

Atomic Structure and Periodicity

Text Chapter 8 & 9

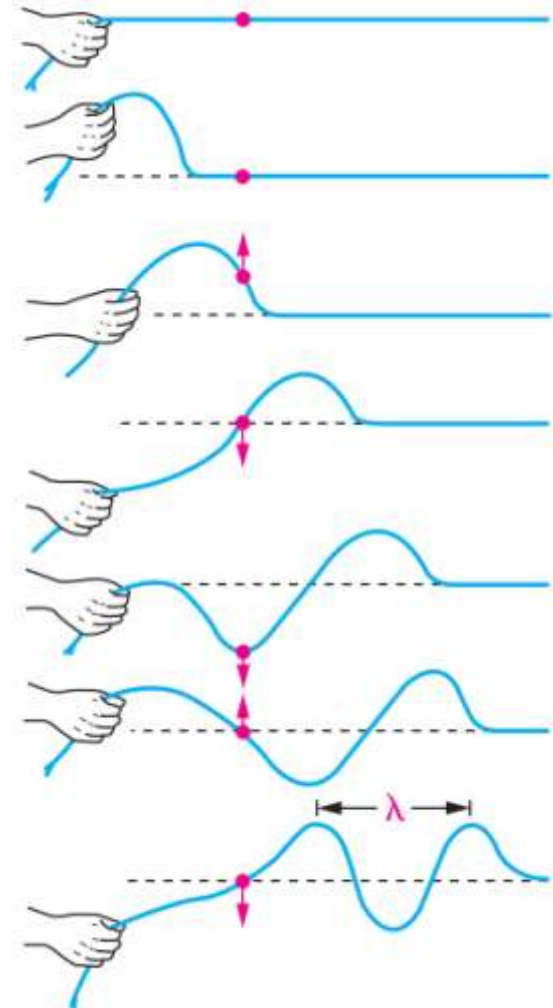
- Nature of light, elementary spectroscopy.
- The quantum theory and the Bohr atom.
- Quantum mechanics; the orbital concept.
- Electron configurations of atoms
- The periodic table: atomic structure & periodic trends

- Our understanding of the electronic structures of atoms will be gained by studying the interactions of *electromagnetic radiation* and *matter*.
- What is Electromagnetic Radiation ?

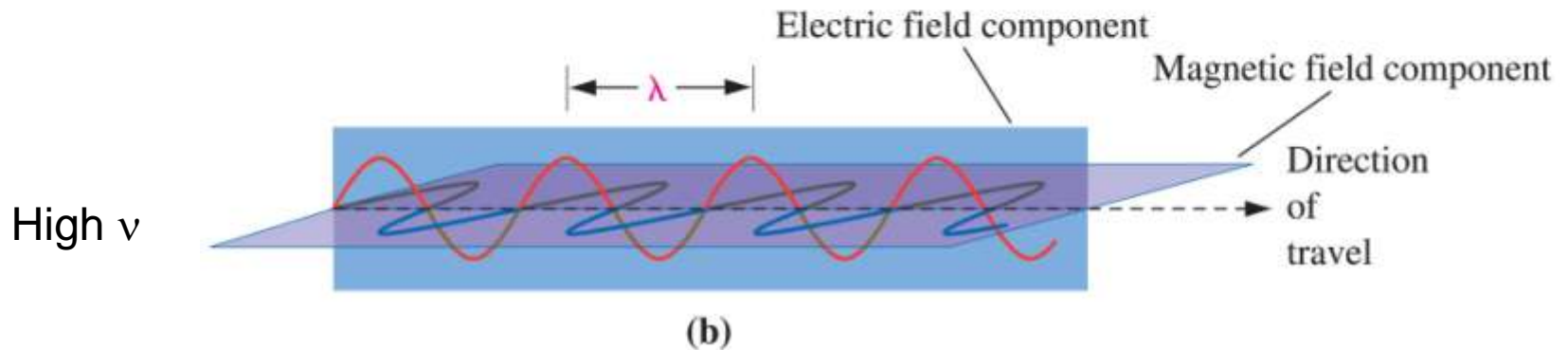
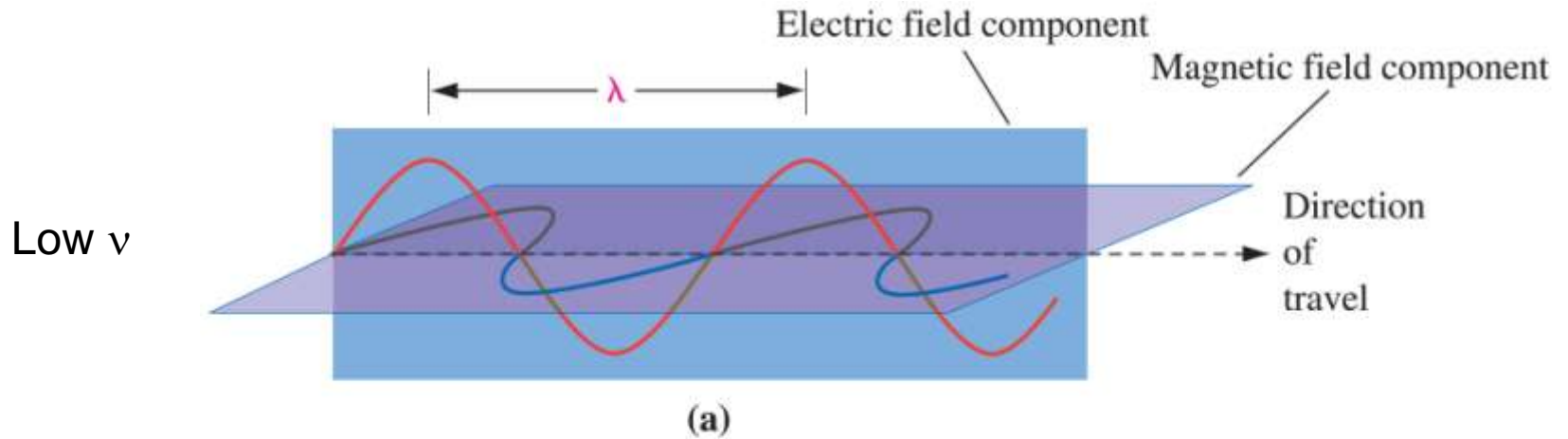
- Electromagnetic Radiation is a type of *energy transmission* in the form of a wave.
- A wave has a periodic variation of some quantity that transmits energy through space or a material medium, e.g. water waves or sound waves.

Electromagnetic Radiation

- As a result of the up-and-down hand motion, waves pass along the rope from left to right.
- The wavelength (λ), the distance between two successive **crests**, is indicated.
- Electric and magnetic fields propagate as waves through empty space or through a medium.
- A wave transmits energy.



EM Radiation



Important characteristics of a wave

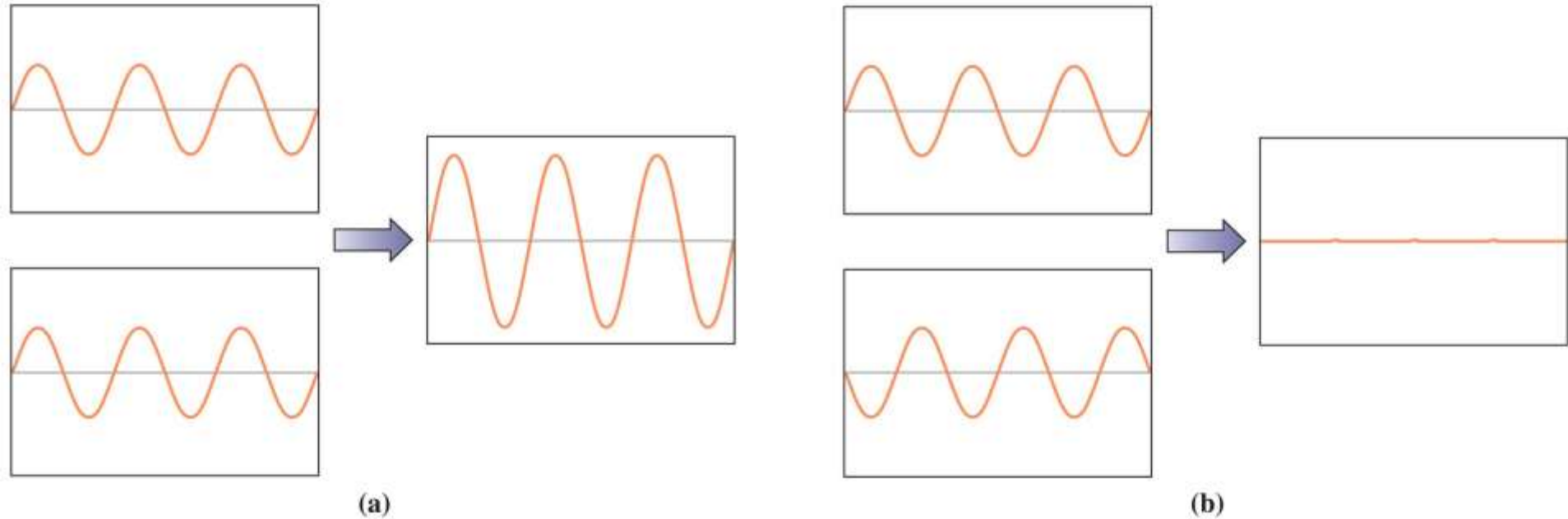
- Frequency (ν) in Hertz—Hz or s^{-1} ; the number of cycles per second.
- Wavelength (λ) in meters—m.
 - cm μm nm pm
(10^{-2} m) (10^{-6} m) (10^{-9} m) (10^{-12} m)
- Velocity (c)— $2.997925 \times 10^8 \text{ m s}^{-1}$.

$$c = \lambda \nu$$

$$\lambda = c/\nu$$

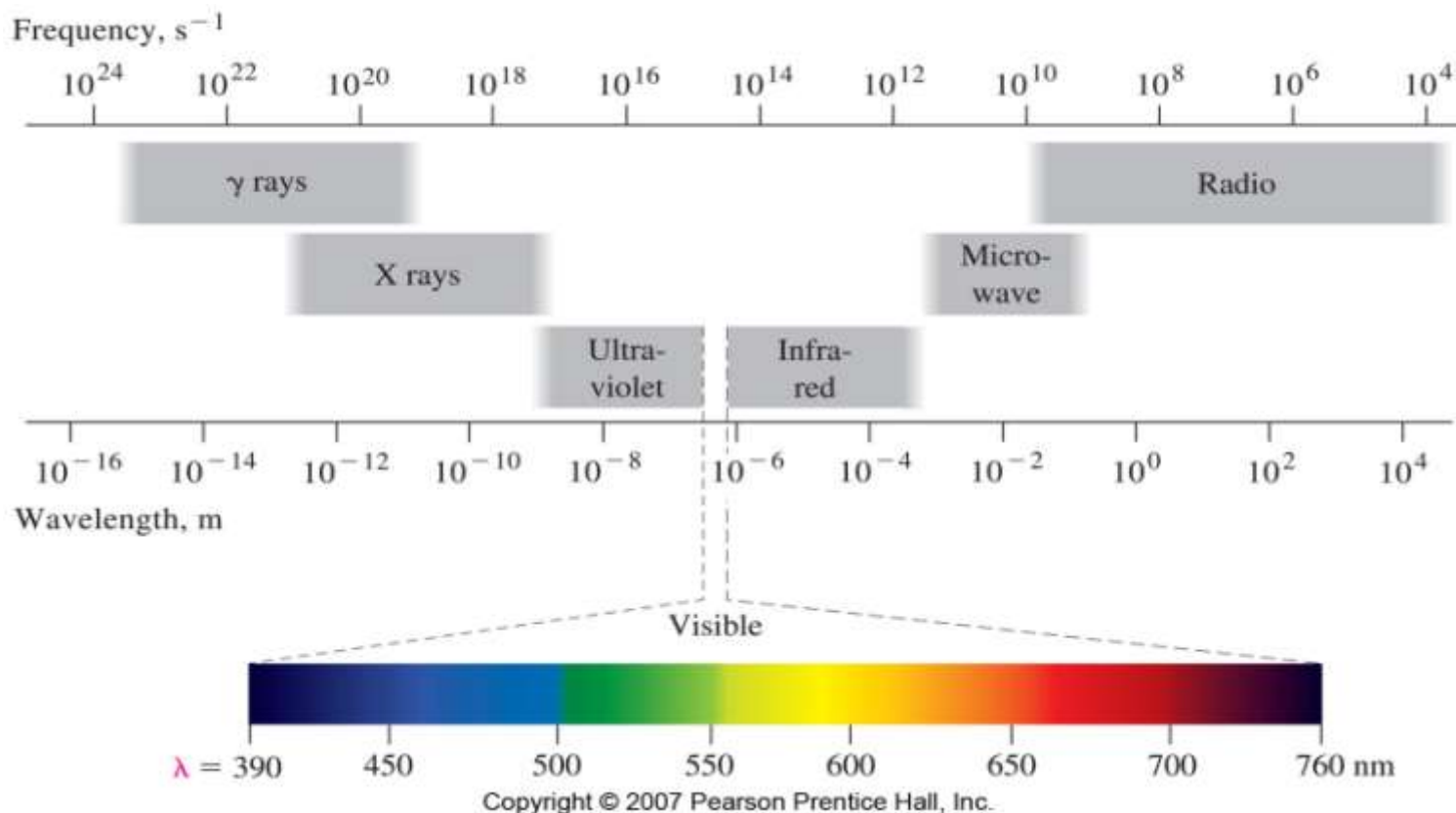
$$\nu = c/\lambda$$

Interference of two waves



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Electromagnetic Spectrum



The visible region, which extends from violet at the shortest wavelength to red at the longest wavelength, is only a small portion of the entire electromagnetic spectrum.

Note *infra~* means "below" and *ultra~* means "above".

Speed of light is different in different media.

Light is refracted (bent) as it passes from air into the glass prism.

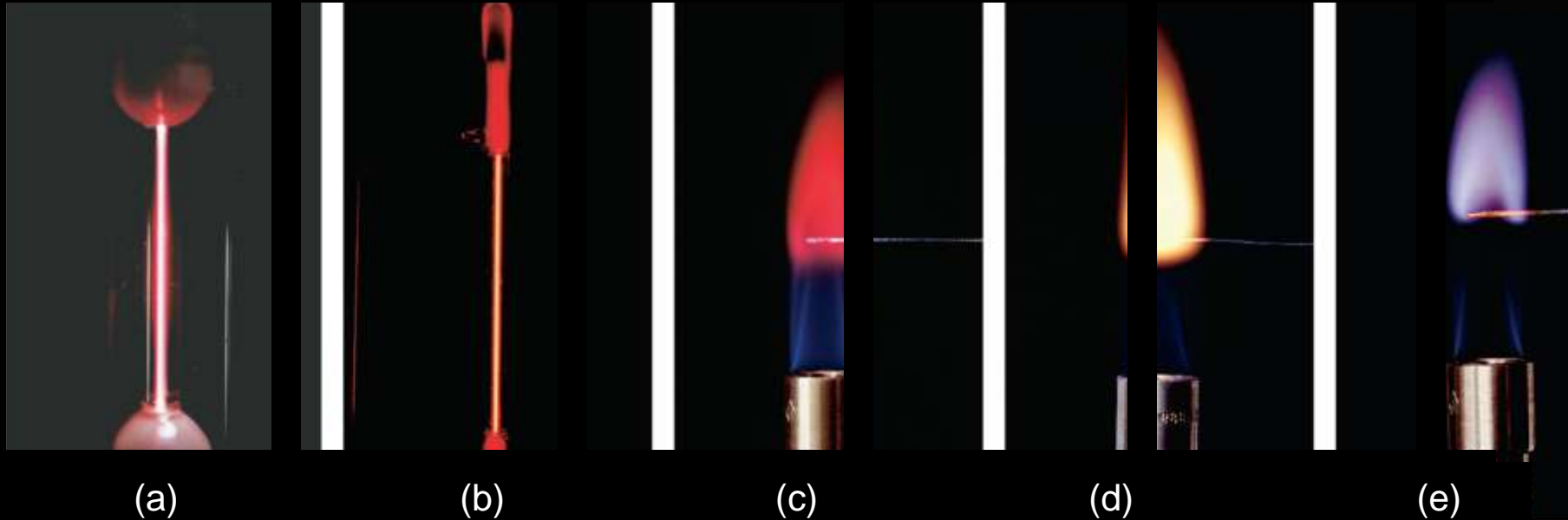
Waves of different wavelengths have different speed in different mediums.

Thus, when a beam of white light is passed through a transparent medium, the wavelength components are refracted differently. The light is dispersed into a band of colors, a *spectrum*.



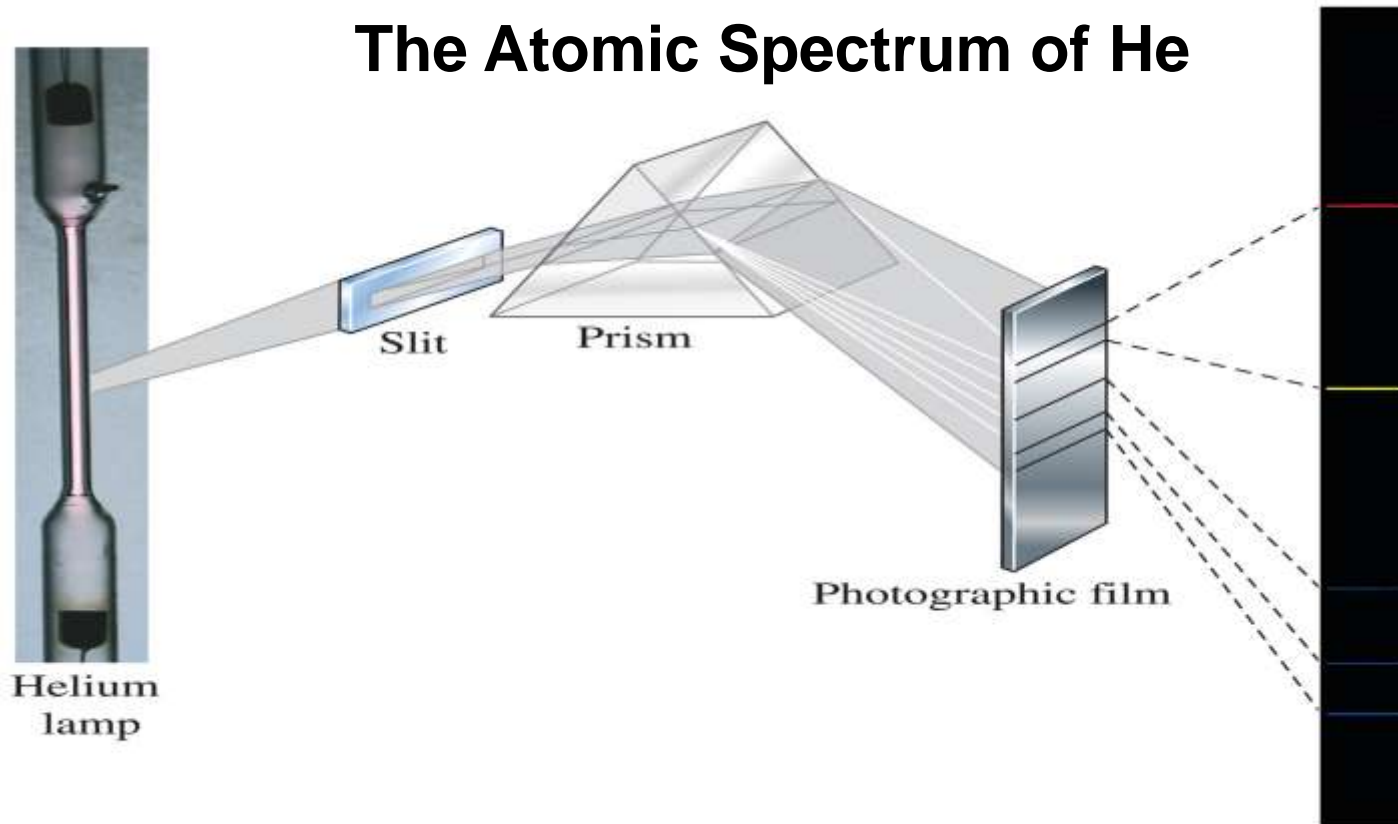
The importance of light to chemistry is that light is a form of energy so by studying light-matter interactions we can detect energy changes in atoms and molecules.

Atomic Spectra



Light emitted by an electric discharge through **(a)** hydrogen gas and **(b)** helium gas. Light emitted when compounds of the alkali metals are excited in the gas flames: **(c)** lithium, **(d)** sodium, and **(e)** potassium. **Such colors can be used to identify the presence of a given element.**

The Atomic Spectrum of He



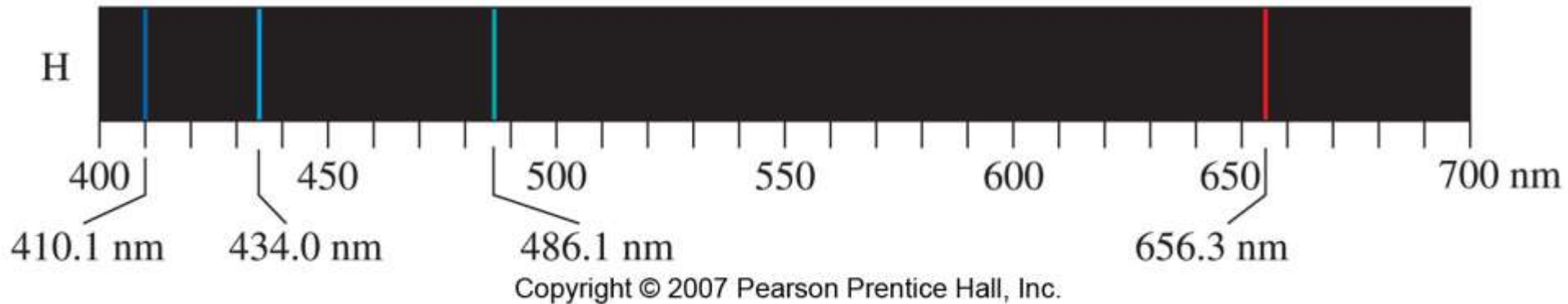
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An electric discharge is passed through a lamp containing He gas at low pressure; **He atoms absorb energy and emit it as light**. The light is passed through a slit and then dispersed by a prism. The colored components of the light are detected.

Each element has its own distinctive line spectrum- a kind of atomic fingerprint.

The line spectrum for hydrogen atoms:

The four lines shown are the only ones visible to the unaided eye.



The fact that atomic spectra consist of only limited numbers of well-defined wavelength lines provides a great opportunity to learn about the structures of atoms. For example, it suggests that only a limited number of energy values are available to excite gaseous atoms.

The Photoelectric Effect

Light striking the surface of certain metals causes ejection of electrons.

$$- \nu > \nu_0$$

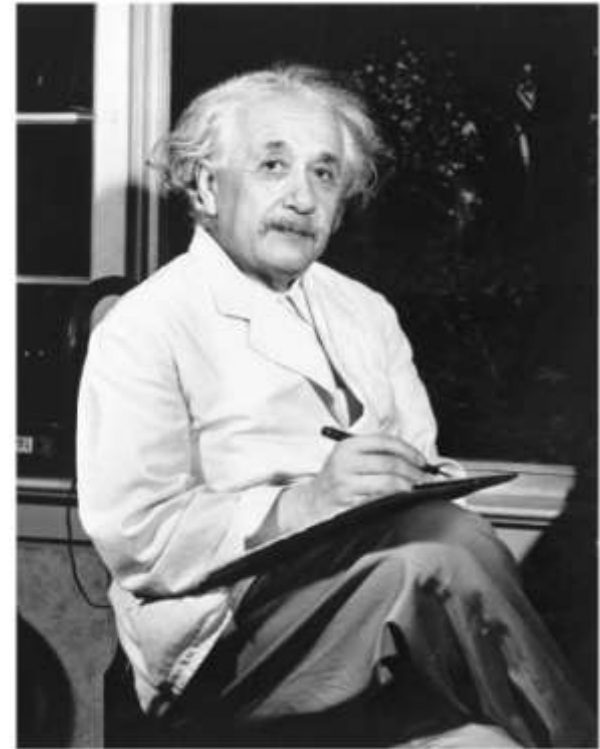
threshold frequency

- $\#e^- \propto I$ (intensity of the incident light)

$$- E_k \propto \nu$$

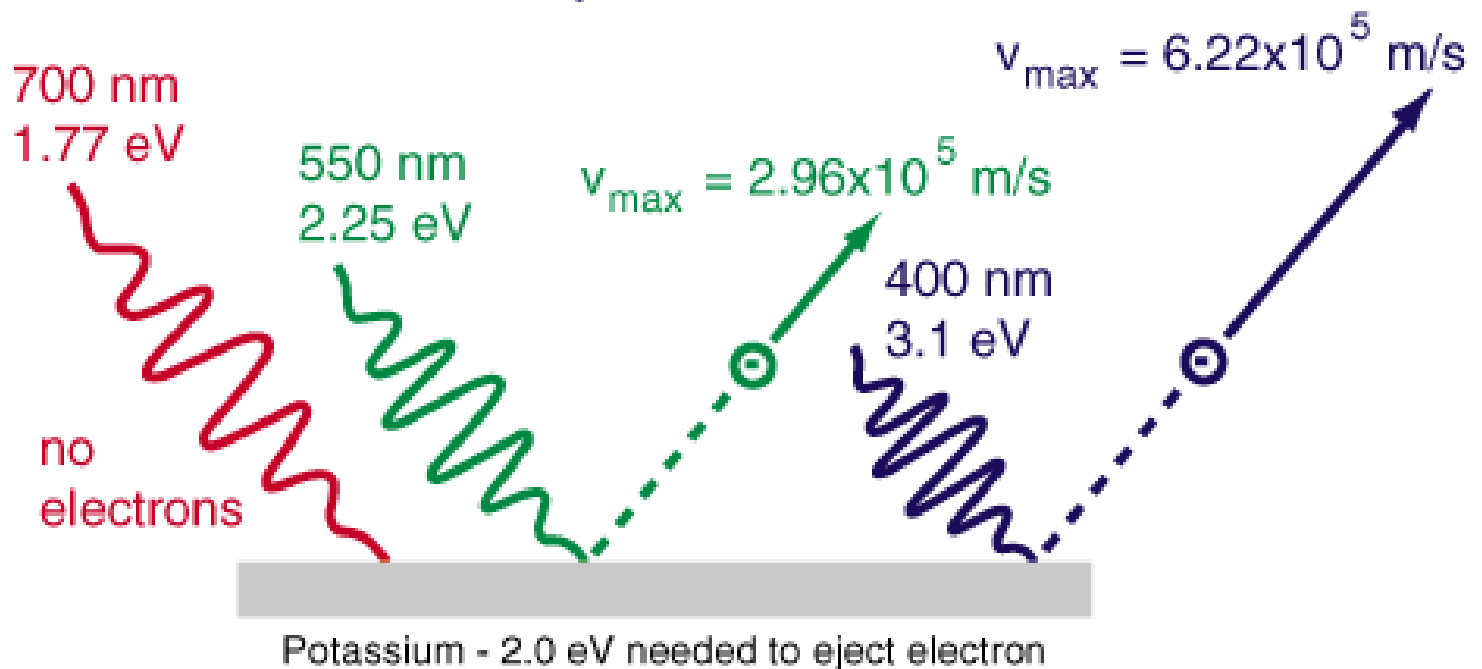
$$E_{\text{photon}} = E_k + E_{\text{binding}}$$

$$h\nu = \frac{1}{2} mu^2 + eV_0$$



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$$E_{\text{photon}} = h\nu$$



Photoelectric effect

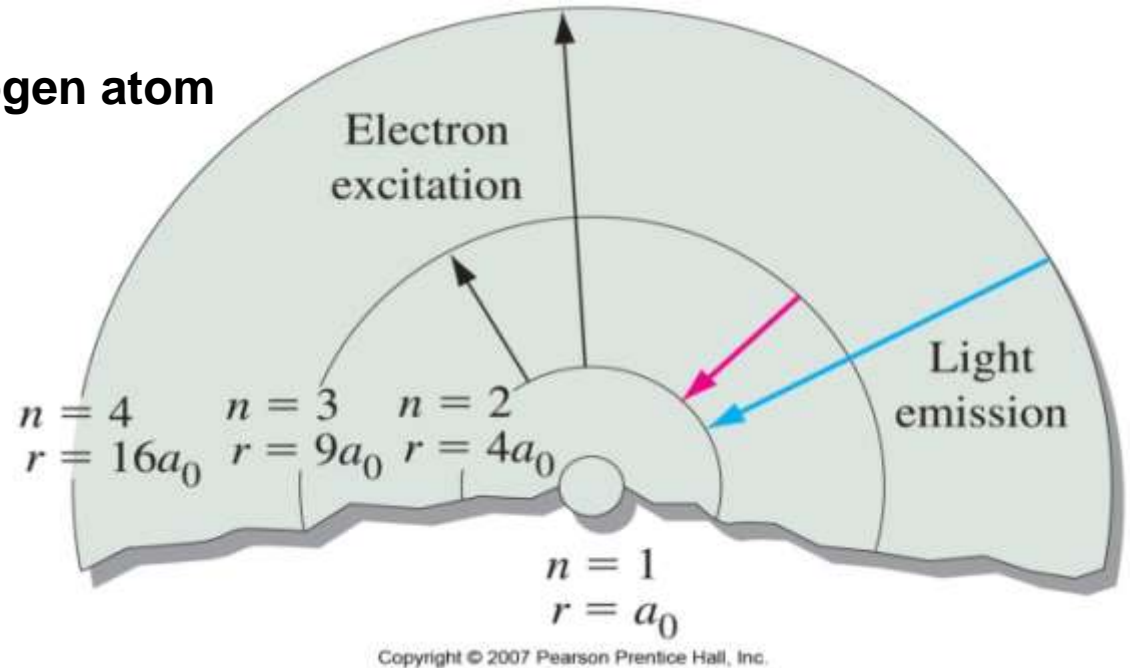
To explain the photoelectric effect, Einstein suggested that light has particle-like properties, which are displayed through photons.

Other phenomena, however, such as the dispersion of light into a spectrum by a prism, are best understood in terms of the wave theory of light. Light, then, appears to have a *dual* nature.

In 1905 Albert Einstein proposed that:

- Light has the ability to eject electrons when striking certain surfaces.
- *So, Light has particle-like qualities.*
- The “particles” of light are called photons.
- Photons have a characteristic energy content given by Planck’s equation: **$E = h \nu$**

Bohr's model of the hydrogen atom



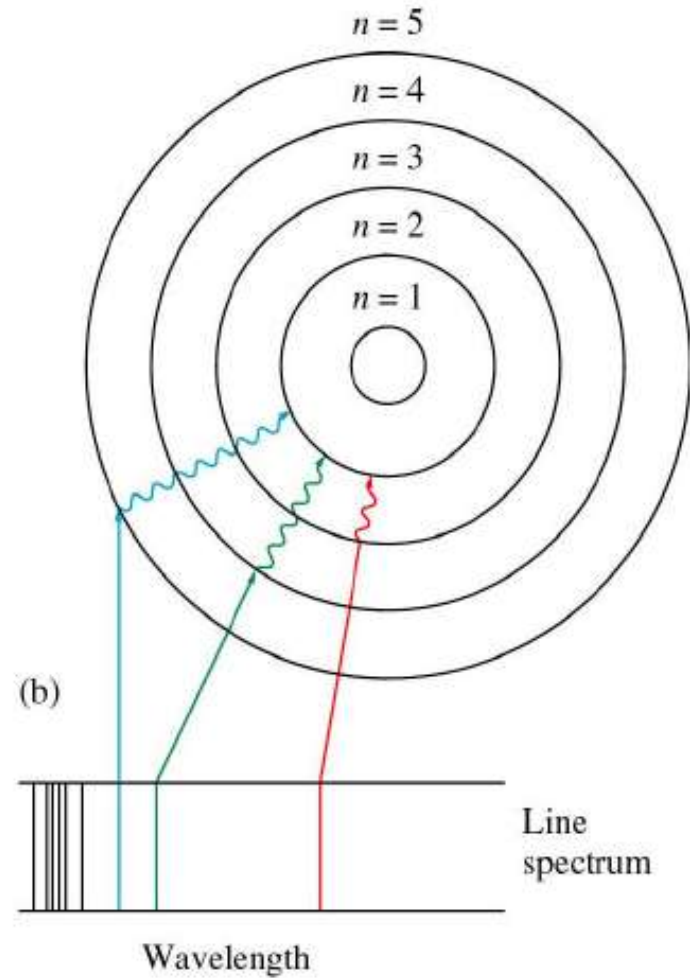
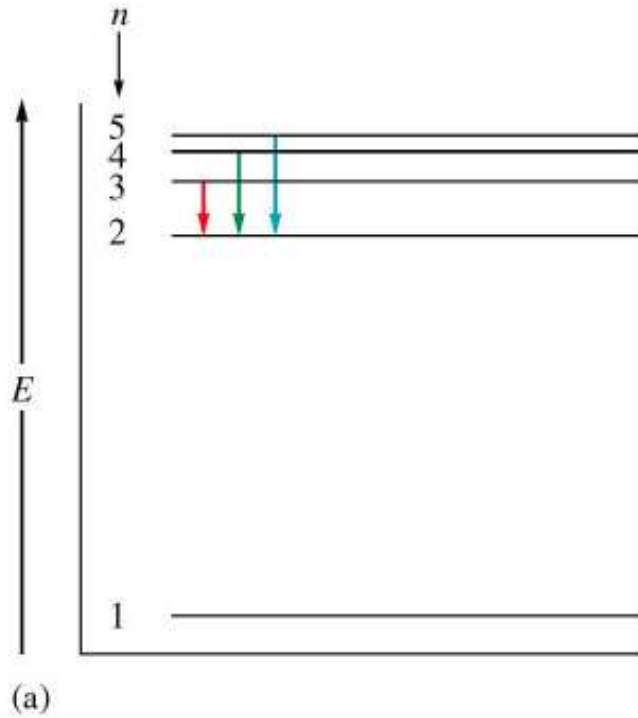
The nucleus is at the center, and the electron is found in one of the discrete (allowed) orbits designated by a quantum number, n , $n=1$, $n=2$, and so on.

Energy packets (quanta) are absorbed or emitted when electrons change stationary states.

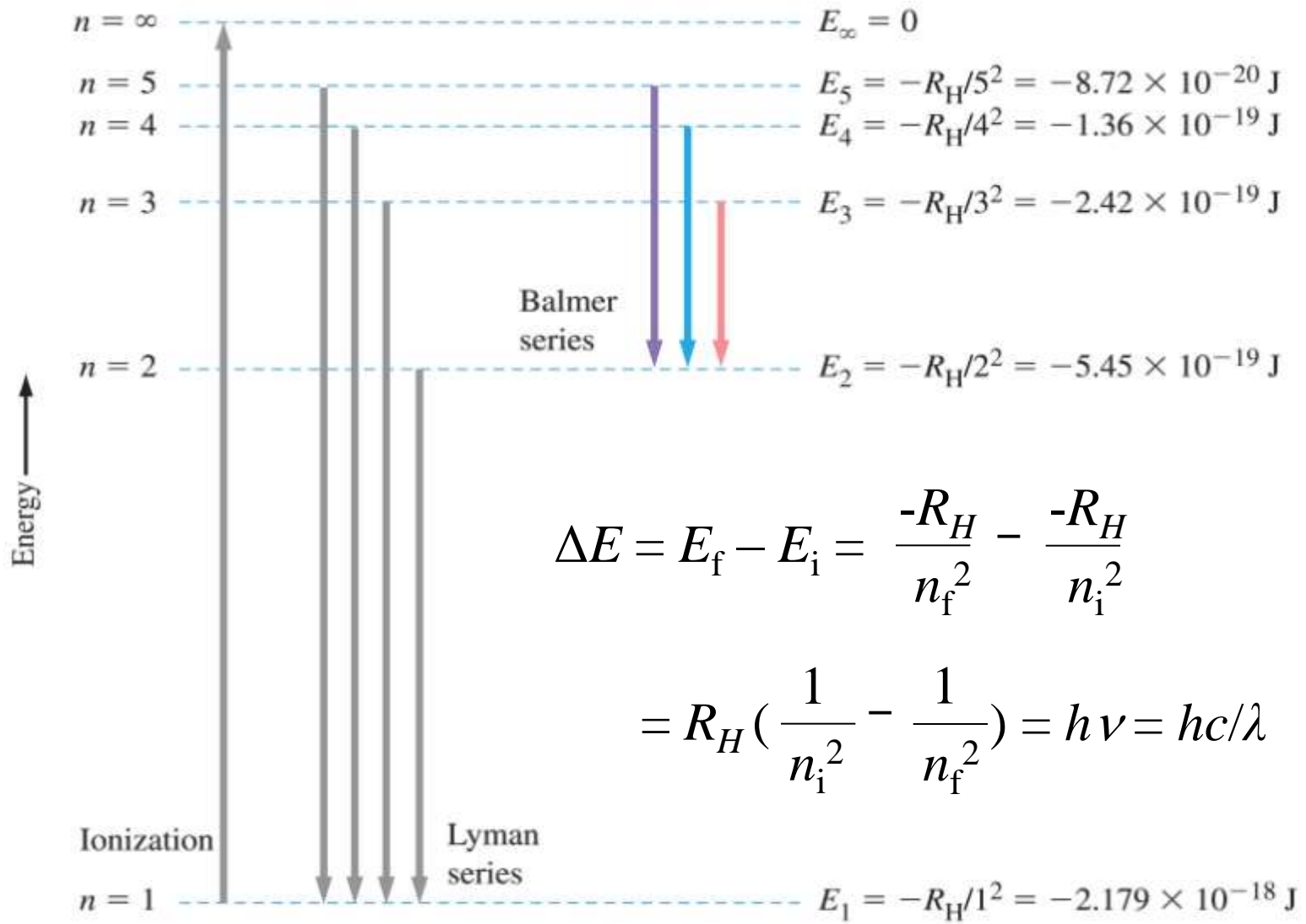
Excitation of the atom raises the electron to higher-numbered orbits, as shown with black arrows.

Light is emitted when the electron falls to a lower-numbered orbit.

Bohr Quantum Model



$$F = k_e \frac{q_1 q_2}{r^2}$$



$$\Delta E = E_f - E_i = \frac{-R_H}{n_f^2} - \frac{-R_H}{n_i^2}$$

$$= R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right) = h\nu = hc/\lambda$$

Energy-level diagram for the hydrogen atom

$$E_n = -R_H \left(\frac{1}{n^2} \right) \quad \text{Where; (Rydberg Constant) } R_H = 2.18 \times 10^{-18} \text{ J}$$

- Developed based on calculated orbits
- Negative sign indicates state is more stable than free electron (relative $E = 0$)
- Lowest energy state is when $n = 1$, this is sometimes called the **ground level**
- When $n > 1$, these levels have higher energy. These states are called **excited states**.

Difference in Energy between states;

$$\Delta E_n = E_f - E_i$$

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

Bohr Theory and the Ionization Energy of Hydrogen

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

As n_f goes to infinity for hydrogen starting in the ground state:

$$h\nu = R_H \left(\frac{1}{n_i^2} \right) = R_H$$

This also works for hydrogen-like species such as He^+ and Li^{2+} .

$$E_n = R_H \left(\frac{-Z^2}{n_i^2} \right) = -Z^2 R_H$$

8-5 Two Ideas Leading to a New Quantum Mechanics

- Wave-Particle Duality
 - Einstein suggested particle-like properties of light could explain the photoelectric effect.
 - Diffraction patterns suggest photons are wave-like.
- deBroglie, 1924
 - Small particles of matter may at times display wavelike properties.



Louis de Broglie
Nobel Prize 1918

Wave-Particle Duality

$$E = mc^2$$

$$h\nu = mc^2$$

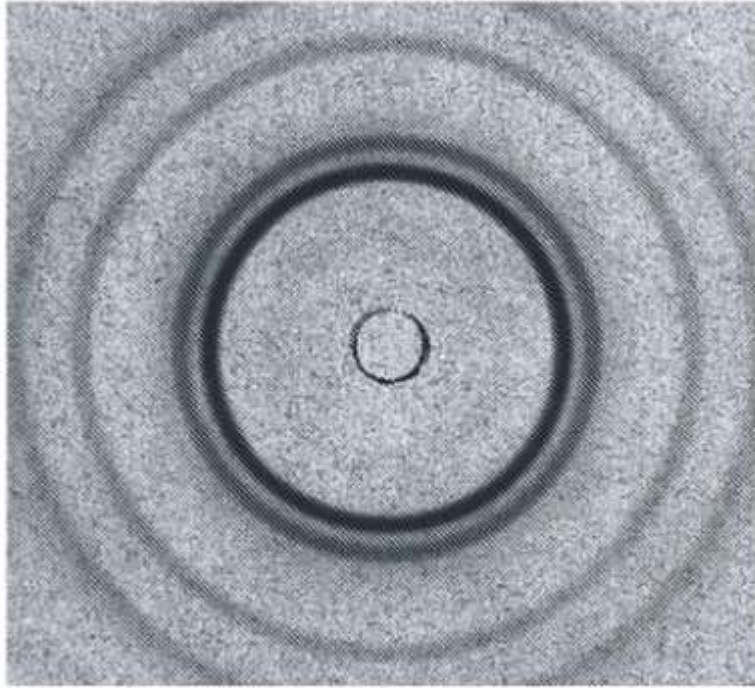
$$h\nu/c = mc = p$$

$$p = h/\lambda$$

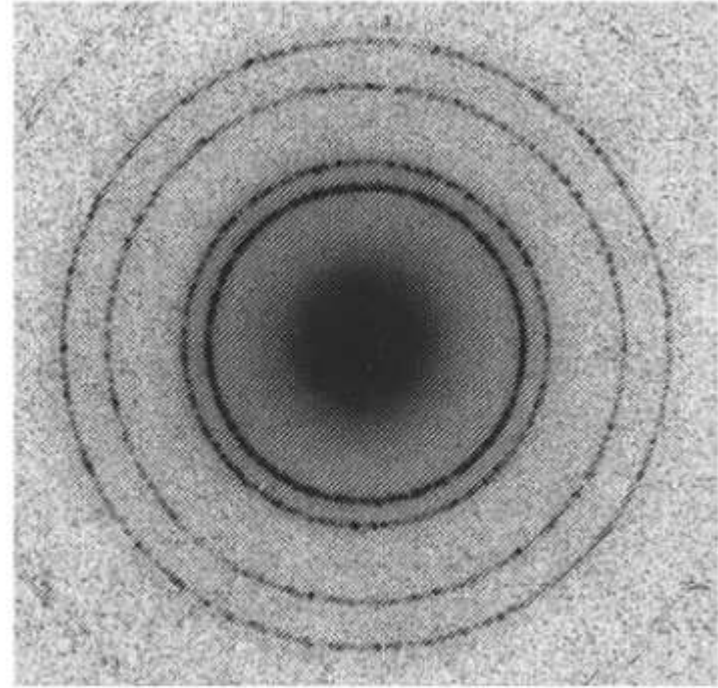
$$\lambda = h/p = h/mu$$

De Broglie called the waves associated with material particles “matter waves.” If matter waves exist for small particles, then beams of particles, such as electrons, should exhibit the characteristic properties of waves, namely diffraction.

Diffraction of electrons by metal foil



(a)



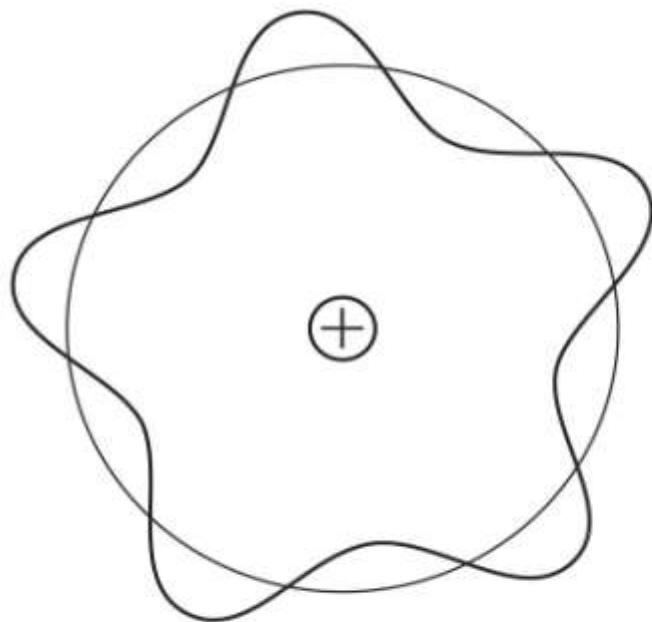
(b)

In 1927, C. J. Davisson and L. H. Germer of the United States showed that a beam of slow electrons is diffracted by a crystal of nickel. In a similar experiment in that same year, G. P. Thomson of Scotland directed a beam of electrons at a thin metal foil. He obtained the same pattern for the diffraction of electrons by aluminum foil as with X-rays of the same wavelength.

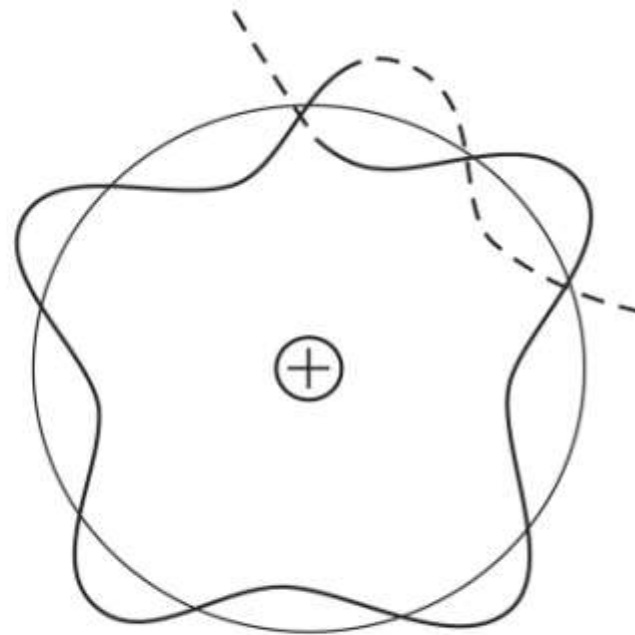
Thomson and Davisson shared the 1937 Nobel Prize in physics for their electron diffraction experiments.

George P. Thomson was the son of J.J. Thomson, who had won the Nobel Prize in physics in 1906 for his discovery of the electron.

It is interesting to note that Thomson the father showed that the electron is a particle, and Thomson the son showed that the electron is a wave. Father and son together demonstrated the **wave–particle duality** of electrons.



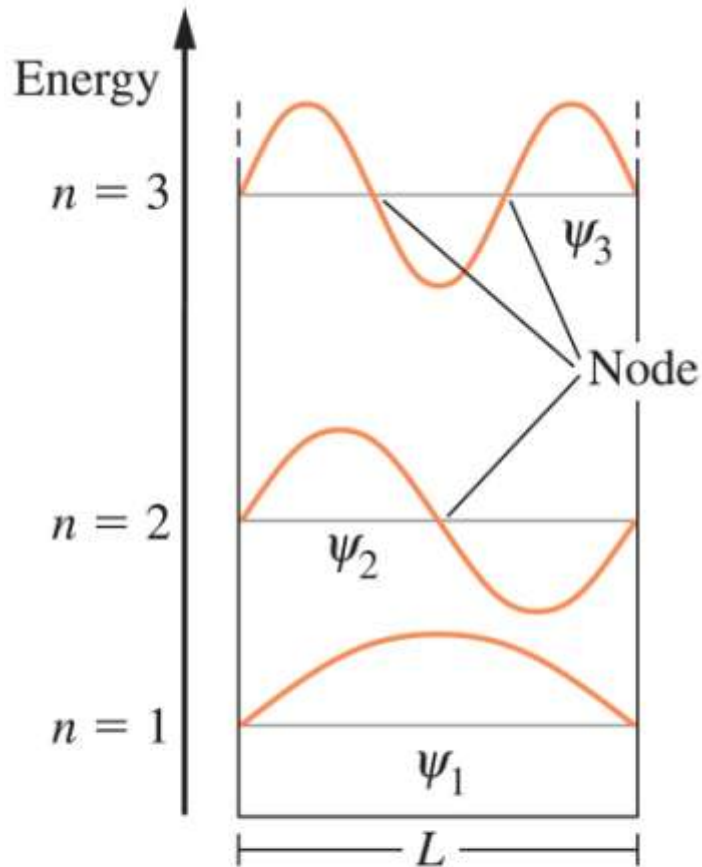
(a)



(b)

FIGURE 8-19
The electron as a matter wave

Particle in Box: Standing Waves, Quantum Particles, and Wave Functions



ψ , psi, the wave function.

Should correspond to a standing wave within the boundary of the system being described.

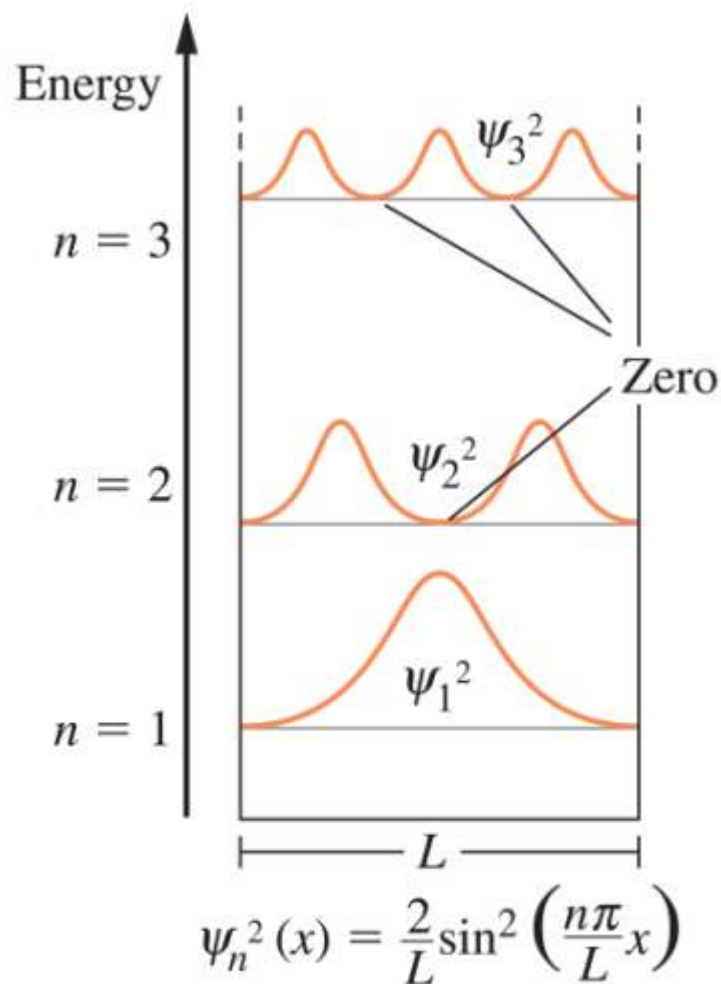
Particle in a box.

$$\psi = \sqrt{\frac{2}{L}} \sin\left(\frac{n\pi}{L} x\right)$$

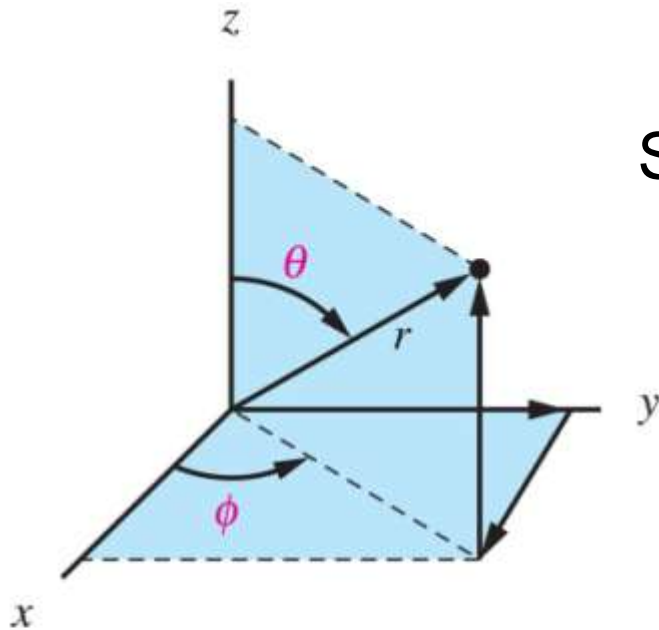
▲ FIGURE 8-20

•The standing waves of a particle in a one-dimensional box

Probability of Finding an Electron



Wave Functions of the Hydrogen Atom



Spherical polar coordinates

$$x^2 + y^2 + z^2 = r^2$$

$$x = r \sin \theta \cos \phi$$

$$y = r \sin \theta \sin \phi$$

$$z = r \cos \theta$$

Schrödinger, 1927

$$E\psi = \mathcal{H}\psi$$

$$\mathcal{H}(x, y, z) \text{ or } \mathcal{H}(r, \theta, \phi)$$

$$\psi(r, \theta, \phi) = R(r) Y(\theta, \phi)$$

$R(r)$ is the radial wave function.

$Y(\theta, \phi)$ is the angular wave function.

▲ FIGURE 8-22

•The relationship between spherical polar coordinates and Cartesian coordinates

Schrödinger, 1927

- Came up with an equation to describe the wave function of the electron in H atom.
- The solutions of Schrödinger's equation give the wave functions called *orbitals*.

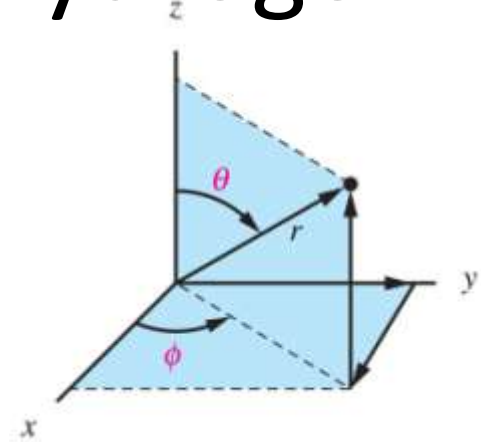
Wave Functions for Hydrogen

- Schrödinger, 1927
- To define a point with respect to nucleus in a sphere we need to define ...

$$\psi(r, \theta, \phi) = R(r) Y(\theta, \phi)$$

$R(r)$ is the **radial wave** function.

$Y(\theta, \phi)$ is the **angular wave function**.



Spherical polar coordinates
 $x^2 + y^2 + z^2 = r^2$
 $x = r \sin \theta \cos \phi$
 $y = r \sin \theta \sin \phi$
 $z = r \cos \theta$

$$\psi(1s) = R(r) \times Y(\theta, \phi) = \frac{2e^{-r/a_0}}{a_0^{3/2}} \times \frac{1}{\sqrt{4\pi}} = \frac{e^{-r/a_0}}{\sqrt{(\pi a_0^3)}}$$

An orbital describes a region in an atom where an electron is likely to be found.

To define an orbital we need to define **three *Quantum Numbers***

These three QNs arise from the solutions of Schrödinger's wave equation for H atom.

9-8 Interpreting and Representing the Orbitals of the Hydrogen Atom.

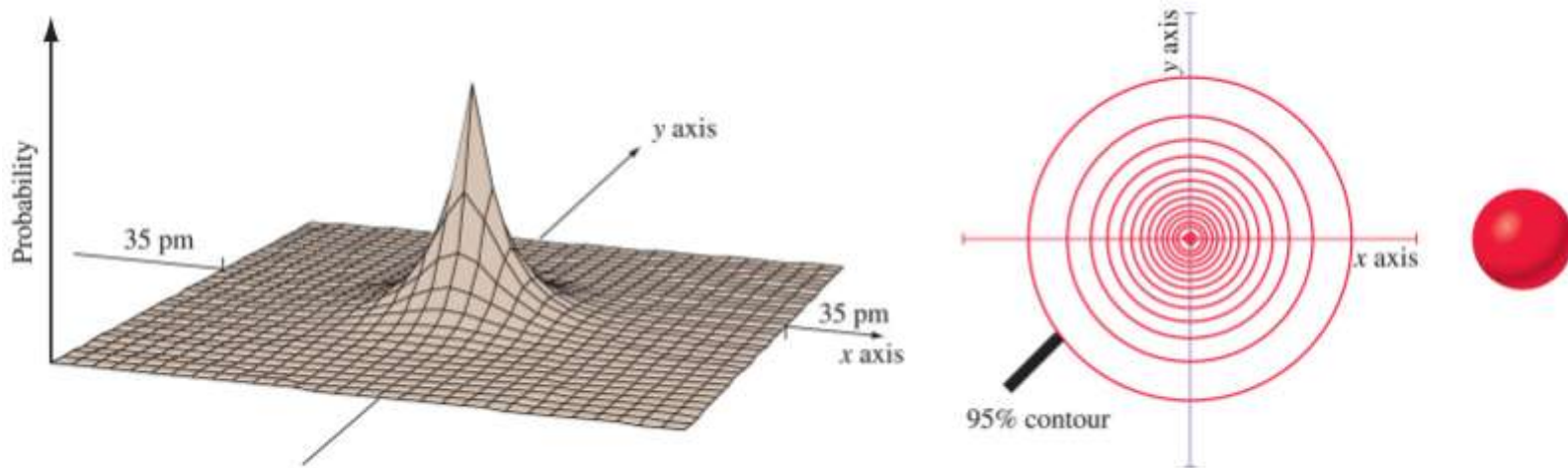
- Represent the probability densities of the orbitals of the hydrogen atom as three dimensional surfaces.
- Each orbital has a distinctive shape.
- Acquire a broad *qualitative* understanding.

TABLE 8.1 The Angular and Radial Wave Functions of a Hydrogen-Like Atom

Angular Part $Y(\theta, \phi)$	Radial Part $R_{n, \ell}(r)$
$Y(s) = \left(\frac{1}{4\pi}\right)^{1/2}$	$R(1s) = 2\left(\frac{Z}{a_0}\right)^{3/2} e^{-\sigma/2}$
	$R(2s) = \frac{1}{2\sqrt{2}}\left(\frac{Z}{a_0}\right)^{3/2} (2 - \sigma)e^{-\sigma/2}$
	$R(3s) = \frac{1}{9\sqrt{3}}\left(\frac{Z}{a_0}\right)^{3/2} (6 - 6\sigma + \sigma^2)e^{-\sigma/2}$
$Y(p_x) = \left(\frac{3}{4\pi}\right)^{1/2} \sin\theta \cos\phi$	$R(2p) = \frac{1}{2\sqrt{6}}\left(\frac{Z}{a_0}\right)^{3/2} \sigma e^{-\sigma/2}$
$Y(p_y) = \left(\frac{3}{4\pi}\right)^{1/2} \sin\theta \sin\phi$	$R(3p) = \frac{1}{9\sqrt{6}}\left(\frac{Z}{a_0}\right)^{3/2} (4 - \sigma)\sigma e^{-\sigma/2}$
$Y(p_z) = \left(\frac{3}{4\pi}\right)^{1/2} \cos\theta$	
$Y(d_{z^2}) = \left(\frac{5}{16\pi}\right)^{1/2} (3\cos^2\theta - 1)$	$R(3d) = \frac{1}{9\sqrt{30}}\left(\frac{Z}{a_0}\right)^{3/2} \sigma^2 e^{-\sigma/2}$
$Y(d_{x^2-y^2}) = \left(\frac{15}{16\pi}\right)^{1/2} \sin^2\theta \cos 2\phi$	
$Y(d_{xy}) = \left(\frac{15}{16\pi}\right)^{1/2} \sin^2\theta \sin 2\phi$	$\sigma = \frac{2Zr}{na_0}$
$Y(d_{xz}) = \left(\frac{15}{4\pi}\right)^{1/2} \sin\theta \cos\theta \cos\phi$	
$Y(d_{yz}) = \left(\frac{15}{4\pi}\right)^{1/2} \sin\theta \cos\theta \sin\phi$	

s orbitals

$$\psi(1s) = R(r) \times Y(\theta, \phi) = \frac{2e^{-r/a_0}}{a_0^{3/2}} \times \frac{1}{\sqrt{4\pi}} = \frac{e^{-r/a_0}}{\sqrt{(\pi a_0^3)}}$$

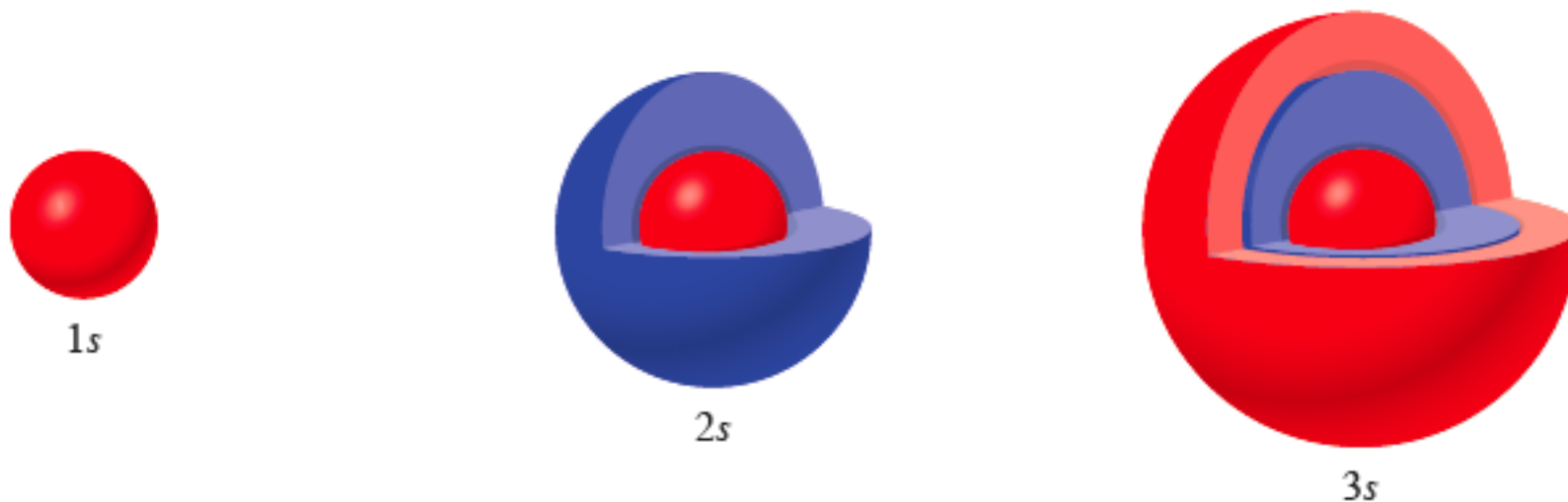


▲ FIGURE 8-24

•Three representations of the electron probability density for the 1s orbital

2s orbitals

$$\psi(2s) = R(r) \times Y(\theta, \phi) = \frac{1}{2\sqrt{2}} \frac{1}{a_0^{3/2}} \left(2 - \frac{r}{a_0} \right) e^{-r/2a_0} \times \frac{1}{\sqrt{4\pi}} = \frac{1}{4} \left(\frac{1}{2\pi a_0^3} \right)^{1/2} \left(2 - \frac{r}{a_0} \right) e^{-r/2a_0}$$



▲ FIGURE 8-24

• Three-dimensional representations of the 95% electron probability density for the 1s, 2s and 3s orbitals

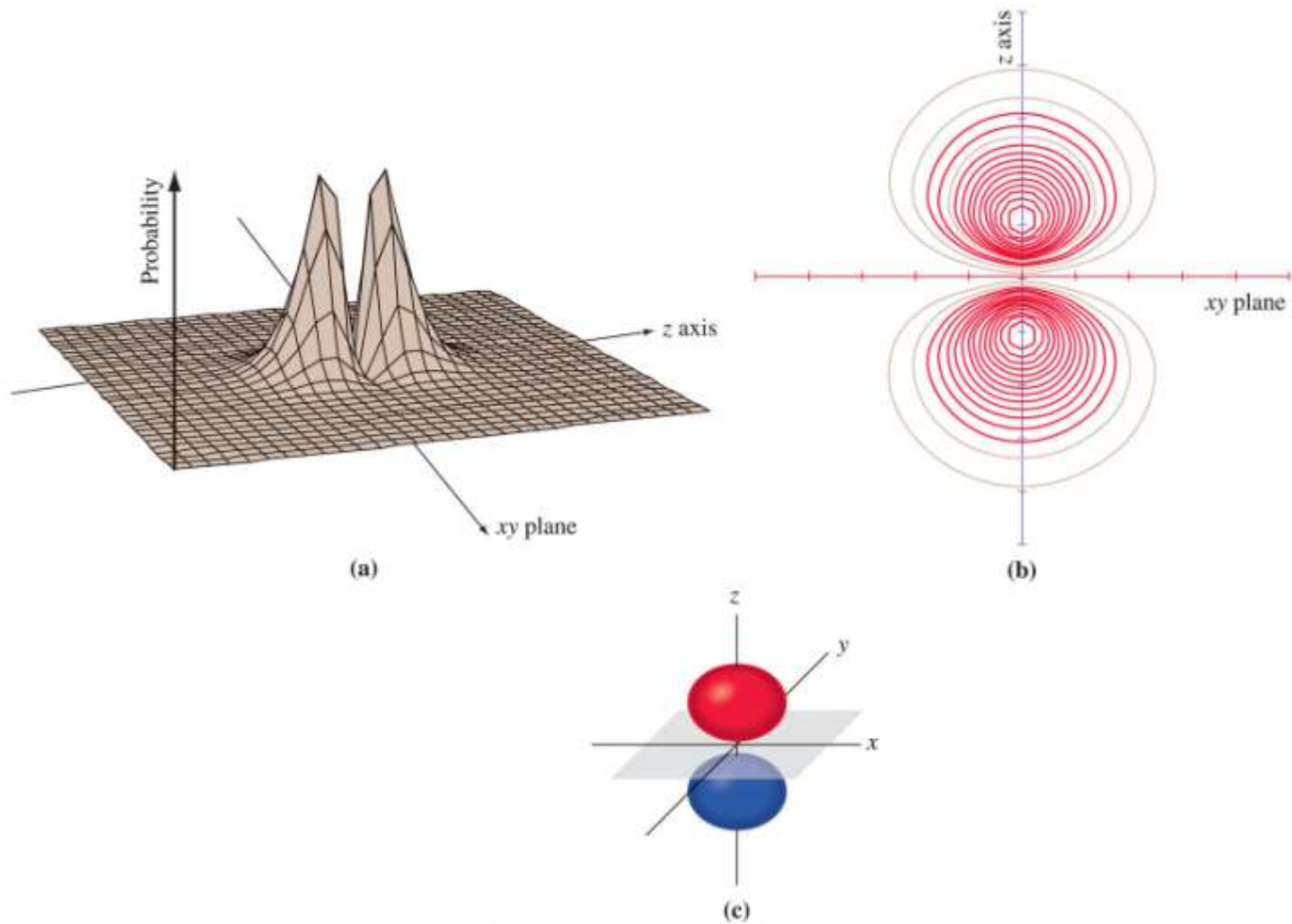


FIGURE 8-27

Three representations of electron probability for a 2p orbital

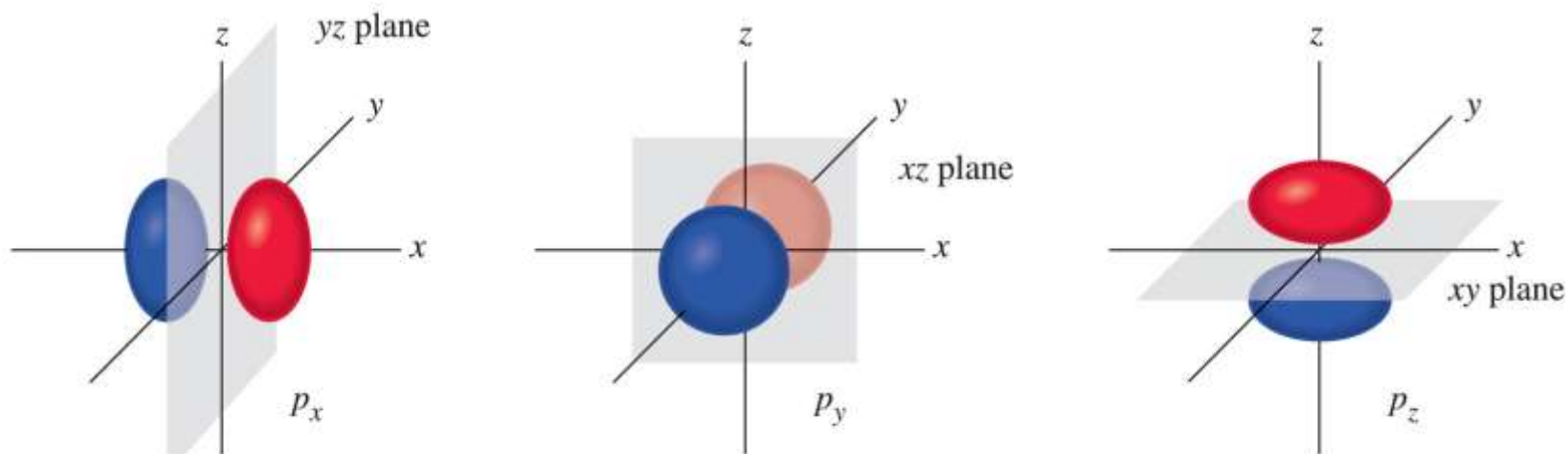
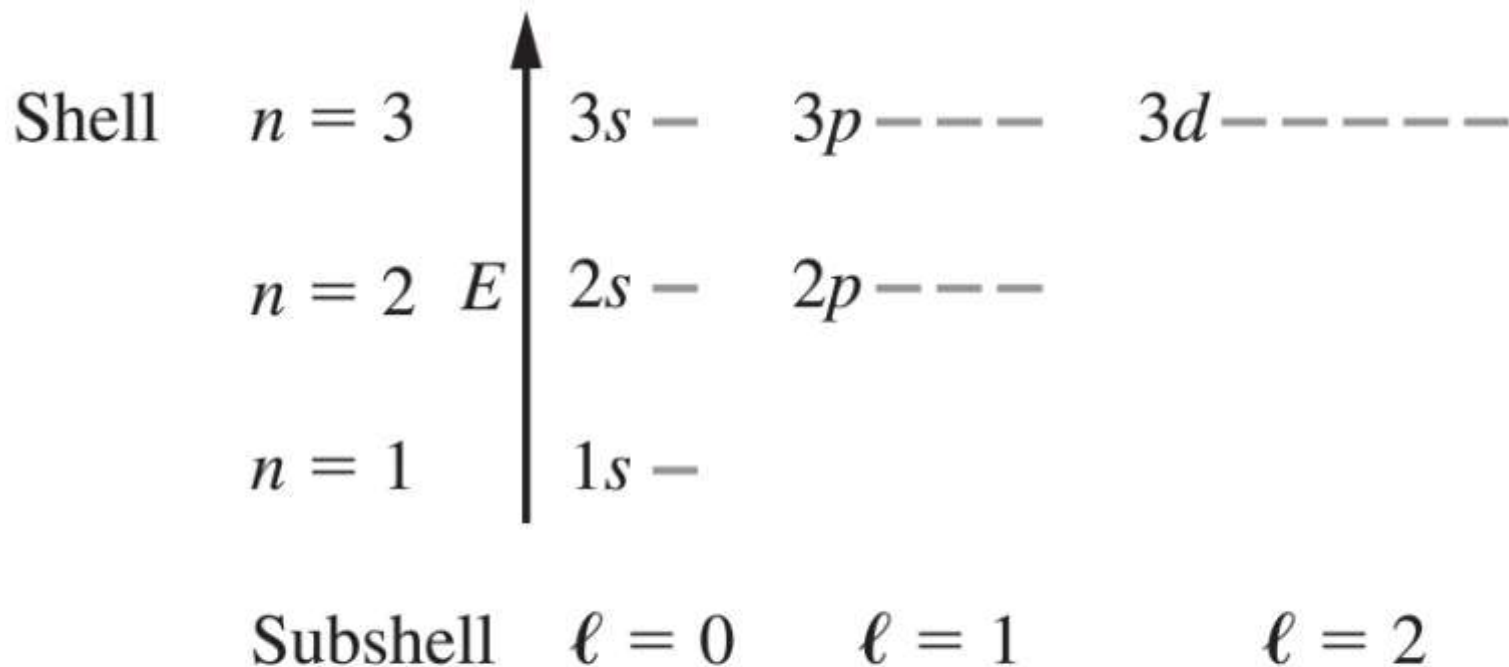


FIGURE 8-28
The three 2p orbitals

Principle Shells and Subshells

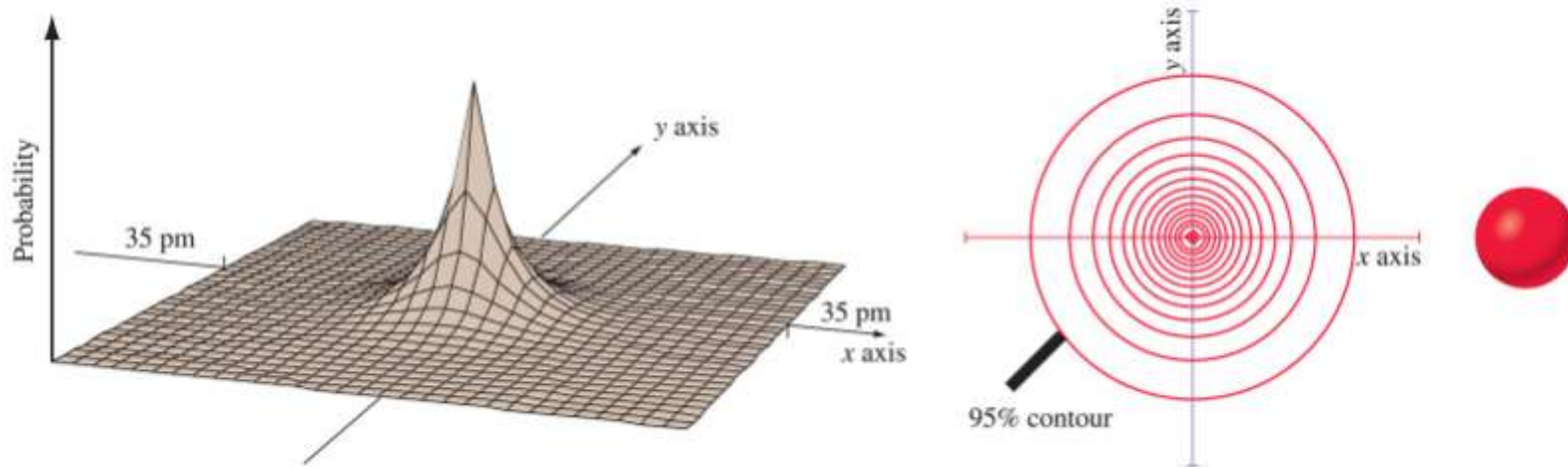
- Principle electronic shell, $n = 1, 2, 3\dots$
 - Angular momentum quantum number, $l = 0, 1, 2\dots(n-1)$
 - $l = 0, s$
 - $l = 1, p$
 - $l = 2, d$
 - $l = 3, f$
- ◆ Magnetic quantum number, $m_l = -l \dots -2, -1, 0, 1, 2\dots+l$

Orbital Energies

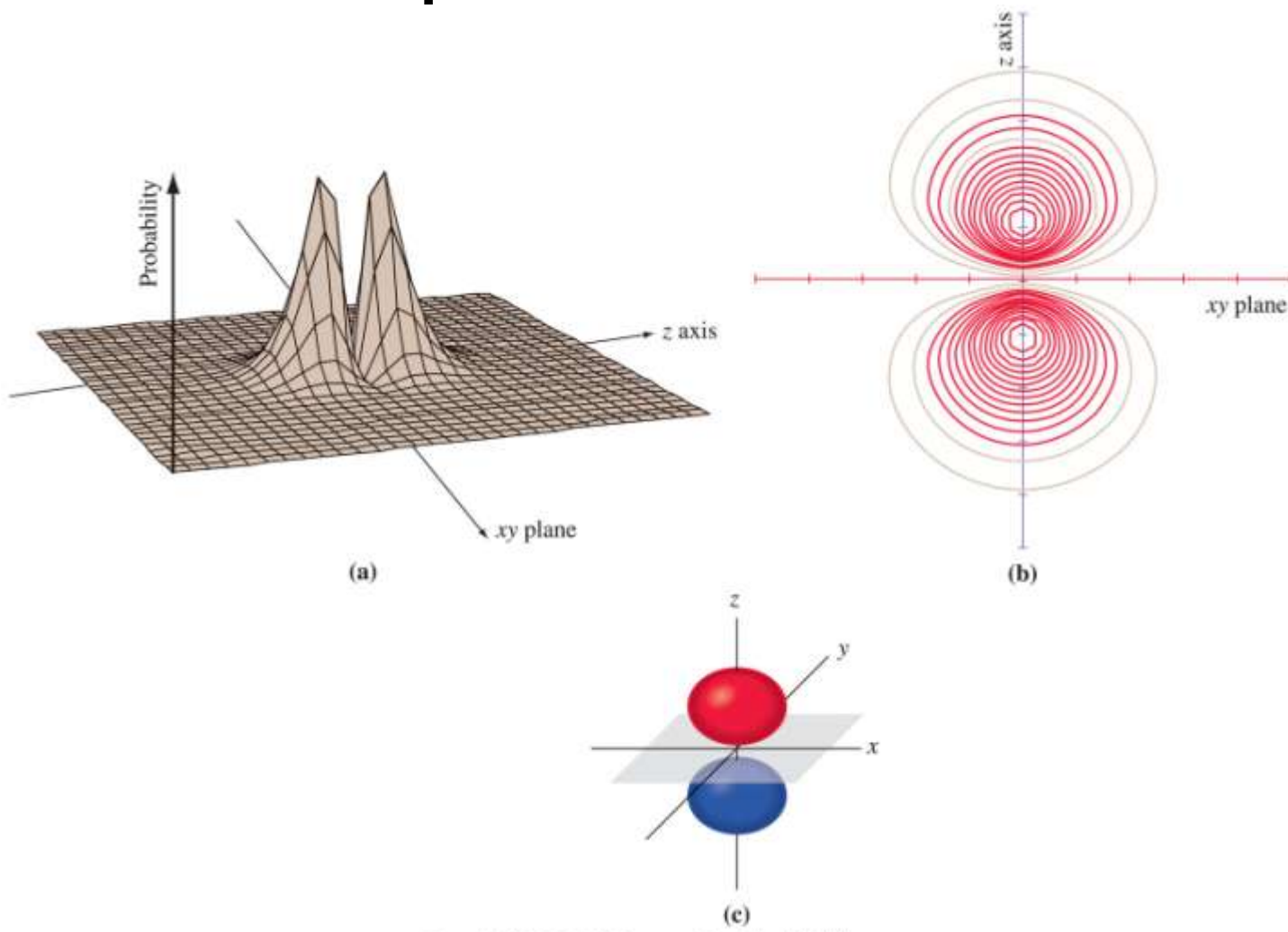


Each subshell is made up of $(2\ell + 1)$ orbitals.

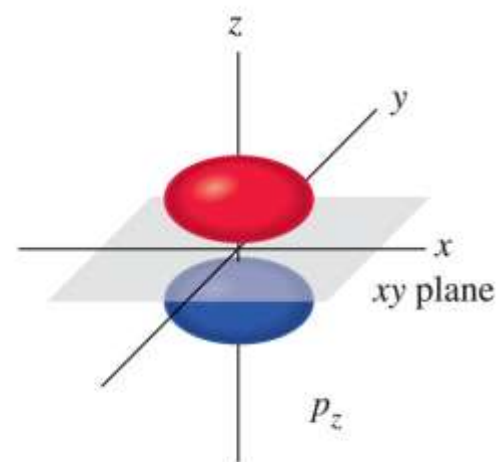
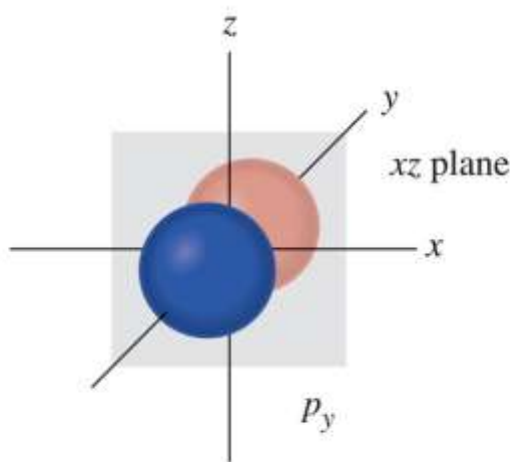
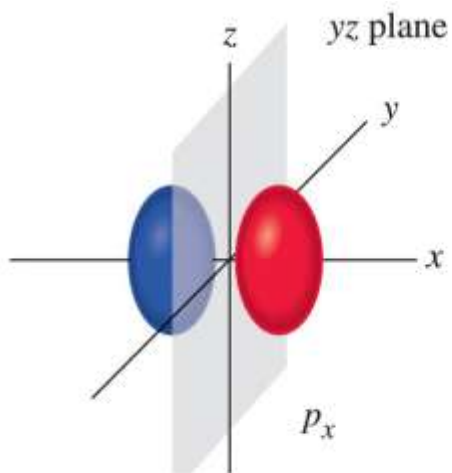
s orbitals



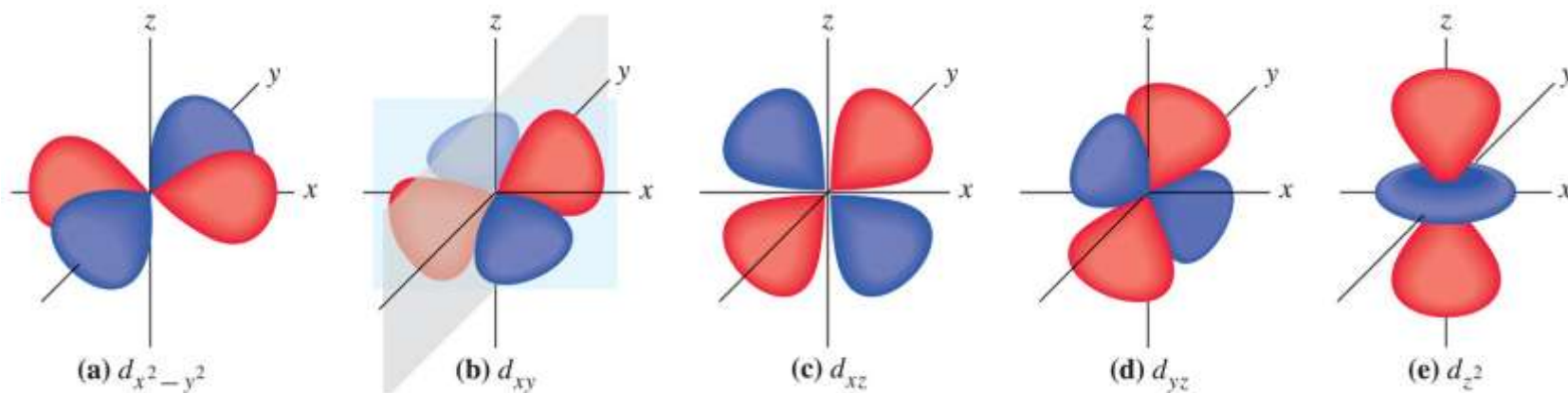
p Orbitals



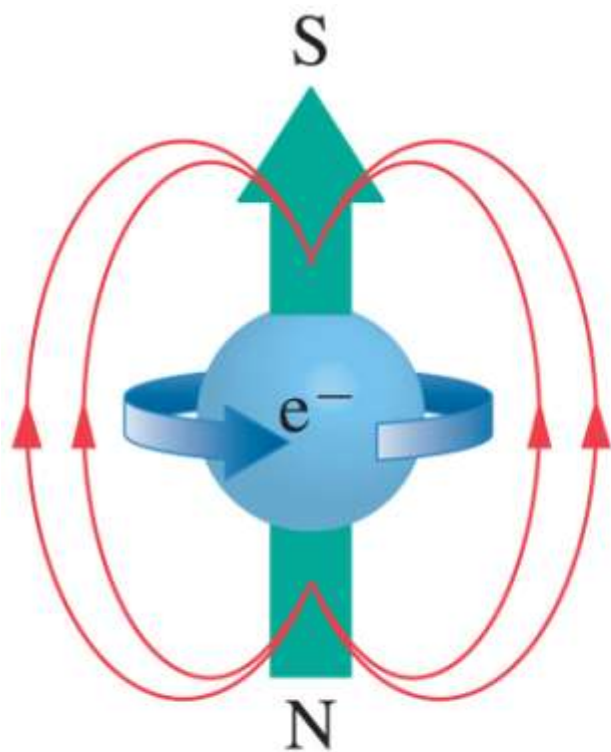
p Orbitals



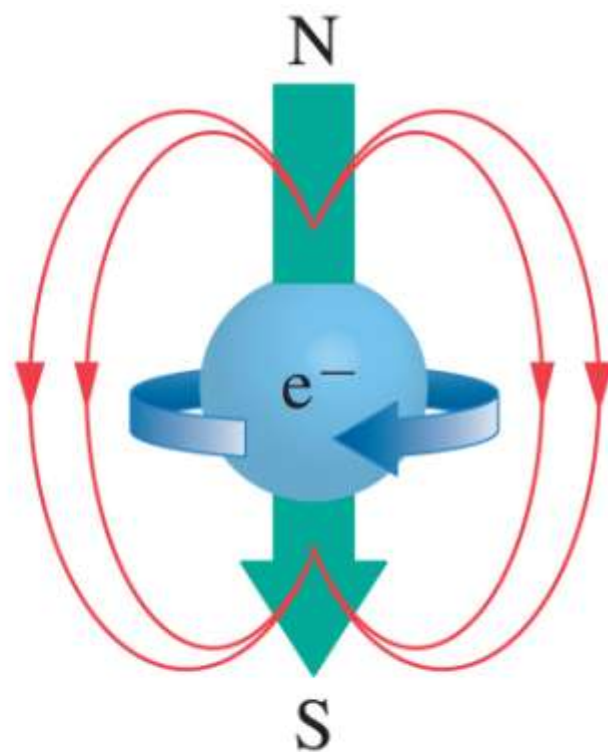
d Orbitals



9-9 Electron Spin: A Fourth Quantum Number



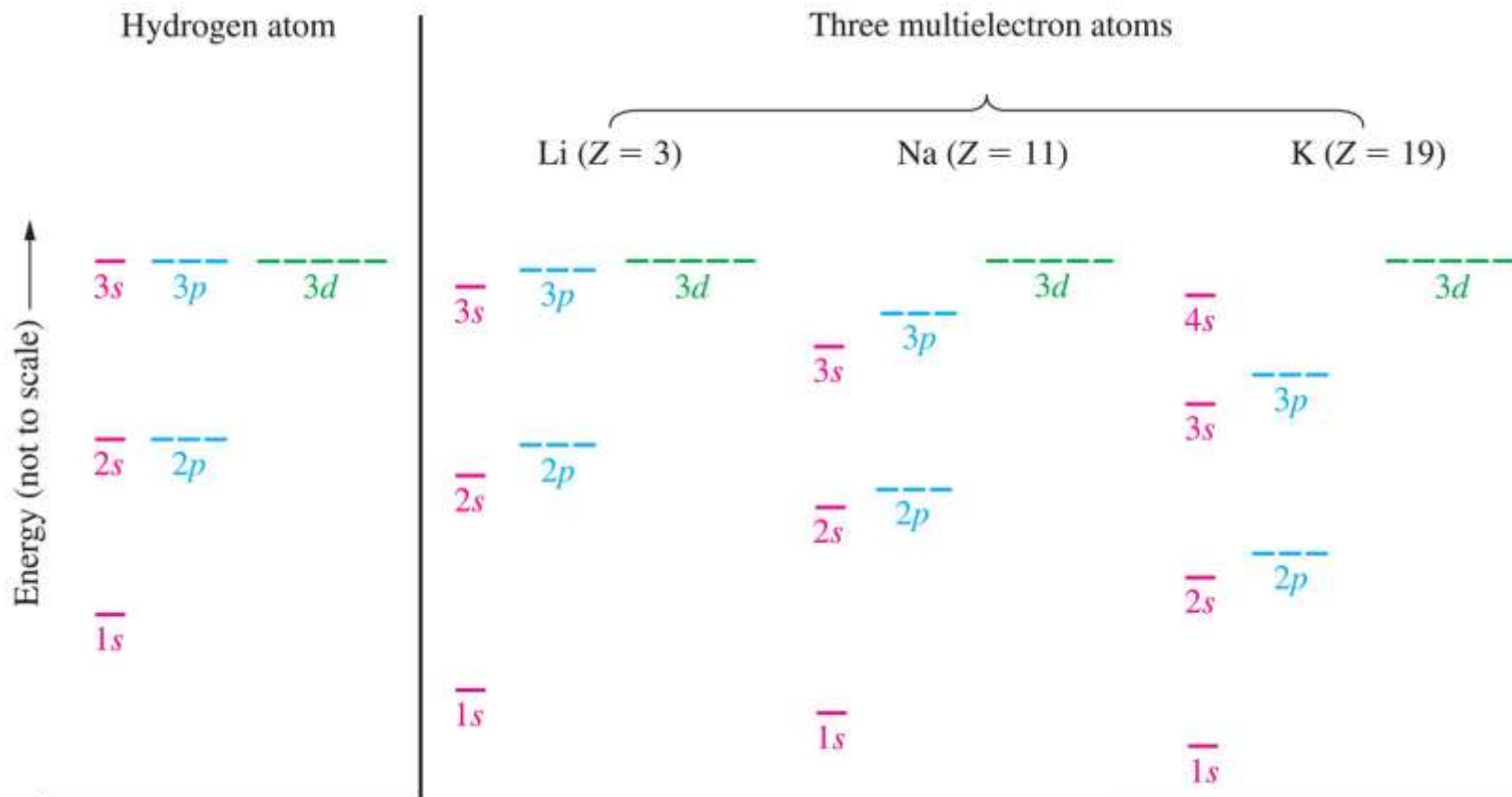
$$m_s = +\frac{1}{2}$$



$$m_s = -\frac{1}{2}$$

8-10 Multi-electron Atoms

- Schrödinger equation was for only one e^- .
- Electron-electron repulsion in multi-electron atoms.
- Hydrogen-like orbitals (by approximation).

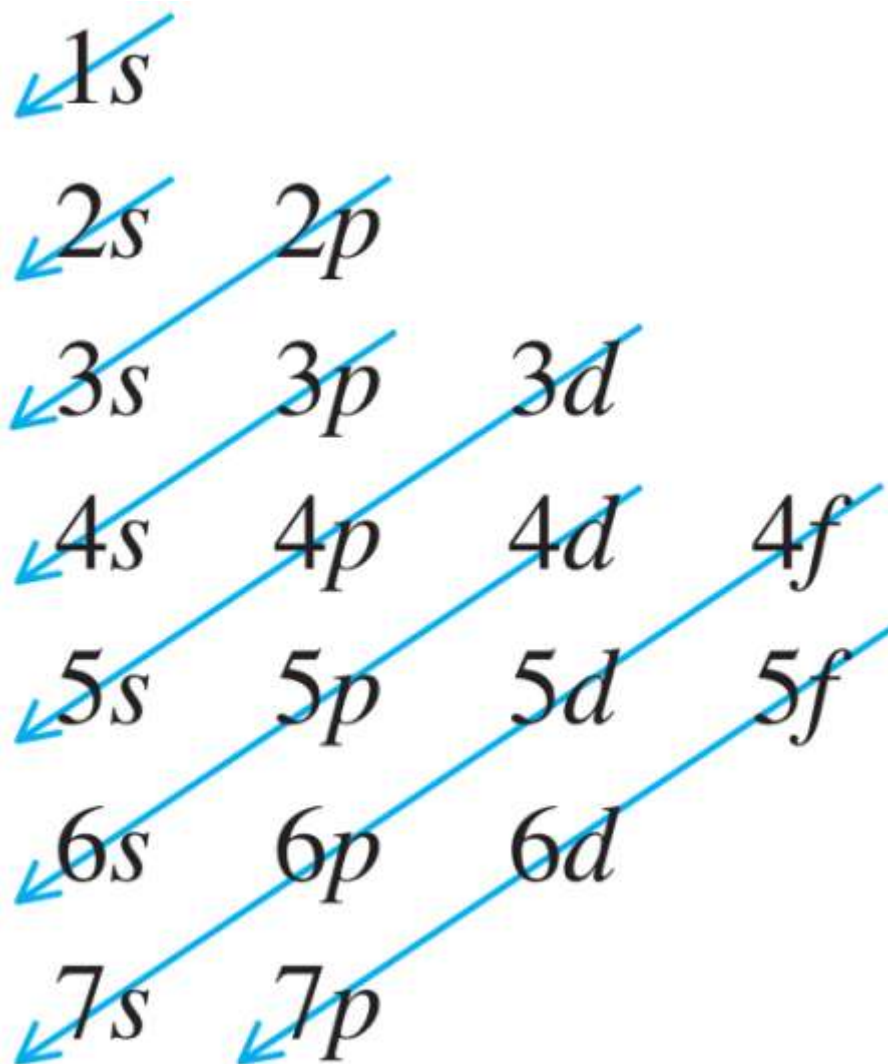


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Orbital energy-level diagram for the first three electronic shells

8-11 Electron Configurations

- **Aufbau process**
 - Electrons occupy orbitals in a way that minimizes the energy of the atom.
- **Pauli exclusion principle**
 - No two electrons can have all four quantum numbers alike.
- **Hund's rule**
- When orbitals of identical energy (degenerate orbitals) are available, electrons initially occupy these orbitals singly.



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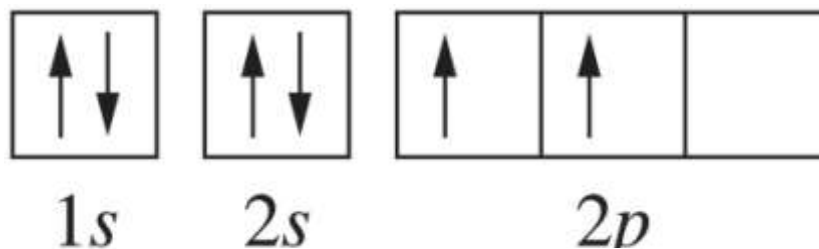
The order of filling of electronic subshells

Representing Electron Configurations

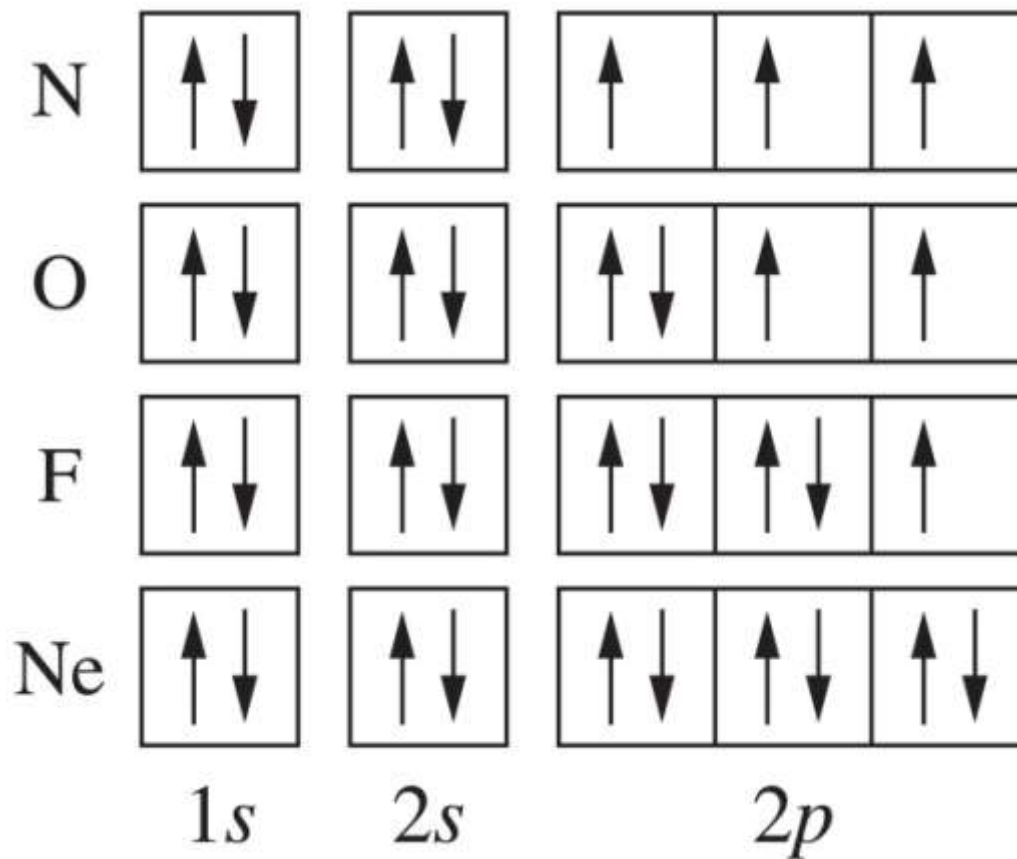
spdf notation (condensed) $1s^2 2s^2 2p^2$

spdf notation (expanded) $1s^2 2s^2 2p_x^1 2p_y^1$

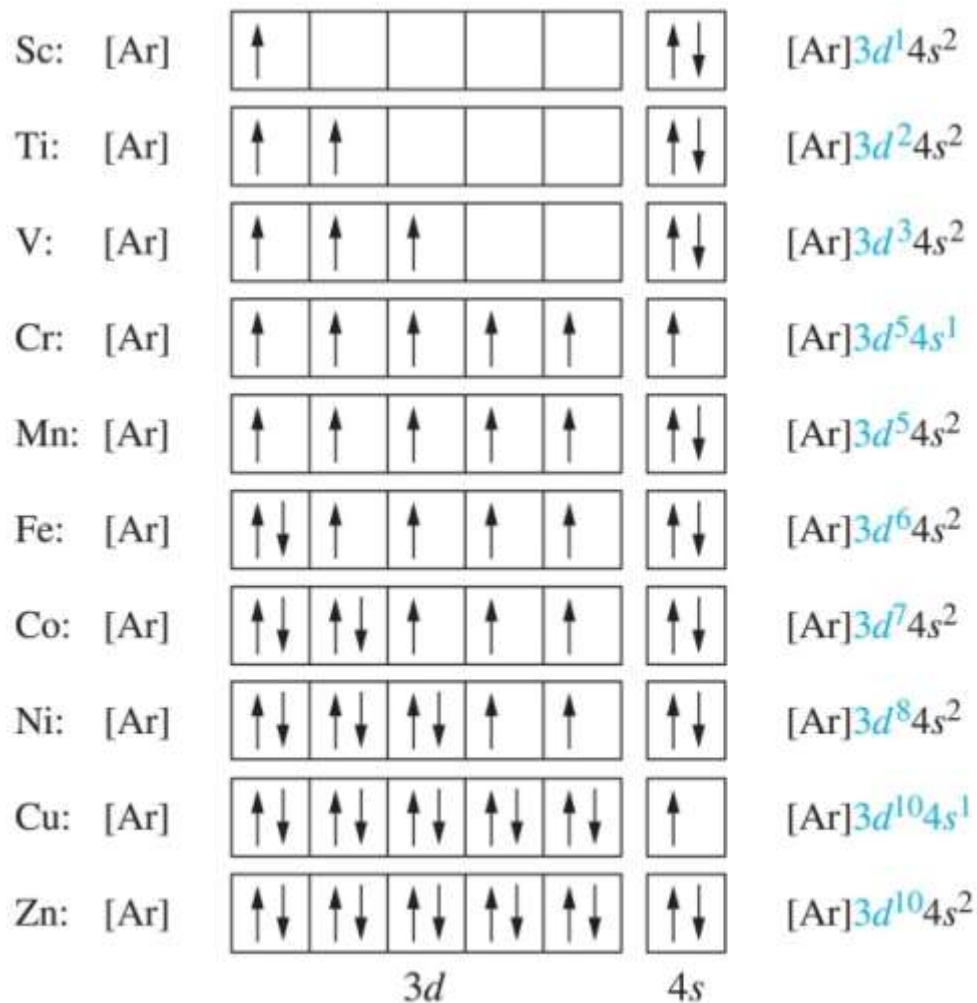
spdf notation



The Aufbau process



The Aufbau Process – Sc through Zn



Electron Configurations and the Periodic Table

Main-group elements

<i>s</i> block																				<i>p</i> block																									
1	2																			13	14	15	16	17	18																				
1s	2s																			3s	4s	5s	6s	7s	2p	3p	4p	5p	6p	1s															
H	He																			Li	Be																			B	C	N	O	F	Ne
		Transition elements																																											
																				<i>d</i> block																									
3	4																			5	6	7	8	9	10	11	12	13	14	15	16	17	18												
11	12	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18																												
Na	Mg	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Al	Si	P	S	Cl	Ar																												
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36																												
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr																												
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54																												
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe																												
55	56	57	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86																												
Cs	Ba	La*	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn																												
87	88	89	104	105	106	107	108	109	110	111	112																																		
Fr	Ra	Ac†	Rf	Db	Sg	Bh	Hs	Mt																																					

Inner-transition elements

														<i>f</i> block													
*														4f													
*	58	59	60	61	62	63	64	65	66	67	68	69	70	71													
*	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu													
†														5f													
†	90	91	92	93	94	95	96	97	98	99	100	101	102	103													
†	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr													

▲ FIGURE 8-38

TABLE 8.2 Electron Configurations of Some Groups of Elements

Group	Element	Configuration
1	H	$1s^1$
	Li	$[\text{He}]2s^1$
	Na	$[\text{Ne}]3s^1$
	K	$[\text{Ar}]4s^1$
	Rb	$[\text{Kr}]5s^1$
	Cs	$[\text{Xe}]6s^1$
	Fr	$[\text{Rn}]7s^1$
17	F	$[\text{He}]2s^22p^5$
	Cl	$[\text{Ne}]3s^23p^5$
	Br	$[\text{Ar}]3d^{10}4s^24p^5$
	I	$[\text{Kr}]4d^{10}5s^25p^5$
	At	$[\text{Xe}]4f^{14}5d^{10}6s^26p^5$
18	He	$1s^2$
	Ne	$[\text{He}]2s^22p^6$
	Ar	$[\text{Ne}]3s^23p^6$
	Kr	$[\text{Ar}]3d^{10}4s^24p^6$
	Xe	$[\text{Kr}]4d^{10}5s^25p^6$
	Rn	$[\text{Xe}]4f^{14}5d^{10}6s^26p^6$

8-12 Electron Configurations and the Periodic Table

Main-group elements

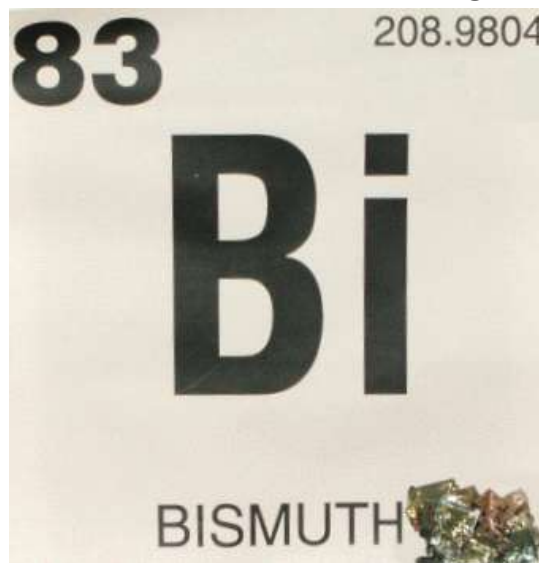
s block																				p block					
1																				18					
1	2																			13	14	15	16	17	18
(1s) → H																						(2p) →			(1s) → He
3	4																			5	6	7	8	9	10
(2s) → Li	Be																			B	C	N	O	F	Ne
11	12																			13	14	15	16	17	18
(3s) → Na	Mg																			Al	Si	P	S	Cl	Ar
		Transition elements																							
		d block																							
19	20	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18								
(4s) → K	Ca	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36								
		(3d) →																		Ga	Ge	As	Se	Br	Kr
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54								
(5s) → Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	49	50	51	52	53	54								
		(4d) →																		In	Sn	Sb	Te	I	Xe
55	56	57	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86								
(6s) → Cs	Ba	La*	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	81	82	83	84	85	86								
		(5d) →																		Tl	Pb	Bi	Po	At	Rn
87	88	89	104	105	106	107	108	109	110	111	112														
(7s) → Fr	Ra	Ac†	Rf	Db	(6d) → Sg	Bh	Hs	Mt																	

Inner-transition elements

f block													
58	59	60	61	62	63	64	65	66	67	68	69	70	71
						(4f) →							
* Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
90	91	92	93	94	95	96	97	98	99	100	101	102	103
† Th	Pa	U	Np	Pu	Am	(5f) →	Bk	Cf	Es	Fm	Md	No	Lr

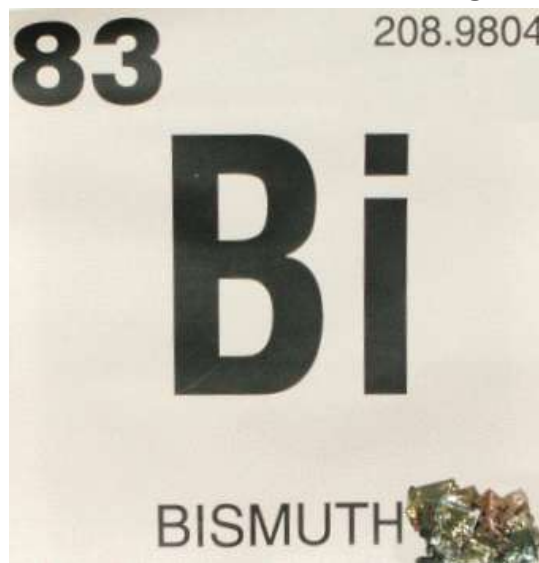
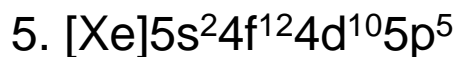
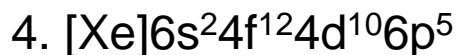
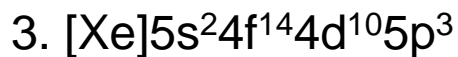
Which of the following is the condensed electron configuration for bismuth?
 [Xe] accounts for 54 electrons.

1. [Xe]6s²4f¹⁴5d¹⁰6p³
2. [Xe]6s²6f¹⁴6d¹⁰6p³
3. [Xe]5s²4f¹⁴4d¹⁰5p³
4. [Xe]6s²4f¹²4d¹⁰6p⁵
5. [Xe]5s²4f¹²4d¹⁰5p⁵



1 1A												18 8A					
1 H 1.00794	2 He 4.00260											13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
3 Li 6.941	4 Be 9.01218											5 B 10.811	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.1797
11 Na 22.9898	12 Mg 24.3050	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 Al 26.9815	14 Si 28.0855	15 P 30.9738	16 S 32.066	17 Cl 35.4527	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.9961	25 Mn 54.9381	26 Fe 55.847	27 Co 58.9332	28 Ni 58.693	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.906	46 Pd 106.42	47 Ag 107.868	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.757	52 Te 127.60	53 I 126.904	54 Xe 131.29
55 Cs 132.905	56 Ba 137.327	57 *La 138.906	72 Hf 178.49	73 Ta 180.948	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.967	80 Hg 200.59	81 Tl 204.383	82 Pb 207.2	83 Bi 208.980	84 Po (209)	85 At (210)	86 Rn (222)

Which of the following is the condensed electron configuration for bismuth?
 [Xe] accounts for 54 electrons.

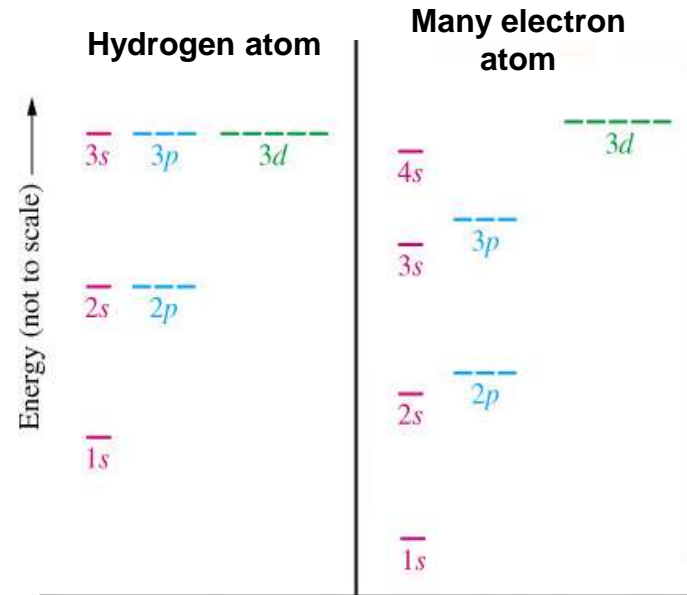


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11 Na 22.9898	12 Mg 24.3050	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 Al 26.9815	14 Si 28.0855	15 P 30.9738	16 S 32.066	17 Cl 35.4527	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.9961	25 Mn 54.9381	26 Fe 55.847	27 Co 58.9332	28 Ni 58.693	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
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55 Cs 132.905	56 Ba 137.327	57 *La 138.906	72 Hf 178.49	73 Ta 180.948	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.967	80 Hg 200.59	81 Tl 204.383	82 Pb 207.2	83 Bi 208.980	84 Po (209)	85 At (210)	86 Rn (222)

- The similar chemical behavior of the elements in a given group in the periodic table is best accounted for by the fact that atoms of these elements have:
 - A) the same number of isotopes
 - B) the same number of electrons
 - C) the same number of electrons in the outermost shell
 - D) similar nuclear structures
 - E) the same number of protons

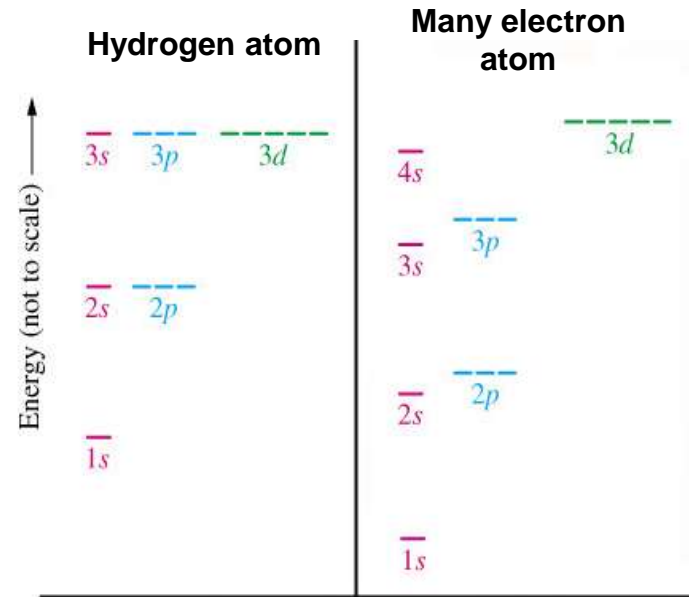
In the hydrogen atom, the orbitals in a given shell are degenerate. For multiple electron atoms what is the main reason for the splitting of, say, the 2s and 2p orbital energies?

1. Higher nuclear charge.
2. The greater distance of the electrons from the nucleus.
3. Penetration of the nucleus by the 2s electron.
4. The presence of an angular node for the 2p orbital.



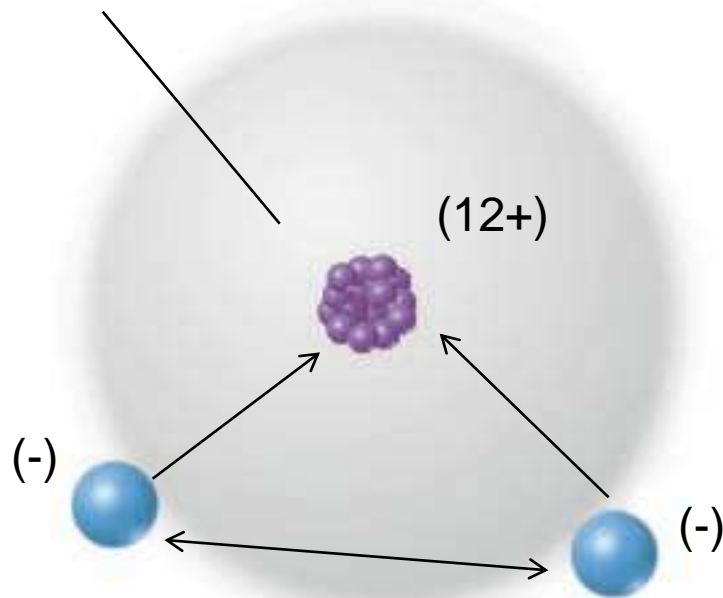
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1. Higher nuclear charge.
2. The greater distance of the electrons from the nucleus.
3. Penetration of the nucleus by the 2s electron.
4. The presence of an angular node for the 2p orbital.



The shielding effect and effective nuclear charge, Z_{eff}

Screen of electron
charge from 10 core
electrons



The core electrons shield or screen the outer shell electrons from the full attractive force of the nucleus.

An effective nuclear charge is defined as the true nuclear charge minus the charge that is screened out by the electrons: $Z_{\text{eff}} = Z - S$