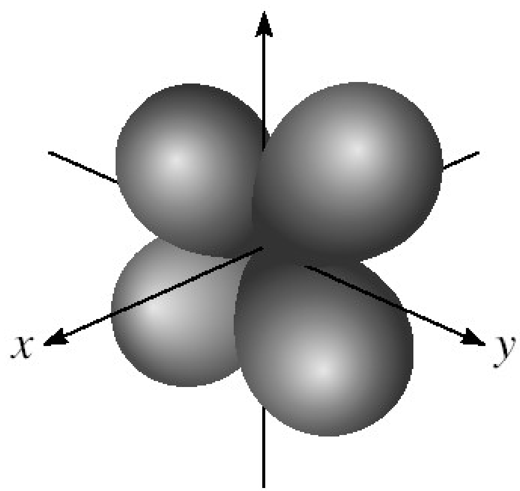
$3d_{z^2}$



$3d_{xy}$



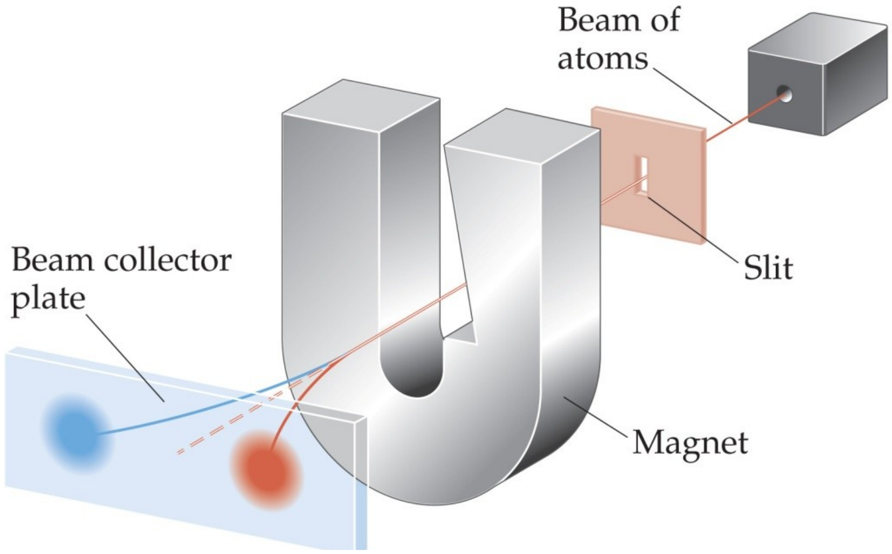
$3d_{xz}$

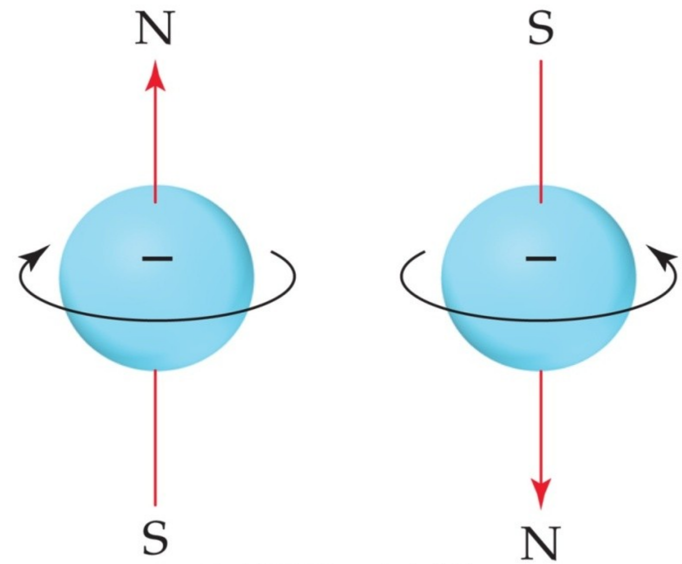


$3d_{yz}$

Spin Quantum Number, m_s

- In the 1920s, it was discovered that two electrons in the same orbital do not have exactly the same energy.
- The “spin” of an electron describes its magnetic field, which affects its energy.





The spin quantum number has only 2 allowed values: +1/2 and −1/2.

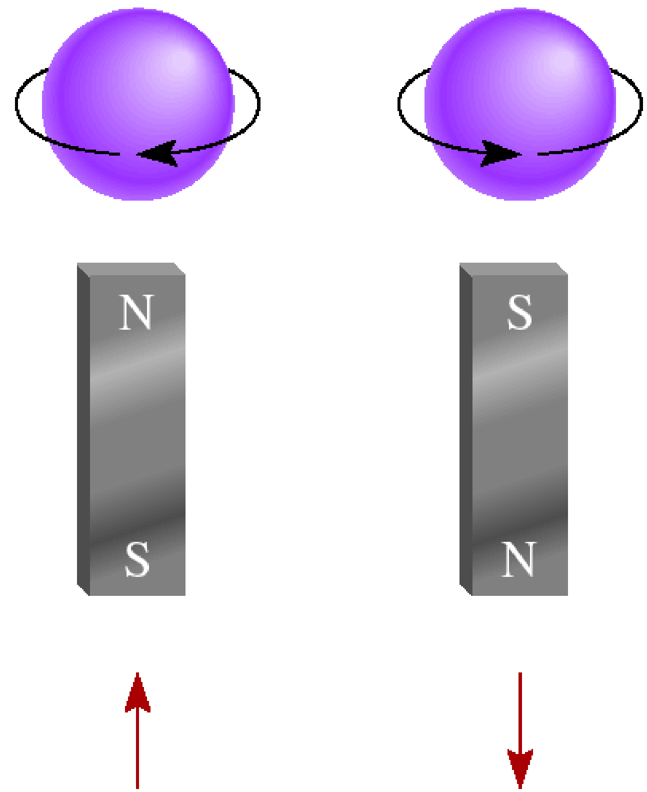
Electron spin: a property of electrons that make it behave as if it were a tiny magnet spinning on its axis

Schrodinger Wave Equation

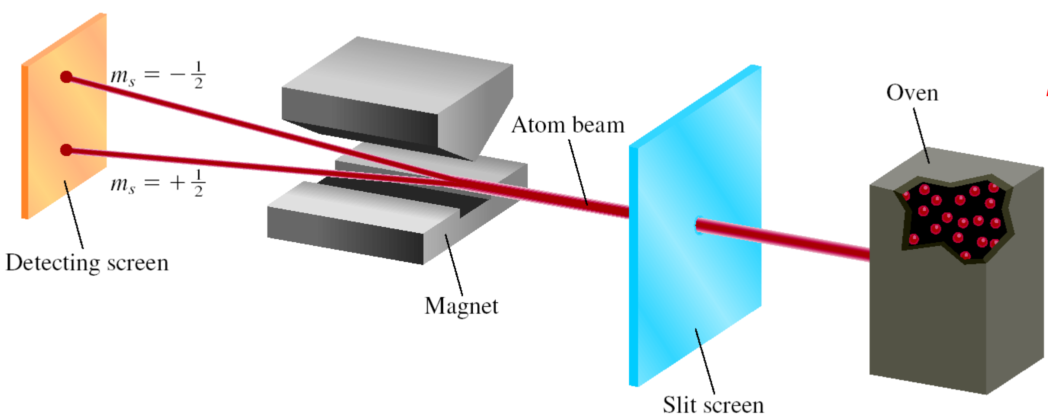
(n, l, m_l, m_s)

spin quantum number m_s

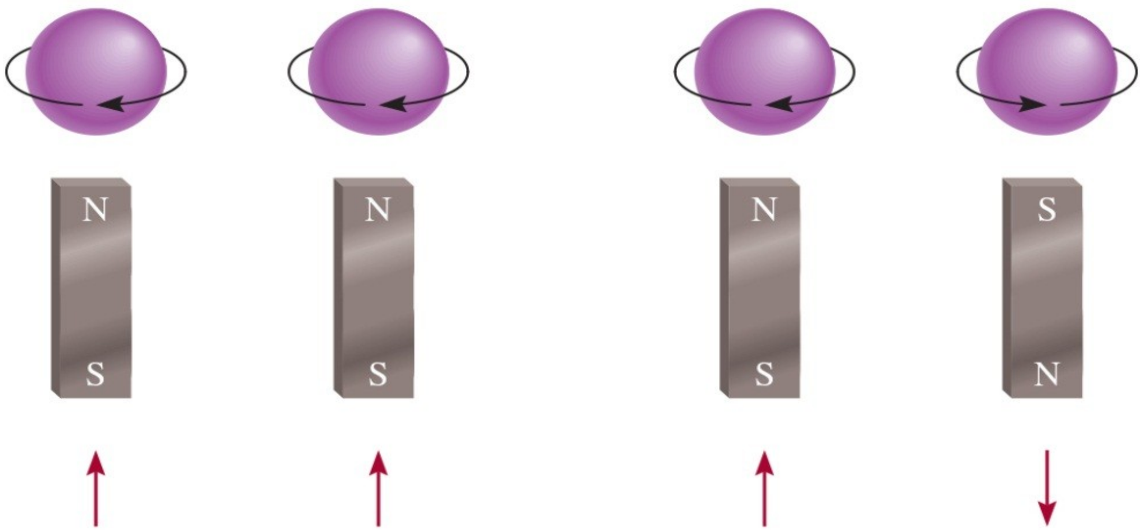
m_s = +½ **or** -½



m_s = +½ m_s = -½

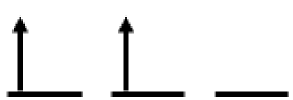


Paramagnetism and Diamagnetism



Paramagnetic

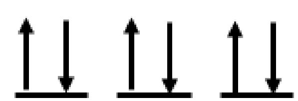
unpaired electrons



2p

Diamagnetic

all electrons paired



2p

Schrodinger Wave Equation

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.

No two electrons in the same atom can have exactly the same energy.

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	$3d_{xy}, 3d_{yz}, 3d_{xz}, 3d_{x^2-y^2}, 3d_{z^2}$
⋮	⋮	⋮	⋮	⋮

Example 7.7 (1 of 2)

What are the allowed values of $\boldsymbol{m_l}$ when $n = 4$ and $\ell = 1$?

Solution

The values of $\boldsymbol{m_l}$ depend only on the value of l, not on the value of n. The values of $\boldsymbol{m_l}$ can vary from $-l$ to l. Therefore, $\boldsymbol{m_l}$ can be −1, 0, or 1.

Check

The values of n and l are fixed, but $\boldsymbol{m_l}$ can have any one of the three values, which correspond to the three *p* orbitals.

Schrodinger Wave Equation

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, **and** m_l

How many electrons can an orbital hold?

If n, l, and m_l are fixed, then $m_s = ½$ or $- ½$

$\psi = (n, l, m_l, ½)$ or $\psi = (n, l, m_l, -½)$

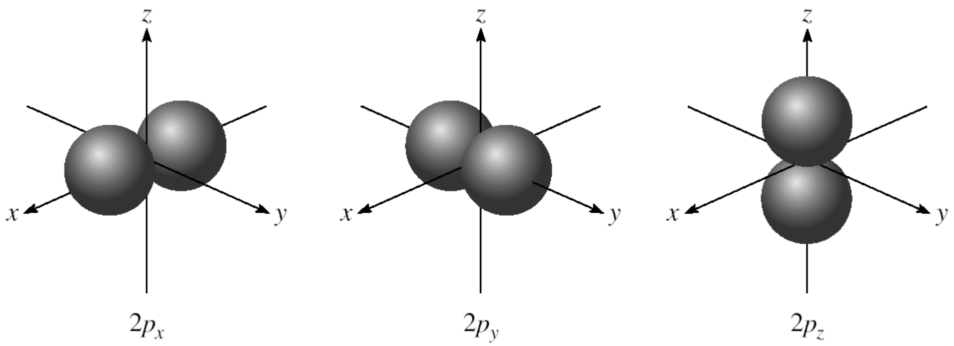
An orbital can hold 2 electrons

How many 2*p* orbitals are there in an atom?

n =2 → 2*p* ← *l* = 1

If l = 1, then m_l = -1, 0, or +1

3 orbitals



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

n =3 → 3*d* ← *l* = 2

If l = 2, then m_l = -2, -1, 0, +1, or +2

5 orbitals which can hold a total of 10 e^-

Example 7.8 (1 of 2)

List the values of n, **l** and m_l for orbitals in the 4d subshell.

Solution

As we saw earlier, the number given in the designation of the subshell is the principal quantum number, so in this case $n = 4$. The letter designates the type of orbital. Because we are dealing with d orbitals, $l = 2$. The values of $\boldsymbol{m_l}$ can vary from $-l$ to l. Therefore, $\boldsymbol{m_l}$ can be −2, −1, 0, 1, or 2.

Check

The values of n and l are fixed for 4d, but $\boldsymbol{m_l}$ can have any one of the five values, which correspond to the five d orbitals.

Example 7.9 (1 of 2)

What is the total number of orbitals associated with the principal quantum number $n = 3$?

Solution

For $n = 3$, the possible values of l are 0, 1, and 2. Thus, there is one 3*s* orbital ($n = 3$, $l = 0$, and $m_l = 0$); there are three 3*p* orbitals ($n = 3$, $l = 1$, and $m_l = -1, 0, 1$); there are five 3*d* orbitals ($n = 3$, $l = 2$, and $m_l = -2, -1, 0, 1, 2$). The total number of orbitals is $1 + 3 + 5 = 9$.

Check

The total number of orbitals for a given value of n is So here we have Can you prove the validity of this relationship?

Example 7.10 (1 of 3)

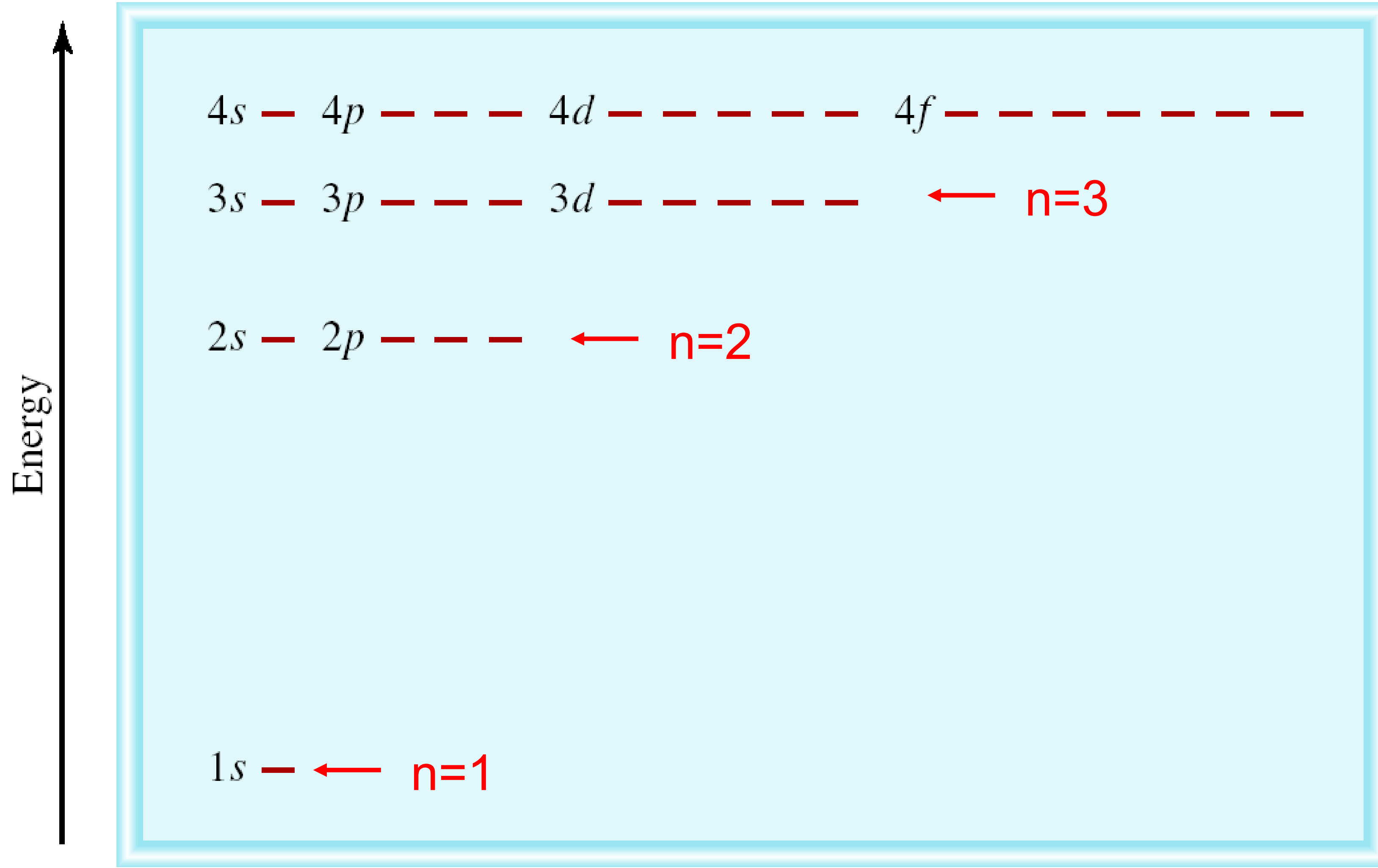
Write the four quantum numbers for an electron in a 3*p* orbital.

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 $l = 1$, there are three values of m_l given by −1, 0, and 1. Because the electron spin quantum number m_s can be either we conclude that there are six possible ways to designate the electron using the (n, l, m_l, m_s) notation.

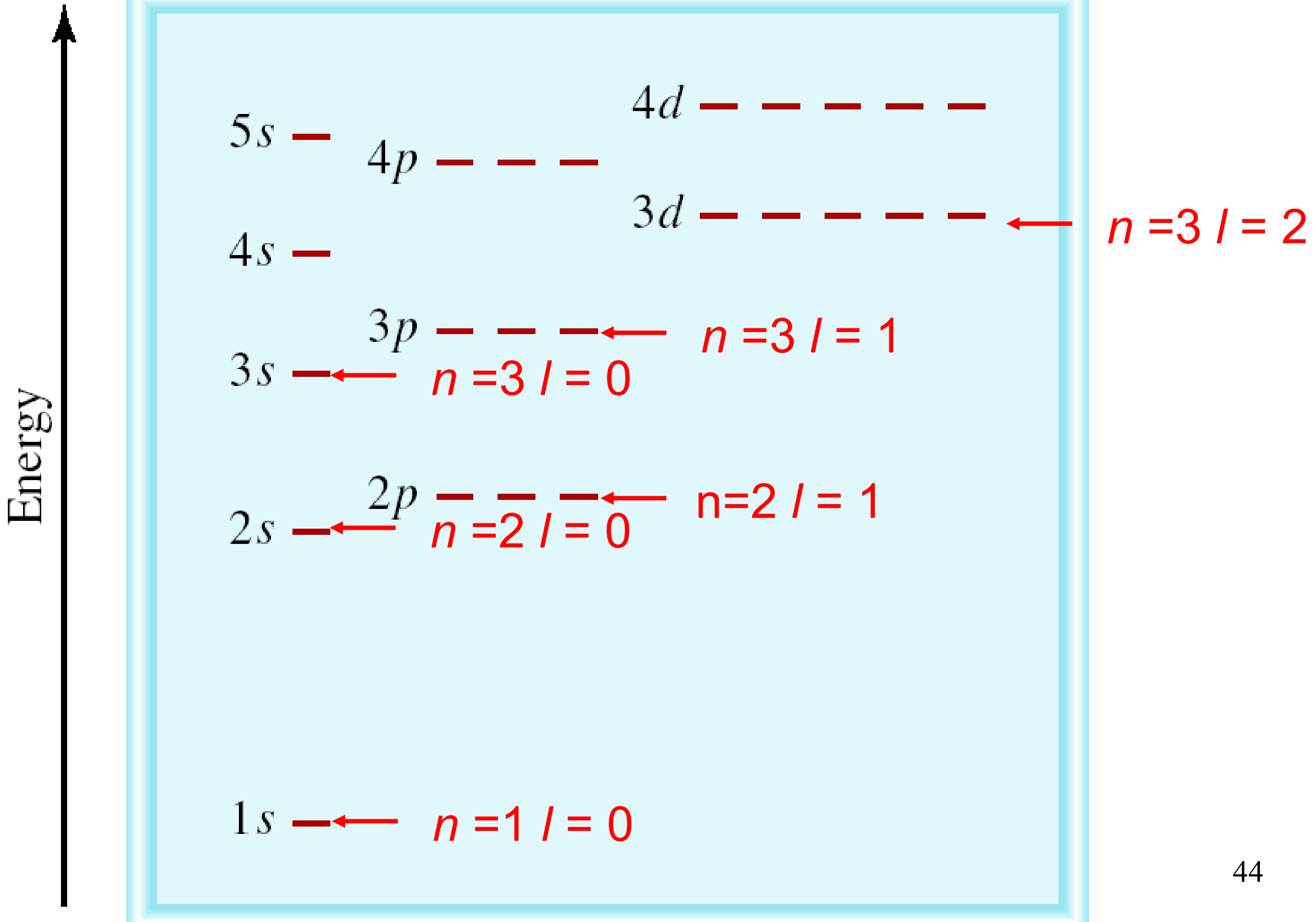
Energy of orbitals in a ***single*** electron atom

Energy only depends on principal quantum number ***n***

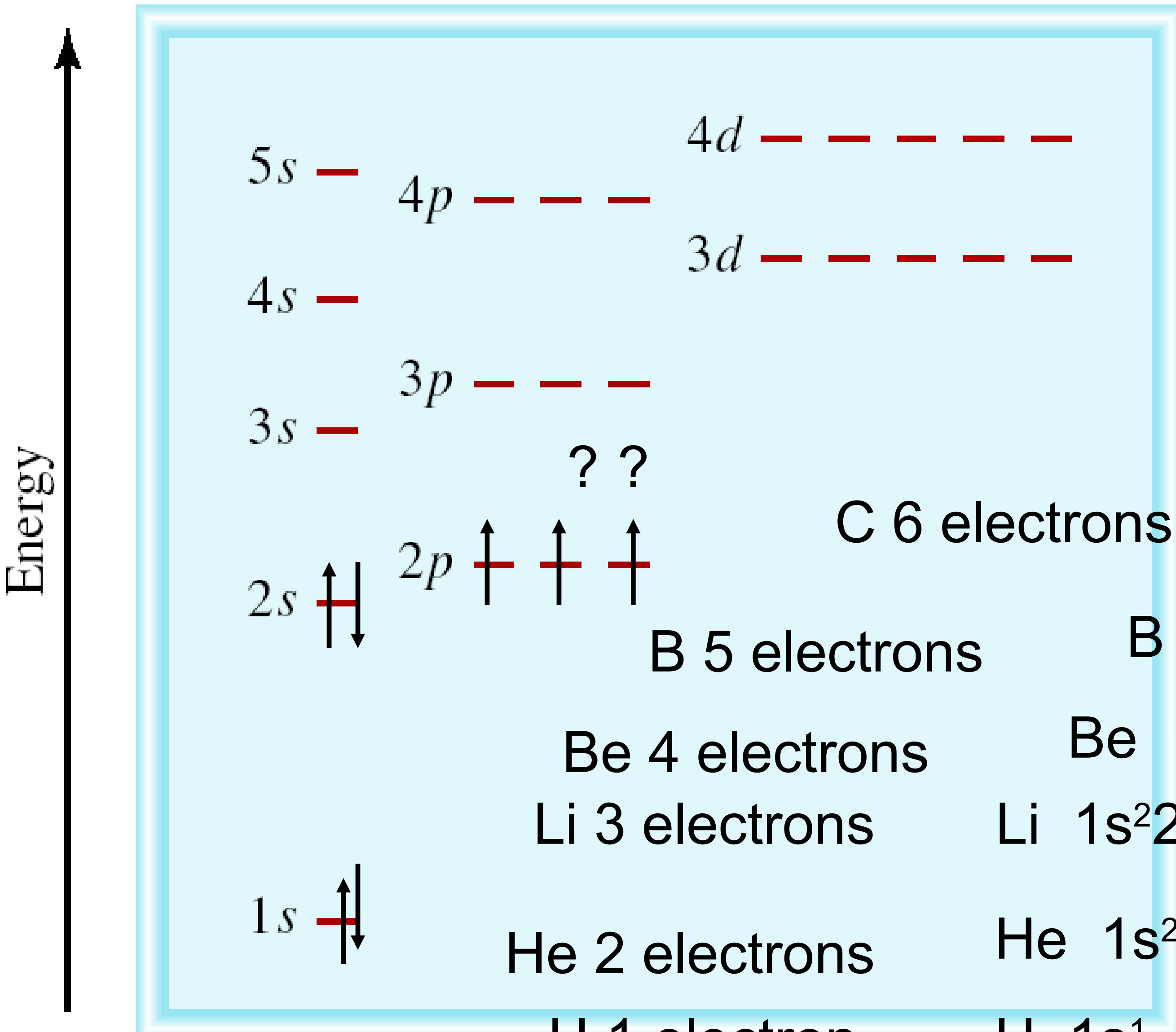


Energy of orbitals in a ***multi***-electron atom

Energy depends on ***n*** and ***l***



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



C 6 electrons

B 5 electrons — B $1s^2 2s^2 2p^1$

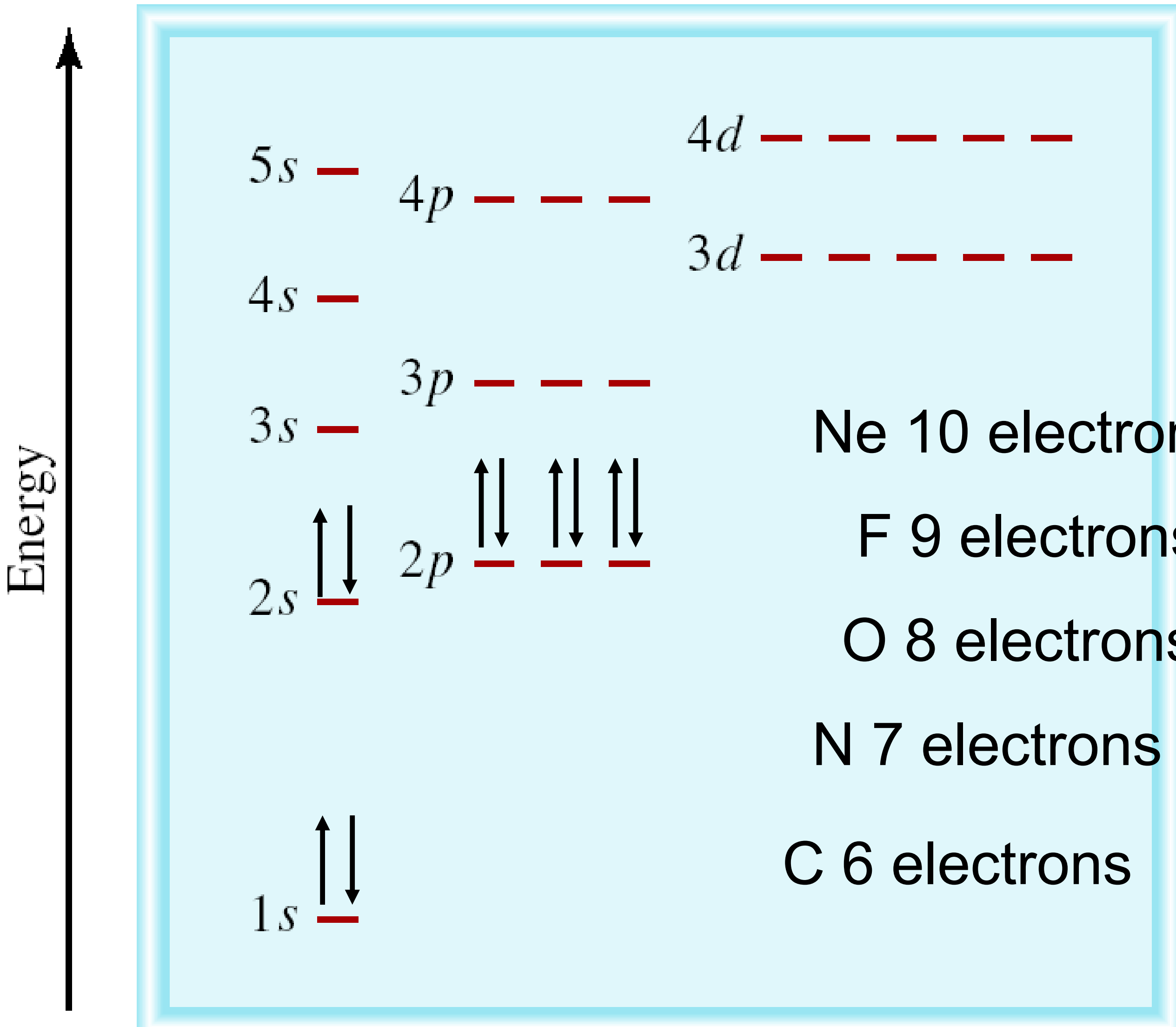
Be 4 electrons — Be $1s^2 2s^2$

Li 3 electrons — Li $1s^2 2s^1$

He 2 electrons — He $1s^2$

H 1 electron — H $1s^1$

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



Ne 10 electrons Ne $1s^22s^22p^6$

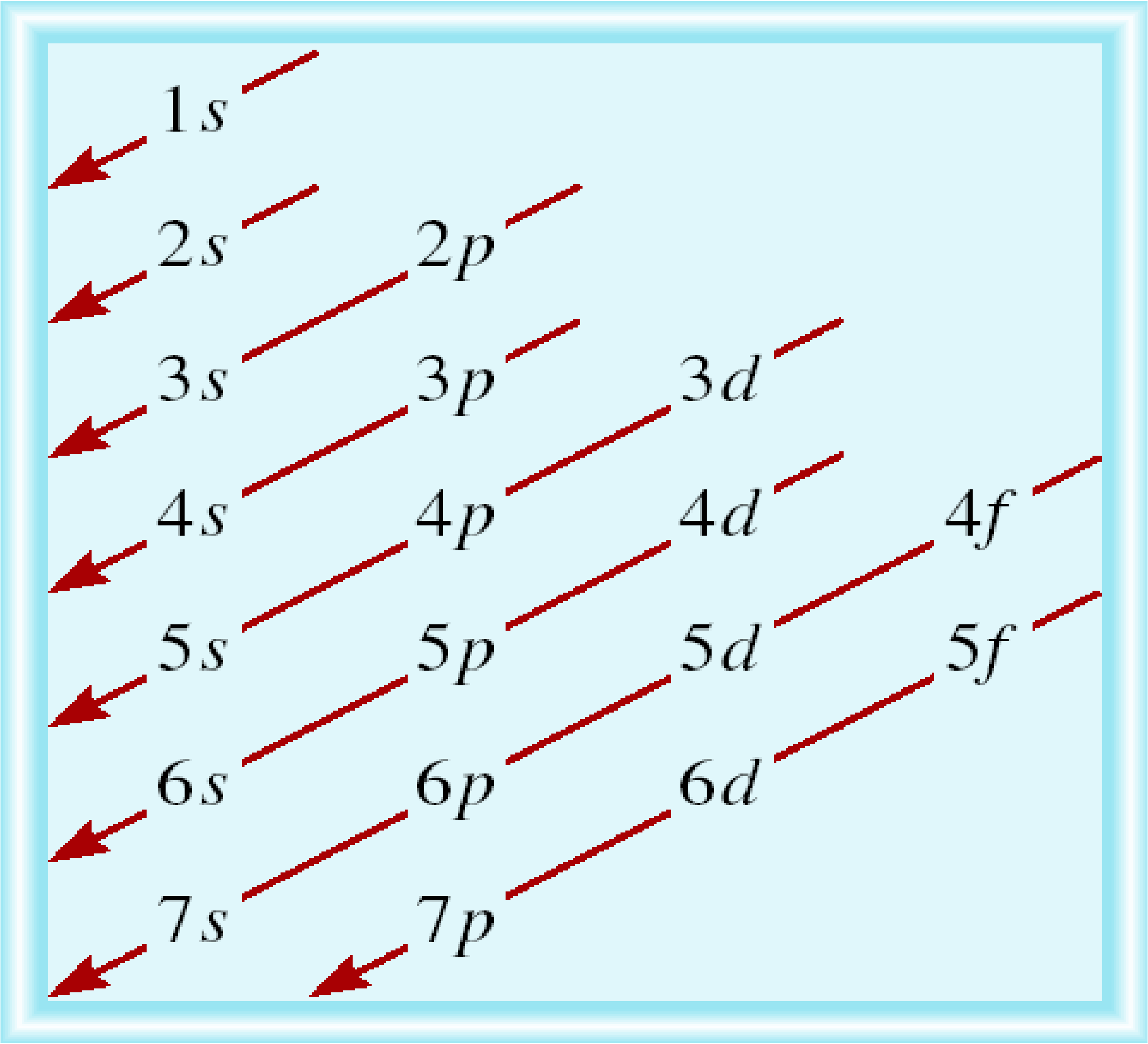
F 9 electrons F $1s^22s^22p^5$

O 8 electrons O $1s^22s^22p^4$

N 7 electrons N $1s^22s^22p^3$

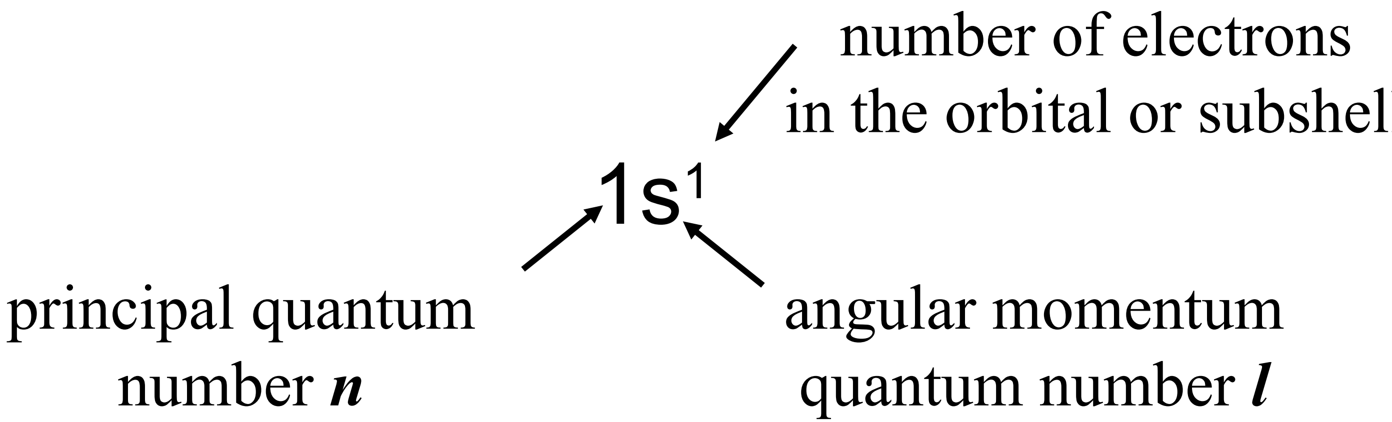
C 6 electrons C $1s^22s^22p^2$

Order of orbitals (filling) in multi-electron atom



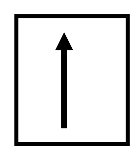
1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s

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



Orbital diagram

H



$1s^1$

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 $[Ne]3s^2$ $[Ne]\ 1s^2 2s^2 2p^6$

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

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

Electron Configurations

- 2nd row elements.

Electron Configurations

- 3rd row elements

Electron Configurations

- 4th row elements

Outermost subshell being filled with electrons

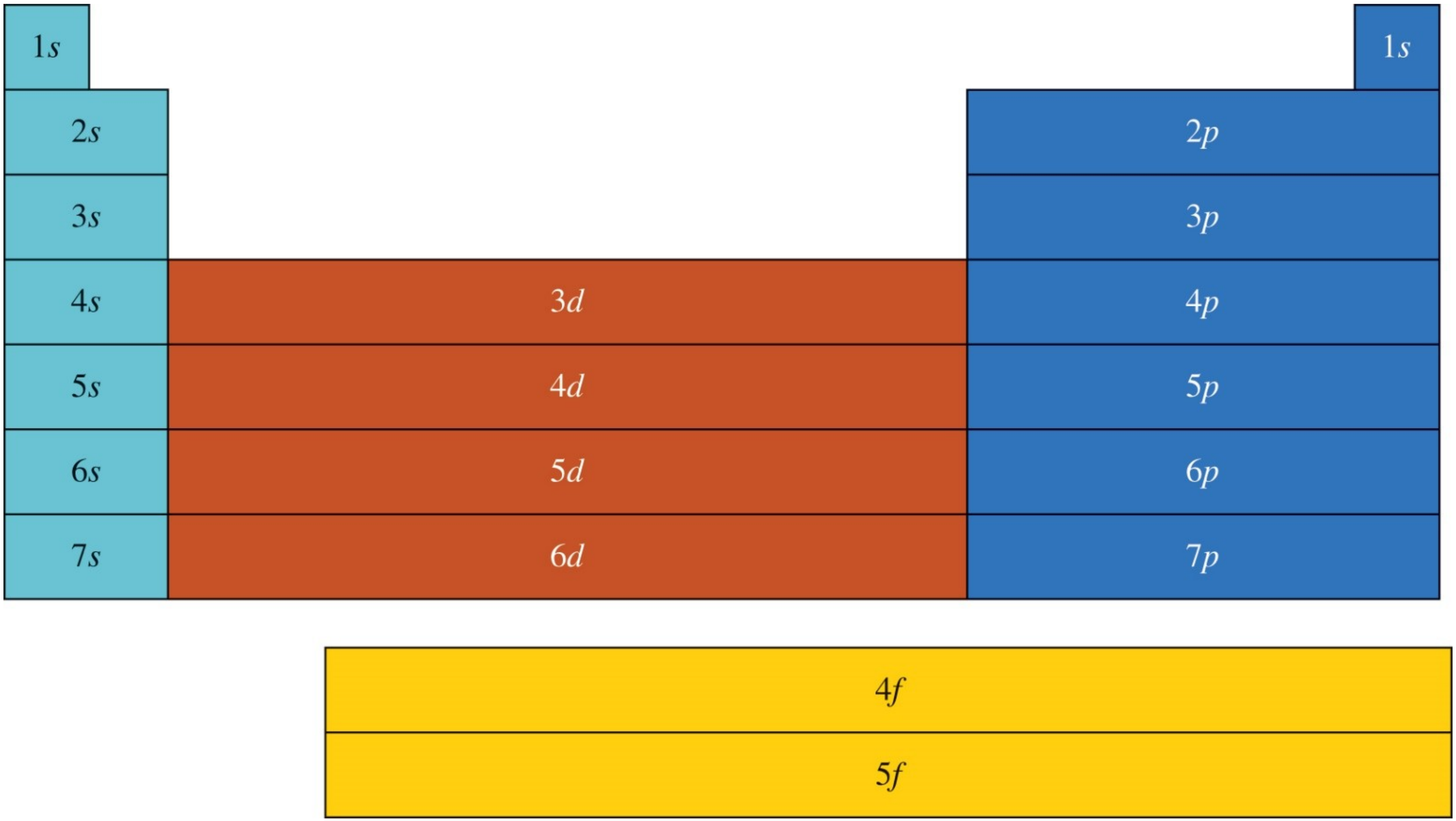


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TABLE 7.3 The Ground-State Electron Configurations of the Elements*

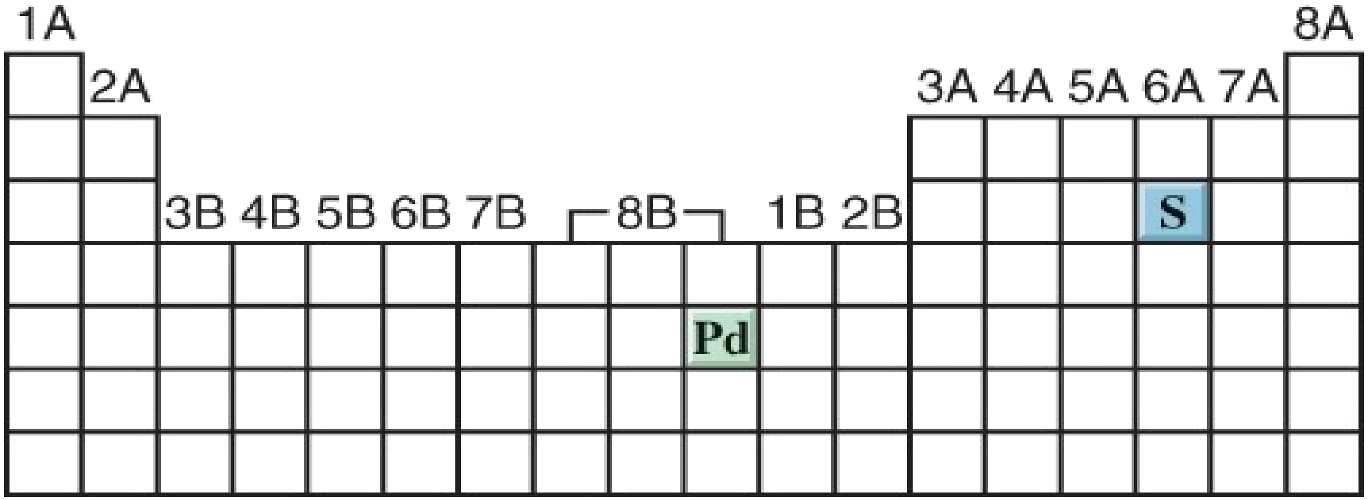
Atomic Number	Symbol	Electron Configuration	Atomic Number	Symbol	Electron Configuration	Atomic Number	Symbol	Electron Configuration
1	H	$1s^1$	38	Sr	$[Kr]5s^2$	75	Re	$[Xe]6s^24f^{14}5d^5$
2	He	$1s^2$	39	Y	$[Kr]5s^24d^1$	76	Os	$[Xe]6s^24f^{14}5d^6$
3	Li	$[He]2s^1$	40	Zr	$[Kr]5s^24d^2$	77	Ir	$[Xe]6s^24f^{14}5d^7$
4	Be	$[He]2s^2$	41	Nb	$[Kr]5s^14d^4$	78	Pt	$[Xe]6s^14f^{14}5d^9$
5	B	$[He]2s^22p^1$	42	Mo	$[Kr]5s^14d^5$	79	Au	$[Xe]6s^14f^{14}5d^{10}$
6	C	$[He]2s^22p^2$	43	Tc	$[Kr]5s^24d^5$	80	Hg	$[Xe]6s^24f^{14}5d^{10}$
7	N	$[He]2s^22p^3$	44	Ru	$[Kr]5s^14d^7$	81	Tl	$[Xe]6s^24f^{14}5d^{10}6p^1$
8	O	$[He]2s^22p^4$	45	Rh	$[Kr]5s^14d^8$	82	Pb	$[Xe]6s^24f^{14}5d^{10}6p^2$
9	F	$[He]2s^22p^5$	46	Pd	$[Kr]4d^{10}$	83	Bi	$[Xe]6s^24f^{14}5d^{10}6p^3$
10	Ne	$[He]2s^22p^6$	47	Ag	$[Kr]5s^14d^{10}$	84	Po	$[Xe]6s^24f^{14}5d^{10}6p^4$
11	Na	$[Ne]3s^1$	48	Cd	$[Kr]5s^24d^{10}$	85	At	$[Xe]6s^24f^{14}5d^{10}6p^5$
12	Mg	$[Ne]3s^2$	49	In	$[Kr]5s^24d^{10}5p^1$	86	Rn	$[Xe]6s^24f^{14}5d^{10}6p^6$
13	Al	$[Ne]3s^23p^1$	50	Sn	$[Kr]5s^24d^{10}5p^2$	87	Fr	$[Rn]7s^1$
14	Si	$[Ne]3s^23p^2$	51	Sb	$[Kr]5s^24d^{10}5p^3$	88	Ra	$[Rn]7s^2$
15	P	$[Ne]3s^23p^3$	52	Te	$[Kr]5s^24d^{10}5p^4$	89	Ac	$[Rn]7s^26d^1$
16	S	$[Ne]3s^23p^4$	53	I	$[Kr]5s^24d^{10}5p^5$	90	Th	$[Rn]7s^26d^2$
17	Cl	$[Ne]3s^23p^5$	54	Xe	$[Kr]5s^24d^{10}5p^6$	91	Pa	$[Rn]7s^25f^26d^1$
18	Ar	$[Ne]3s^23p^6$	55	Cs	$[Xe]6s^1$	92	U	$[Rn]7s^25f^36d^1$
19	K	$[Ar]4s^1$	56	Ba	$[Xe]6s^2$	93	Np	$[Rn]7s^25f^46d^1$

Example 7.13 (1 of 4)

Write the ground-state electron configurations for

(a)sulfur (S)

(b)palladium (Pd), which is diamagnetic.



Example 7.13 (2 of 4)

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

This leaves us 6 electrons to fill the 3*s* subshell and partially fill the 3*p* subshell. Thus, the electron configuration of S is

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