Chapter 6

Electronic Structure of Atoms

The number & arrangement of e^- in an atom is responsible for its chemical behavior

I) The Wave Nature of Light

A) Electromagnetic Radiation

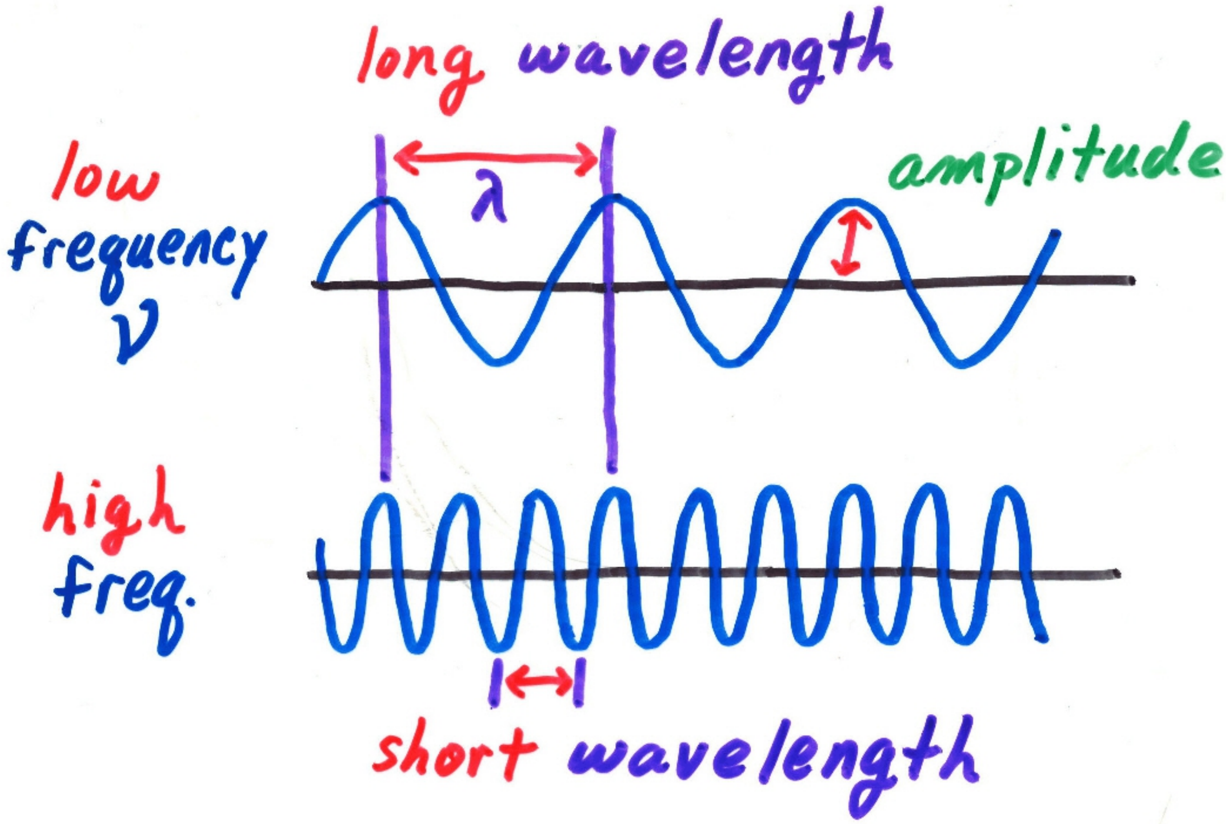
Radiant Energy

light, X-rays, UV, microwaves, etc.

All move at the speed of light,

$$\mathbf{c} = 2.99792 \times 10^8 \text{ m/s}$$

have wavelike characteristics



λ, wavelength	distance between successive peaks
ν, frequency	number of complete wavelengths or cycles which pass a given point per second
amplitude	height of peak - related to intensity of radiation

long wavelength → low frequency

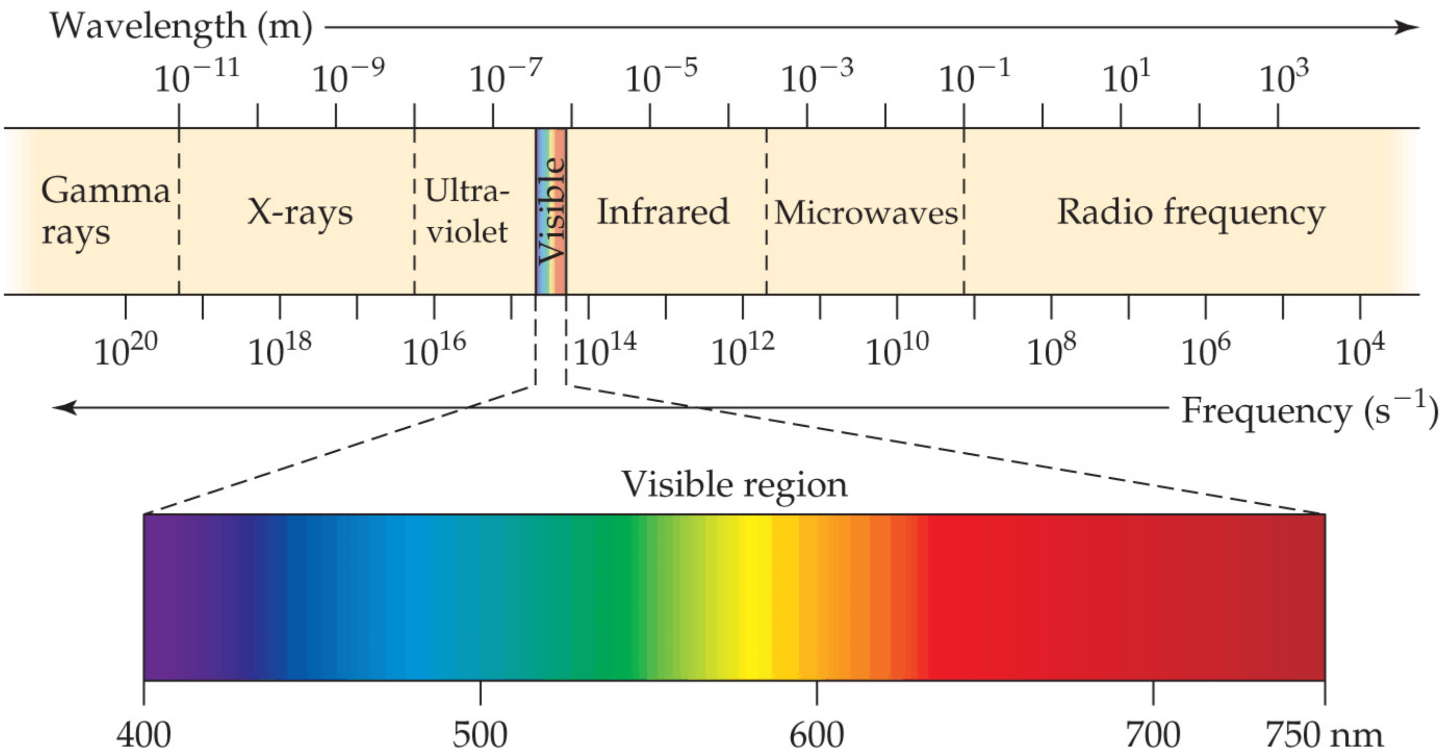
short wavelength → high frequency

$$\nu \propto \frac{1}{\lambda}$$

$$\nu = \frac{c}{\lambda} \quad \text{or} \quad \nu \cdot \lambda = c$$

<u>units for ν</u>

s^{-1} ; cycles/s ; hertz, Hz



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	X- rays	visible	IR	microwave	radio
λ (m)	10^{-9}	10^{-7}	10^{-5}	10^{-2}	10^{2}
ν (s^{-1})	10^{17}	10^{15}	10^{13}	10^{10}	10^{6}

II) Quantized Energy and Photons

A) Plank's Theory

Energy changes are quantized

- discrete energy changes

$$\Delta E = n\,h\,\nu \qquad n = 1, 2, 3, 4, \ldots$$

Planck's constant

$h = 6.63 \times 10^{-34}$ J•s,

Smallest increment of energy, at a given frequency, is termed a quantum of energy

B) Photoelectric Effect

A minimum freq. of light shining on a metal surface causes it to emit e^-

Einstein: energy is a stream of particle like energy packets called photons

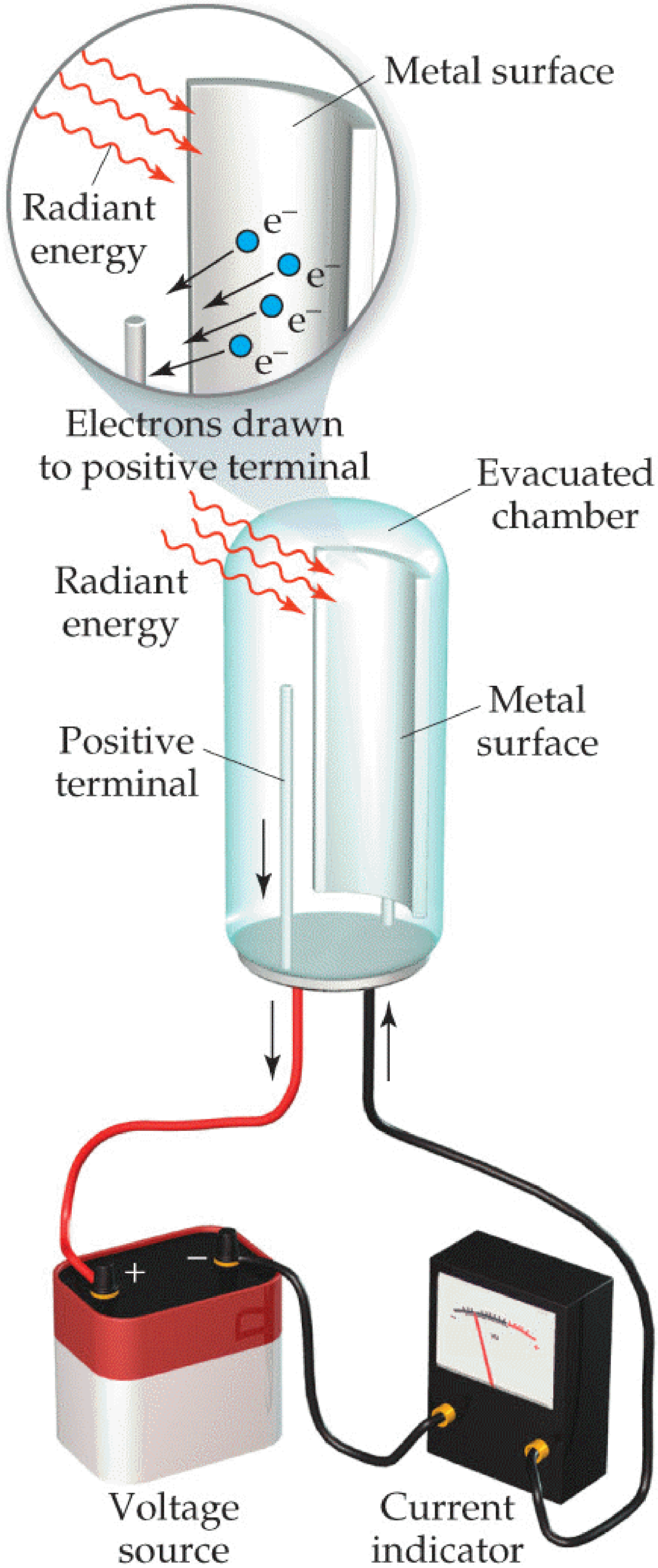
- radiant energy is quantized

$$E_{photon} = h\,\nu = \frac{h\,c}{\lambda}$$

high ν (low λ) $\Rightarrow$ high E

low ν (high λ) $\Rightarrow$ low E

Note : duality of light - behaves both as a wave and particle



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1) Ex: A laser emits a signal with a wavelength of 351 nm. Calculate the energy of a photon of this radiation.

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

$$= \frac{(6.63 \times 10^{-34}\ \text{Js})(3.00 \times 10^{8}\ \text{m/s})}{3.51 \times 10^{-7}\ \text{m}}$$

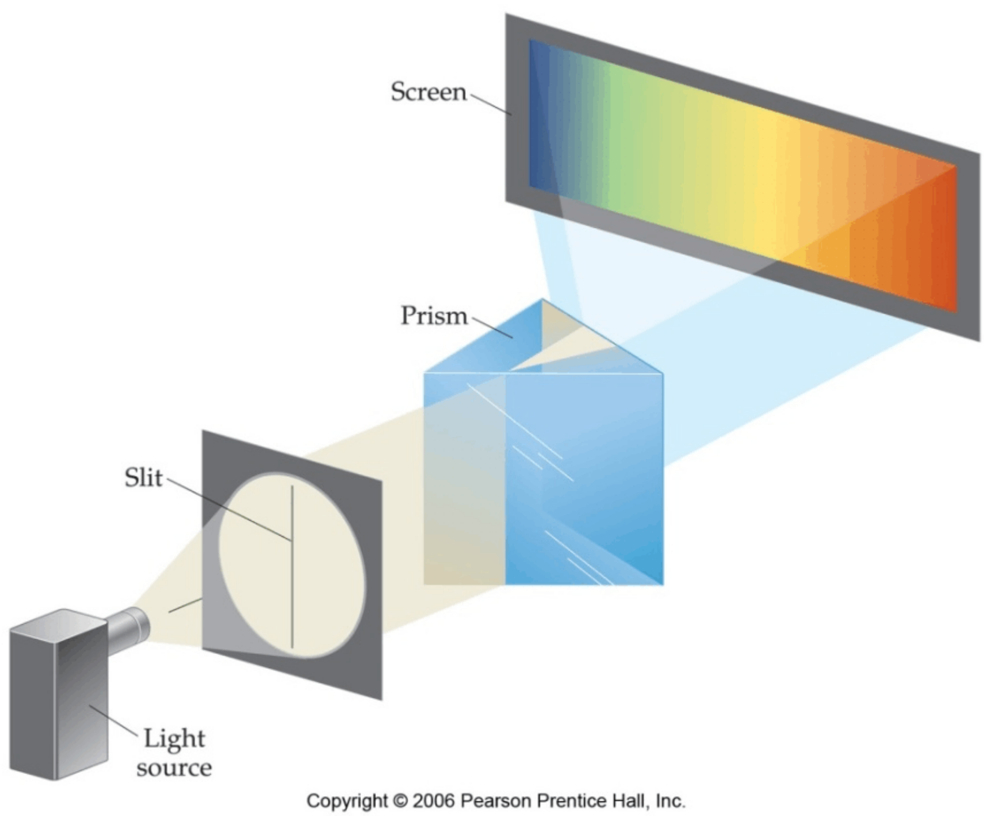
$$= 5.67 \times 10^{-19}\ \text{J}$$

III) Line Spectra and the Bohr Model

A) Line Spectra

1) White light passing through a prism results in band called a

continuous spectrum (rainbow)



2) monochromatic light

Light with a single wavelength

- lasers

3) Line Spectra

discharge tube - atom absorbs energy & it can later emit it as light

Passed through a prism see a series of narrow colored lines (specific λ's)

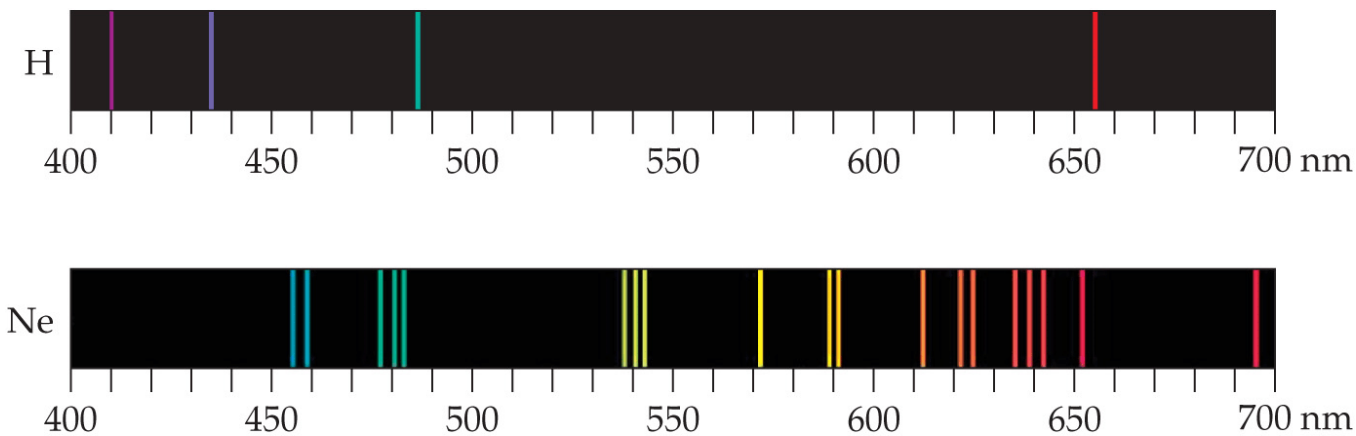
Line Spectrum

Each line associated with a particular energy and color

Different elements give different & distinctive line spectra

- characteristic of a particular element

- use to identify elements



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B) Rydberg Equation

Wavelengths of lines in hydrogen spectrum given by,

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

$$n_2 > n_1$$

Rydberg Constant

$$R_H = 1.097 \times 10^7 \ m^{-1}$$

B) Bohr Model

1) Energy Levels & Orbits

e^- is restricted to certain energy levels corresponding to spherical orbits, w. certain radii, about the nucleus

$$r = n^2 a_0$$

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

n = principle quantum number

$n = 1, 2, 3, \ldots, \infty$

Bohr radius:

$$a_0 = 5.292 \times 10^{-11} \text{ m} = 0.5292 \text{ Å}$$

$$hc \bullet R_H = 2.180 \times 10^{-18} \text{ J}$$

a) Ground State

e^- in $n = 1$ orbit
closest to nucleus

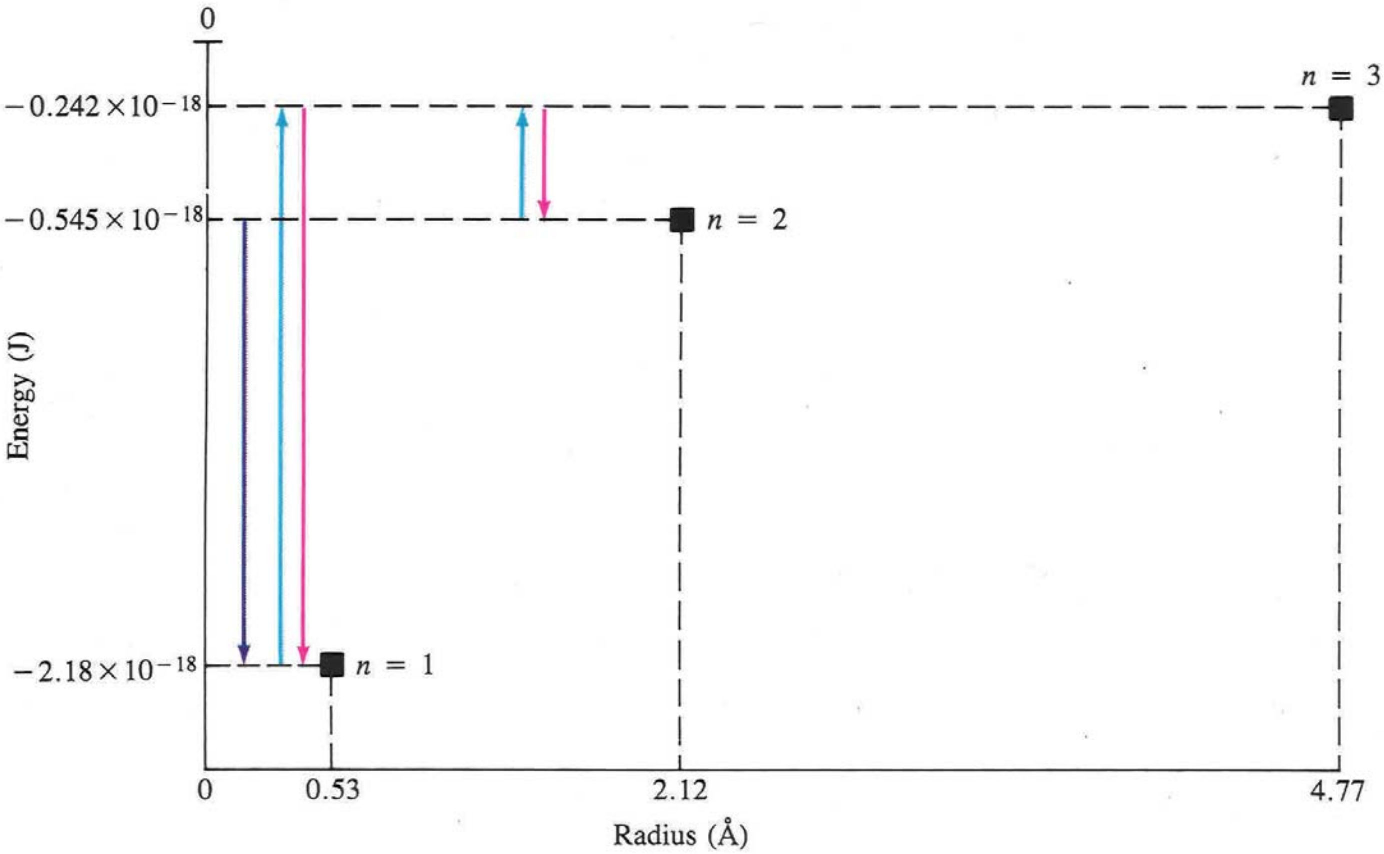
largest value of $1/n^2$

most negative E

* Lowest energy level

Note: most neg. E represents
most stable state

Radii and energies of Bohr orbits 1-3



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b) Excited States

$n > 1$

higher energy

less neg. E, less stable

inc. distance from nucleus

$r \propto n^2$

c) Zero-Point of Energy

$n = \infty$

e^- completely separated from nucleus

$$E_\infty = -hc \cdot R_H \left(\frac{1}{\infty}\right) = 0$$

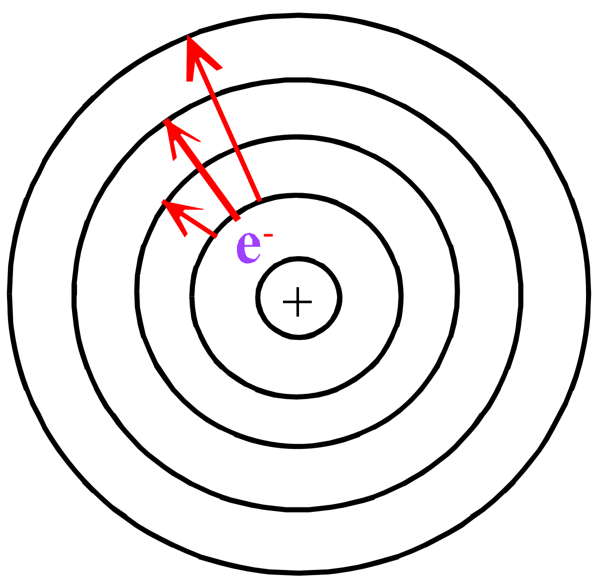
2) Energy Transitions

a) Absorption of Energy

e^- absorbs energy

- jumps to higher energy levels, farther from nucleus

- Excited State

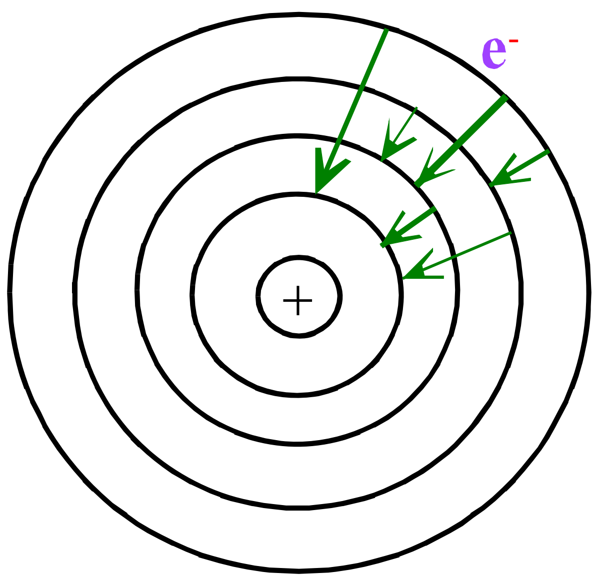


b) Emission of Energy - light

e^- "falls" to lower level

- emits the energy diff. as a quantum of light,

a photon



$$E_{photon} = -\Delta E_{emission} = h\,\nu = \frac{h\,c}{\lambda}$$

c) Energy Changes

Energy diff. between orbits

$$\Delta E = E_f - E_i = \frac{-h\,\mathbf{c} \bullet R_H}{n_f^2} - \frac{-h\,\mathbf{c} \bullet R_H}{n_i^2}$$

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

$$\Delta E = -2.180 \times 10^{-18}\ \mathrm{J} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

1) $n_f > n_i$

$\Delta E > 0$, E inc.

Absorption

2) $n_f < n_i$

$\Delta E < 0$, E dec.

Emission

3) $n_f = \infty$

complete removal of e^-

Ionization

$$H(g) \longrightarrow H^+(g) + e^-$$

$n_i = 1$ $n_f = \infty$

$$\Delta E = h\,c \bullet R_H \left(\frac{1}{1^2}\right) = 2.180 \times 10^{-18}\ J$$

d) Energy of a Photon

Energy of a photon emitted when e^- "drops" to a lower energy level is related to freq. (wavelength) of radiation

$$E_{photon} = -\Delta E_{em} = h\nu = \frac{hc}{\lambda}$$

$$\nu = c \cdot R_H \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

or

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

e) Ex : Calc. the wavelength of a line in the visible spectrum for which $n_i = 3$.

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

Balmer Series (visible):

$$n_f = 2$$

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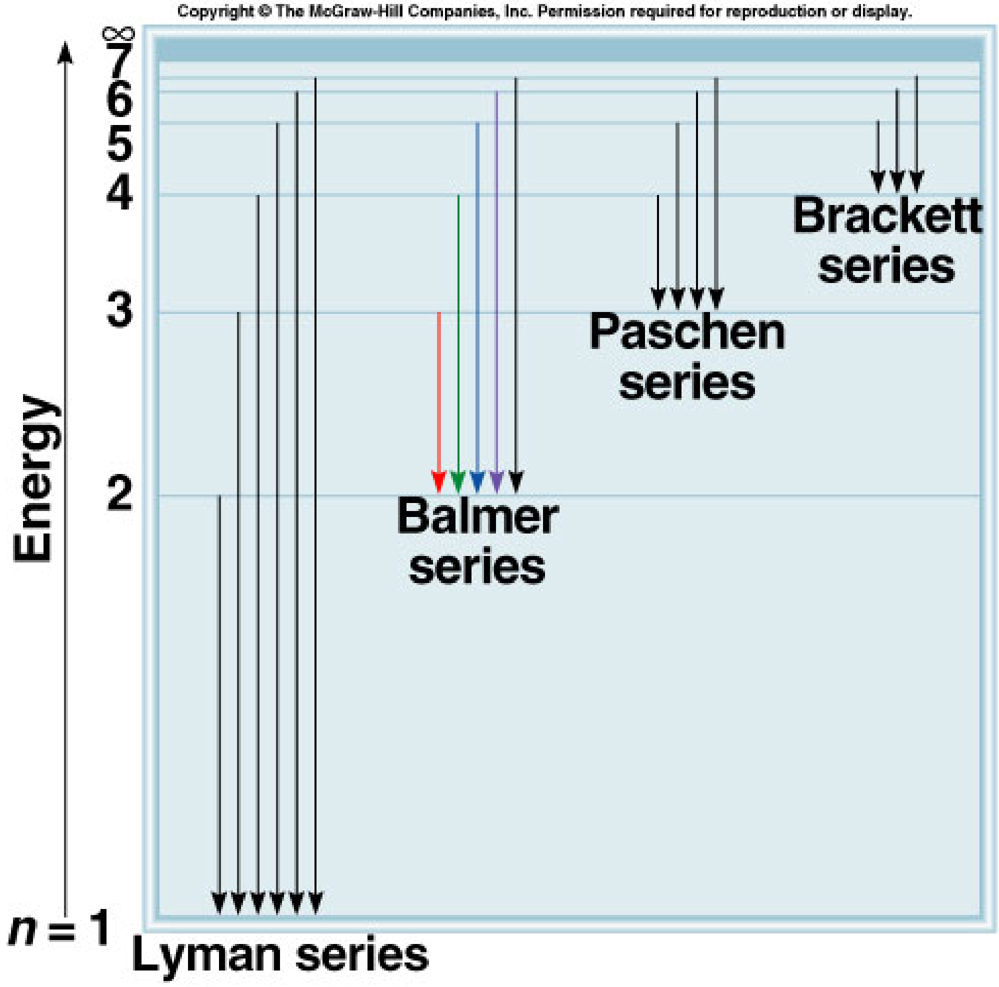


IV) Wave Behavior of Matter

A) de Broglie

Matter should have wave prop.

For photons:

$$E_{photon} = h\,\nu = \frac{h\,c}{\lambda}$$

From Einstein:

$$E = m\,c^2$$

$$\lambda = \frac{h}{m\,c}$$

wavelength for photon traveling at c with an effective mass, m

B) de Broglie Wavelength for Particles

$$\lambda = \frac{h}{m\,v}$$

v = velocity of the particle

h (6.63×10^{-34} J•s) is extremely small so λ is too small for macroscopic particles.

λ can only be detected for particles w. very small mass,

i.e. e^- ($m = 9.11 \times 10^{-28}$ g)

1) Ex 1: Calculate the de Broglie wavelength for a 907.2 kg car moving at a speed of 96.6 km/hr.

$$\lambda = \frac{(6.63 \times 10^{-34}\ \text{J}\bullet\text{s})}{(907.2\ \text{kg})\ (26.\underline{\mathbf{8}}3\ \text{m/s})}$$

$$= 2.72 \times 10^{-38}\ \text{m}$$

2) Ex 2: Calculate the de Broglie wavelength for an electron moving at a speed of 3×10^6 m/s.

$$\lambda = \frac{(6.63 \times 10^{-34}\ \text{J}\bullet\text{s})}{(9.11 \times 10^{-31}\ \text{kg})\ (3 \times 10^{6}\ \text{m/s})}$$

$$= 2.43 \times 10^{-10}\ \text{m} \qquad (0.243\ \text{nm})$$

X-rays

C) Heisenberg Uncertainty Principle

The wave-particle duality of matter makes it impossible to precisely measure both the position and momentum of an object.

Δx = uncertainty in position

Δp = uncertainty in momentum (mv)

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

Limit on simultaneously measuring position and momentum (speed).

V) Quantum Mechanics

Impose wave properties on e^-

A) Schrödinger's Wave Equation

Total energy of H-atom is sum of K.E. and P.E.

Time-Independent Sch. Eqn.:
(in one dimension)

$$-\frac{\hbar^2}{2m}\frac{d^2\psi(x)}{dx^2} + V(x)\psi(x) = E\psi(x)$$

K.E.   P.E.   Total E

$\hbar = h/2\pi$

1) Wave Functions

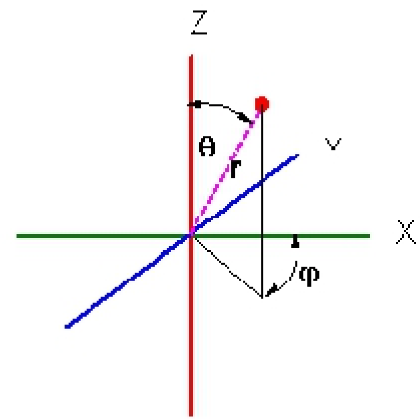
Get a series of solutions to the wave eqn.

wave functions, ψ

Each ψ corresponds to a specific energy & describes a region about the nucleus, an orbital, in which an e^- w. that energy may be found

ψ has no direct physical meaning

The general equation for ψ for H in polar coordinates is:



Polar coordinate system

$$\Psi_{(n,l,m,r,\theta,\phi)} = \sqrt{\frac{(n-l-1)!}{n((n+l)!)^3}}\left(\frac{1}{na_0}\right)^{\frac{3}{2}+l} r^l e^{\left(\frac{-r}{na_0}\right)}(n+l)!\left(\sum_{i=0}^{n-l-1}\frac{(-1)^i (n=l)!\left(\frac{2r}{na_0}\right)^i}{(n-l-1-i)!(2l+1+i)!i!}\right)\bullet$$

$$\sqrt{(2l+1)(l-|m|)!(l+|m|)!}\,\sin(\theta)^{|m|}\left(\sum_{i=0}^{l-|m|}\frac{(l!)^2(\cos\theta-1)^{(l-|m|-i)}(\cos\theta+1)^i}{(l-i)!i!(l-|m|-i)!(|m|+i)!}\right)e^{\mathrm{Im}\,\phi}\left(\frac{1}{\sqrt{\pi}\,l!}\right)$$

Can only determine the probability of finding e^- in a certain region of space at a given instant,

ψ^2 probability density

Electron density

Greater where e^- spends more of its time.

Probability of finding an e^- is high in regions of high e^- density

B) Orbitals & Quantum Numbers

ψ represents an orbital and has 3 characteristic quantum numbers associated with it,

n	ℓ	m_ℓ
energy and distance from nucleus	shape	orientation of an orbital

The first 3 arise naturally from the solution of the Sch. Eqn.

There is a 4th quantum no.

m_s : spin

1) Principal quantum number, n

Determines:

- energy level
- average distance from nucleus
- Identifies the shell

$n = 1, 2, 3, 4, ...$

Larger n ⇛ farther shell is from nucleus & higher energy

Max. no. of e^- in shell = $2n^2$

$n = 1$	$2(1)^2$	$= 2\ e^-$
$n = 2$	$2(2)^2$	$= 8\ e^-$
$n = 3$	$2(3)^2$	$= 18\ e^-$

2) Azimuthal q. n. , ℓ

(Angular Momentum q.n.)

identifies subshell (energy sublevels)

defines shape of orbital

subshells in a shell = n

ℓ = 0, 1, 2, ... (n-1)

Subshell designated by letters:

ℓ =	0	1	2	3	4	....
	s	p	d	f	g	
$\#e^-$ in subshell	2	6	10	14	18	
$2(2\ell + 1)$						

If $n = 4$ ℓ = 0, 1, 2, 3

4s 4p 4d 4f

n^{th} shell (n)	no. of subshells (= n)	designation	max # e^- by subshell	total # e^- in shell ($2n^2$)
1	1	1s	2	2
2	2	2s, 2p	2+6	8
3	3	3s, 3p, 3d	2+6+10	18
4	4	4s, 4p, 4d, 4f	2+6+10+14	32

3) Magnetic q.n., m_ℓ

Describes orientation of orbital in space

$m_\ell = +\ell, ..., 0, ..., -\ell$

integer values from $+\ell$ to $-\ell$

possible values for m_ℓ = # orbitals in a subshell

$(2\ell + 1)$ orbitals in a subshell

Total # orbitals in shell n = n^2

orbital contains a max. of 2 e^-

max. # e^- in subshell = $2(2\ell + 1)$

a) Ex:

$\ell = 0$ $m_\ell = 0$; s subshell has 1 orbital

ℓ	subshell	# orbitals	max # e^- in subshell
0	s		
1	p		
2	d		
3	f		
4	g		
5	h		

TABLE 6.2 • Relationship among Values of *n, l*, and m_l through *n* = 4

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

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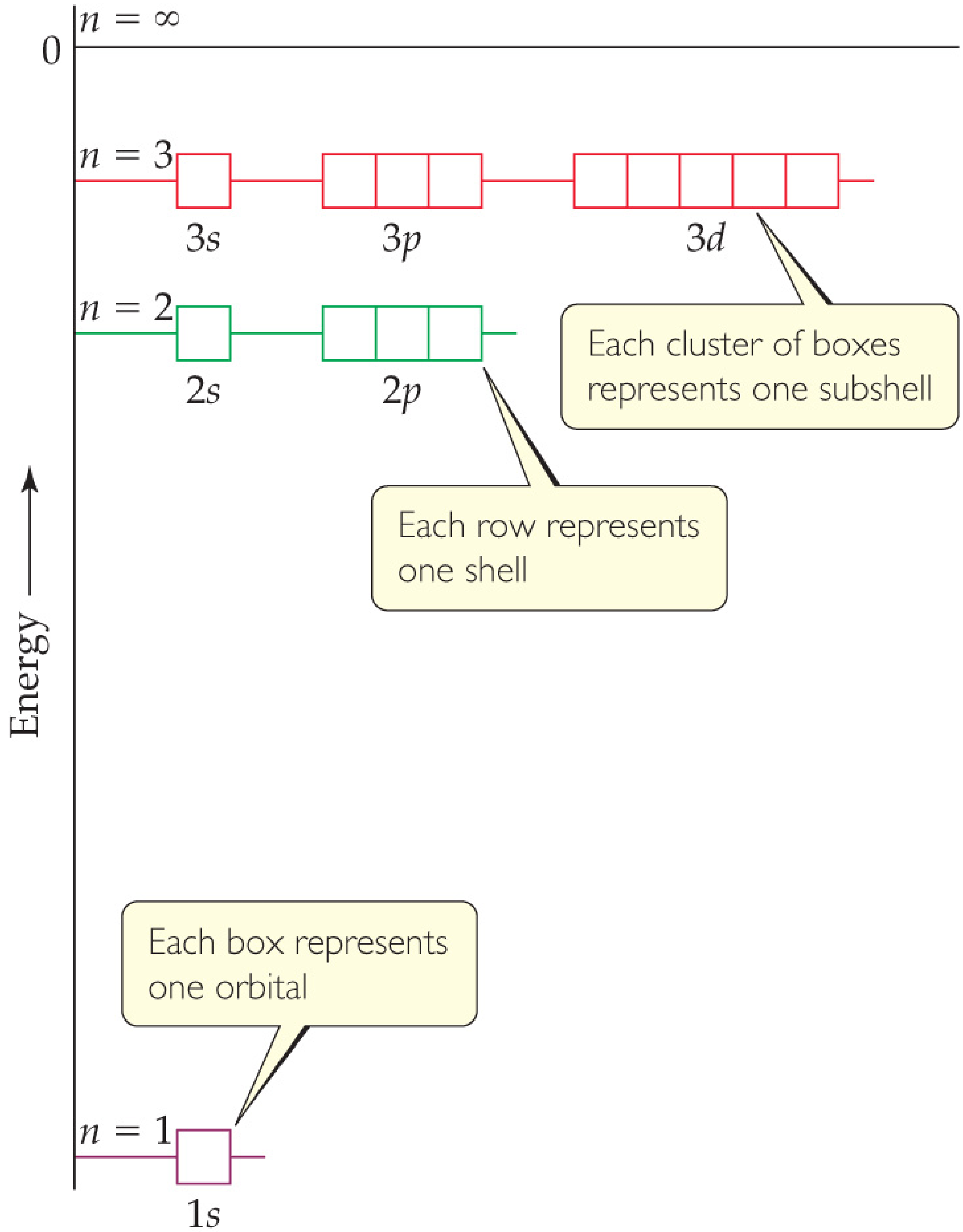
Energy Levels in the H atom



$n = 1$ shell has one orbital
$n = 2$ shell has two subshells composed of four orbitals
$n = 3$ shell has three subshells composed of nine orbitals

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VI) Representations of Orbitals

ψ has no direct physical meaning

ψ^2 probability density
(electron density)

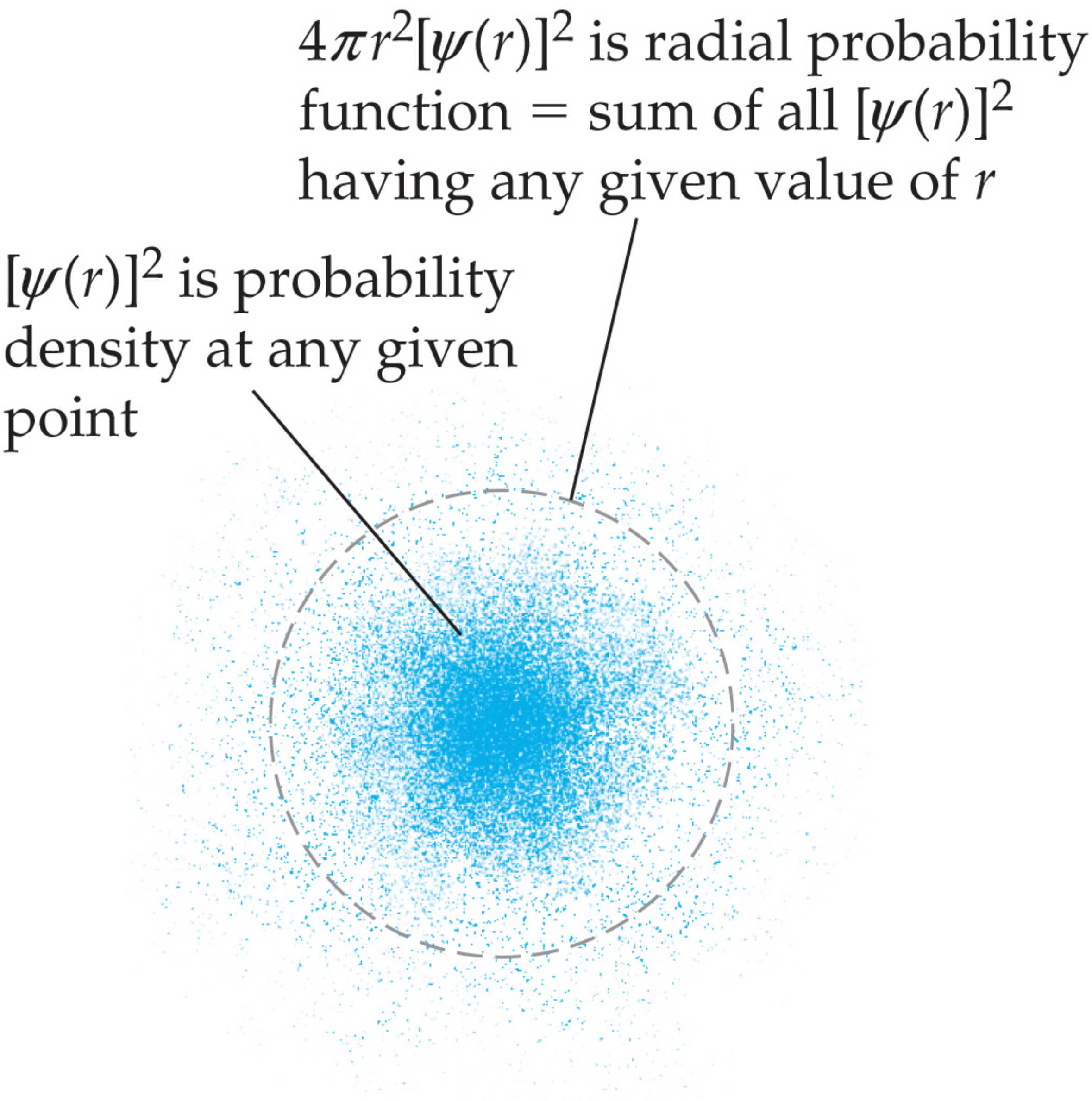
probability of finding e^- at a given point in space

$(4\pi r^2)\,\psi^2$ radial probability density

probability of finding e^- at a specific distance, r, from the nucleus

A) s orbitals

$\ell = 0$ All s orb. are spherical



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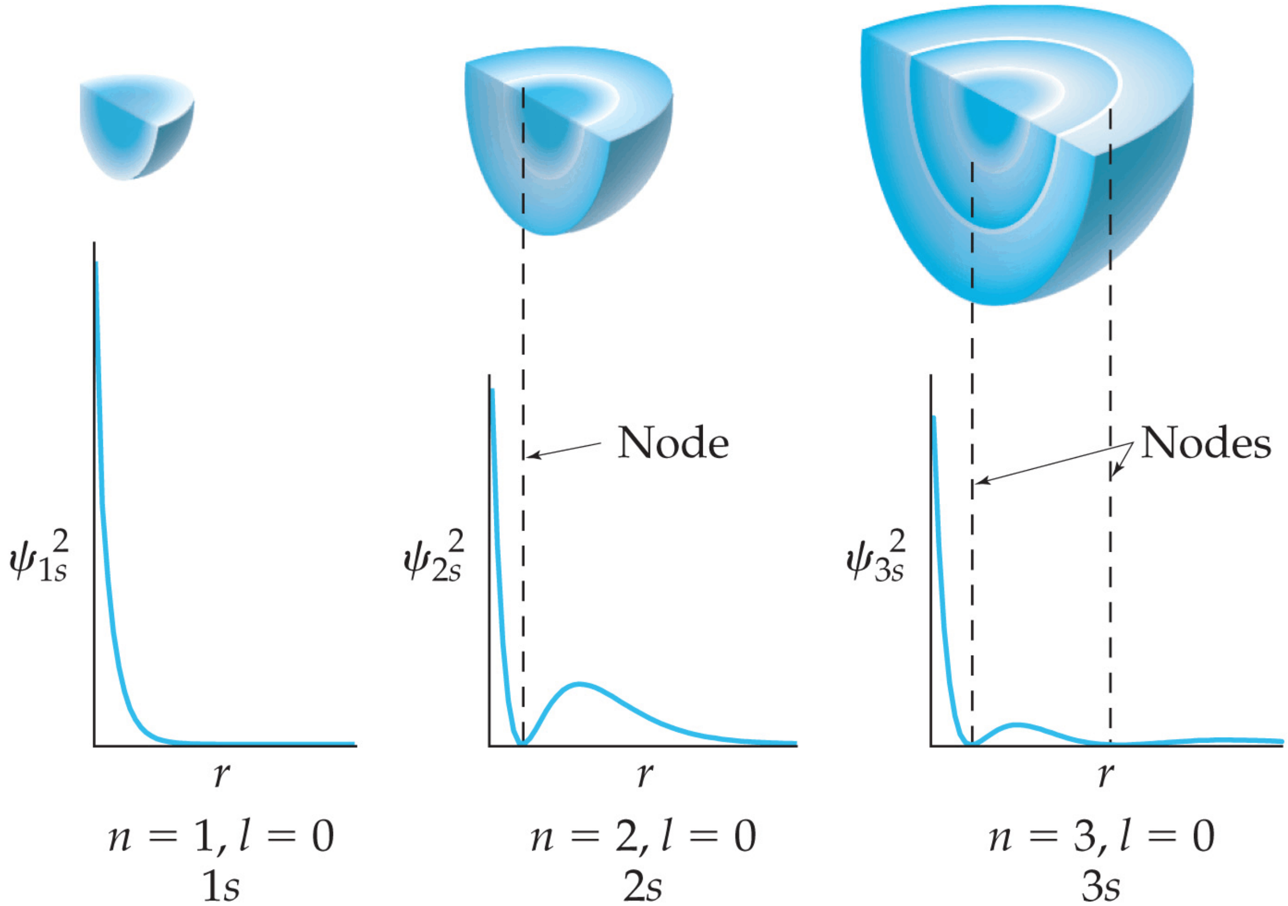

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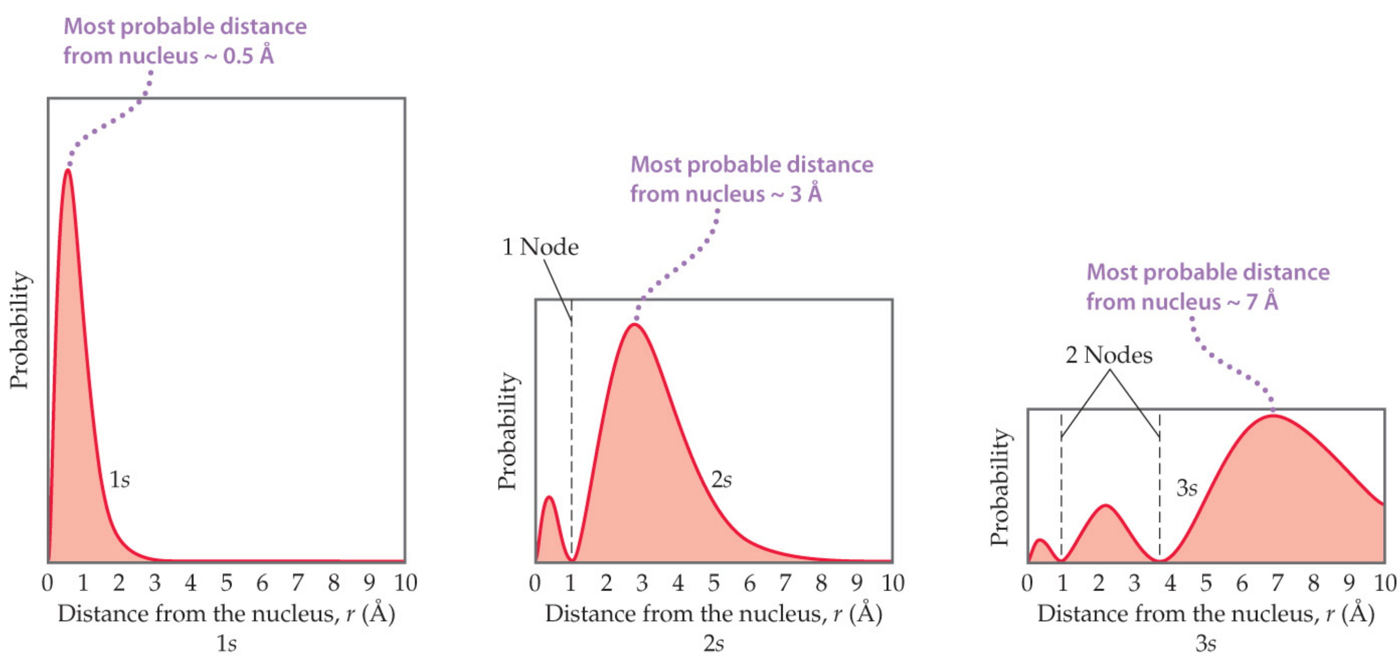

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1) 3 trends from radial prop. dist.

a) Number of peaks inc. w. inc. n

peaks = n

most probable distance further out & peaks get larger as move further from nucleus

b) Number of nodes inc. w. inc. n

points where the prob. is zero

nodes = n − 1

spherical nodes = n − ℓ − 1

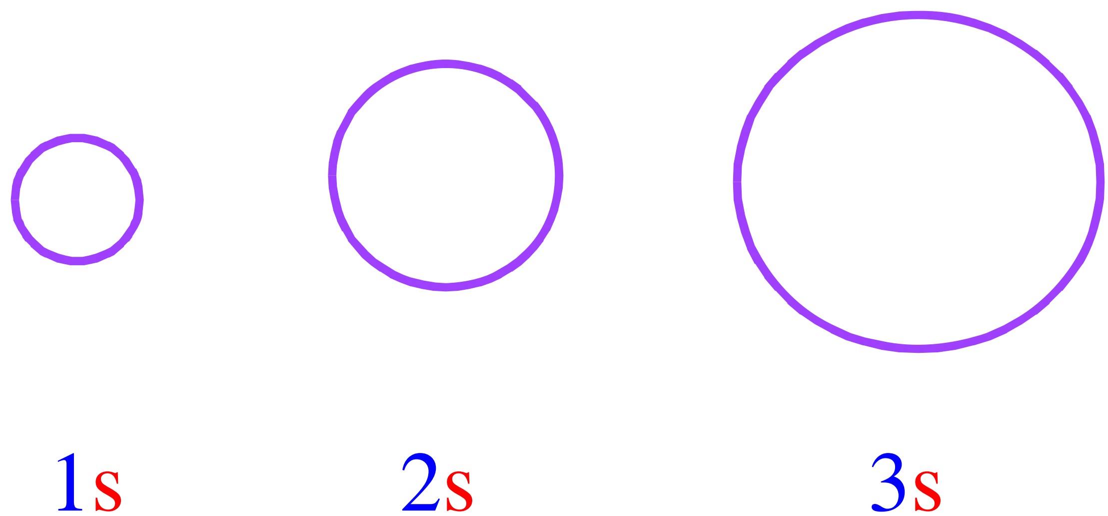
angular nodes = ℓ

c) e^- density spreads out w. inc. n

2) Contour Representation

represent a volume of space in which there is a high probability of finding the e^-

usually 90%



e^- in orb. of higher n will be greater avg. distance from nucleus

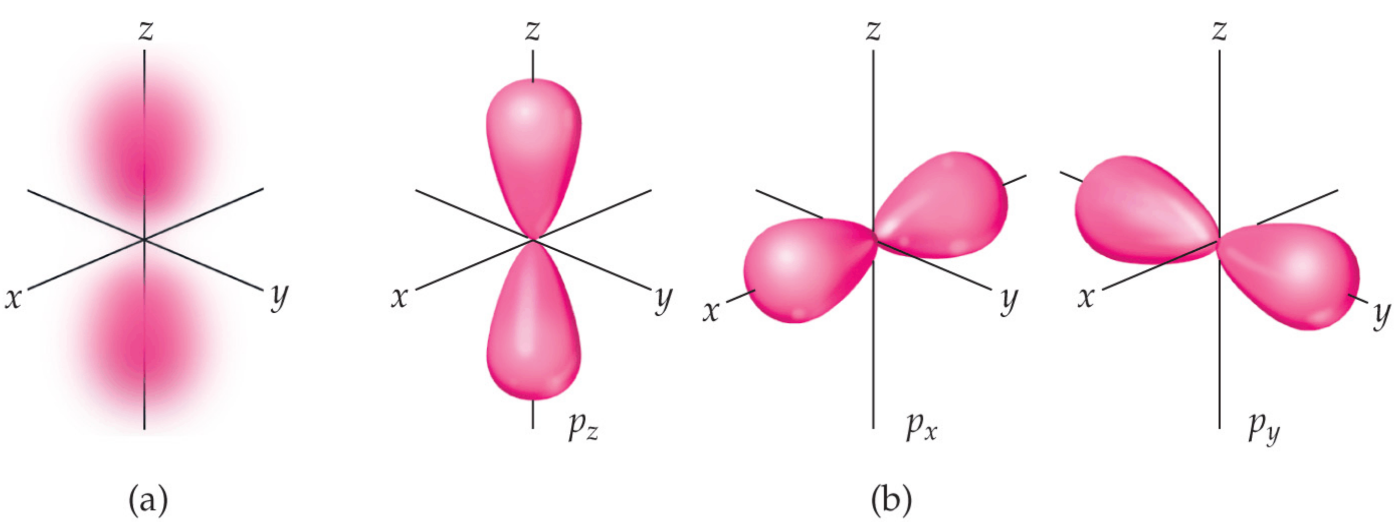
B) p orbitals

All p orbitals have 2 lobes pointing in opposite directions

dumbbell or teardrop

The 3 p orbs in a subshell differ in their orientation in space

- at right angles to each other



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VII) Many-Electron Atoms

H atom has only 1 e^-

E_{orb} depends on n and is determined by attraction between positive proton and negative e^- and average distance between them

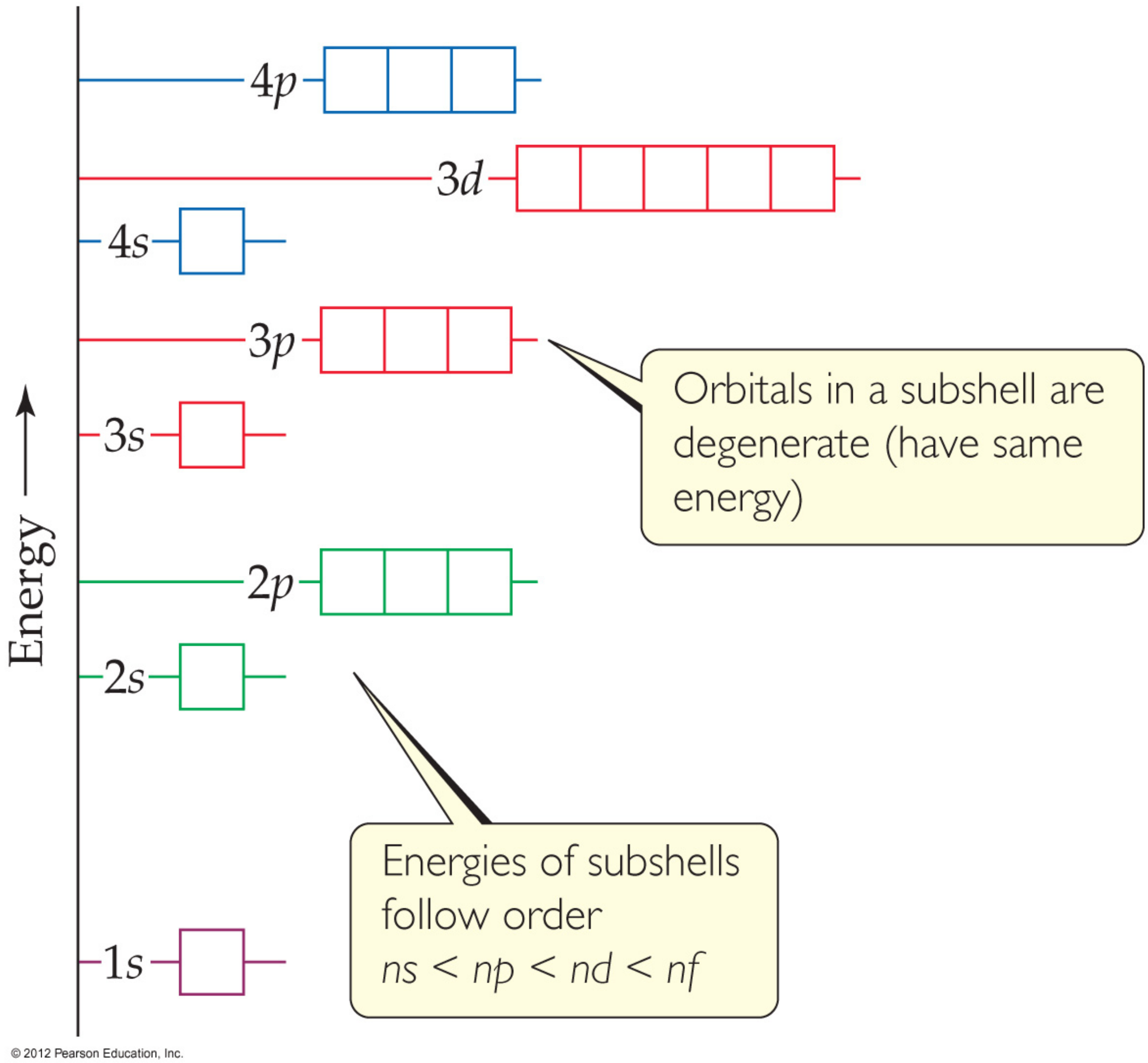
Many-e^- atoms:

Add e^- - e^- repulsions to E & diff. e^--nucleus attractions

Causes subshells to have diff. E

E_{orb} now depends on n and ℓ

E of orbitals w/in subshell still degenerate


Energy
4p
3d
4s
3p
3s
2p
2s
1s
Orbitals in a subshell are degenerate (have same energy)
Energies of subshells follow order
ns < np < nd < nf
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A) Electron Spin

e^- "spins" about its own axis

- spinning charge generates a magnetic field

e^- only spin in either of 2 directions

quantized

electron spin q.n., m_s

+1/2	−1/2
up	down
↿	⇂

B) Pauli Exclusion Principle

No 2 e^- in an atom can have same set of 4 quantum no.'s

$$n, \quad \ell, \quad m_\ell, \quad m_s$$

Look at 1s orbital

$n = 1$, $\ell = 0$, $m_\ell = 0$

can have only 2 e^- w. diff. values of m_s, $+1/2$ or $-1/2$

Limits max . # e^- in orbital to 2

- MUST have opposite spins

C) Summary of Quantum Numbers

1) Shell number, n

n = 1, 2, 3, 4, ...

energy level & avg. distance

Period no. ⇒ highest n

Max # e^- in shell = $2n^2$

2) Subshell, ℓ (shape of orbital)

subshells in shell = n

ℓ = 0, 1, 2, ... (n-1)

s, p, d, f, g, h....

e^- in subshell = $2(2\ell + 1)$

3) Orbitals, m_ℓ (orientation)

$m_\ell = +\ell, \ldots, 0, \ldots, -\ell$

 # orb. in shell = n^2

$(2\ell + 1)$ orbitals in a subshell

max. # e^- in subshell = $2\,(2\ell + 1)$

4) Spin, m_s

$+1/2$ (↑) $-1/2$ (↓)

Subshell letters, # orbitals & max # e^- in subshell

ℓ =	0	1	2	3	4	5
subshell letters	s	p	d	f	g	h ···
# orbitals in subshell	1	3	5	7	9	11 ···
max # e^- in subshell	2	6	10	14	18	22 ···

VIII) Electron Configurations

Orbitals filled in order of inc. energy until all e^- have been used

$$n\ell^{x} \quad x \leftarrow \#\ e^- \text{ in subshell}$$

n ↗ shell; ℓ ↖ subshell

$_{1}H$ $1s^1$

$_{2}He$ $1s^2$

$_{6}C$ 1s 2s 2p

A) Ex: Consider sulfur, $_{16}S$: 16 e^-

$_{16}S$ $1s^2\ 2s^2\ 2p^6\ 3s\ \ 3p$

valence shell (outer shell) ← 3s 3p

$_{16}S$ is in 3^{rd} period ; $n_{max} = 3$

$_{16}S$ is in group VI A, 6 e^- in outer or valence shell

valence e^- ⇒ e^- in outer or valence shell

core e^- ⇒ e^- in inner shells

Note: For representative elements

Period no. ⇒ n value of valence shell

Group no. ⇒ # of valence e^-

Elements in a group have similar chemical and physical properties

- same valence shell e^- configuration

e^- in outer shell are ones involved in chemical reactions

B) Shorthand Electron Configuration

Focus attention on valence shell e^-

$_{16}S$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

completed subshells ⇒ [Ne]

noble gas from previous period

$[Ne]\ 3s^2\ 3p^4$

1) Ex: $_{6}C$

$1s^2\ 2s^2\ 2p^2$ ⇒ [ ]

C) Orbital Diagrams

A dash ___ indicates an orbital

Use arrows, ↿ or ⇂ to indicate e^-
with up or down spin

$_1$H 1s^1 $\underset{1s}{\underline{\upharpoonleft}}$; $_2$He 1s^2 $\underset{1s}{\underline{\upharpoonleft\downharpoonright}}$

$_5$B : 5 e^-

1s 2s 2p

[ ] $\underset{2s}{\underline{\quad\quad}}$ $\underbrace{\underline{\quad\quad}\ \underline{\quad\quad}\ \underline{\quad\quad}}_{2p}$

single e^- in an orbital, $\underline{\uparrow}$, unpaired

paramagnetic substance

- unpaired e^-'s
- attracted by magnetic field

2 e^- in same orbital, $\underline{\uparrow\downarrow}$, paired

Diamagnetic substance

- all e^- paired
- not attracted by magnetic field

D) Hund's Rule

$_6C$: 6 e^-

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

3 possible orbital diagrams:

[He] ↿⇂ (2s) ___ ___ ___ (2p) paired

[He] ↿⇂ (2s) ___ ___ ___ (2p) unpaired, diff. spin

[He] ↿⇂ (2s) ___ ___ ___ (2p) unpaired, same spin

Hunds Rule: e^- occupy diff. orbitals of a subshell until all are singly occupied before e^- pairing occurs.

E) Electron-Dot Symbols

Represent e^- in the s & p orb. of the valence shell as dots arranged around the symbol of the element.

There are 4 s & p orb. & 4 positions about the symbol

- treat like orb. diagrams

s

p E p

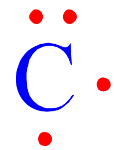
p

Note: only real useful for representative elements

A) Ex's: Draw e^- dot symbols

1) $_6C$

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



2) $_{12}Mg$

[Ne] $3s^2$ Mg

3) $_{16}S$

[Ne] $3s^2\ 3p^4$ S

[Ne] ___ ___ ___ ___

3s 3p

IX) Electron Conf & Periodic Table

Look at $_{32}Ge$

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^2$

n = 4 5s __

4p __ __ __

3d __ __ __ __ __

4s __

n = 3 3p __ __ __

3s __

n = 2 2p __ __ __

2s __

n = 1 1s __

What was happening?

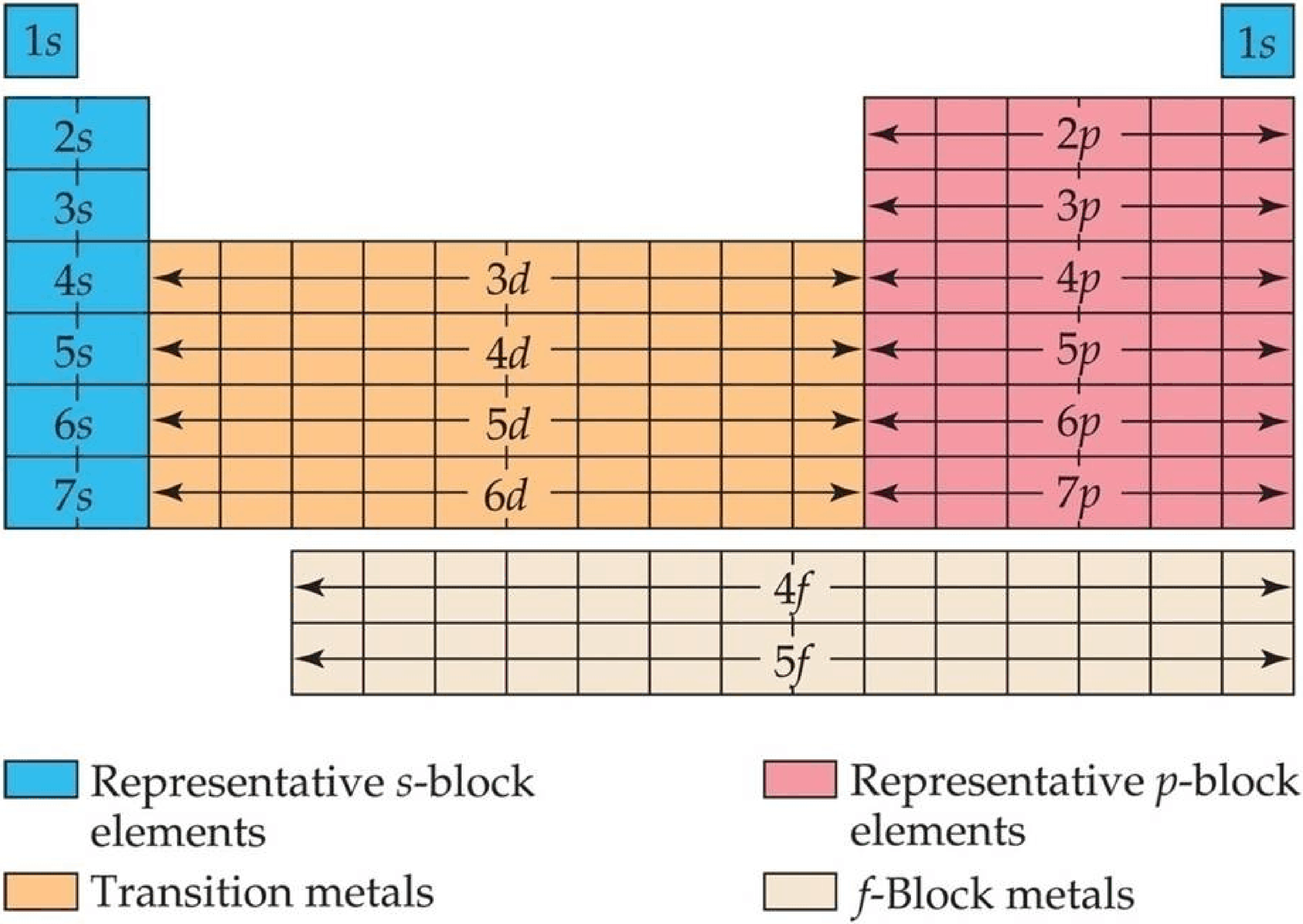
left, filling s orb — $2e^-$, 2 columns

right, filling p orb. — $6e^-$, 6 columns

center, filling d orb — $10e^-$, 10 columns

Period no. ⇛ n value of s & p subshells of valence shell

Group no. ⇛ # of valence e^-



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A) Ex's:

1) $_{16}S$

Period no.	Group no.
3	VI A

2) $_{34}Se$

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB			IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	1.008 **H** 1																	4.003 **He** 2
2	6.941 **Li** 3	9.012 **Be** 4											10.81 **B** 5	12.011 **C** 6	14.007 **N** 7	15.999 **O** 8	18.998 **F** 9	20.179 **Ne** 10
3	22.990 **Na** 11	24.305 **Mg** 12											26.98 **Al** 13	28.09 **Si** 14	30.974 **P** 15	32.06 **S** 16	35.453 **Cl** 17	39.948 **Ar** 18
4	39.098 **K** 19	40.08 **Ca** 20	44.96 **Sc** 21	47.88 **Ti** 22	50.94 **V** 23	52.00 **Cr** 24	54.94 **Mn** 25	55.85 **Fe** 26	58.93 **Co** 27	58.69 **Ni** 28	63.546 **Cu** 29	65.38 **Zn** 30	69.72 **Ga** 31	72.59 **Ge** 32	74.92 **As** 33	78.96 **Se** 34	79.904 **Br** 35	83.80 **Kr** 36
5	85.47 **Rb** 37	87.62 **Sr** 38	88.91 **Y** 39	81.22 **Zr** 40	92.91 **Nb** 41	95.94 **Mo** 42	98 **Tc** 43	101.07 **Ru** 44	102.91 **Rh** 45	106.42 **Pd** 46	107.87 **Ag** 47	112.41 **Cd** 48	114.82 **In** 49	118.69 **Sn** 50	121.75 **Sb** 51	127.60 **Te** 52	126.90 **I** 53	131.39 **Xe** 54
6	132.91 **Cs** 55	137.33 **Ba** 56	138.91 **La** 57	178.39 **Hf** 72	180.95 **Ta** 73	183.85 **W** 74	186.21 **Re** 75	190.23 **Os** 76	192.22 **Ir** 77	195.08 **Pt** 78	196.97 **Au** 79	200.59 **Hg** 80	204.38 **Tl** 81	207.2 **Pb** 82	208.98 **Bi** 83	209 **Po** 84	210 **At** 85	222 **Rn** 86
7	223 **Fr** 87	226.03 **Ra** 88	227.03 **Ac** 89	261 **Rf** 104	262 **Ha** 105	263 **Sg** 106	262 **Ns** 107	265 **Hs** 108	266 **Mt** 109	269 110	272 111	277 112						

6	Lanthanide Series	140.12 **Ce** 58	140.91 **Pr** 59	144.24 **Nd** 60	145 **Pm** 61	150.36 **Sm** 62	151.96 **Eu** 63	157.25 **Gd** 64	158.93 **Tb** 65	162.50 **Dy** 66	164.93 **Ho** 67	167.26 **Er** 68	168.93 **Tm** 69	173.04 **Yb** 70	173.04 **Lu** 71
7	Actinide Series	232.04 **Th** 90	231.04 **Pa** 91	238.03 **U** 92	237.05 **Np** 93	**Pu** 94	**Am** 95	**Cm** 96	**Bk** 97	**Cf** 98	**Es** 99	**Fm** 100	**Md** 101	**No** 102	**Lr** 103

A PERIODIC CHART OF THE ELEMENTS
(Based on ^{12}C)

3) $_{43}Tc$

4) $_{82}Pb$

B) Exceptions

$_{24}Cr$ expect [Ar] $4s^2\ 3d^4$

find [Ar] $4s^1\ 3d^5$

$_{29}Cu$ expect [Ar] $4s^2\ 3d^9$

find [Ar] $4s^1\ 3d^{10}$

Reason: 4s and 3d are very close in energy. (Can act like degenerate orb)

½ filled & filled subshells are more stable.