The Scope of Inorganic Chemistry

Medicine
- MRI
- X-ray contrast imaging
- drugs (arthritis, cancer,...

Biochemistry/Biology
- metalloproteins
- metalloenzymes
- O_2 binding
- catalysis
- ion transport

Organic Chemistry
- organometallics
- metal compounds in synthesis/catalysis

INORGANIC CHEMISTRY
- new compounds
- geometrical and electronic structures
- reactivity

Materials Science
- electrical and magnetic properties of solids
- solid state structures
- semiconductors
- superconductors (high T_c)

Geology/Geochemistry
- synthesis and structure of minerals
- stellar evolution
- astrochemistry

Organometallic Chemistry
- new compounds
- structures
- catalysis

If some universal catastrophe were to engulf the world, and humankind Could retain one scientific concept in order to rebuild civilization, what would that one concept be?

Response for physicists (Richard Feynman in “Six Easy Pieces”):

***The modern idea of atoms*.**

Response for chemists:

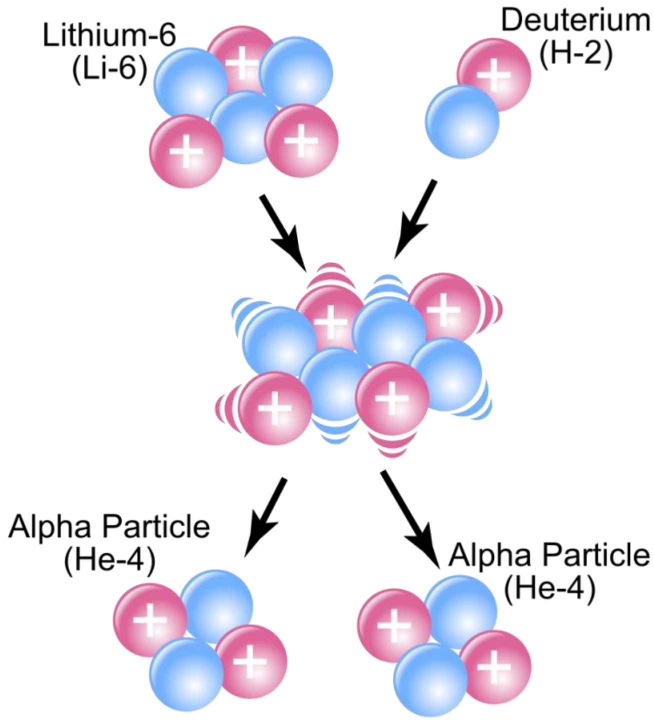
The periodic table

The periodic table encapsulates the concept of elements, organizes physical and chemical trends of substances, and compares the structure of the different atoms— All in a very small space.

Inorganic Chemists think they own the Periodic Table.

So, How did it all start?

An example of nuclear FUSION and the creation of an unstable isotope of Be

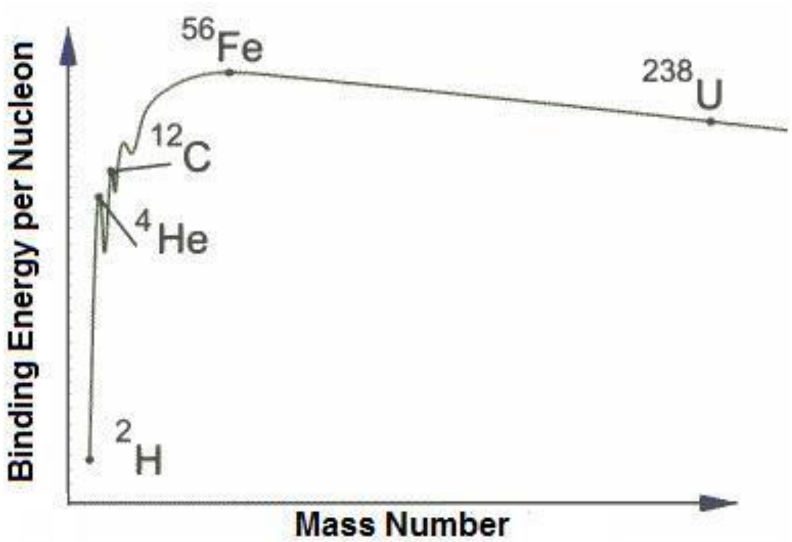


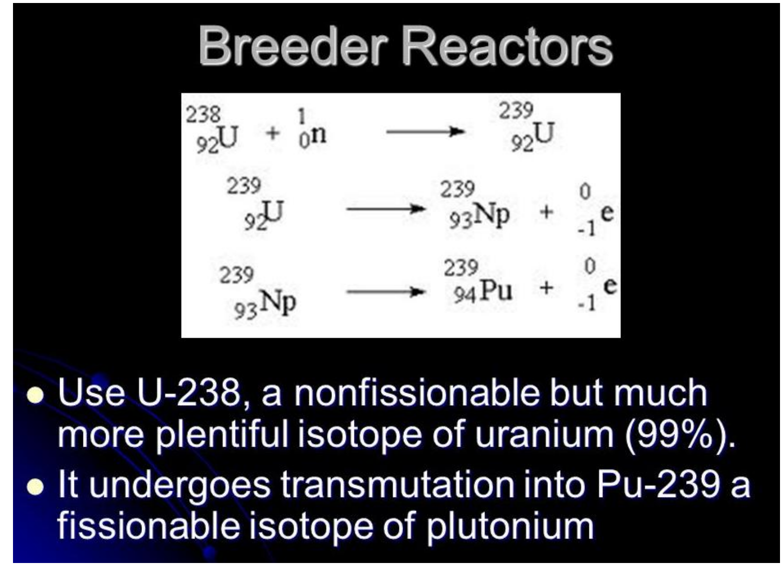
Lithium-6 – Deuterium Reaction

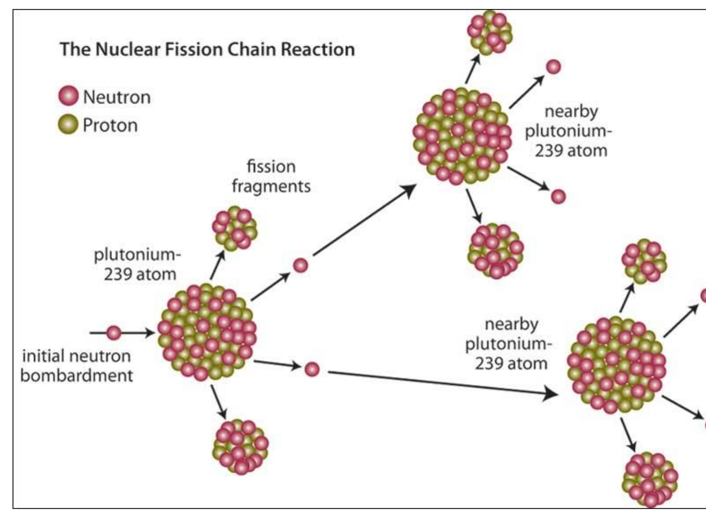
${}_{8}^{4}Be$ with $t_{1/2}$ of ca. 10^{-18} sec

$$ {}^{6}_{3}Li \quad + \quad {}^{2}_{1}H \longrightarrow 2\ {}^{4}_{2}He \ + \ \text{Energy} $$

Stable vs. Unstable or Fissionable Nuclei







Project no. 1. Nuclear Reactions; Web Elements

WebElements™ Periodic Table

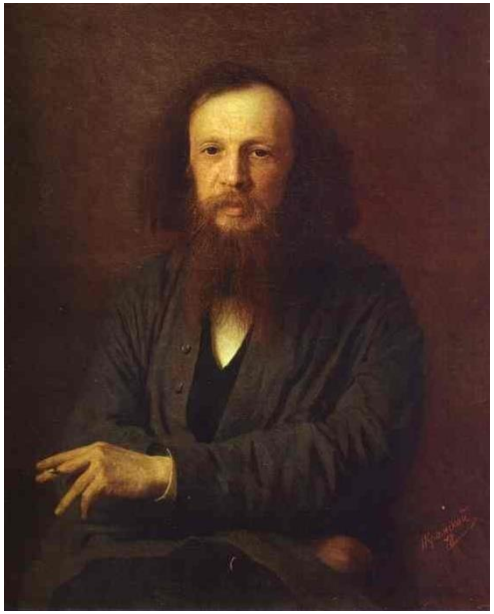
Group / Period	1	2		3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																		2 He
2	3 Li	4 Be												5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg												13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca		21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr		39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	*	71 Lu	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	**	103 Lr	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Uub	113 Uut	114 Uuq	115 Uup	116 Uuh	117 Uus	118 Uuo

*Lanthanoids	*	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb
**Actinoids	**	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No

Mass Spectra computer:

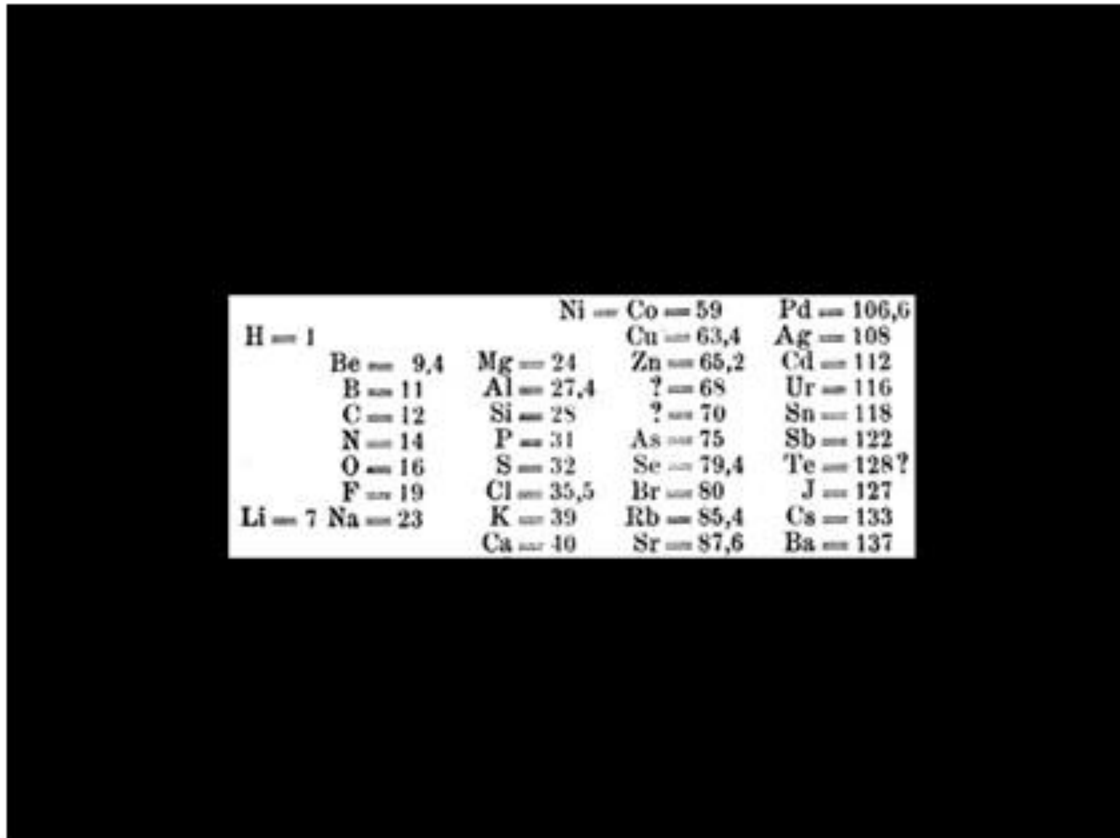
http://www.sisweb.com/mstools/isotope.htm

http://www.sisweb.com/mstools/isotope.htm



Dimitri Mendeleev

1834-1907

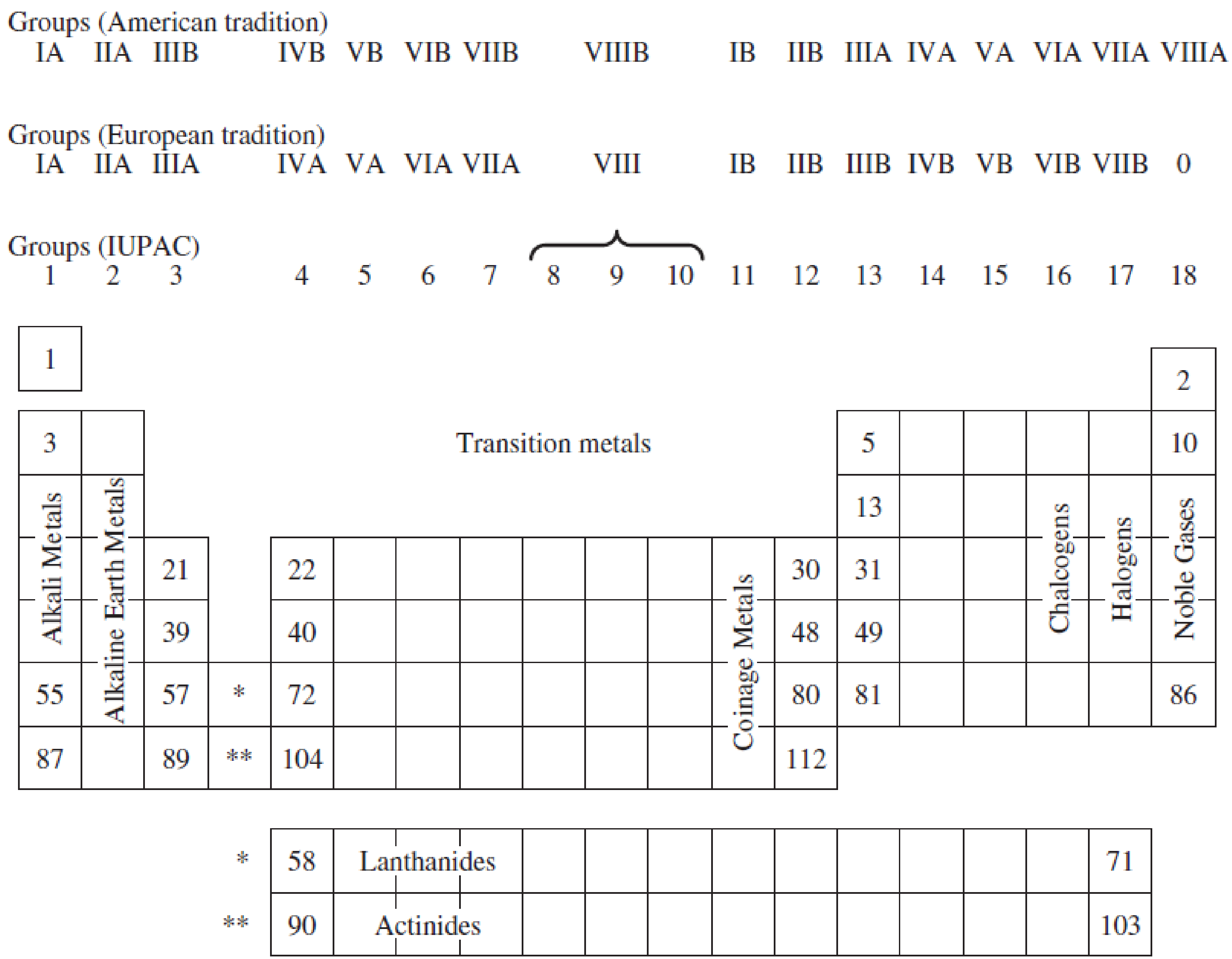


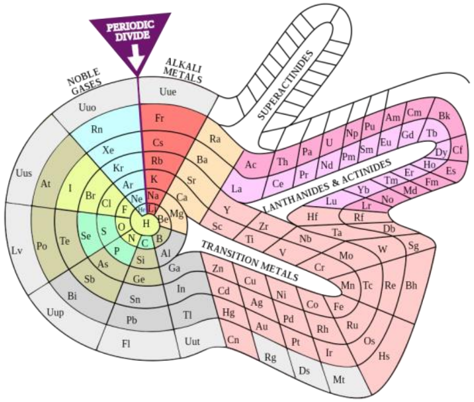
			Ni = Co = 59	Pd = 106,6
H = 1			Cu = 63,4	Ag = 108
	Be = 9,4	Mg = 24	Zn = 65,2	Cd = 112
	B = 11	Al = 27,4	? = 68	Ur = 116
	C = 12	Si = 28	? = 70	Sn = 118
	N = 14	P = 31	As = 75	Sb = 122
	O = 16	S = 32	Se = 79,4	Te = 128?
	F = 19	Cl = 35,5	Br = 80	J = 127
Li = 7	Na = 23	K = 39	Rb = 85,4	Cs = 133
		Ca = 40	Sr = 87,6	Ba = 137

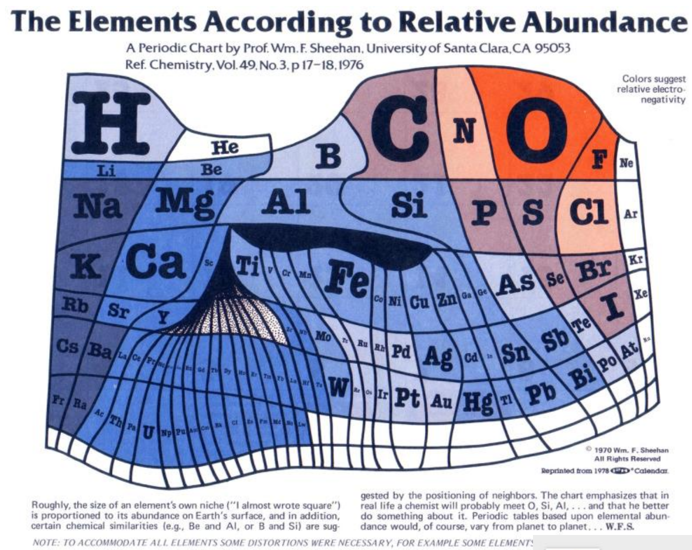
$$^{A}_{Z}E$$

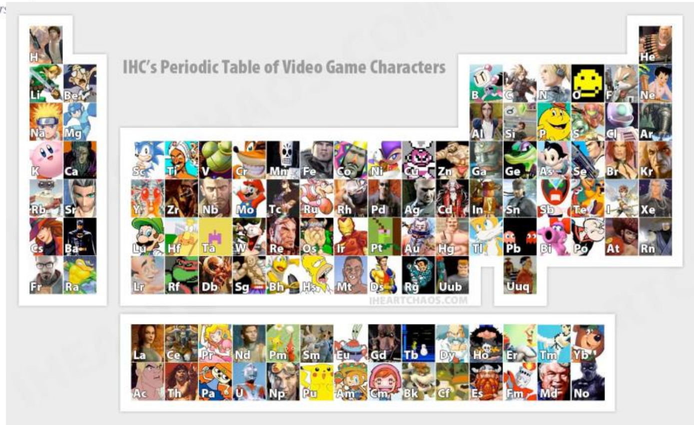
Z = No. protons in nucleus, Atomic number

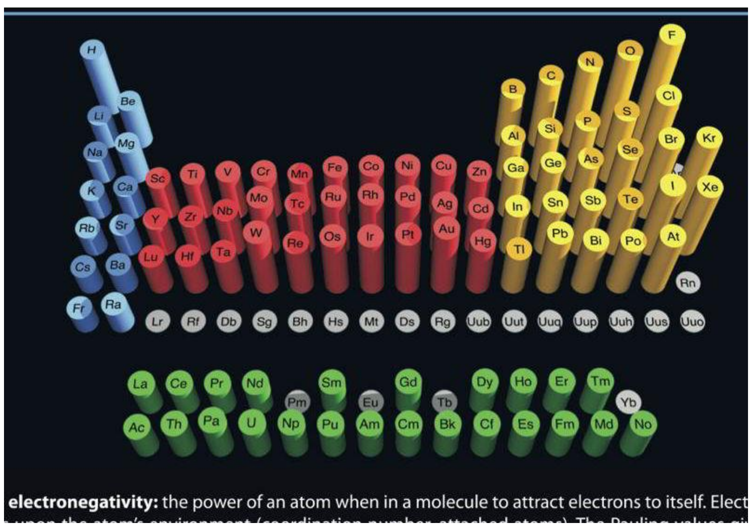
A = Mass number; no. of protons + neutrons in nucleus

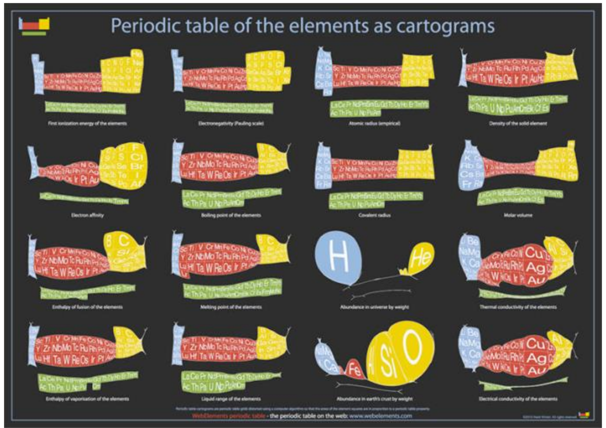

Groups (American tradition)
IA IIA IIIB IVB VB VIB VIIB VIIIB IB IIB IIIA IVA VA VIA VIIA VIIIA
Groups (European tradition)
IA IIA IIIA IVA VA VIA VIIA VIII IB IIB IIIB IVB VB VIB VIIB 0
Groups (IUPAC)
1 2 3 4 5 6 7 8 9 10 11 12 13 14 15 16 17 18
1
2
3
5
10
Transition metals
13
Alkali Metals
Alkaline Earth Metals
21
22
30
31
Chalcogens
Halogens
Noble Gases
Coinage Metals
39
40
48
49
55
57
*
72
80
81
86
87
89
**
104
112
*
58
Lanthanides
71
**
90
Actinides
103


PERIODIC DIVIDE
NOBLE GASES
ALKALI METALS
SUPERACTINIDES
LANTHANIDES & ACTINIDES
TRANSITION METALS


The Elements According to Relative Abundance
A Periodic Chart by Prof. Wm. F. Sheehan, University of Santa Clara, CA 95053
Ref. Chemistry, Vol. 49, No. 3, p 17–18, 1976
Colors suggest relative electro-negativity
© 1970 Wm. F. Sheehan
All Rights Reserved
Roughly, the size of an element's own niche ("I almost wrote square") is proportioned to its abundance on Earth's surface, and in addition, certain chemical similarities (e.g., Be and Al, or B and Si) are suggested by the positioning of neighbors. The chart emphasizes that in real life a chemist will probably meet O, Si, Al, . . . and that he better do something about it. Periodic tables based upon elemental abundance would, of course, vary from planet to planet . . . W.F.S.
NOTE: TO ACCOMMODATE ALL ELEMENTS SOME DISTORTIONS WERE NECESSARY, FOR EXAMPLE SOME ELEMENTS


IHC's Periodic Table of Video Game Characters


electronegativity: the power of an atom when in a molecule to attract electrons to itself. Elect


Periodic table of the elements as cartograms

Color	**Metal Flame Colors**
Red	*Carmine*: Lithium compounds. Masked by barium or sodium. *Scarlet or Crimson*: Strontium compounds. Masked by barium. *Yellow-Red*: Calcium compounds. Masked by barium.
Yellow	Sodium compounds, even in trace amounts. A yellow flame is not indicative of sodium unless it persists and is not intensified by addition of 1% NaCl to the dry compound.
White	*White-Green*: Zinc
Green	*Emerald*: Copper compounds, other than halides. Thallium. *Blue-Green*: Phosphates, when moistened with H_2SO_4 or B_2O_3. *Faint Green*: Antimony and NH_4 compounds. *Yellow-Green*: Barium, molybdenum.
Blue	*Azure*: Lead, selenium, bismuth, $CuCl_2$ and other copper compounds moistened with hydrochloric acid. *Light Blue*: Arsenic and come of its compounds. *Greenish Blue*: $CuBr_2$, antimony
Violet	Potassium compounds other than borates, phosphates, and silicates. Masked by sodium or lithium. *Purple-Red*: Potassium, rubudium, and/or cesium in the presence of sodium when viewed through a blue glass.

Atomic Emission (Spectroscopy)

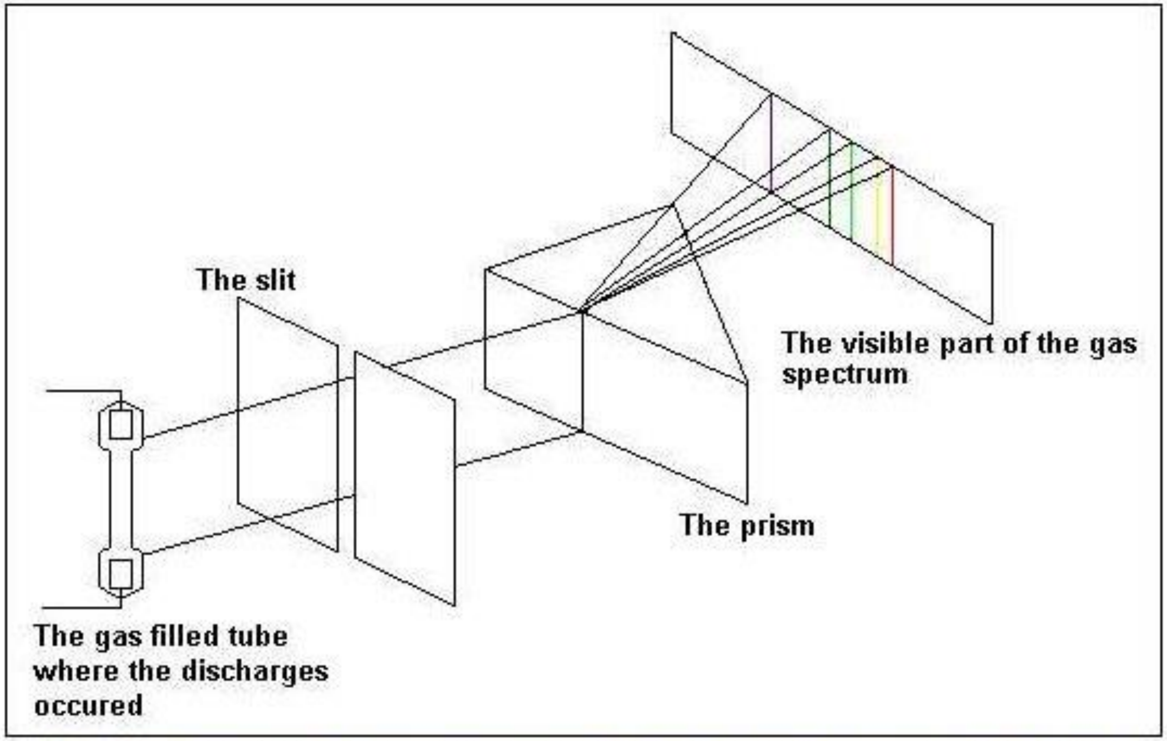
- An emission spectrum requires first the addition of energy to a material.
- The addition of energy promotes electrons of that material from the ground state to the excited state.
- As the electrons "fall" from the excited state to the ground state, they emit the energy they absorbed in the form of electromagnetic radiation (heat, light, etc.)

Comments

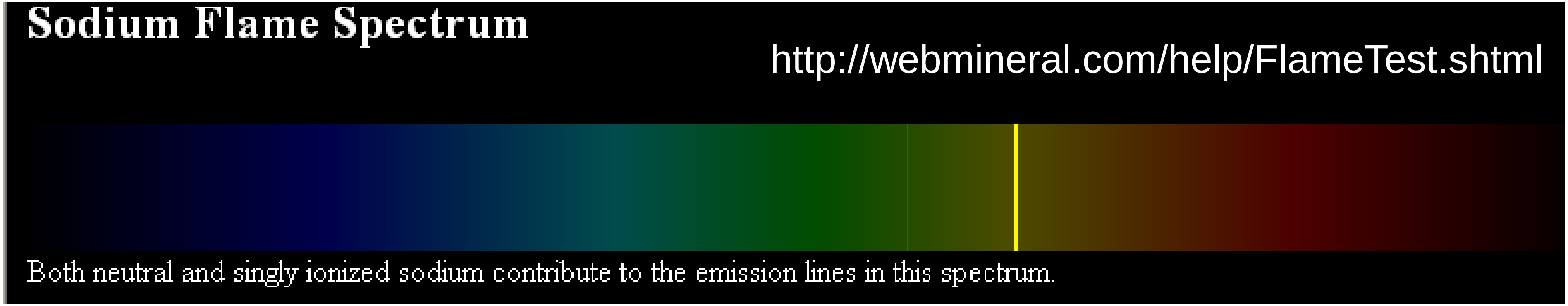
- Atomic emission is used in street lamps, fluorescent lights, and neon signs.
- Two common street lamps using this are the mercury lamp and the sodium lamp.
- "Neon" signs frequently implement the emission spectra of other gases such as argon and krypton.
- Very sophisticated instrumental techniques such as "flame photometry" and "atomic absorption" are based on the principles of atomic emission.

Continuous and Line Spectra

Dark Side of the Moon
Pink Floyd

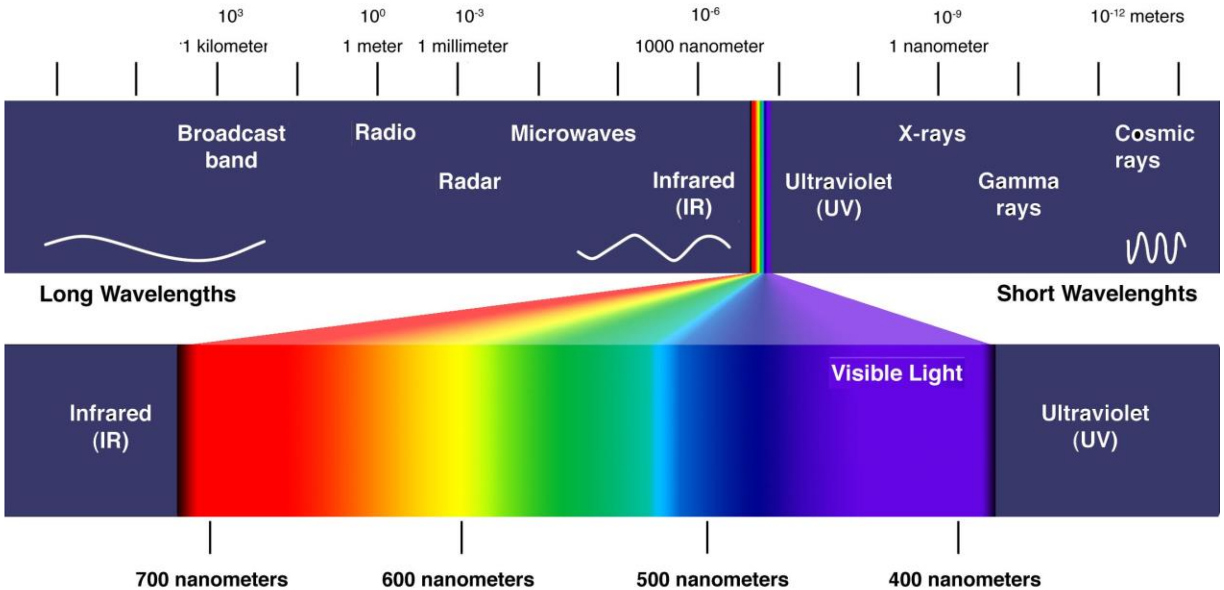


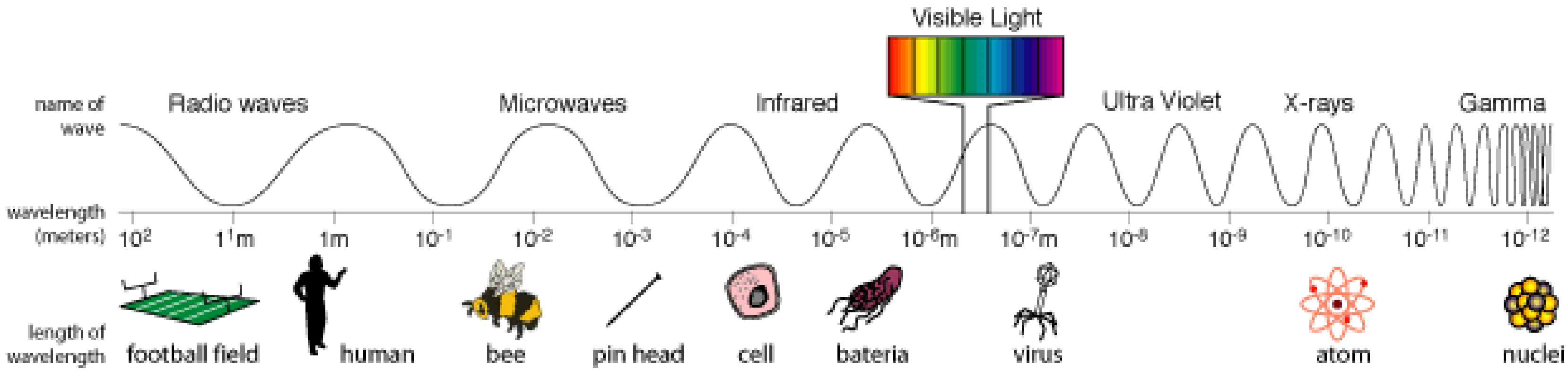
The discharges in the low pressure gas filled tube are the sources of the light which undergo refraction on the prism. We see the line spectrum of the gas.



http://webmineral.com/help/FlameTest.shtml

Electromagnetic Radiation Spectrum

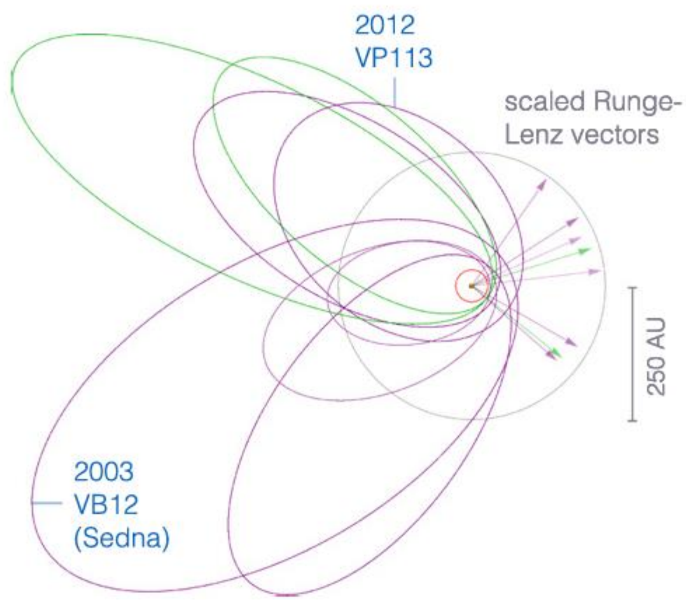




In the news: January 20, 2016

EVIDENCE FOR A DISTANT GIANT PLANET IN THE SOLAR SYSTEM

KONSTANTIN BATYGIN AND MICHAEL E. BROWN
Division of Geological and Planetary Sciences, California Institute of Technology, Pasadena, CA 91125, USA; kbatygin@gps.caltech.edu
Received 2015 November 13; accepted 2016 January 10; published 2016 January 20



Cumulatively, the presented results offer credence to the hypothesis that the observed structure of the distant Kuiper Belt can be explained by invoking perturbations from an unseen planetary mass companion to the solar system. Simultaneously, the suggestive nature of the results should be met with a healthy dose of skepticism, given the numerous assumptions made in the construction of our simple analytical model. In

Lecture 2 362 January 22, 2016

Paradigm Shift:
Development of Current Atomic Theory

Emission Spectra

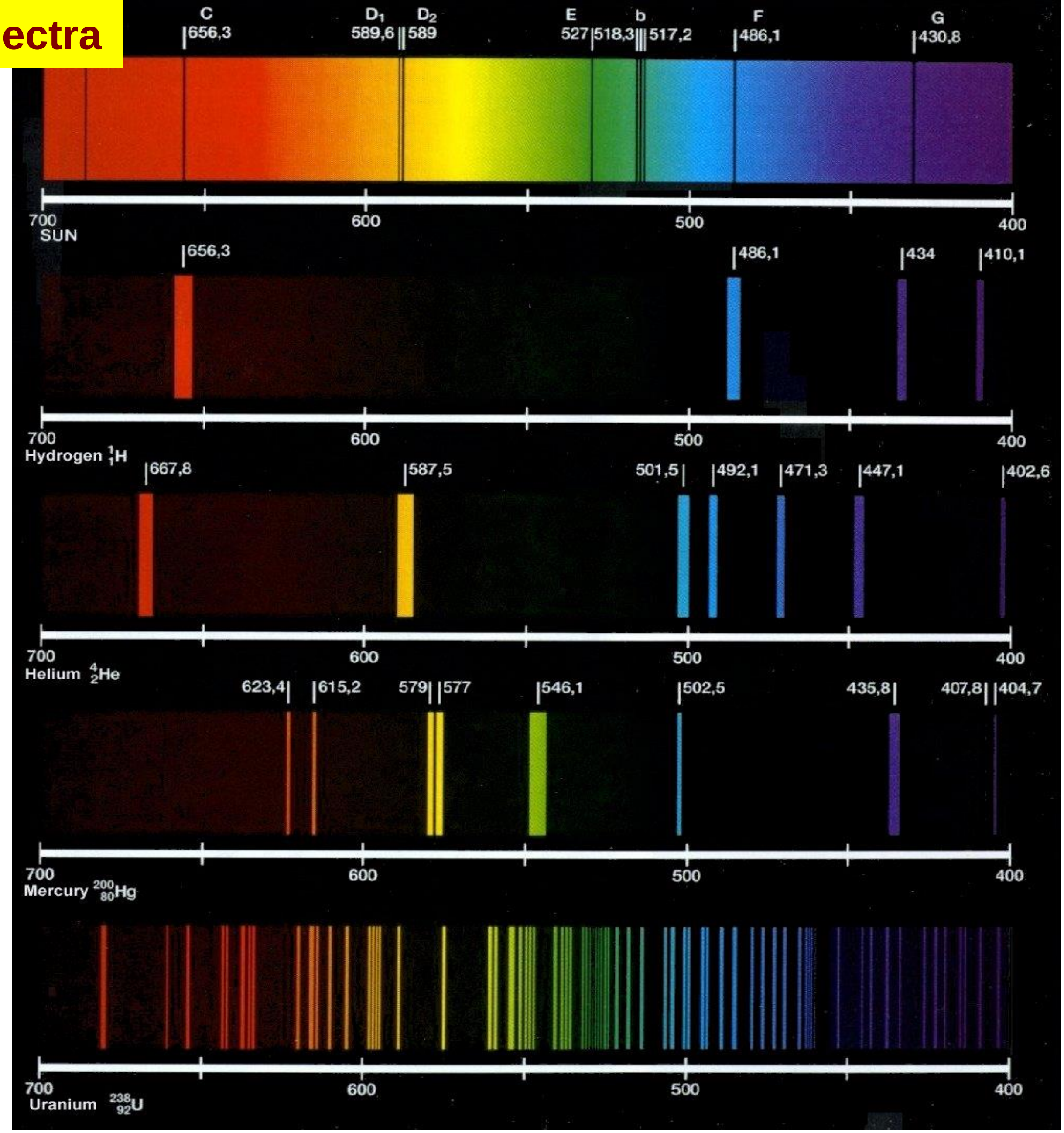
Sun

Hydrogen

Helium

Mercury

Uranium



Complexity increases With size of Element

What do these lines mean?????

700 nm 600 nm 500 nm 400 nm

A Bit of History

TABLE 2.1 Discoveries in Atomic Structure

1896	A. H. Becquerel	Discovered radioactivity of uranium
1897	J. J. Thomson	Showed that electrons have a negative charge, with charge/mass $= 1.76 \times 10^{11}$ C/kg
1909	R. A. Millikan	Measured the electronic charge as 1.60×10^{-19} C; therefore, mass of electron $= 9.11 \times 10^{-31}$ kg
1911	E. Rutherford	Established the nuclear model of the atom: a very small, heavy nucleus surrounded by mostly empty space
1913	H. G. J. Moseley	Determined nuclear charges by X-ray emission, establishing atomic numbers as more fundamental than atomic masses

$$^{A}_{Z}E$$

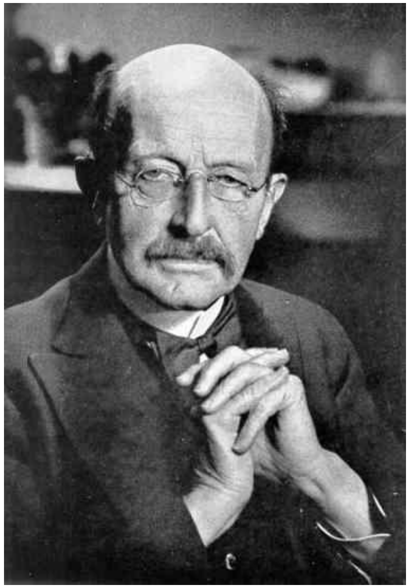
Z = No. protons in nucleus, Atomic number

A = Mass number; no. of protons + neutrons in nucleus

Marie Curie—1867-1934

So how to connect the physical properties of elements to the Periodic Table? Physicists! The current model of the atom belongs to Physicists!

DeBroglie



Planck

Schrodinger

Einstein

Heisenberg

Bohr

Pauli

Niels Bohr and wife Margrethe around 1930

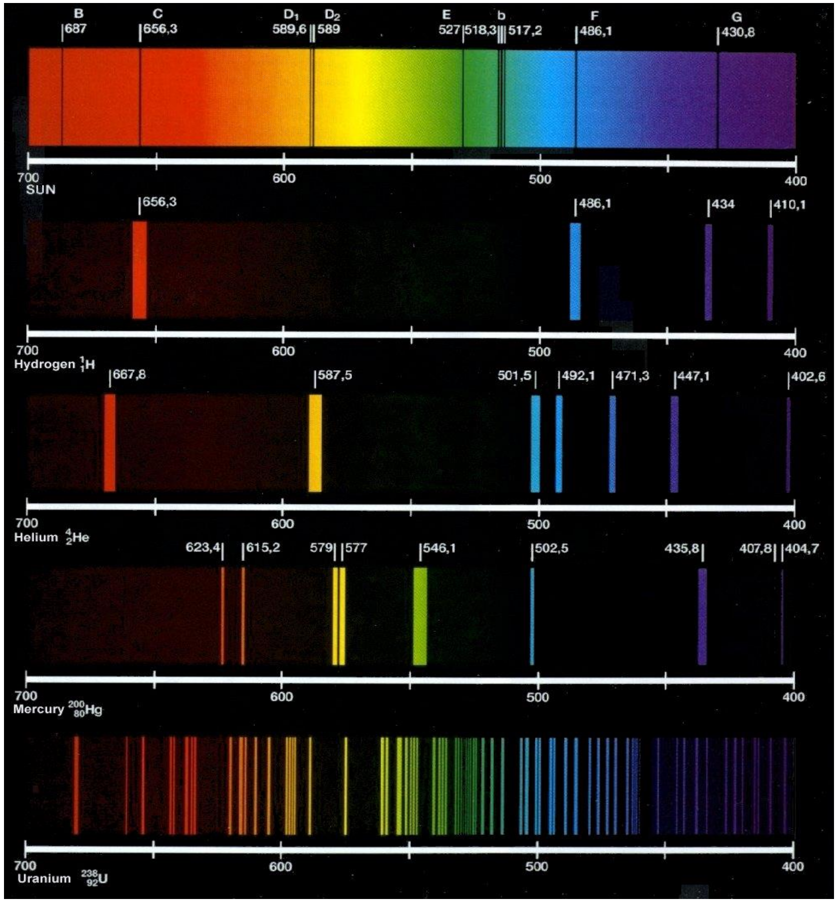
Taken from John L. Heilbron's "History: The Path to the quantum Atom", Nature 498, 27-30, (06 June 2013)

Taken from John L. Heilbron's "History: The Path to the quantum Atom", Nature 498, 27-30, (06 June 2013)

To develop his model, Bohr followed an analogy to the radiation theory of Max Planck (right). ". . . Bohr had developed a doctrine of multiple partial truths, each of which contained some bit of reality, and all of which together might exhaust it. "There exist so many different truths I can almost call it my religion that I think that everything that is of value is true."

Lecture 2

Sun
Hydrogen
Helium
Mercury
Uranium

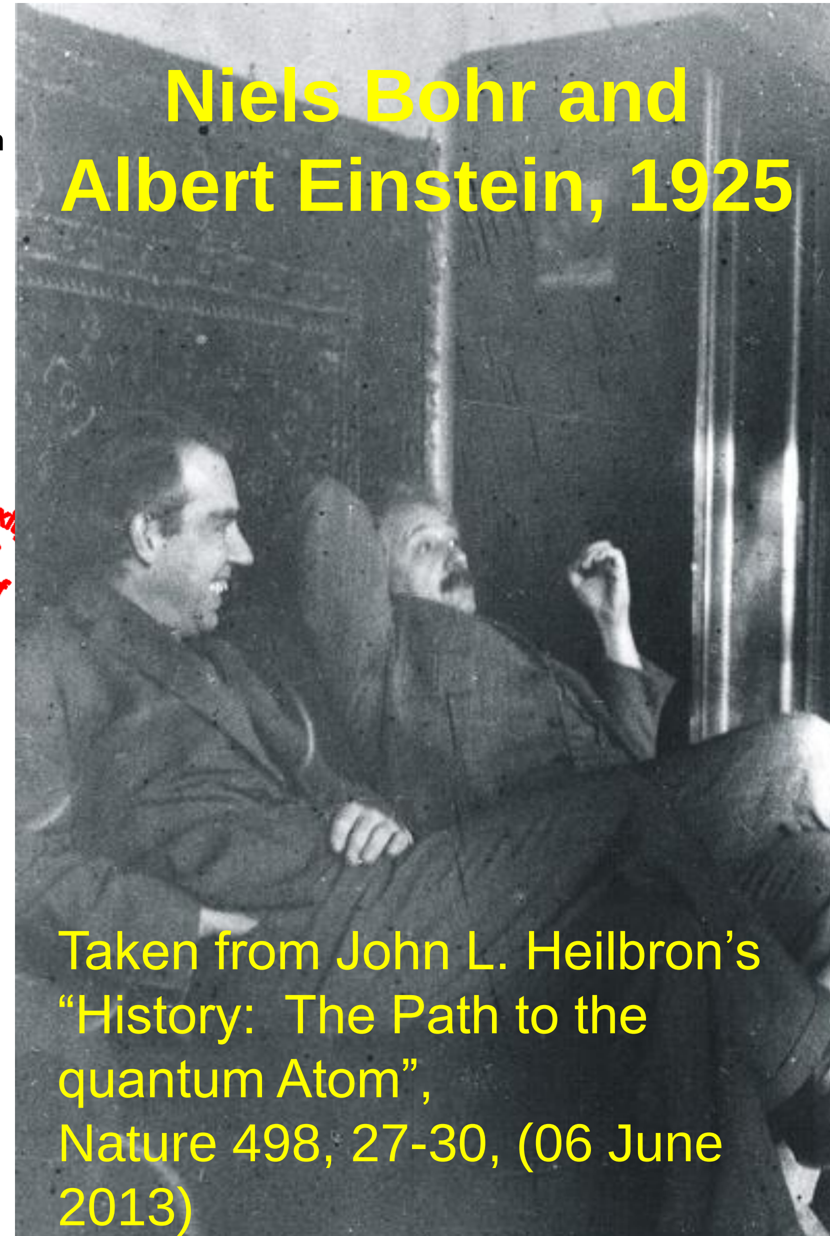


Emission Spectra

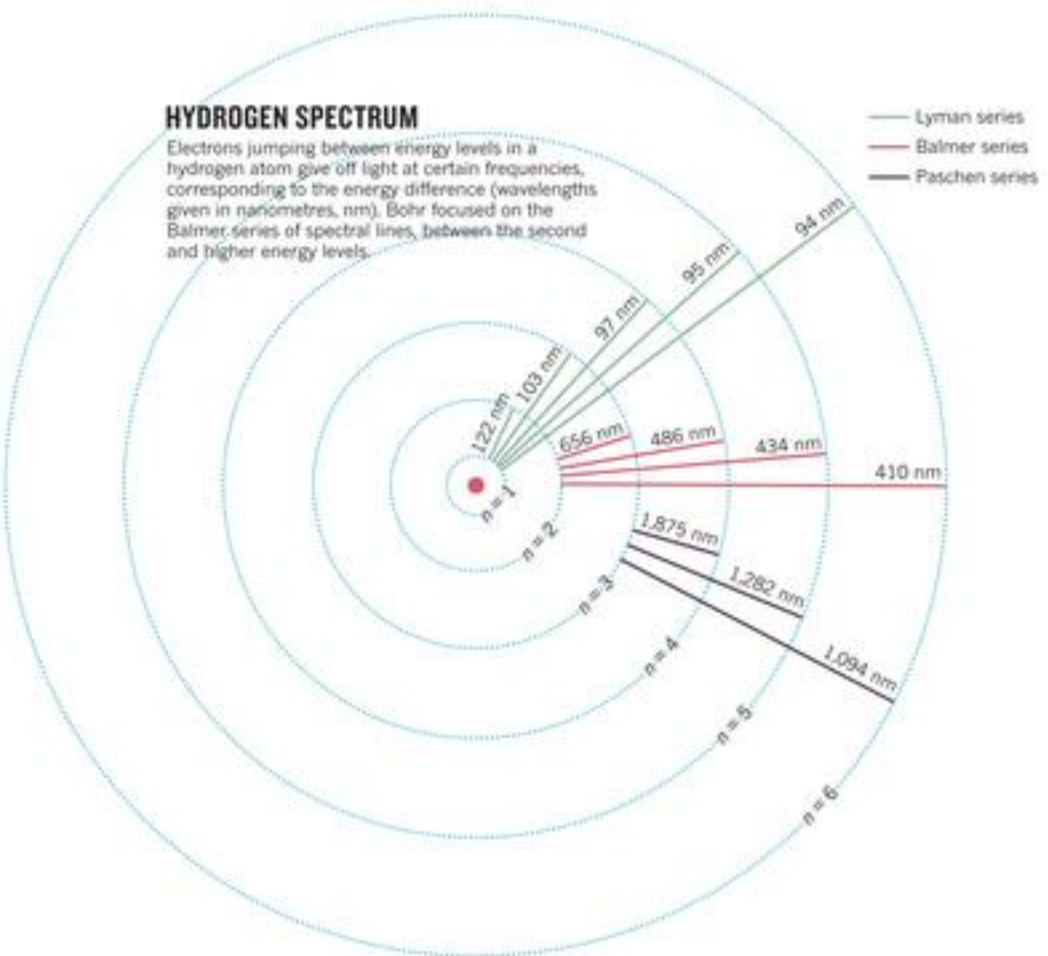
Complexity increases With size of Element

What do these lines mean?????

700nm 600nm 500nm 400nm



The Bohr Atom: electrons in concentric rings



Taken from John L. Heilbron's "History: The Path to the quantum Atom",
Nature 498, 27-30, (06 June 2013)

The Balmer formula expresses the frequencies of some lines in the spectrum of hydrogen in simple algebra:

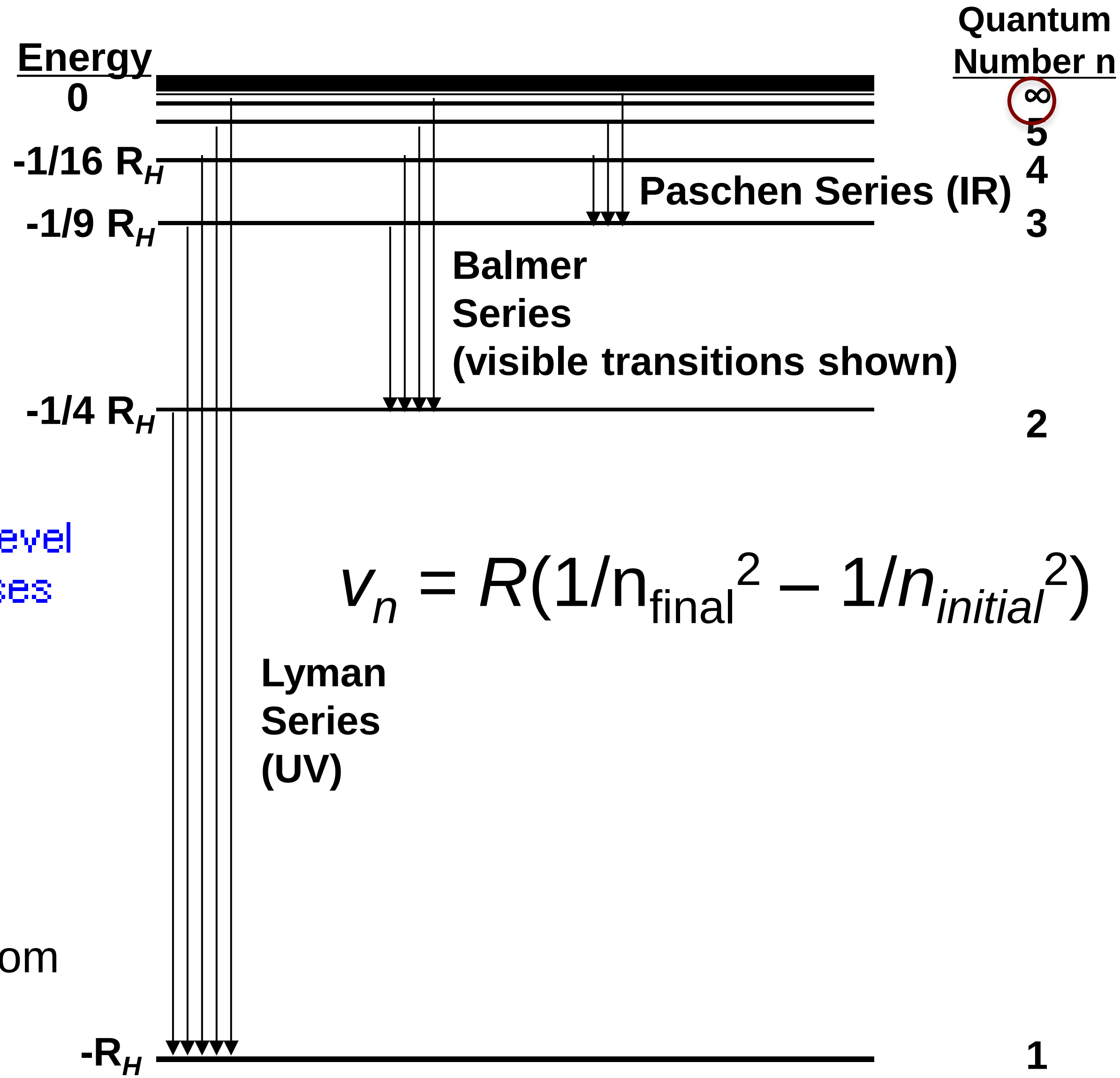
$v_n = R(1/2^2 – 1/n^2)$

where v_n is the *n*th Balmer line and *R* is the universal Rydberg constant for frequency, named in honour of the Swedish spectroscopist Johannes Rydberg, who generalized Balmer's formula to apply to elements beyond hydrogen.

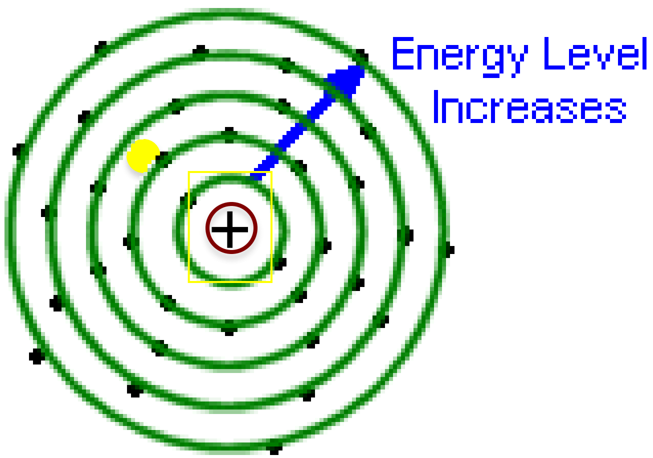
Each level can accommodate $2n^2$ electron:

Periodic Table Rows

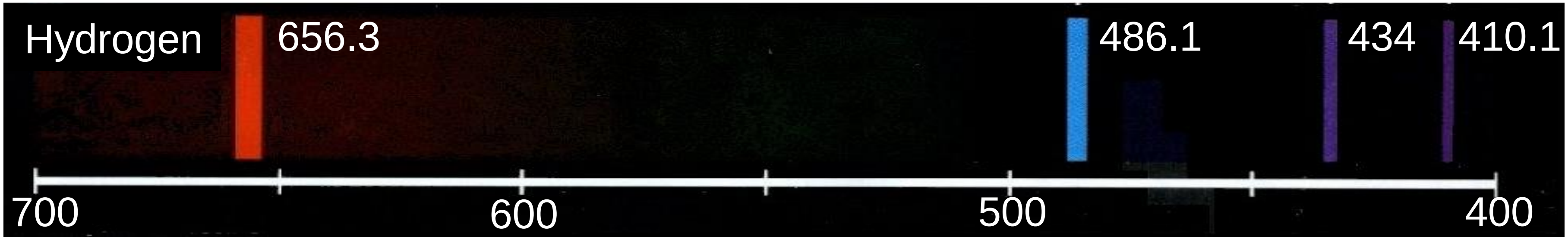
The Hydrogen Atom Spectrum and Energy Levels

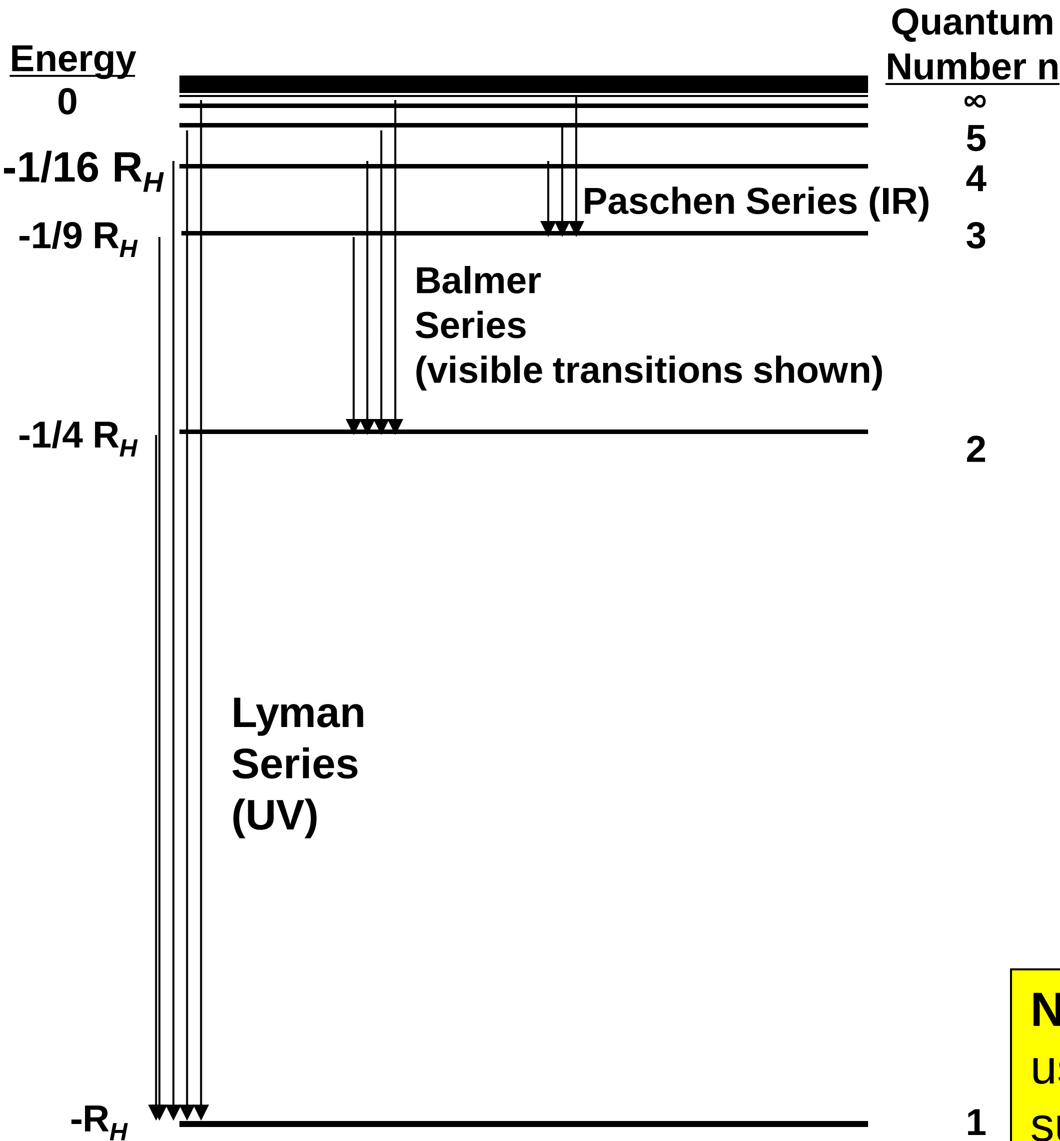


$$\nu_n = R(1/n_{final}^2 - 1/n_{initial}^2)$$



Niels Bohr's view of the atom





For Hydrogen:

$$E = \frac{-R_H}{n^2}$$

Rydberg constant for hydrogen, R_H

$$R_H = \frac{m_e e^4}{8\,\varepsilon_o^2 h^2} = 2.179 \times 10^{-18} J$$

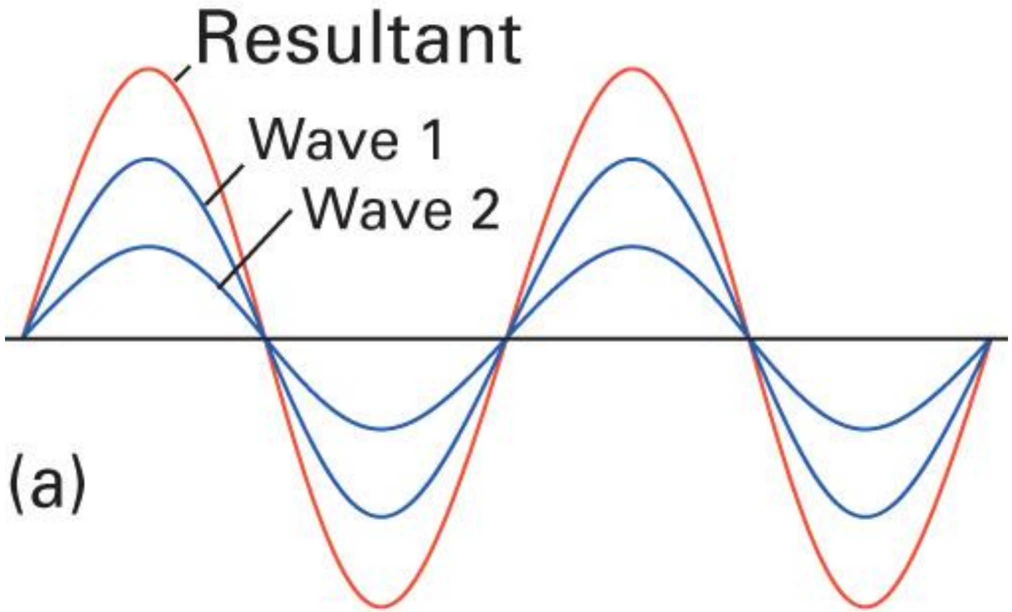
$= 13.6$ eV

General equation for Rydberg constant for any element

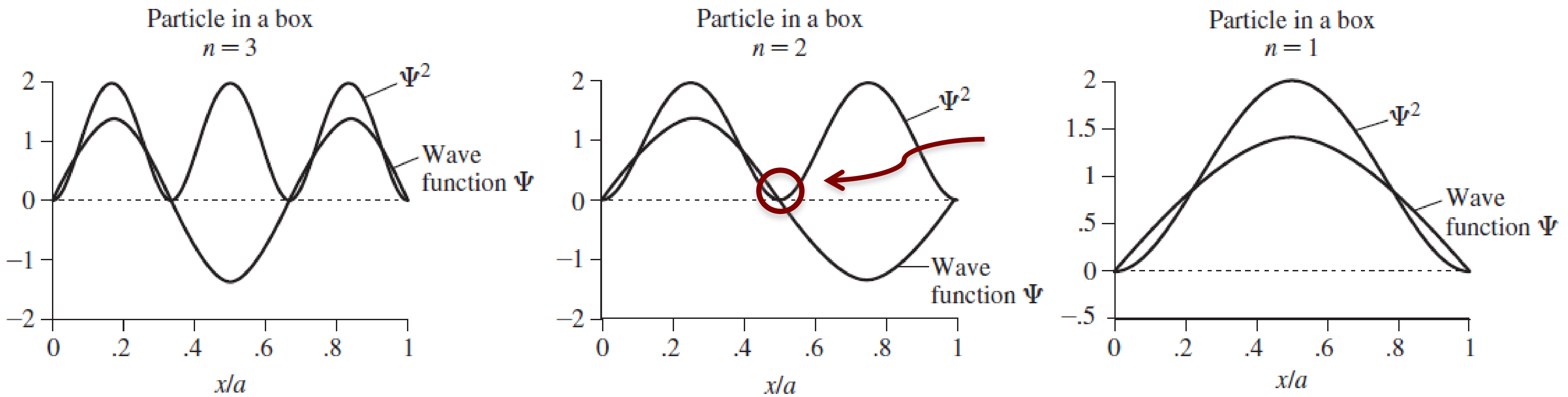
$$R = \frac{-\mu Z^2 e^4}{8\,\varepsilon_o^2 h^2}$$

Note: The predicted emission spectra using the Rydberg constant was only successful for simple elements such as H and failed for heavier atoms due to the limitations of the Bohr view of the atom. ***This led to the foundations of quantum mechanics***.

Properties of waves: Addition for reinforcement or cancellation



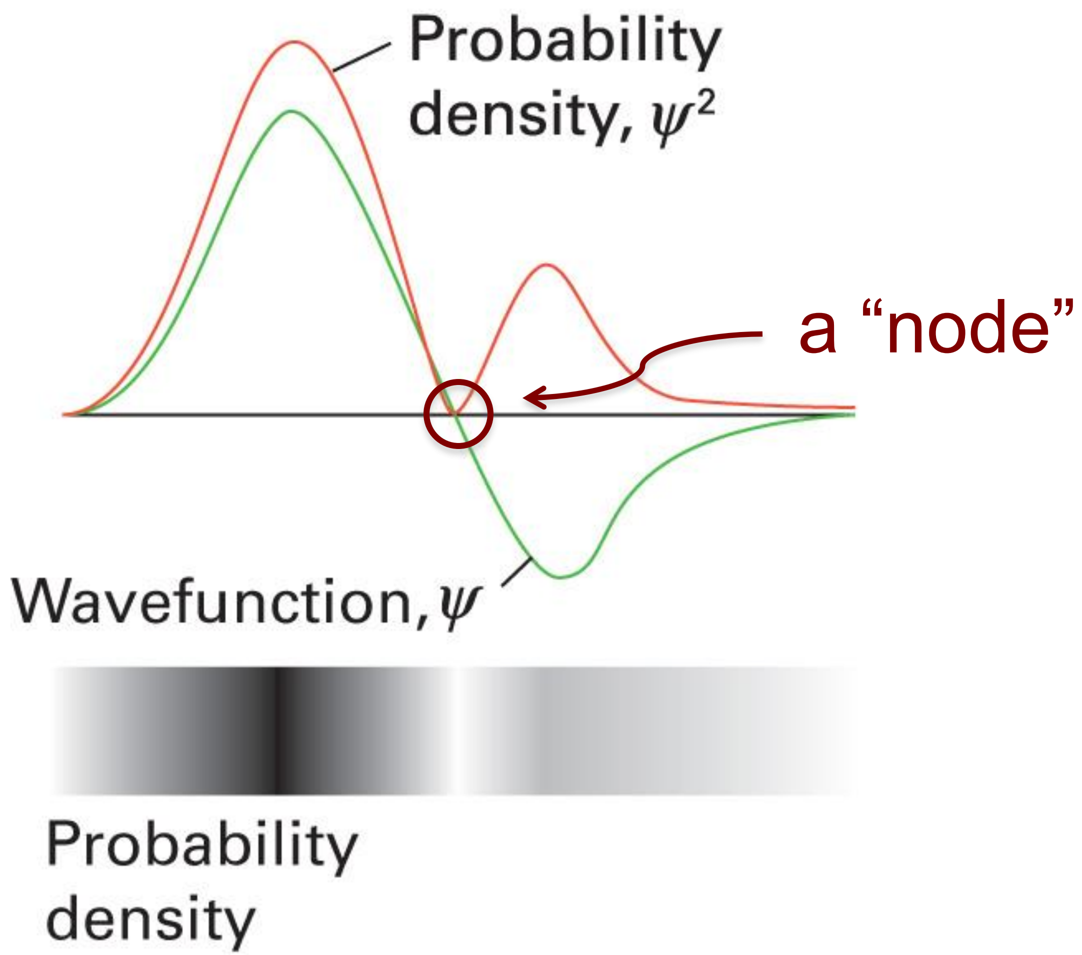
Properties of waves: Squared = amplitude
Boundaries => Restrictions on values



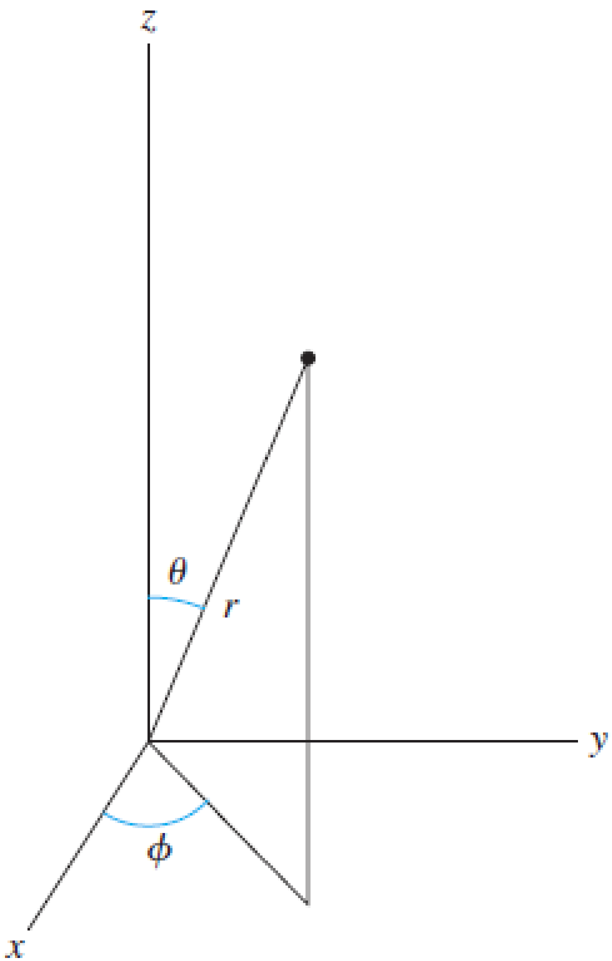
a “node”

Time-independent Schrödinger equation (general—one dimension)

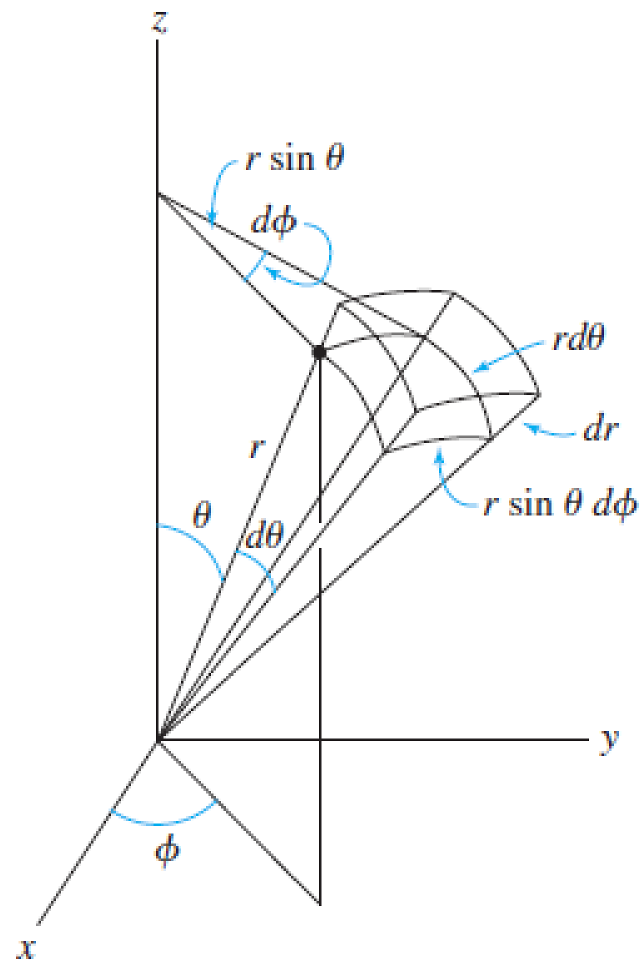
$$E\psi = H\psi$$



Need for Spherical Coordinates and Volume Elements



Spherical coordinates



Volume element

Need both radial and angular functions

1.4

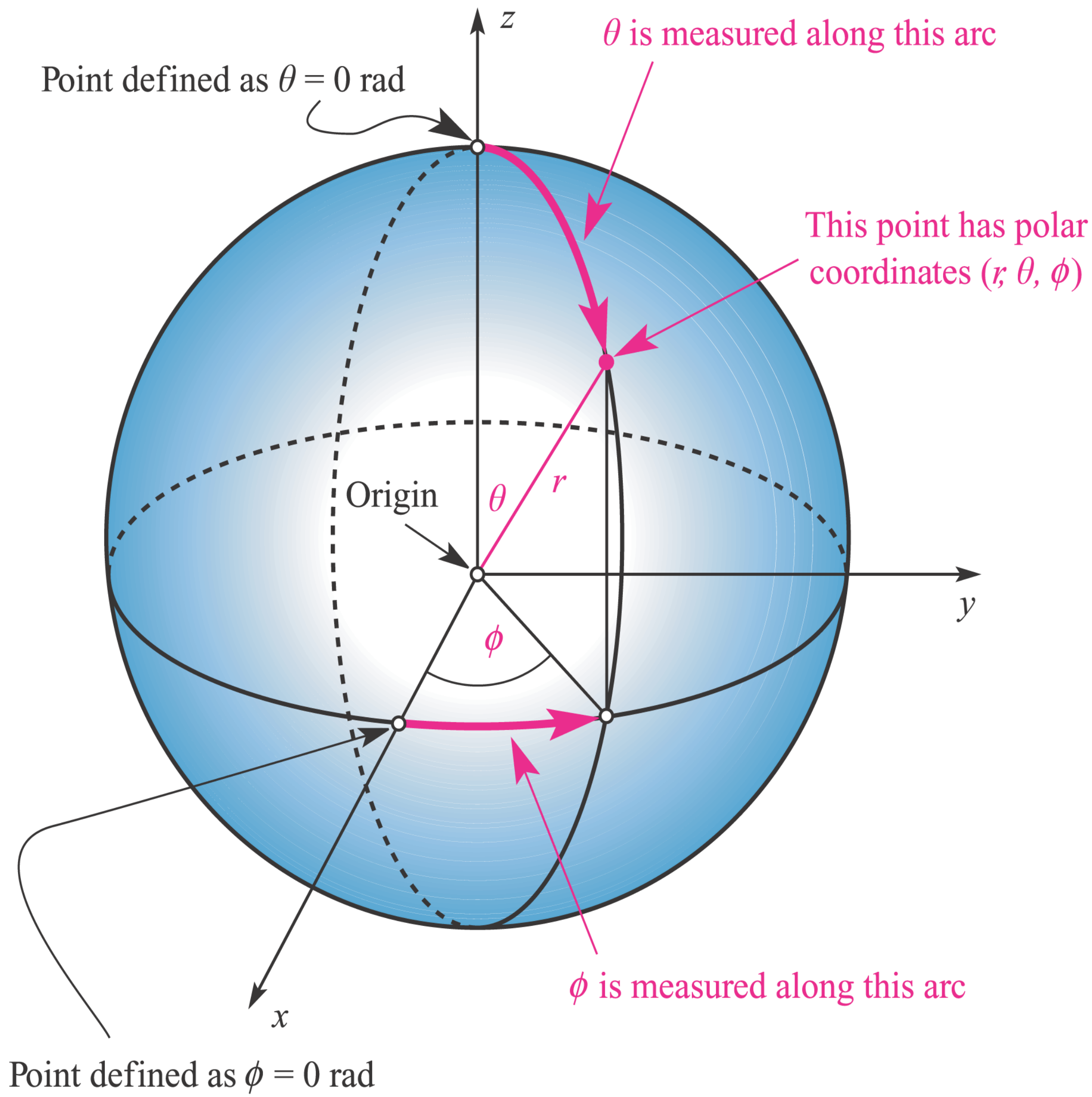


Fig. 1.4 Definition of the polar coordinates (r, θ, ϕ) for a point shown here in pink; r is the radial coordinate and θ and ϕ are angular coordinates. θ and ϕ are measured in radians (rad). Cartesian axes (x, y and z) are also shown.

Summarizing: Solutions Required Quantum Numbers

Quantum Numbers

Atomic orbital	n	l	m_l	Radial part of the wavefunction, $R(r)$[†]	Angular part of wavefunction, $A(\theta, \phi)$
$1s$	1	0	0	$2e^{-r}$	$\frac{1}{2\sqrt{\pi}}$
$2s$	2	0	0	$\frac{1}{2\sqrt{2}}(2-r)\,e^{-r/2}$	$\frac{1}{2\sqrt{\pi}}$
$2p_x$	2	1	$+1$	$\frac{1}{2\sqrt{6}} r\,e^{-r/2}$	$\frac{\sqrt{3}(\sin\theta\cos\phi)}{2\sqrt{\pi}}$
$2p_z$	2	1	0	$\frac{1}{2\sqrt{6}} r\,e^{-r/2}$	$\frac{\sqrt{3}(\cos\theta)}{2\sqrt{\pi}}$
$2p_y$	2	1	-1	$\frac{1}{2\sqrt{6}} r\,e^{-r/2}$	$\frac{\sqrt{3}(\sin\theta\sin\phi)}{2\sqrt{\pi}}$

[†] For the $1s$ atomic orbital, the formula for $R(r)$ is actually $2(\frac{Z}{a_0})^{\frac{3}{2}} e^{-Zr/a_0}$ but for the hydrogen atom, $Z = 1$ and $a_0 = 1$ atomic unit. Other functions are similarly simplified.

Table 1.2 Solutions of the Schrödinger equation for the hydrogen atom which define the $1s$, $2s$ and $2p$ atomic orbitals. For these forms of the solutions, the distance r from the nucleus is measured in atomic units.

Quantum Numbers

n is the *principal quantum number*, indicates the size of the orbital, has all positive integer values of 1 to ∞(infinity)

ℓ is the *angular momentum quantum number*, represents the shape of the orbital, has integer values of (n – 1) to 0

m_ℓ is the *magnetic quantum number*, represents the spatial direction of the orbital, can have integer values of -ℓ to 0 to ℓ

m_s is the *spin quantum number*, has little physical meaning, can have values of either +1/2 or -1/2

ℓ (angular momentum)	orbital
0	s
1	p
2	d
3	f

Other terms: electron configuration, noble gas configuration, valence shell

Pauli Exclusion principle: no two electrons can have all four of the same quantum numbers in the same atom (Every electron has a unique set.)

Hund's Rule: when electrons are placed in a set of degenerate orbitals, the *ground state* has as many electrons as possible in different orbitals, and with *parallel spin*.

Aufbau (Building Up) Principle: the ground state electron configuration of an atom can be found by putting electrons in orbitals, starting with the lowest energy and moving progressively to higher energy.

While n is the principal energy level, the ℓ value also has an effect

Heavier elements, more chances
For close energy levels

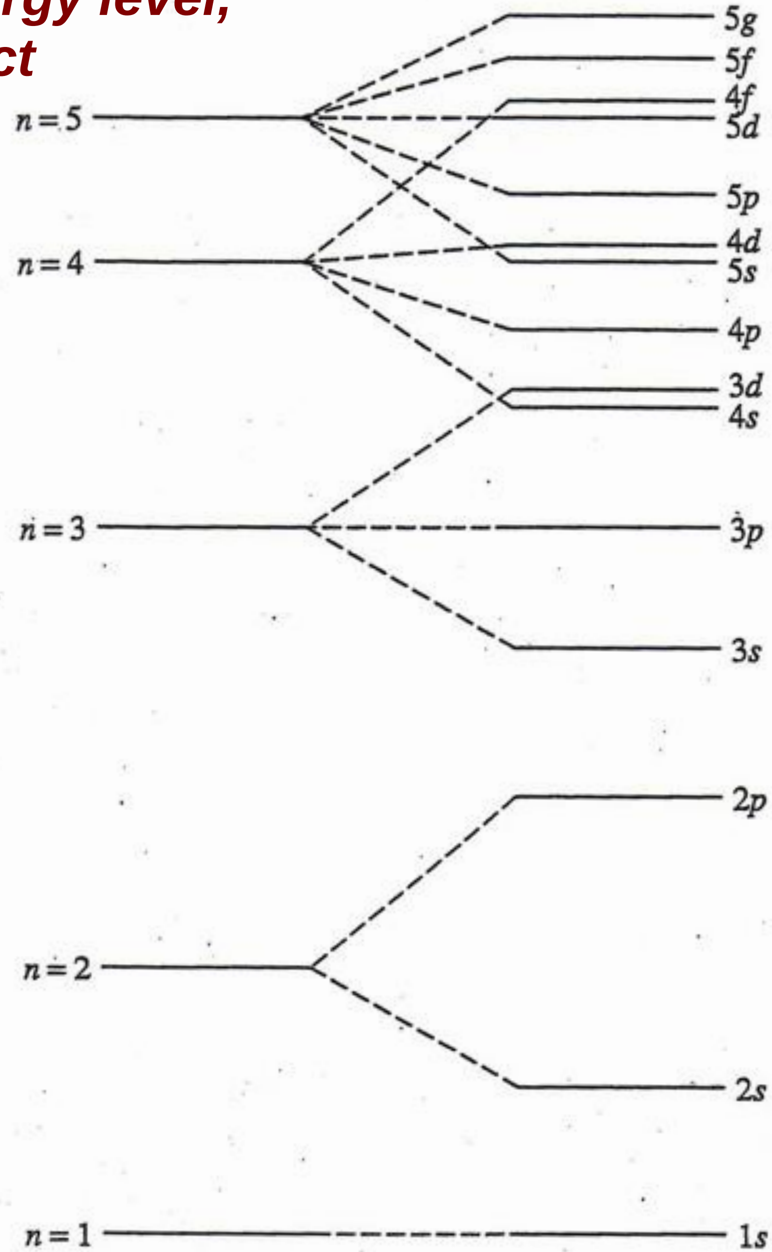
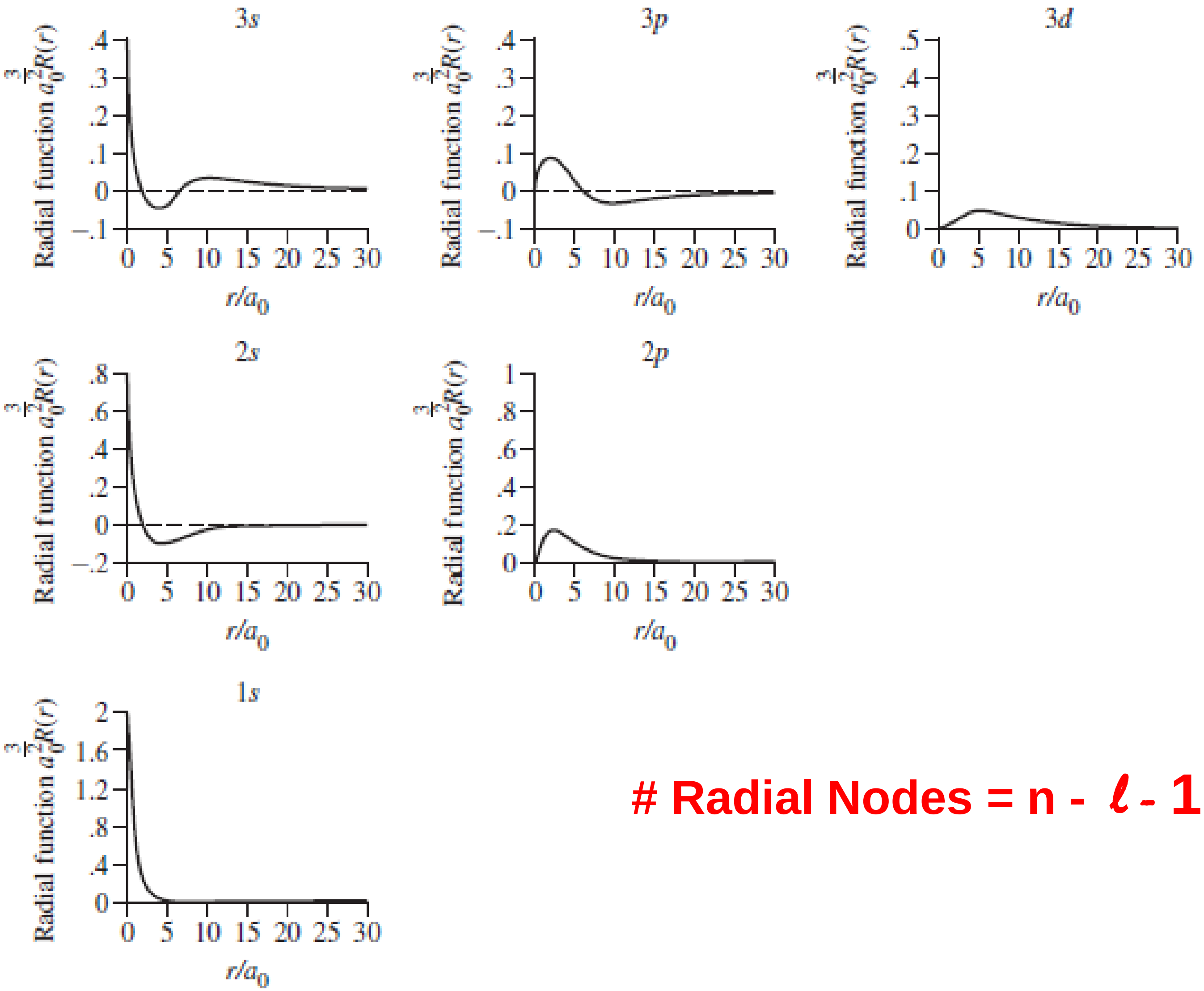


FIGURE 2-10 Energy Level Splitting and Overlap. The differences between the upper levels are exaggerated for easier visualization.

Radial Wave Functions and Nodes

Radial Wave Functions



 # Radial Nodes = n - ℓ - 1

Radial Probability Functions and Nodes

Radial Probability Functions

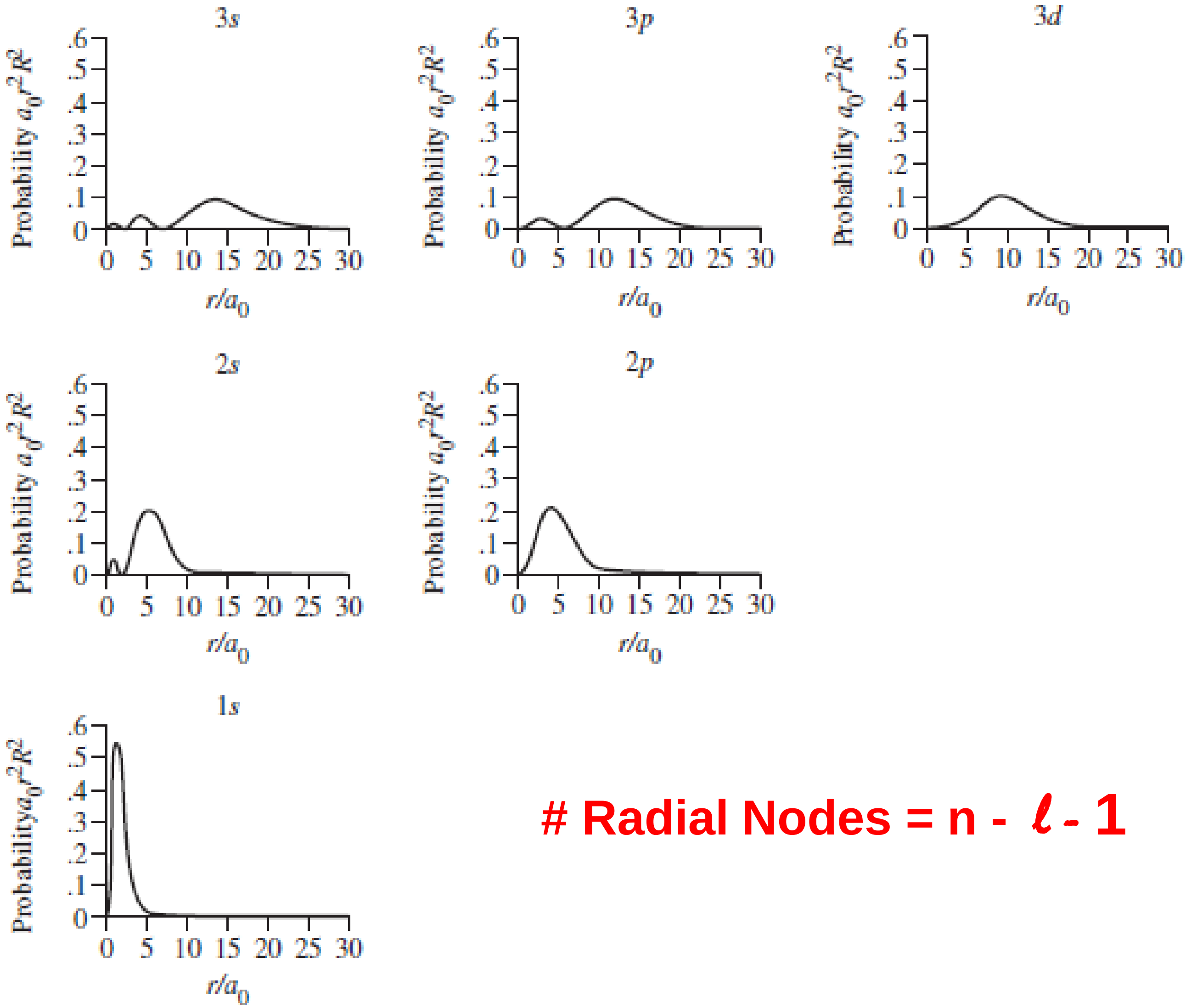


FIGURE 2.7 Radial Wave Functions and Radial Probability Functions.

 # Radial Nodes = n - ℓ - 1

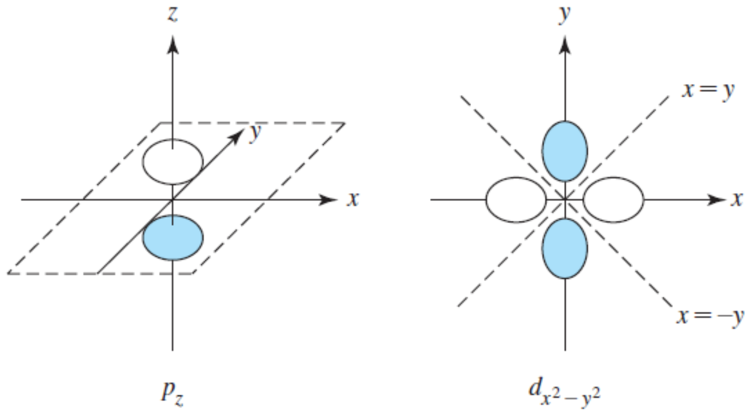
TABLE 2.5 Nodal Surfaces

Angular Nodes [$Y(\theta, \phi) = 0$]	
Examples (number of angular nodes)	
s orbitals	0
p orbitals	1 plane for each orbital
d orbitals	2 planes for each orbital except d_{z^2}
	1 conical surface for d_{z^2}

Radial Nodes [$R(r) = 0$]					
Examples (number of radial nodes)					
1*s*	0	2*p*	0	3*d*	0
2*s*	1	3*p*	1	4*d*	1
3*s*	2	4*p*	2	5*d*	2

 # Radial Nodes = n - ℓ - 1

 # Angular Nodes = ℓ

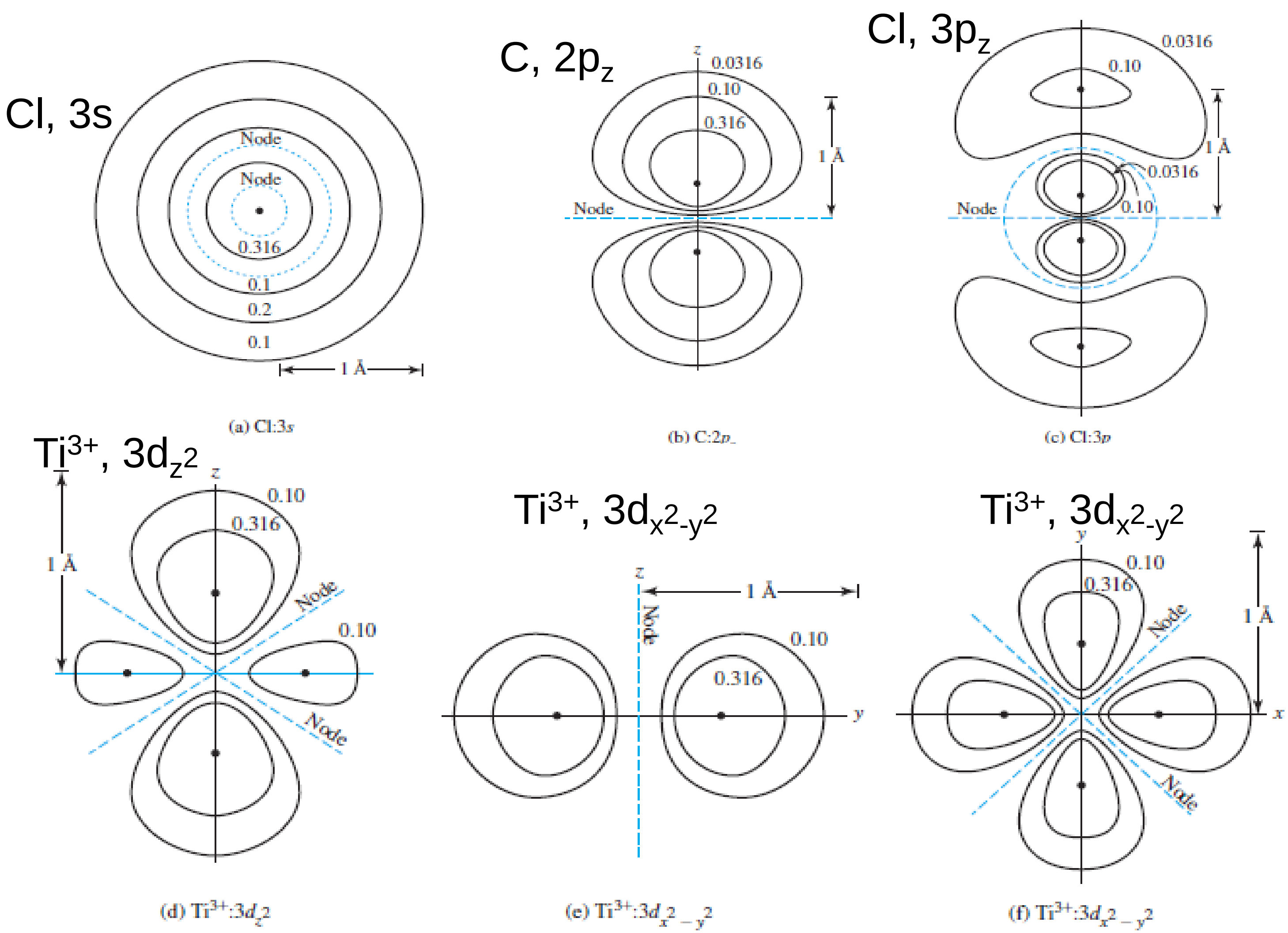


Summary:
Radial Nodes = n - ℓ - 1
Angular Nodes = ℓ

Total # Nodes = n - 1

Where are the nodes in orbitals???

(MFT, Figure 2.8 shows both radial and angular)

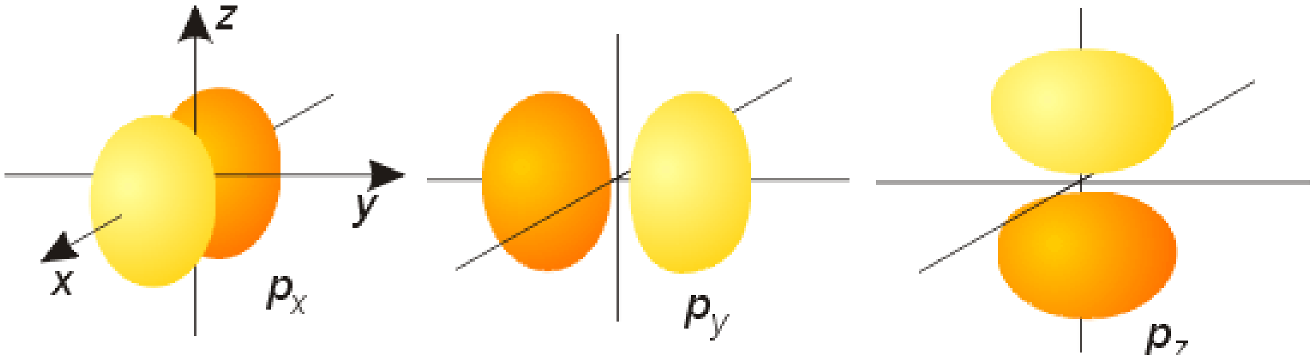


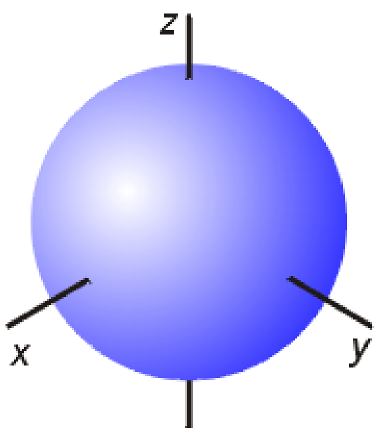
Note the orientation of the viewer: down z or x or y axes

Orbitals and Shapes/Electron Distribution

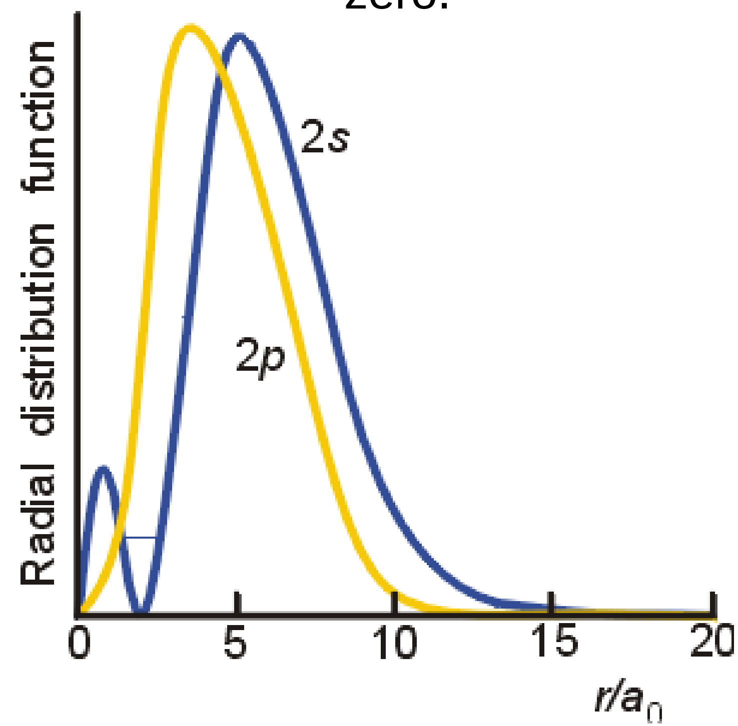
Each p-orbital has two lobes with positive and negative values (phases) of the wavefunction either side of the nucleus separated by a nodal plane where the wavefunction is zero.

The p-orbitals

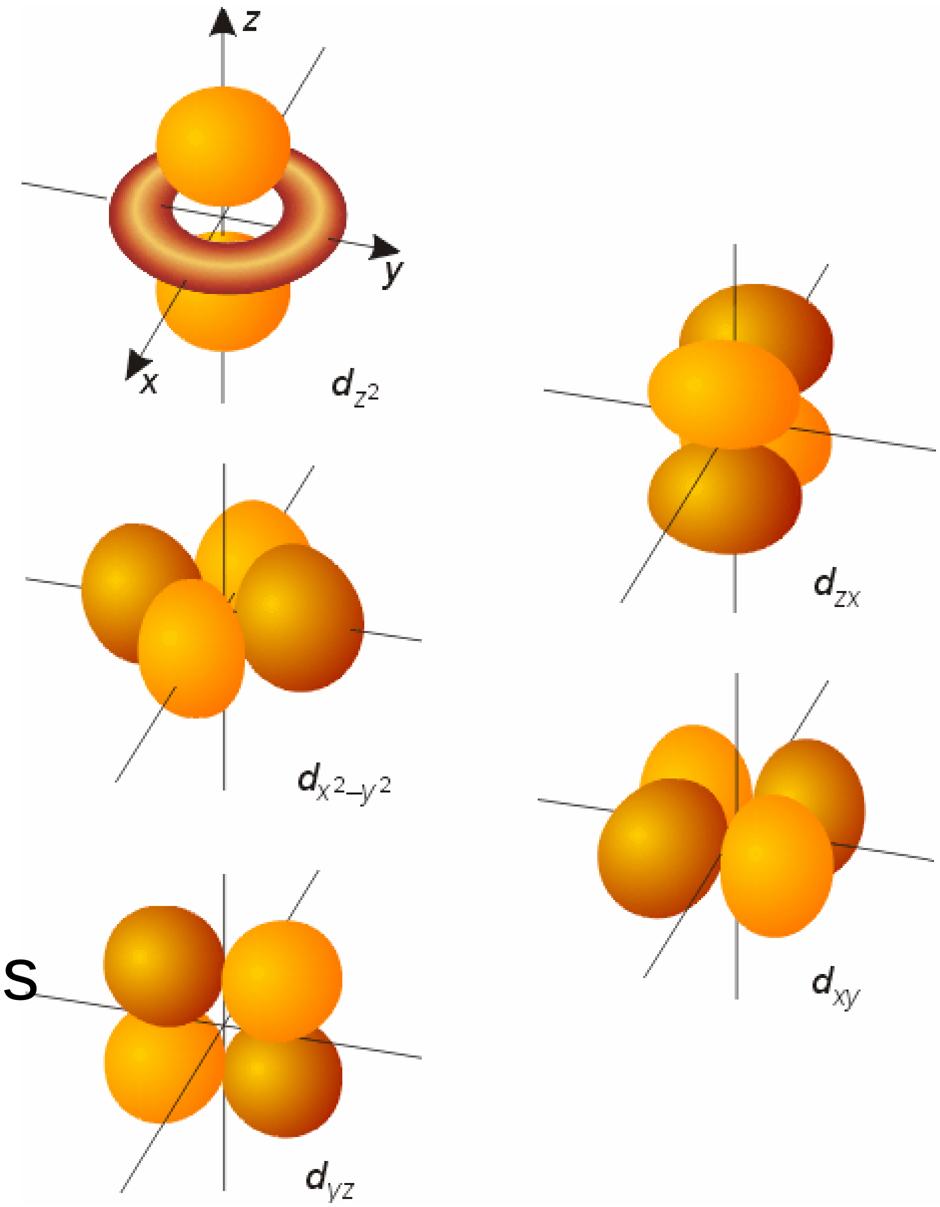




The s-orbital

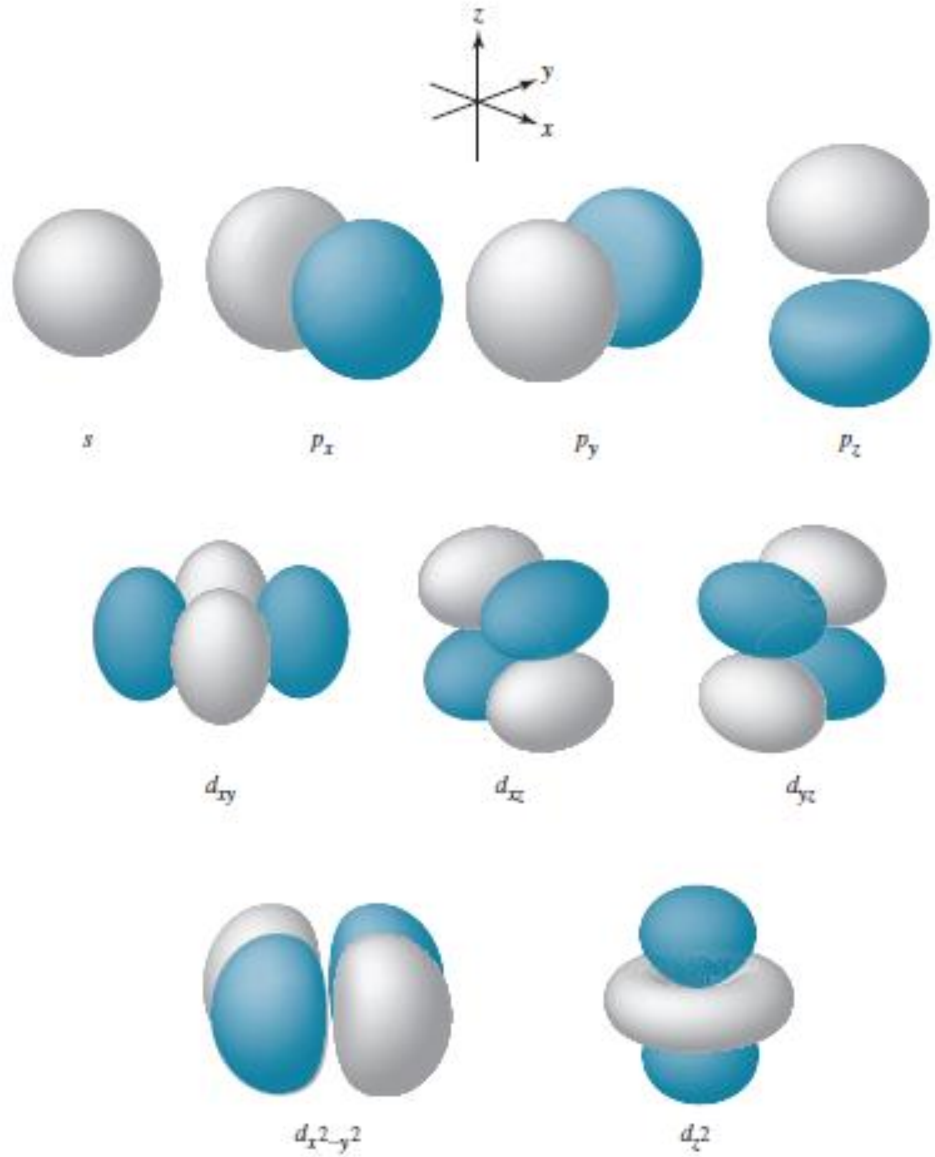


The d-orbitals



Two angular nodes

As pictured in MFT, Figure 2.6



Quantum Numbers

n is the *principal quantum number*, indicates the size of the orbital, has all positive integer values of 1 to ∞(infinity)

ℓ (angular momentum)	orbital
0	s
1	p
2	d
3	f

ℓ is the *angular momentum quantum number*, represents the shape of the orbital, has integer values of n-1 to 0

m_ℓ is the *magnetic quantum number*, represents the spatial direction of the orbital, can have integer values of -ℓ to 0 to ℓ

Other terms: electron configuration, noble gas configuration, valence shell

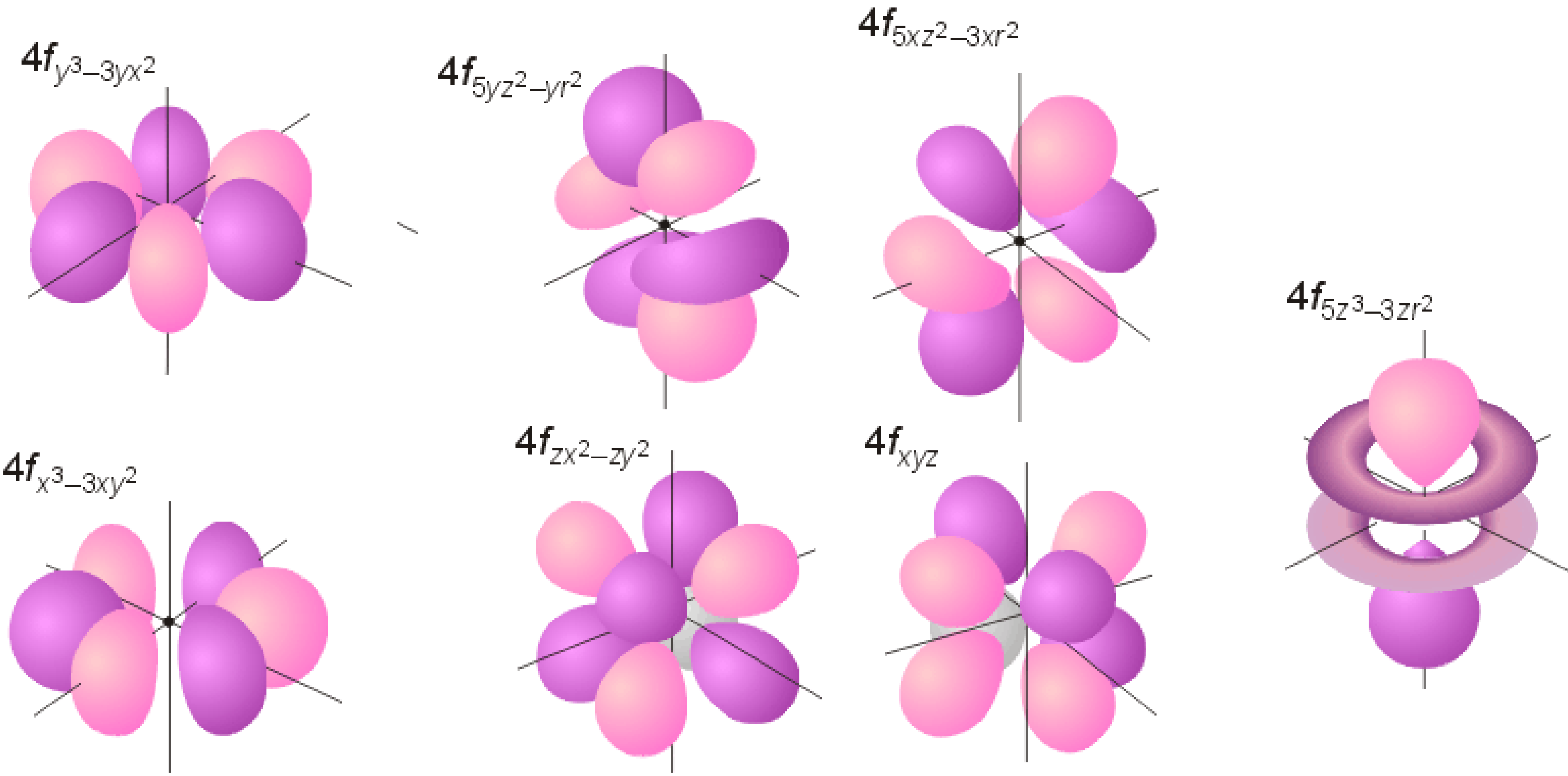
m_s is the *spin quantum number*, has little physical meaning, can have values of either +1/2 or -1/2

Pauli Exclusion principle: no two electrons can have all four of the same quantum numbers in the same atom

Hund's Rule: when electrons are placed in a set of degenerate orbitals, the *ground state* has as many electrons as possible in different orbitals, and with *parallel spin*.

Aufbau (Building Up) Principle: the ground state electron configuration of an atom can be found by putting electrons in orbitals, starting with the lowest energy and moving progressively to higher energy.

The f-orbitals



Quantum Numbers

n is the *principal quantum number*, indicates the size of the orbital, has all positive integer values of 1 to ∞(infinity)

ℓ (angular momentum)	orbital
0	s
1	p
2	d
3	f

ℓ is the *angular momentum quantum number*, represents the shape of the orbital, has integer values of n-1 to 0

m_ℓ is the *magnetic quantum number*, represents the spatial direction of the orbital, can have integer values of -ℓ to 0 to ℓ

m_s is the *spin quantum number*, has little physical meaning, can have values of either +1/2 or -1/2

Other terms: electron configuration, noble gas configuration, valence shell

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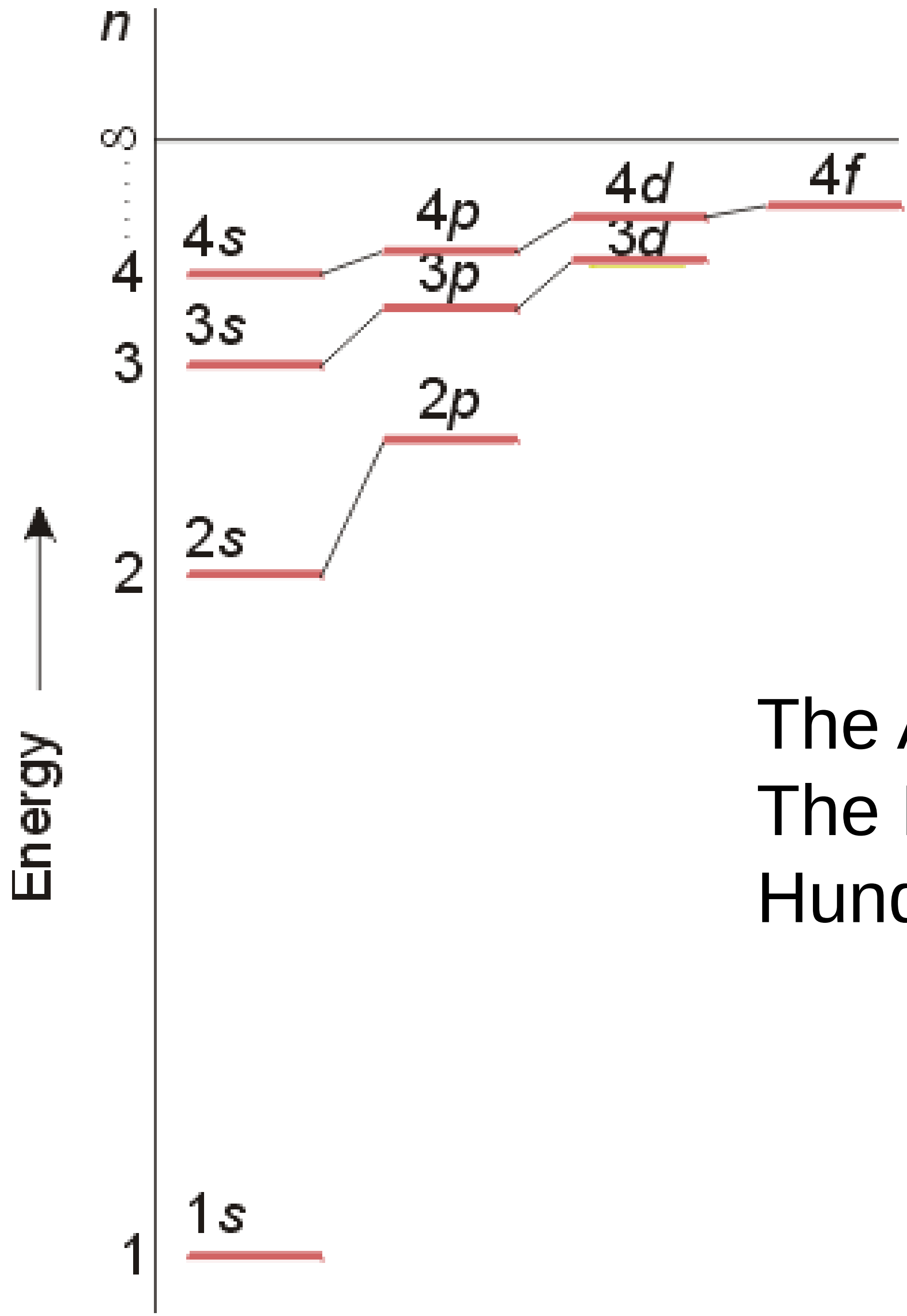
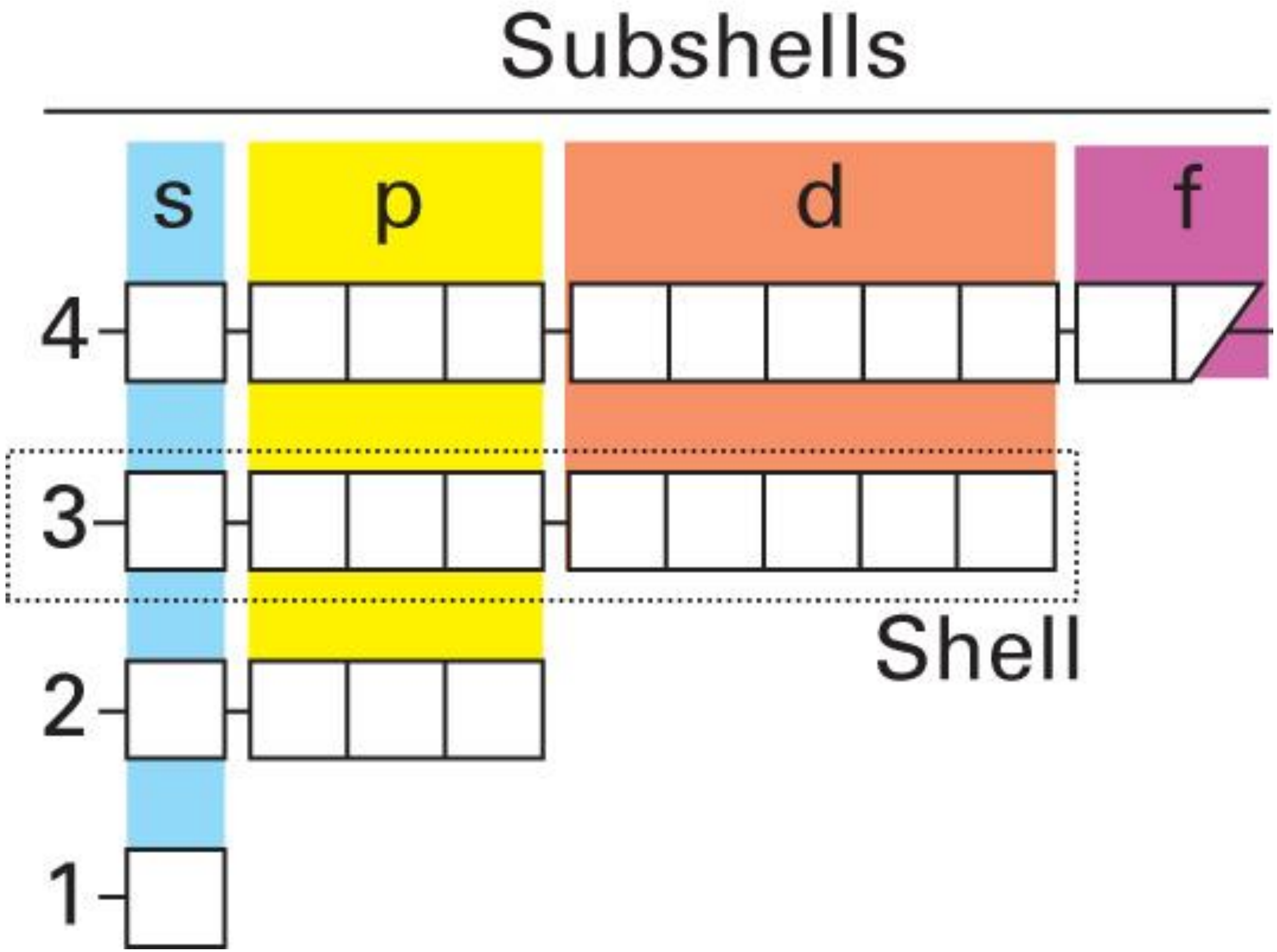


Energy Levels for Electron Configurations

The Aufbau Principle
The Pauli Exclusion Principle
Hund's Rules

Box Diagrams



While n is the principal energy level,
the ℓ value also has an effect

Heavier elements, more chances
For close energy levels

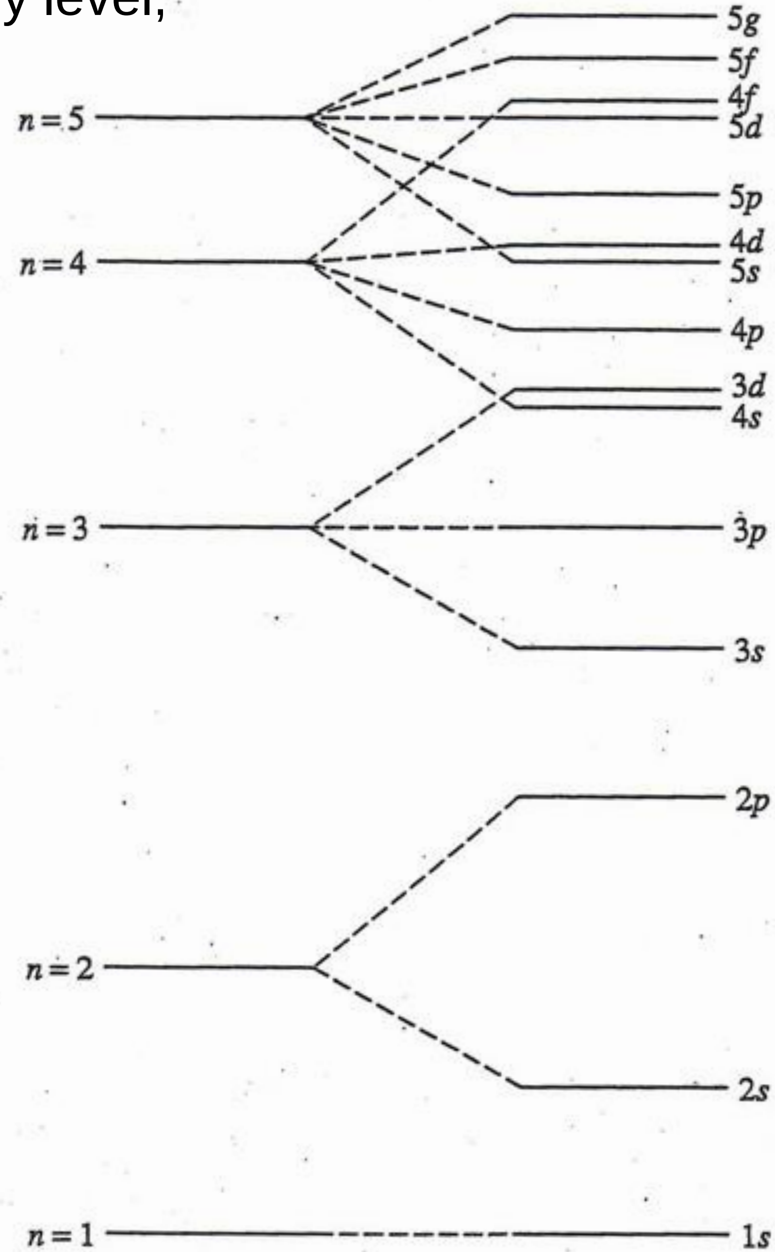
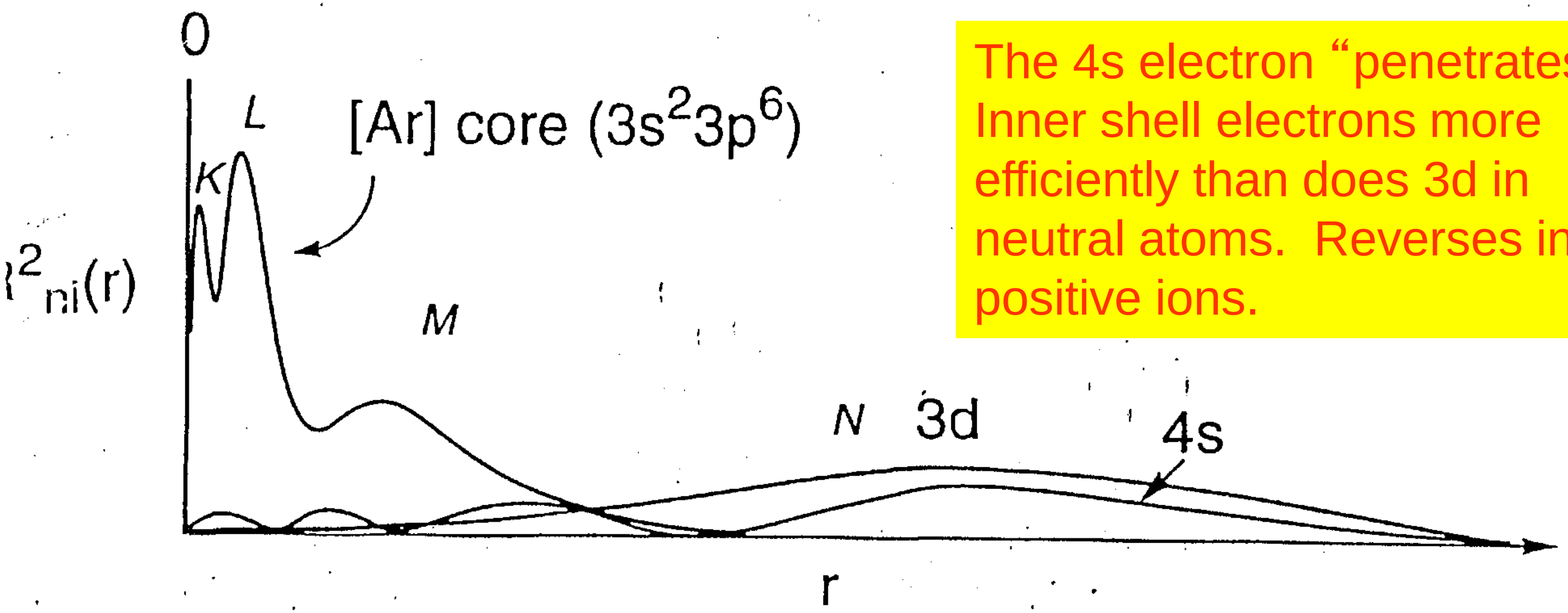


FIGURE 2-10 Energy Level Splitting and Overlap. The differences between the upper levels are exaggerated for easier visualization.

Screening:

Interpenetration of 4s and 3d orbitals at K:



The 4s electron "penetrates" Inner shell electrons more efficiently than does 3d in neutral atoms. Reverses in positive ions.

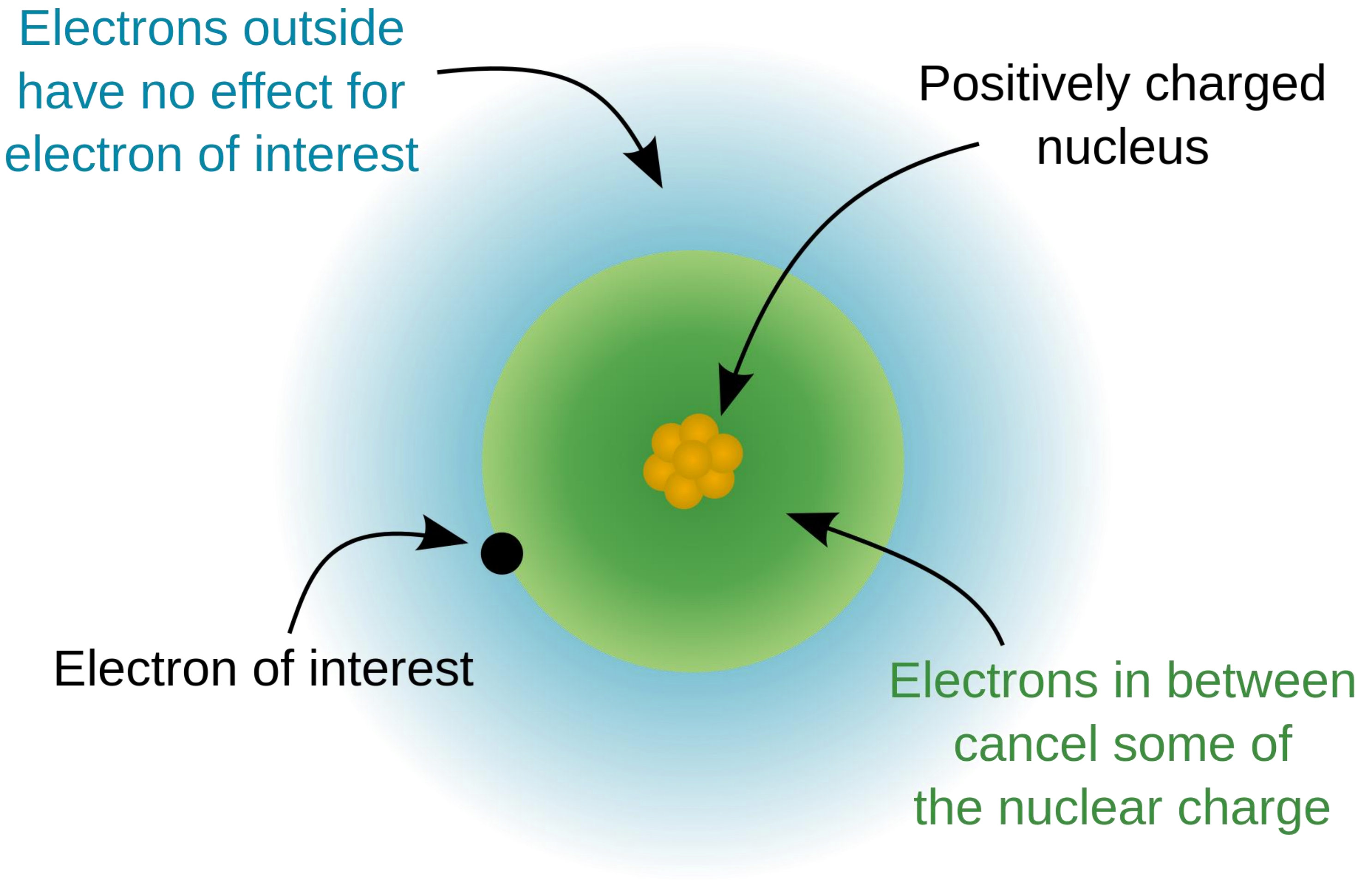
order of orbital filling, Z = 19-36, for neutral atoms:

—— 4s — 3d —— 4p →

for positive ions, the order of 4s and 3d is reversed:

—— 3d — 4s —— 4p →

How to handle atoms larger than H? Effective Nuclear Charge or Z_{eff}



Slater's Rules for Calculating Z_{eff}

1) Write the electron configuration for the atom as follows:

(1*s*)(2*s*,2*p*)(3*s*,3*p*) (3*d*) (4*s*,4*p*) (4*d*) (4*f*) (5*s*,5*p*)

2) Any electrons to the right of the electron of interest contributes no shielding. (Approximately correct statement.)

3) All other electrons in the same group as the electron of interest shield to an extent of 0.35 nuclear charge units

4) If the electron of interest is an *s* or *p* electron: All electrons with one less value of the principal quantum number shield to an extent of 0.85 units of nuclear charge. All electrons with two less values of the principal quantum number shield to an extent of 1.00 units.

5) If the electron of interest is an *d* or *f* electron: All electrons to the left shield to an extent of 1.00 units of nuclear charge.

6) Sum the shielding amounts from steps 2 - 5 and subtract from the nuclear charge value to obtain the effective nuclear charge.

Slater's Rules: Examples

Calculate Z_{eff} for a valence electron in fluorine.

$(1s^2)(2s^2,2p^5)$

Rule 2 does not apply; therefore, for a valence electron the shielding or screening is (0.35 • 6) + (0.85 • 2) = 3.8

$$Z_{eff} = 9 - 3.8 = 5.2$$

Calculate Z_{eff} for a 6s electron in Platinum.

$(1s^2)(2s^2,2p^6)(3s^2,3p^6)\ (3d^{10})\ (4s^2,4p^6)\ (4d^{10})\ (4f^{14})\ (5s^2,5p^6)\ (5d^8)\ (6s^2)$

Rule 2 does not apply, and the shielding is:

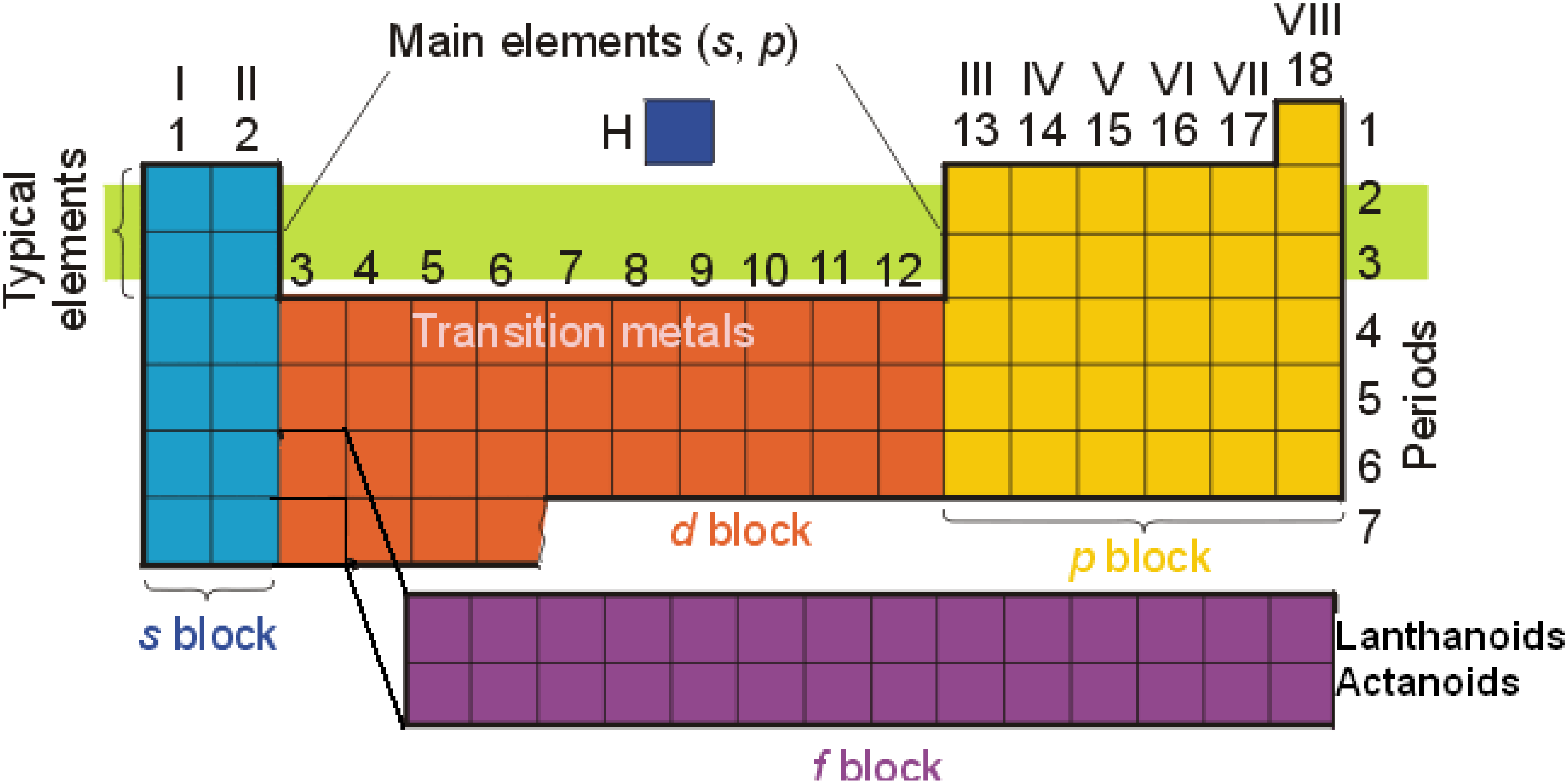
$$(0.35 \cdot 1) + (0.85 \cdot 16) + (60 \cdot 1.00) = 73.95$$

W. H. FREEMAN

Table 1.2 Effective nuclear charge, Z_{eff}

	H							**He**
Z	1							2
1s	1.00							1.69
	Li	**Be**	**B**	**C**	**N**	**O**	**F**	**Ne**
Z	3	4	5	6	7	8	9	10
1s	2.69	3.68	4.68	5.67	6.66	7.66	8.65	9.64
2s	1.28	1.91	2.58	3.22	3.85	4.49	5.13	5.76
2p			2.42	3.14	3.83	4.45	5.10	5.76
	Na	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**	**Ar**
Z	11	12	13	14	15	16	17	18
1s	10.63	11.61	12.59	13.57	14.56	15.54	16.52	17.51
2s	6.57	7.39	8.21	9.02	9.82	10.63	11.43	12.23
2p	6.80	7.83	8.96	9.94	10.96	11.98	12.99	14.01
3s	2.51	3.31	4.12	4.90	5.64	6.37	7.07	7.76
3p			4.07	4.29	4.89	5.48	6.12	6.76

The Periodic Table



Special Names Group 1: Alkali Metals
Group 2: Alkaline-earth metals
Group 16: Chalcogens
Group 17: Halogens
Group 18: Noble gases

http://www.youtube.com/watch?v=SmwlzwGMMwc

Electron Configurations and Screening

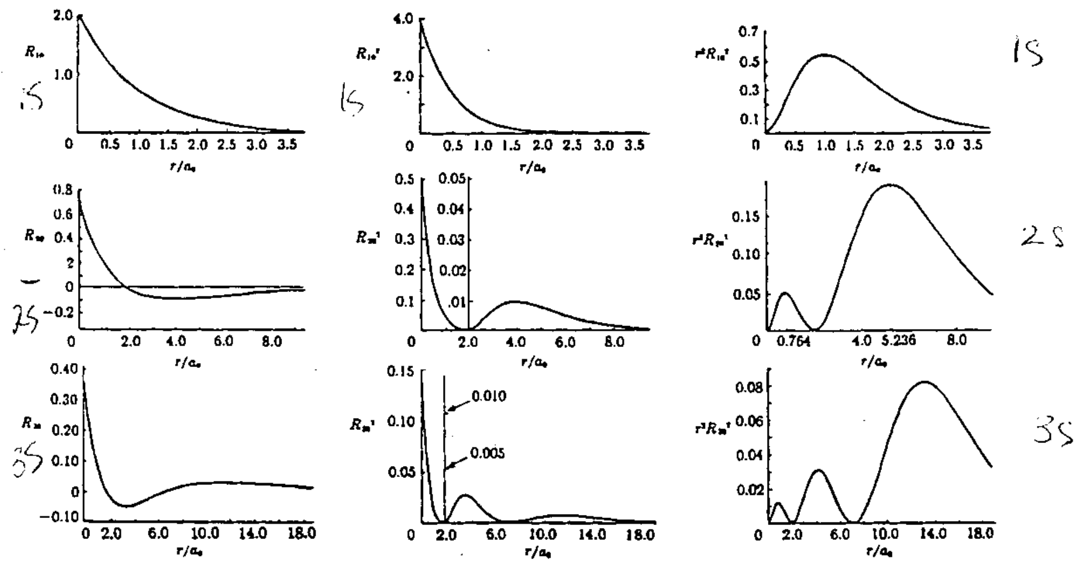
H → Ar energy order:

1s < 2s < 2p < 3s < 3p

I'd like to make two points:

Radial Wave Functions (hydrogenic)

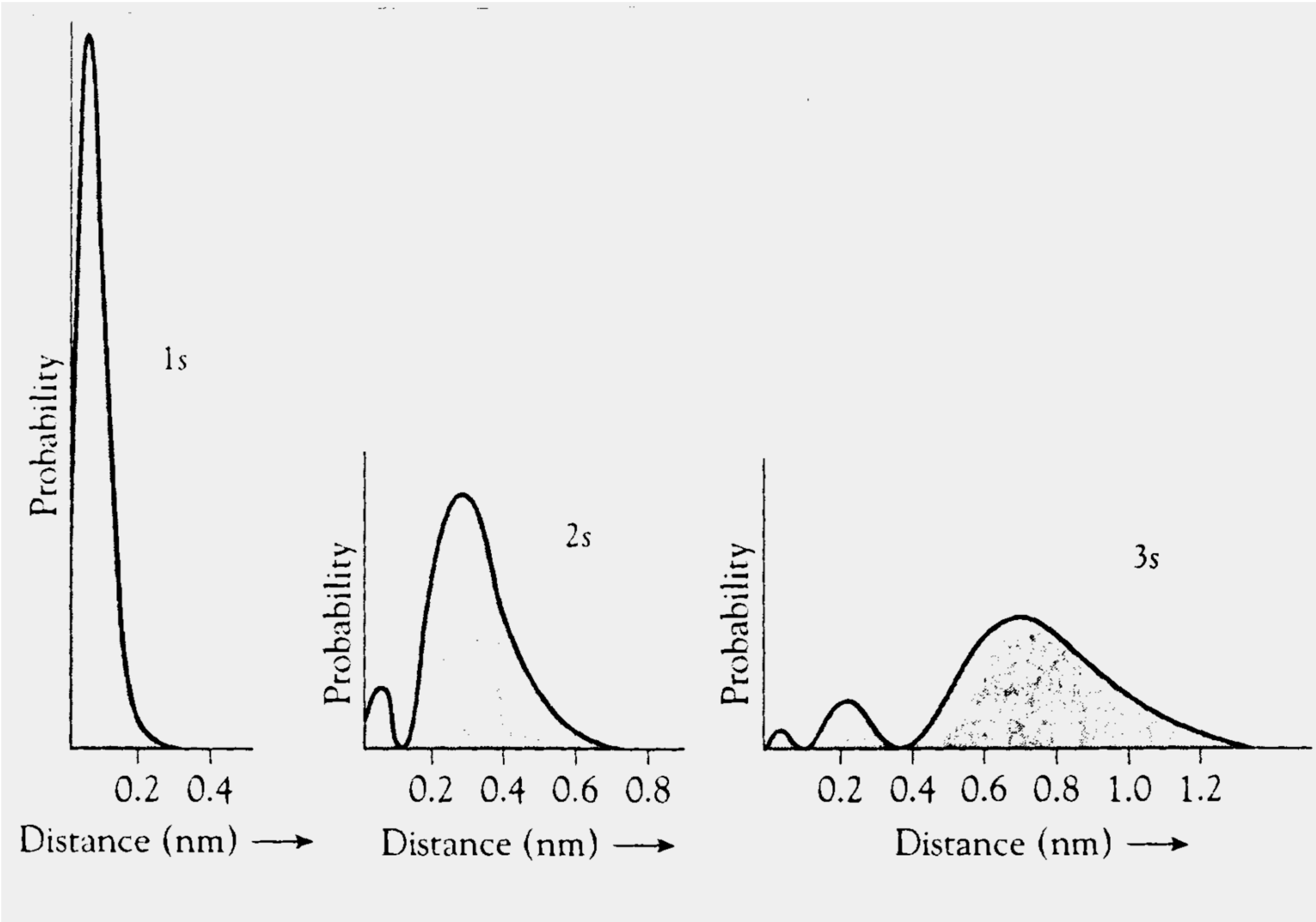
$R_{nl}(r)$ | $R^2_{nl}(r)$ | $r^2R^2_{nl}(r)$



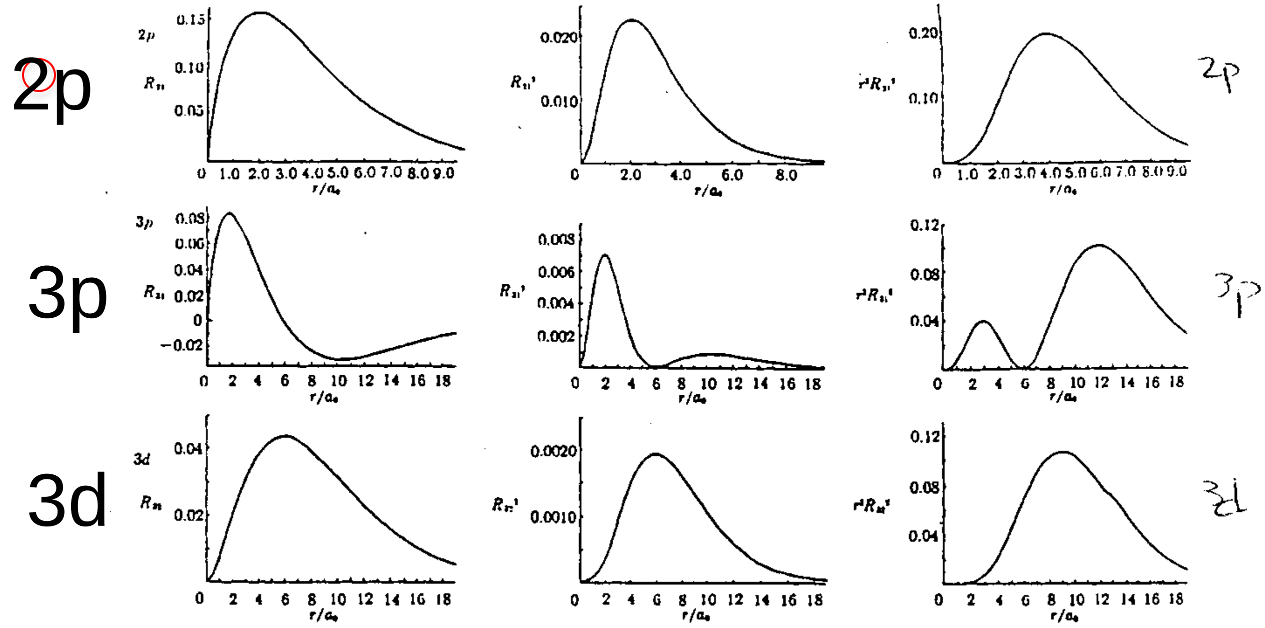
$R^2_{nl}(r)$: radial probability function

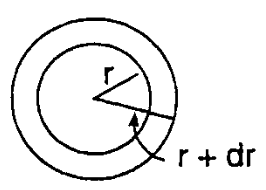
- s orbitals (l = 0) have non-zero probability at nucleus
- node: distance at which, or surface on which, the electron described by that orbital is not found

Figure 1.7 The variation of the radial density distribution function with distance from the nucleus for electrons in the 1s, 2s, and 3s orbitals of a hydrogen atom.



Radial Wave Functions (hydrogenic)





$$V = \frac{4}{3}\pi r^3$$

$$dV = \underbrace{4\pi r^2}_{\text{surface area}} dr$$

$$R^2_{nl}(r)dV = 4\pi r^2 R^2_{nl}(r)dr = S(r)$$

S(r): surface density function

- gives the probability of finding e^-_{nl} in a spherical shell of thickness dr at a distance r from the nucleus http://www.youtube.com/watch?v=lhTSfOZUNLo

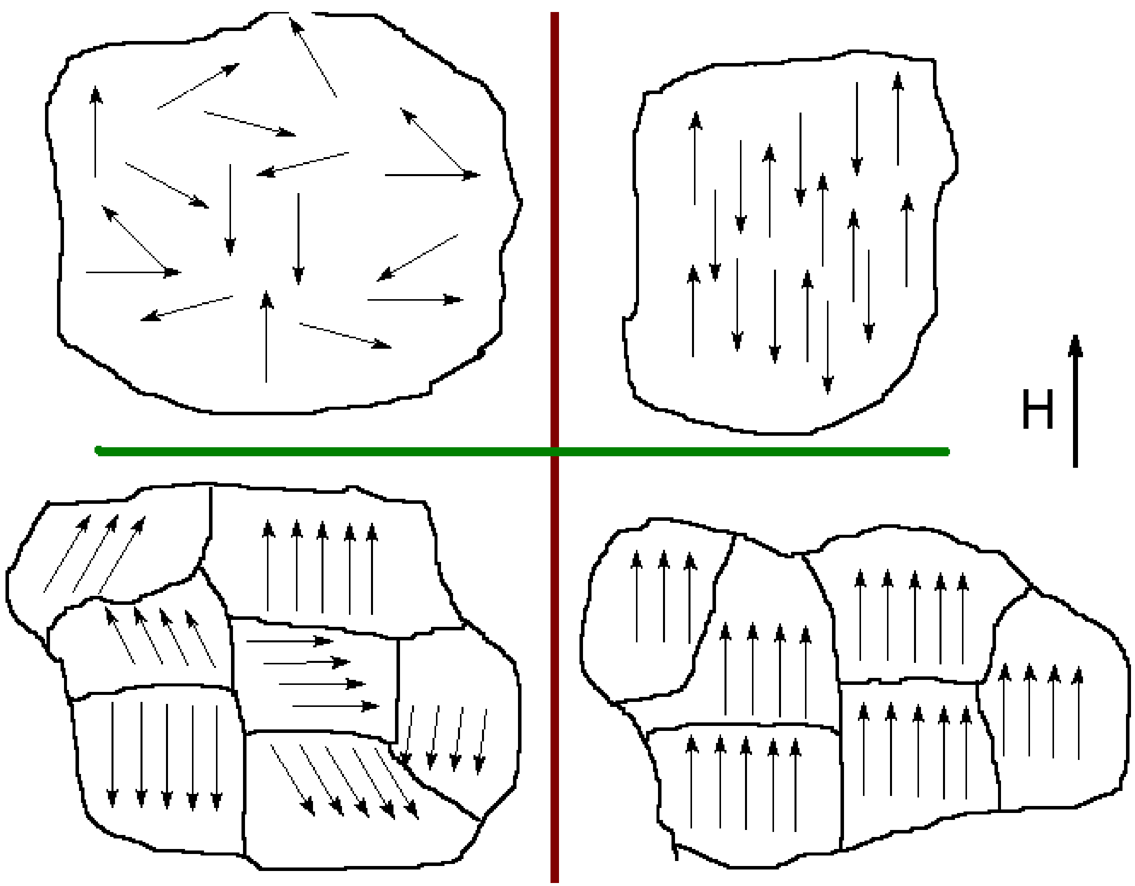
Three types of Magnetic Behavior

Paramagnetism: atoms, molecules, and solids with unpaired electrons are attracted in a magnetic field

Diamagnetic: substances with no unpaired electrons which are weakly repelled in a magnetic field

Ferro-magnetism: the unpaired electons are aligned with their neighbors even in the absence of a magnetic field

Magnetic domains: the groups of mutually aligned spins in a ferromagnetic substance



Ferro-magnet In the absence of a magnetic field

Ferro-magnet In the presence of a magnetic field

Figure 1.4 The possible sets of quantum numbers for *n* = 1 and *n* = 2.

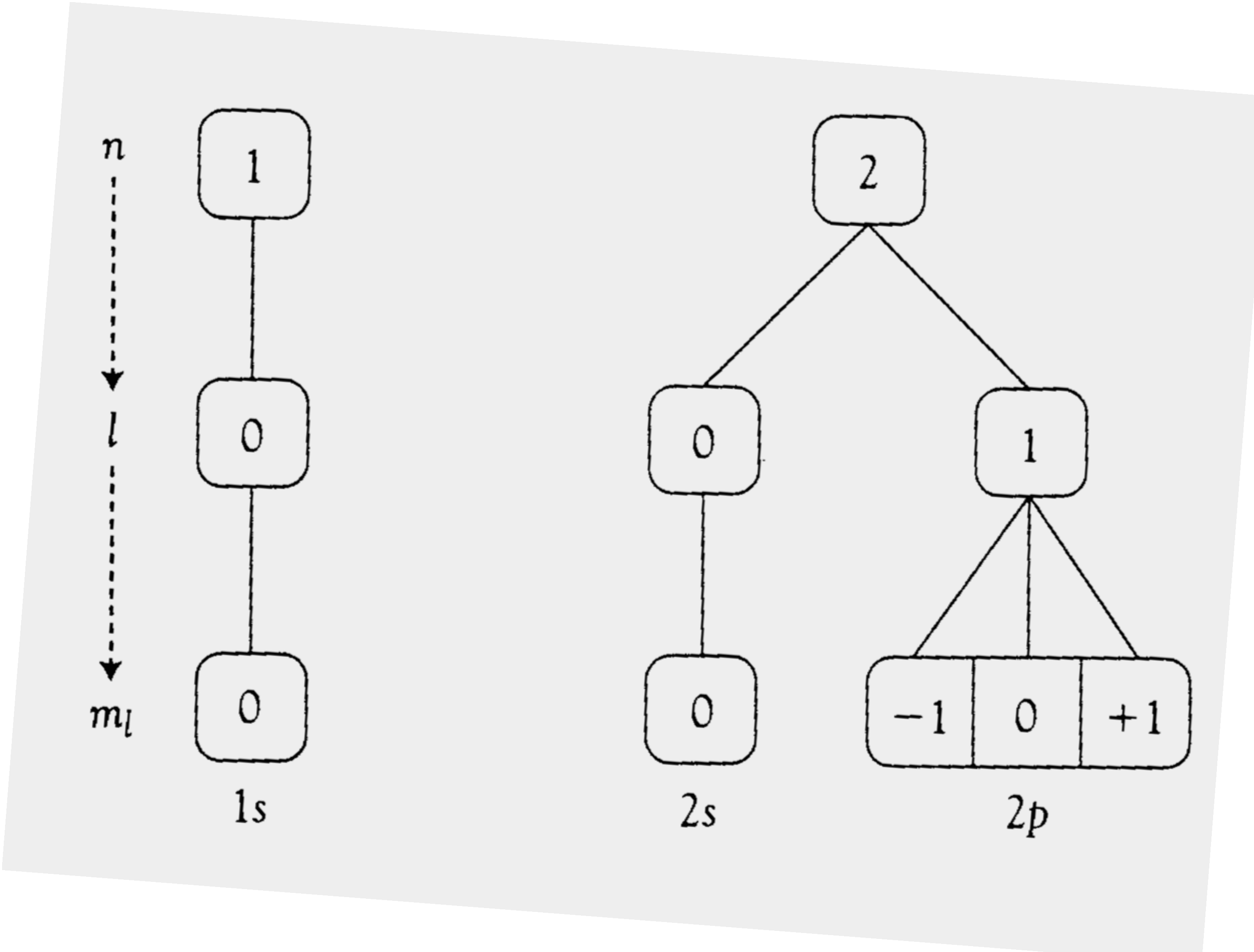
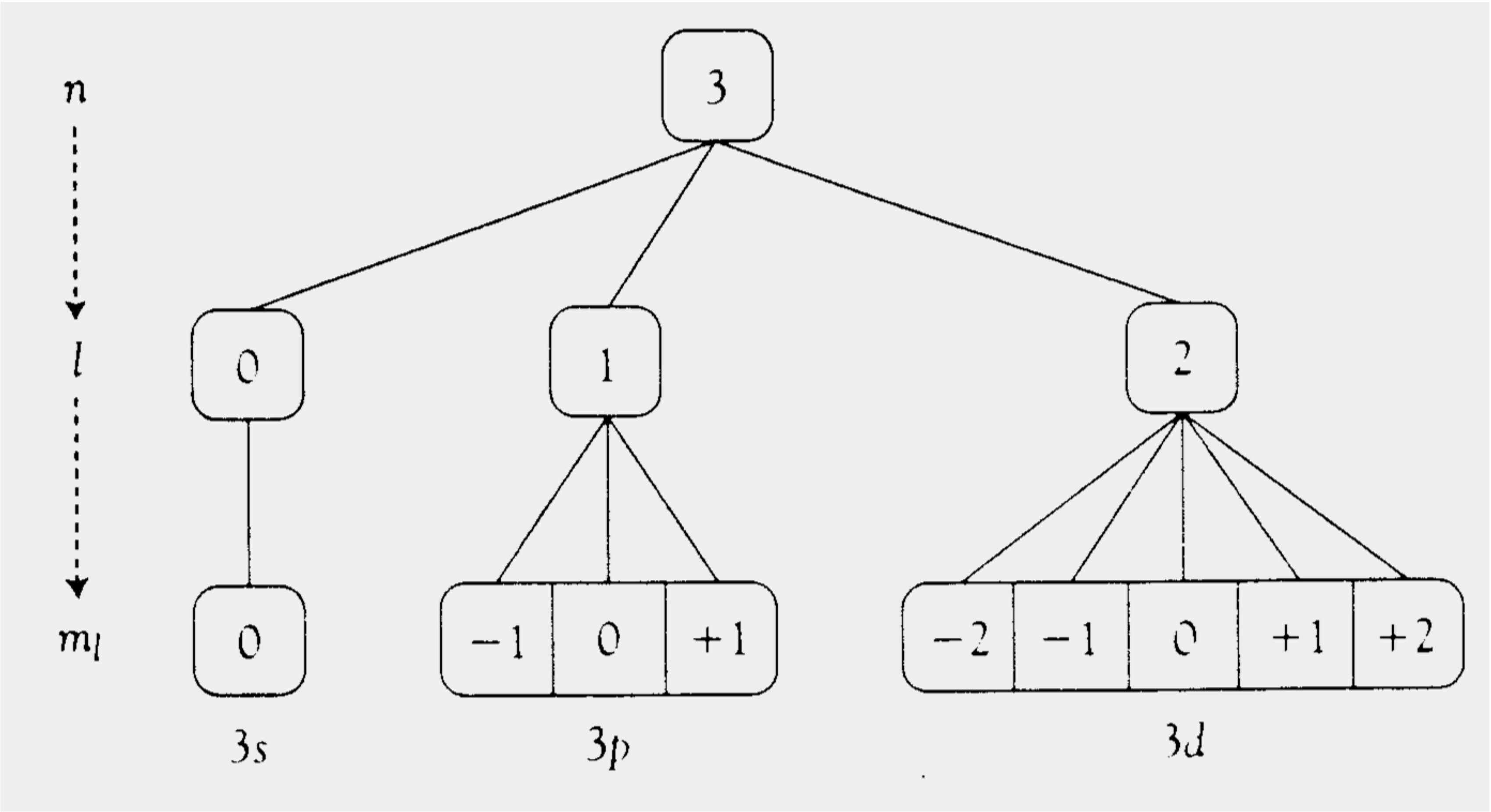


Figure 1.5 The possible sets of quantum numbers for n = 3.



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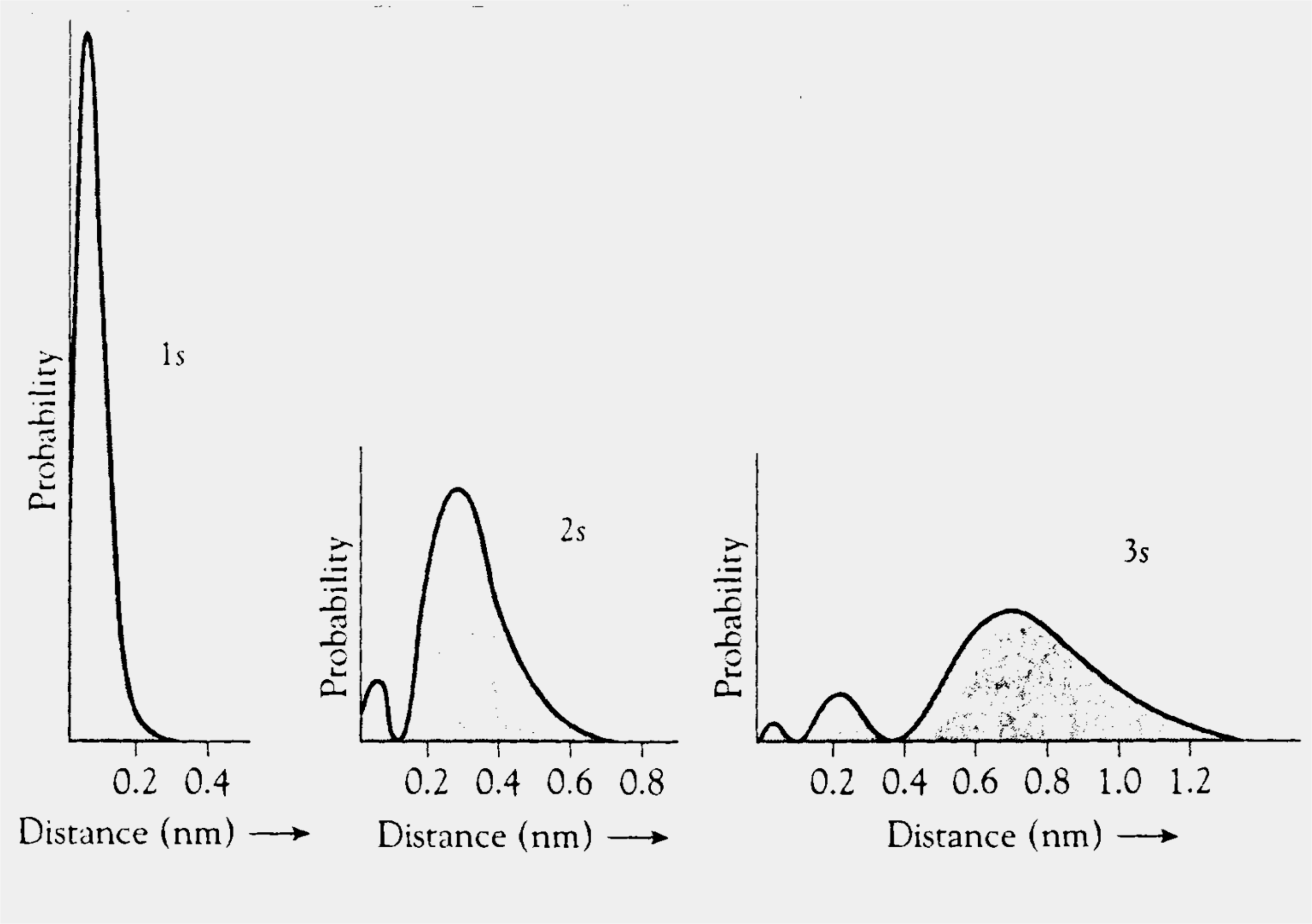
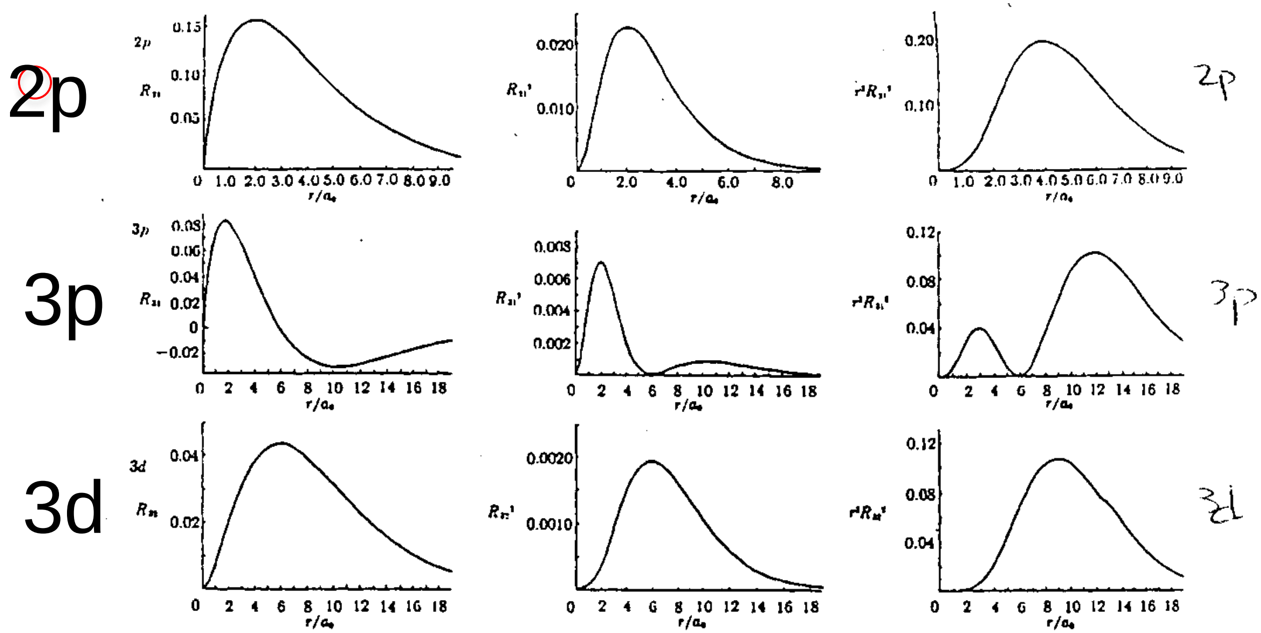
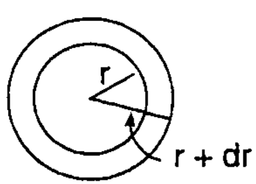


Radial Wave Functions (hydrogenic)



sphere of radius r



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$$dV = \underbrace{4\pi r^2}_{\text{surface area}} dr$$

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S(r): surface density function

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