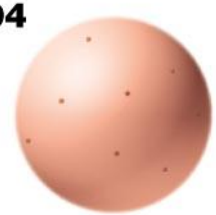
Review Models of the Atom

1803

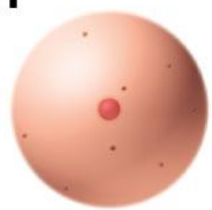
Dalton proposes the **indivisible** unit of an element is the atom.

1904



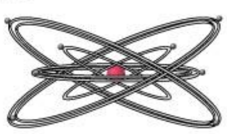
Thomson discovers electrons, believed to reside within a sphere of uniform positive charge (the “plum-pudding model).

1911



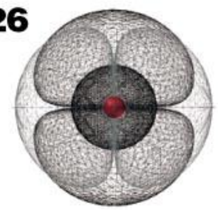
Rutherford demonstrates the existence of a positively charged nucleus that contains nearly all the mass of an atom.

1913



Bohr proposes fixed circular orbits around the nucleus for electrons.

1926



In the current model of the atom electrons occupy regions of spac (orbitals) around the nucleus determined by their energies.

Rutherford's Gold Foil Expt.

Observation

1. Most α particles pass through undeflected.

1. Some α particles are slightly deflected.

1. Few particles are reflected back (i.e. do not pass through).

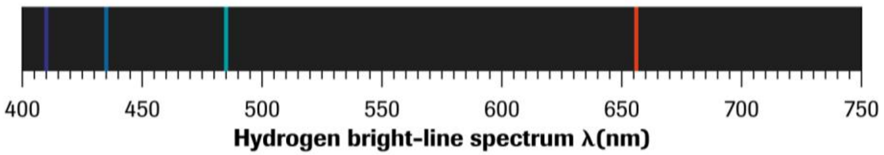
Inference

1. Atom is mostly empty space.
2. Positively charged part of atom must be small.
3. Core of atom is very small, dense and positively charged.

Bohr Model and Emission Line Spectra of Hydrogen

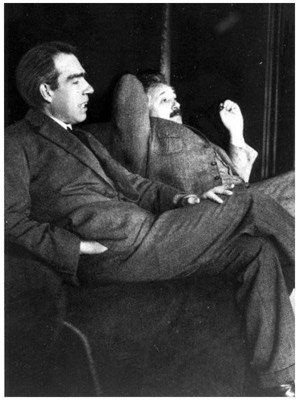
Observation

- 4 distinct bands of colour



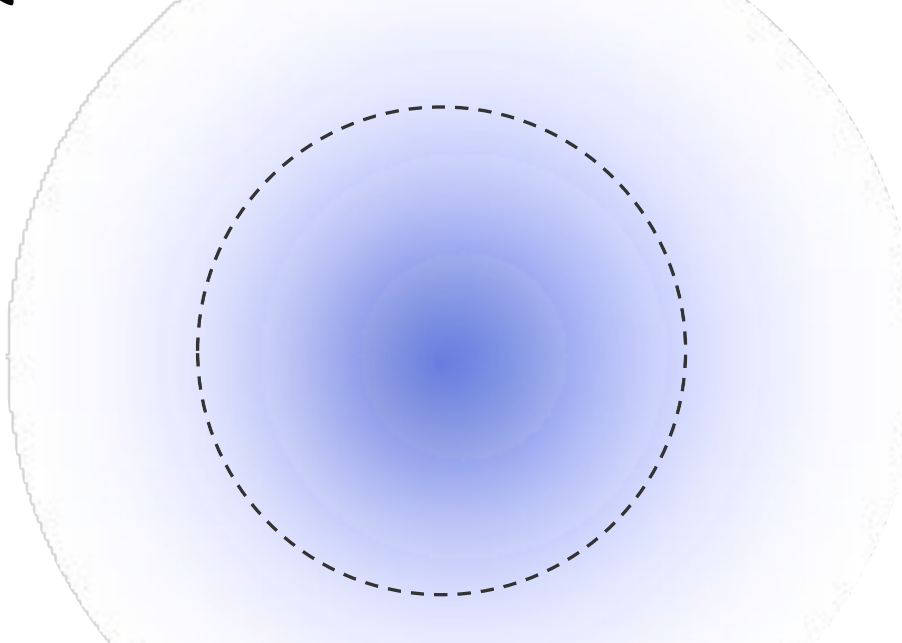
Inference

- e^- can only posses certain energies (quantized)
- Certain wavelengths of light are absorbed by atom as e^- transitions from lower to higher energy level (orbit).
- As e^- returns to lower energy level, photons of light are released.
- If energy released corresponds to a wavelength of visible light, a specific band of colour will be observed.



Niels Bohr &
Albert Einstein

Quantum Mechanical Model



Modern atomic theory describes the electronic structure of the atom as the *__probability__ of finding electrons within certain regions of space* *(__orbitals__)*.

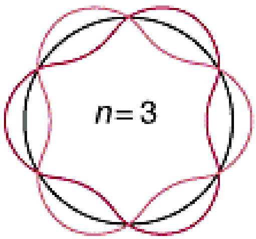
Electrons as Waves

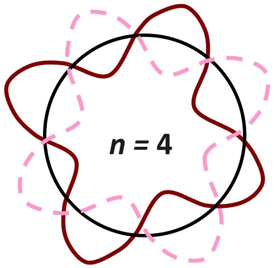
Louis de Broglie ~1924

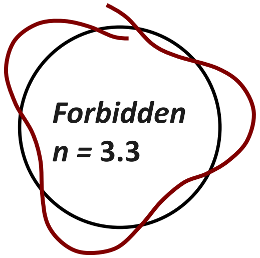
- Energy is quantized. photons –> discrete packets of light (particle-like)

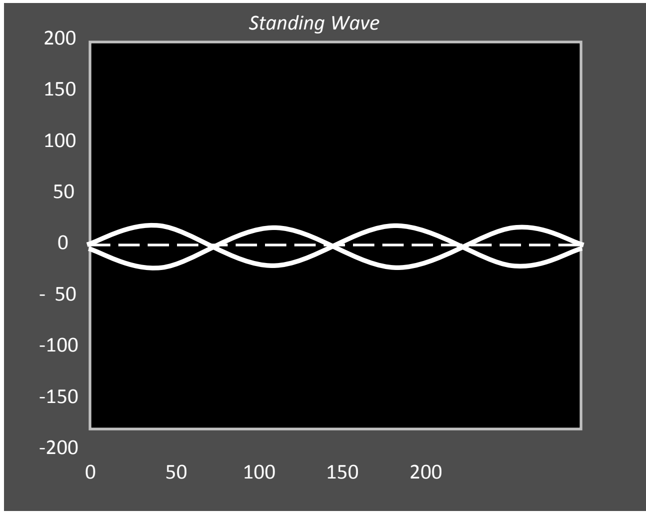
  (E=hν   Planck)

Standing waves









- Matter has wavelike properties. (deBroglie)

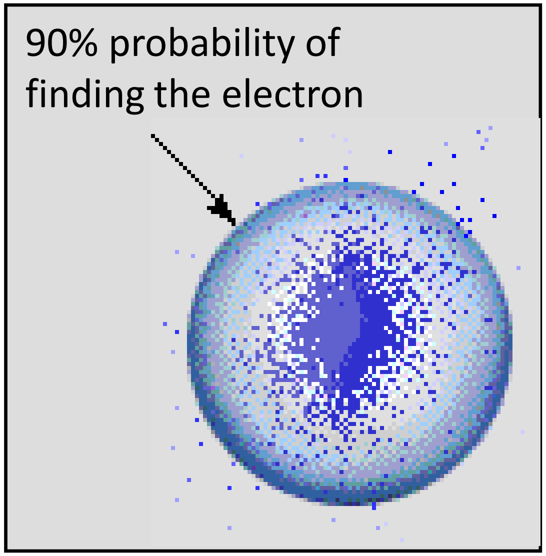
Adapted from work by Christy Johannesson www.nisd.net/communicationsarts/pages/chem

Schrodinger's Wave Equation

- Used to determine the probability of finding the electron at any given distance from the nucleus
- Orbital "electron cloud"

  -region in space where the electron is likely to be found

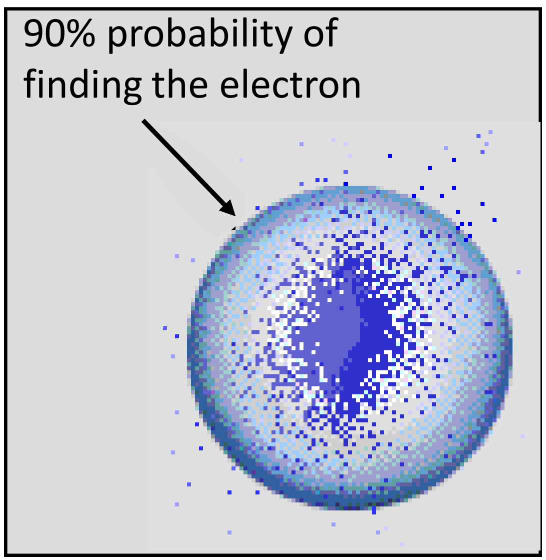
$$\Psi_{1s} = \frac{1}{\sqrt{\pi}}\left(\frac{Z}{a_0}\right)^{3/2} e^{-\sigma}$$



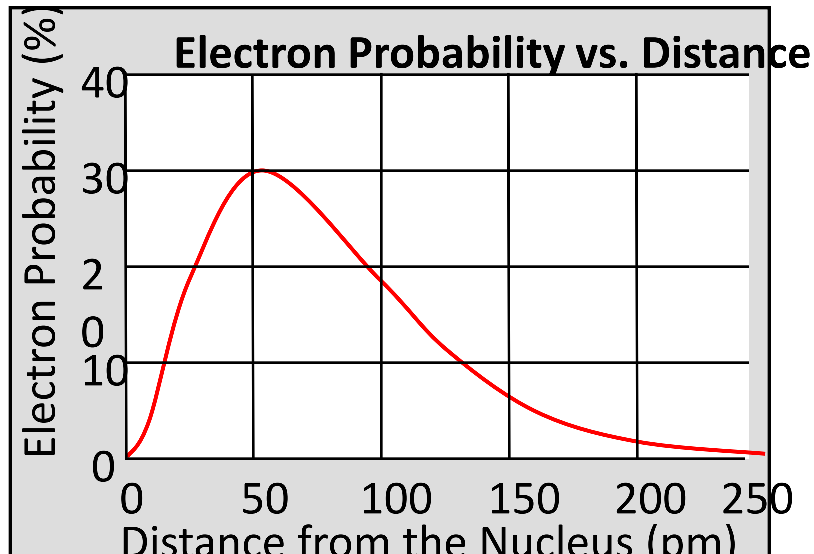
Orbital

Quantum Mechanics

- **Orbital** ("electron cloud")
  - Region in space where there is 90% probability of finding an electron



Orbital

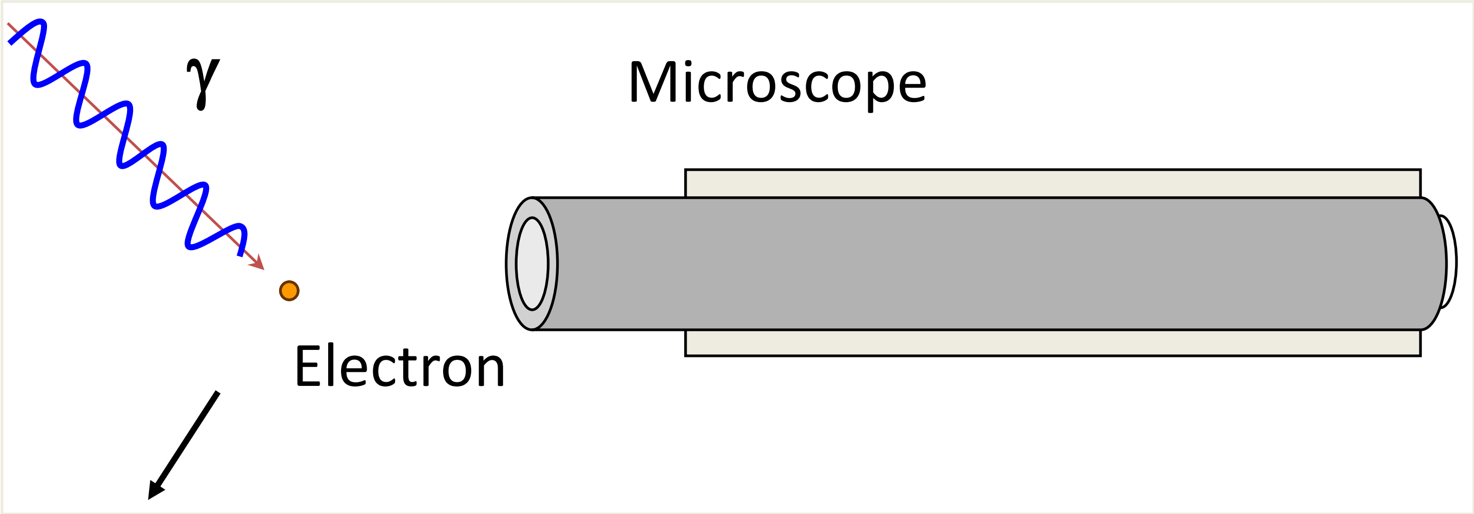


Courtesy Christy Johannesson www.nisd.net/communicationsarts/pages/chem

p. 200 Crazy Enough?

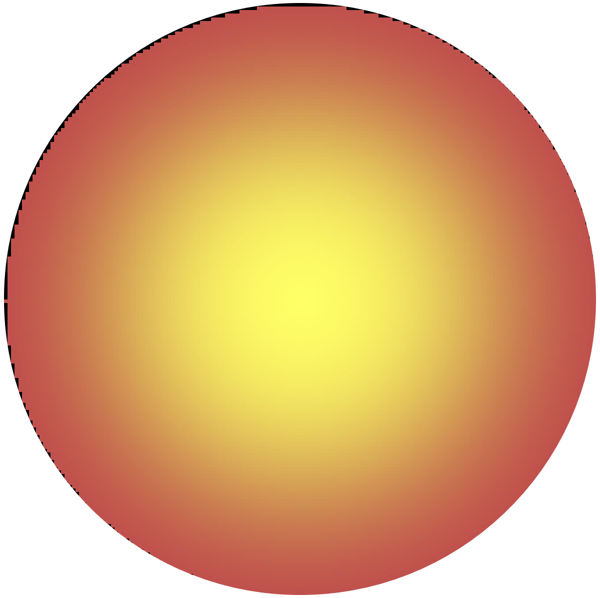
Werner Heisenberg
~1926

- Heisenberg Uncertainty Principle
  - Impossible to know both the *velocity* and *position* of an electron at the same time

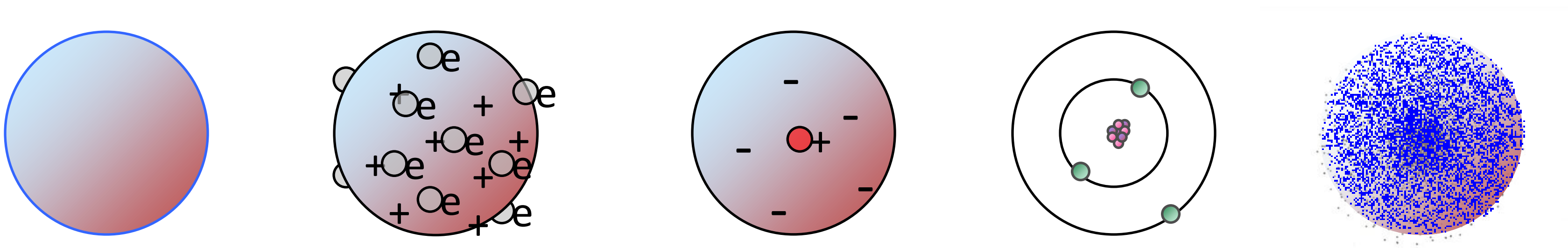


Modern View

- The atom is mostly empty space
- Two regions
  - Nucleus
    - protons and neutrons
  - Electron cloud
    - region where you might find an electron



Models of the Atom



Greek model (400 B.C.)

Thomson's plum-pudding model (1897)

Rutherford's model (1909)

Bohr's model (1913)

Charge-cloud model (present)

1803 John Dalton pictures atoms as tiny, indestructible particles, with no internal structure.

1897 J.J. Thomson, a British scientist, discovers the electron, leading to his "plum-pudding" model. He pictures electrons embedded in a sphere of positive electric charge.

1911 New Zealander Ernest Rutherford states that an atom has a dense, positively charged nucleus. Electrons move randomly in the space around the nucleus.

1913 In Niels Bohr's model, the electrons move in spherical orbits at fixed distances from the nucleus.

1926 Erwin Schrödinger develops mathematical equations to describe the motion of electrons in atoms. His work leads to the electron cloud model.

1800 1805 ………………… 1895 1900 1905 1910 1915 1920

1904 Hantaro Nagaoka, a Japanese physicist, suggests that an atom has a central nucleus. Electrons move in orbits like the rings around Saturn.

1924 Frenchman Louis de Broglie proposes that moving particles like electrons have some properties of waves. Within a few years evidence is collected to support his idea.

1932 James Chadwick, a British physicist, confirms the existence of neutrons, which have no charge. Atomic nuclei contain neutrons and positively charged protons.

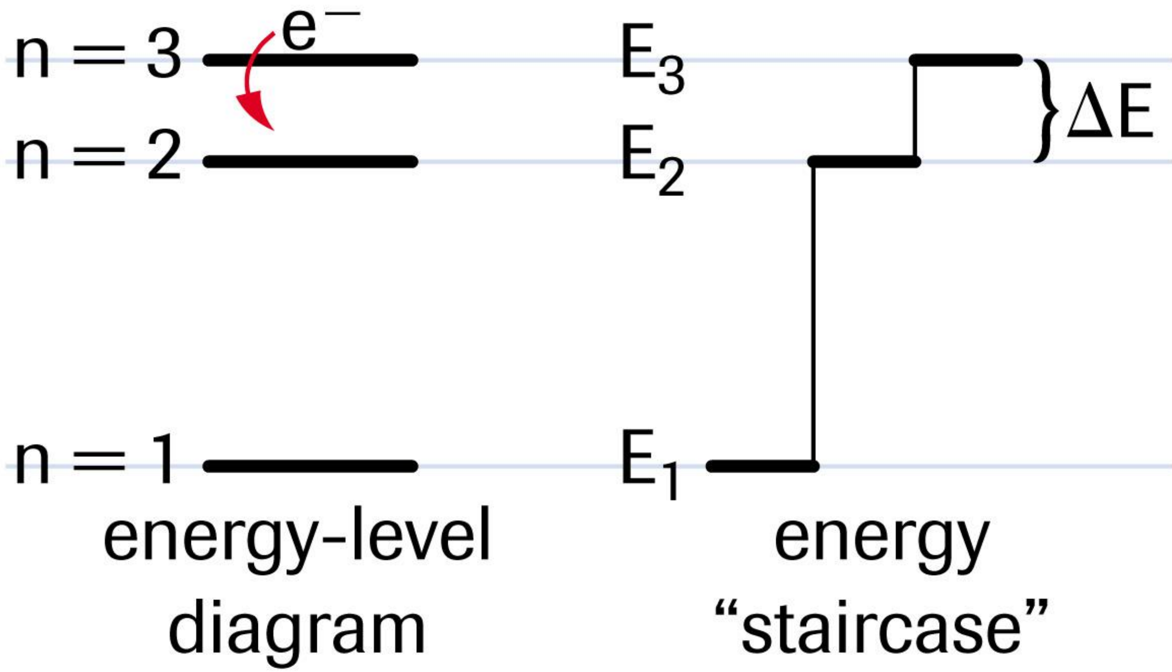
Dorin, Demmin, Gabel, Chemistry The Study of Matter, 3rd Edition, 1990, page 125

Homework

1. Describe the current model of the atom.
2. Describe the contributions of deBroglie, Schrodinger and Heisenberg to the development of the modern view of the atom.

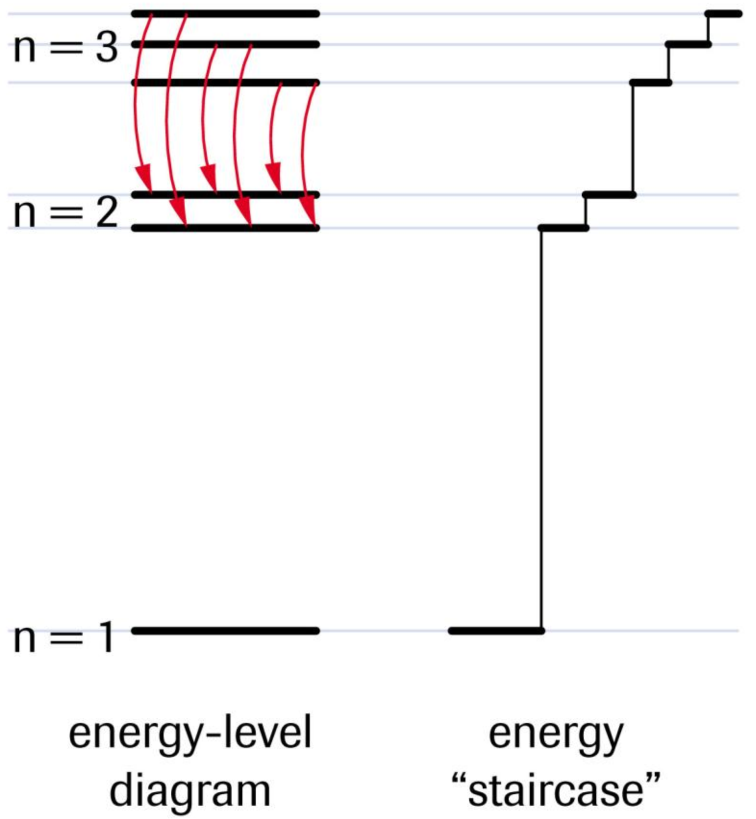
Explaining emission spectrum:

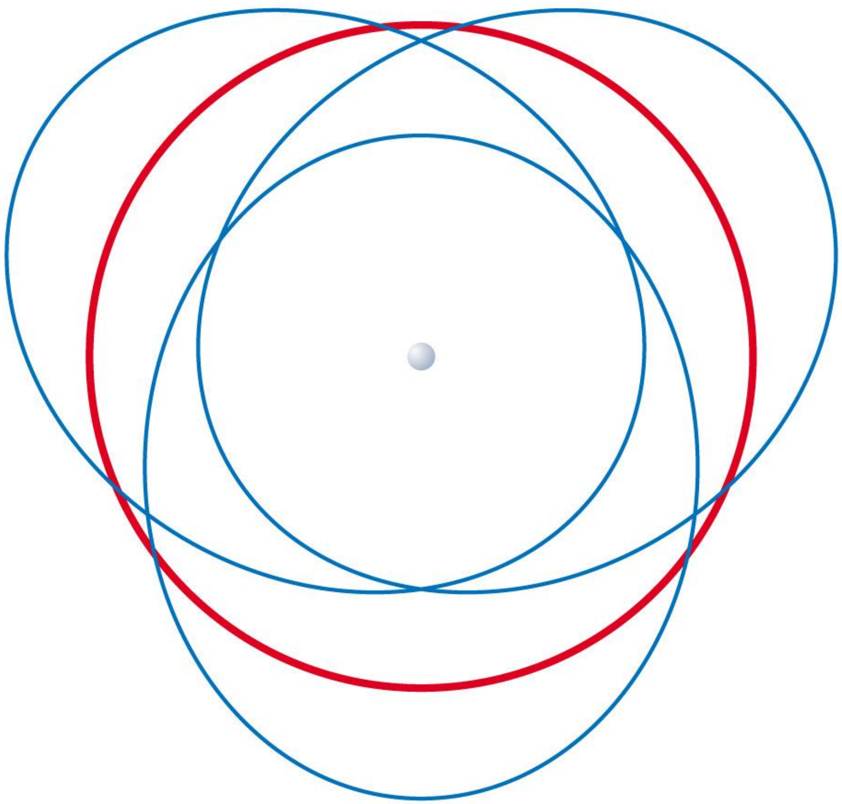
ΔE is released as a photon of light.



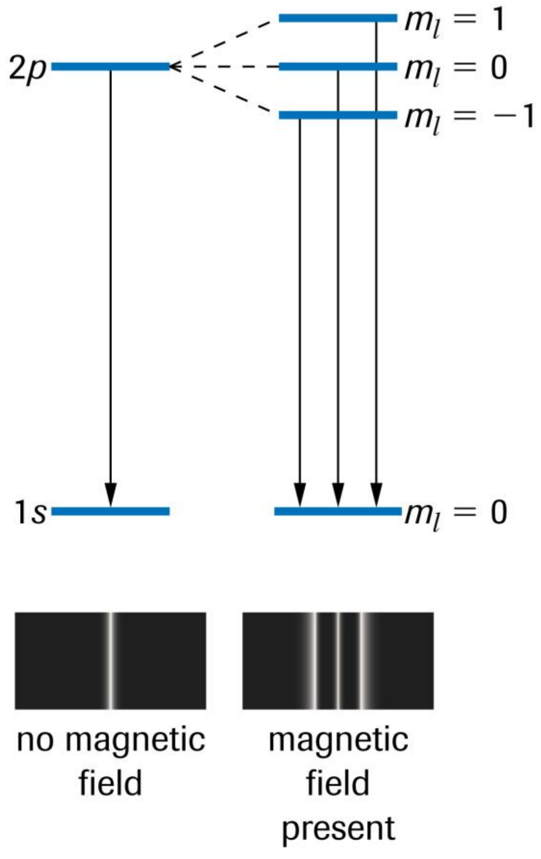
Sommerfeld (1915)

-main lines of H spectrum are actually composed of more lines
-introduced the idea of sublevels which is linked to the orbital shape



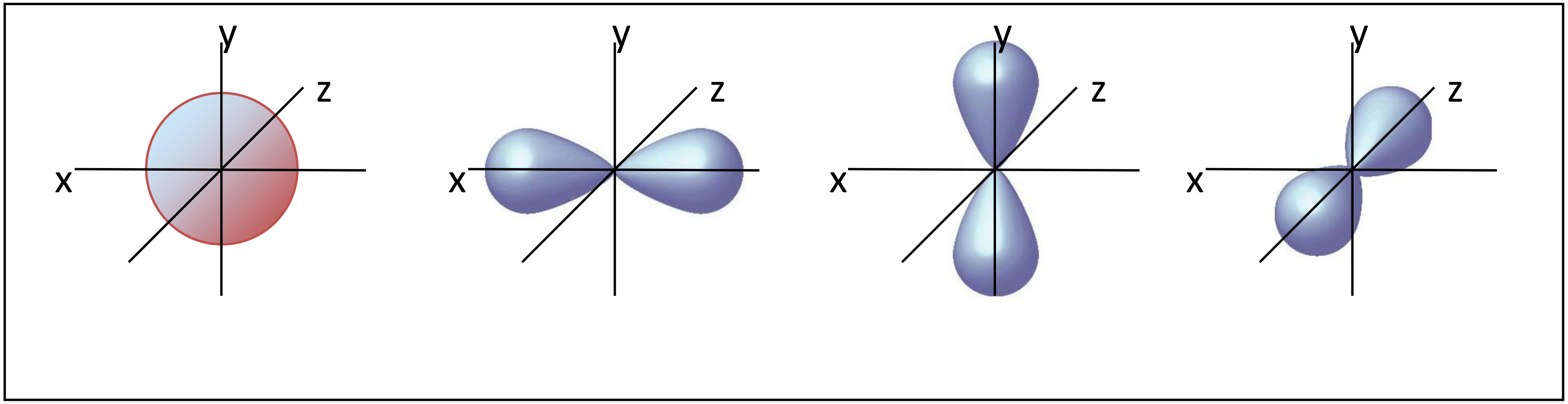


Zeeman effect - Single lines split into multiple lines in the presence of a magnetic field. Orbits can exist at various angles and their energies are different when the atom is near a strong magnet.

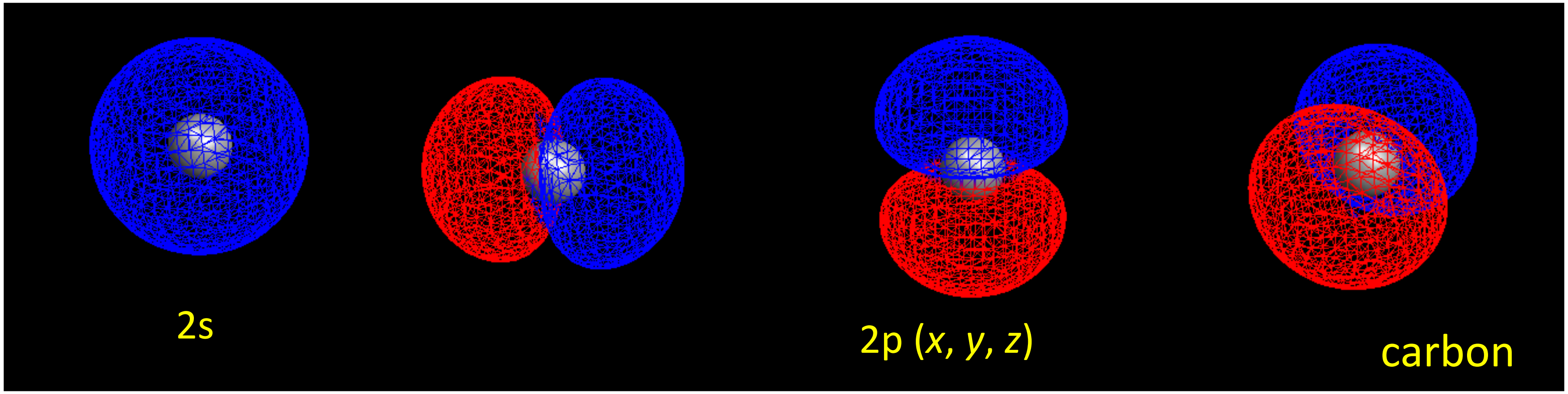


Stern-Gerlach experiment – a beam of neutral atoms is separated into two groups by passing them through a magnetic field

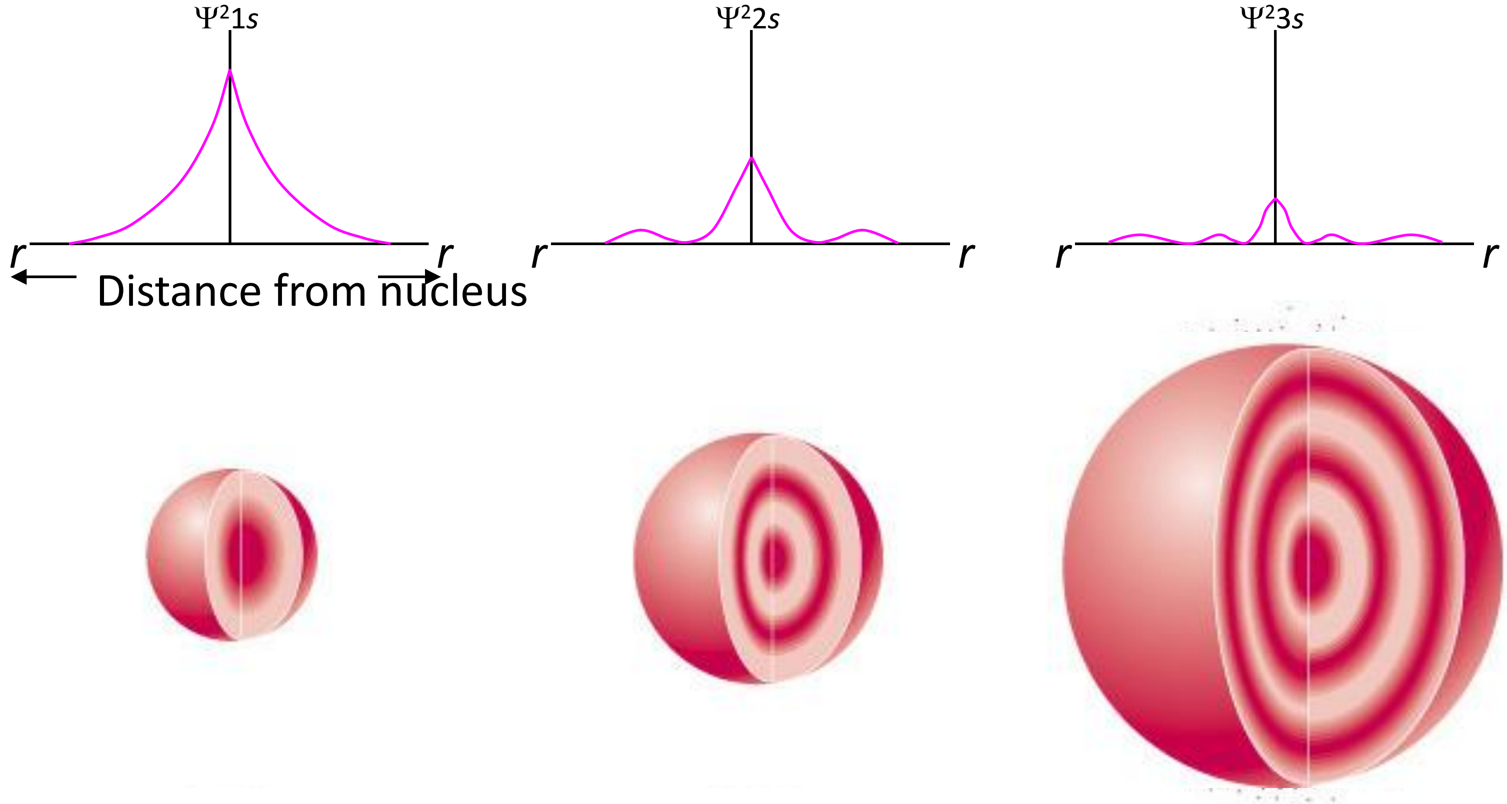
Orbitals



s p_x p_z p_y

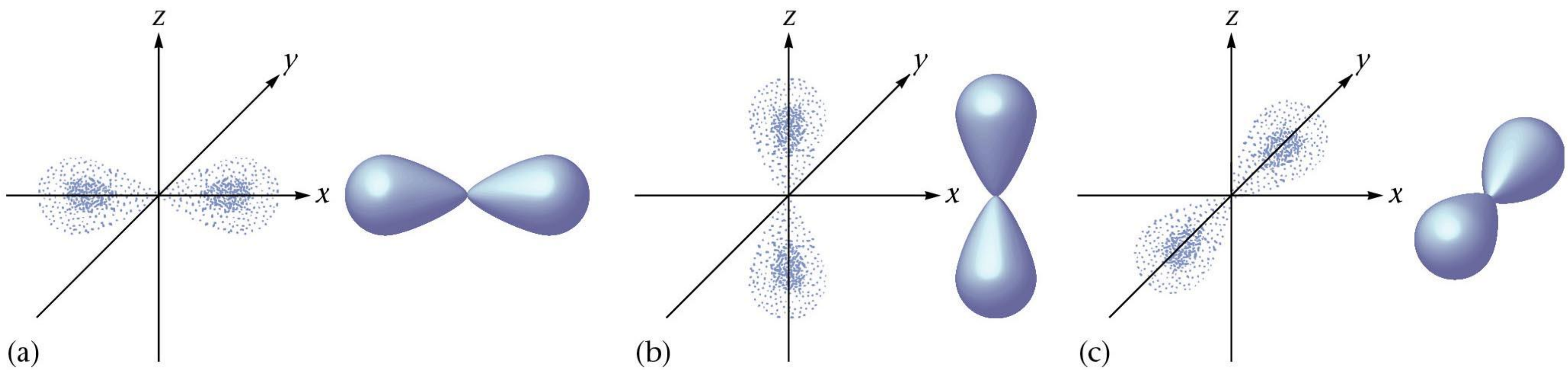


Mark Wirtz, Edward Ehrat, David L. Cedeno*



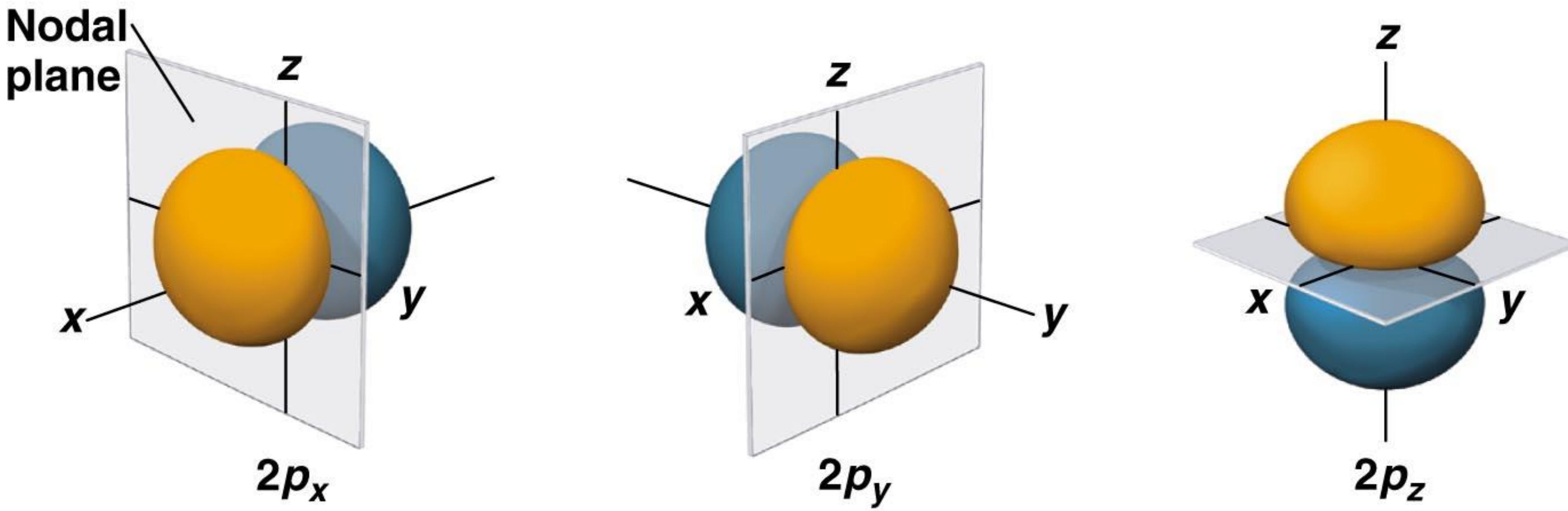
(a) 1*s* **(b)** 2*s* **(c)** 3*s*

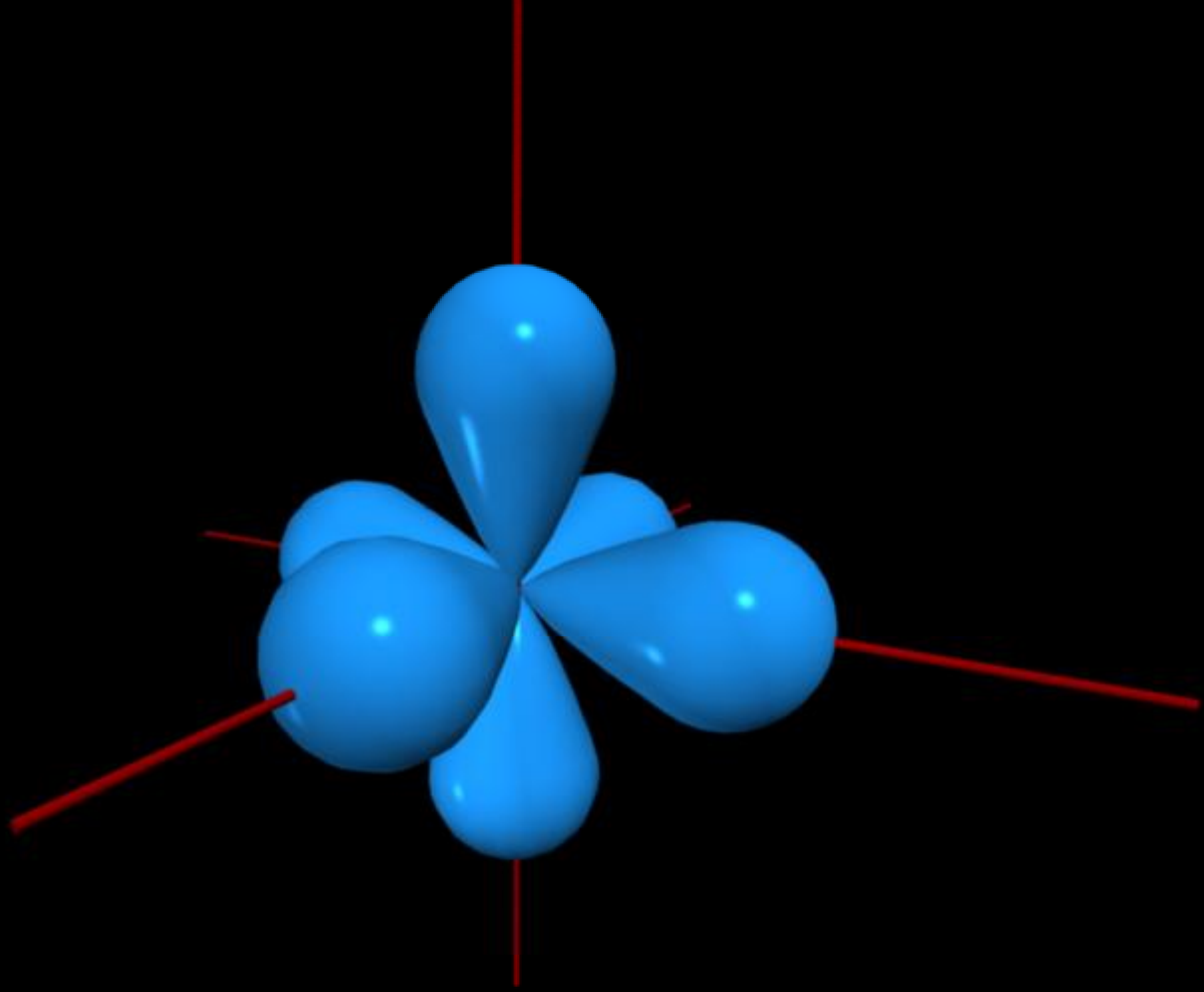
p-Orbitals



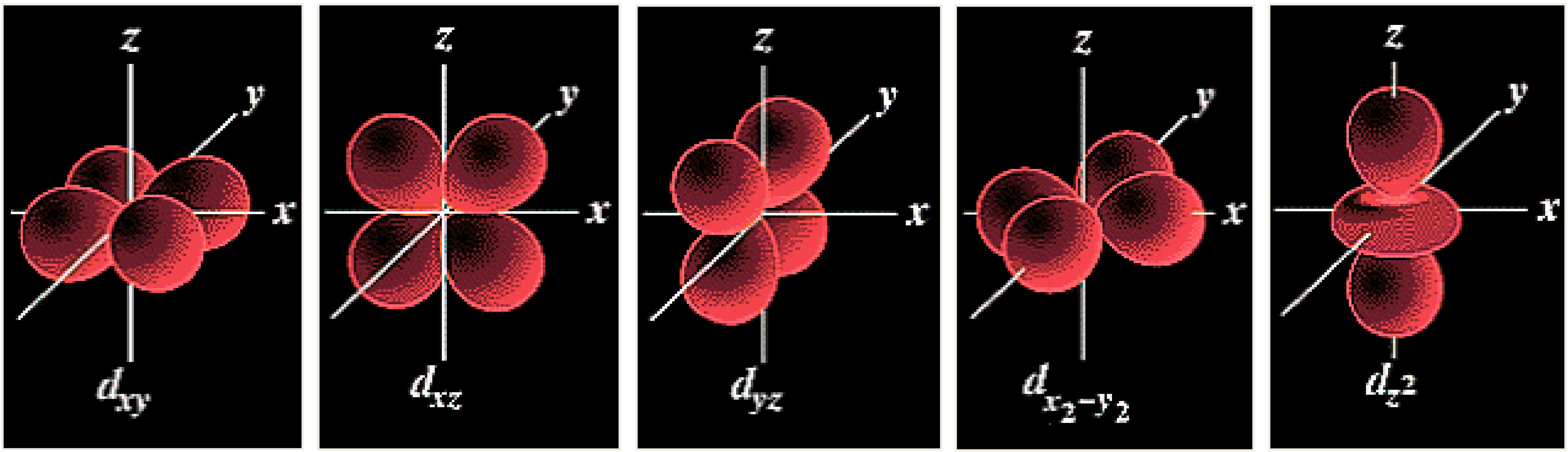
p_x p_z p_y

Zumdahl, Zumdahl, DeCoste, World of Chemistry2002, page 335


Nodal
plane
z
x
y
2p_x
z
x
y
2p_y
z
x
y
2p_z



dOrbitals

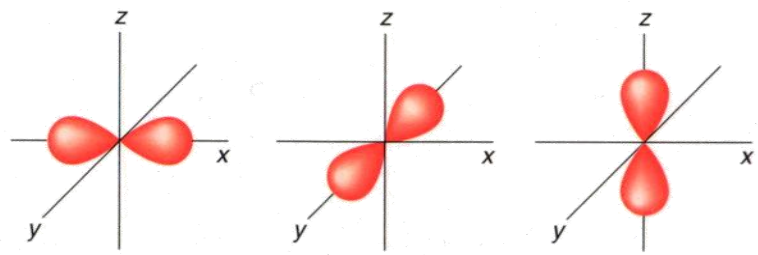


Courtesy Christy Johannesson www.nisd.net/communicationsarts/pages/chem

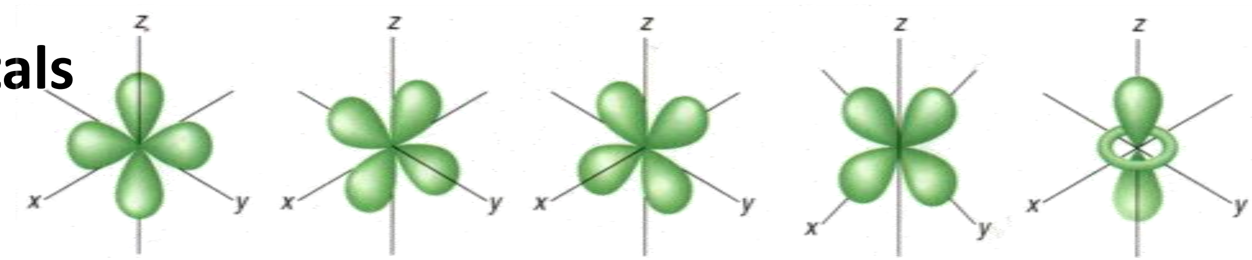
Shapes of s, p, and d-Orbitals

s **orbital**

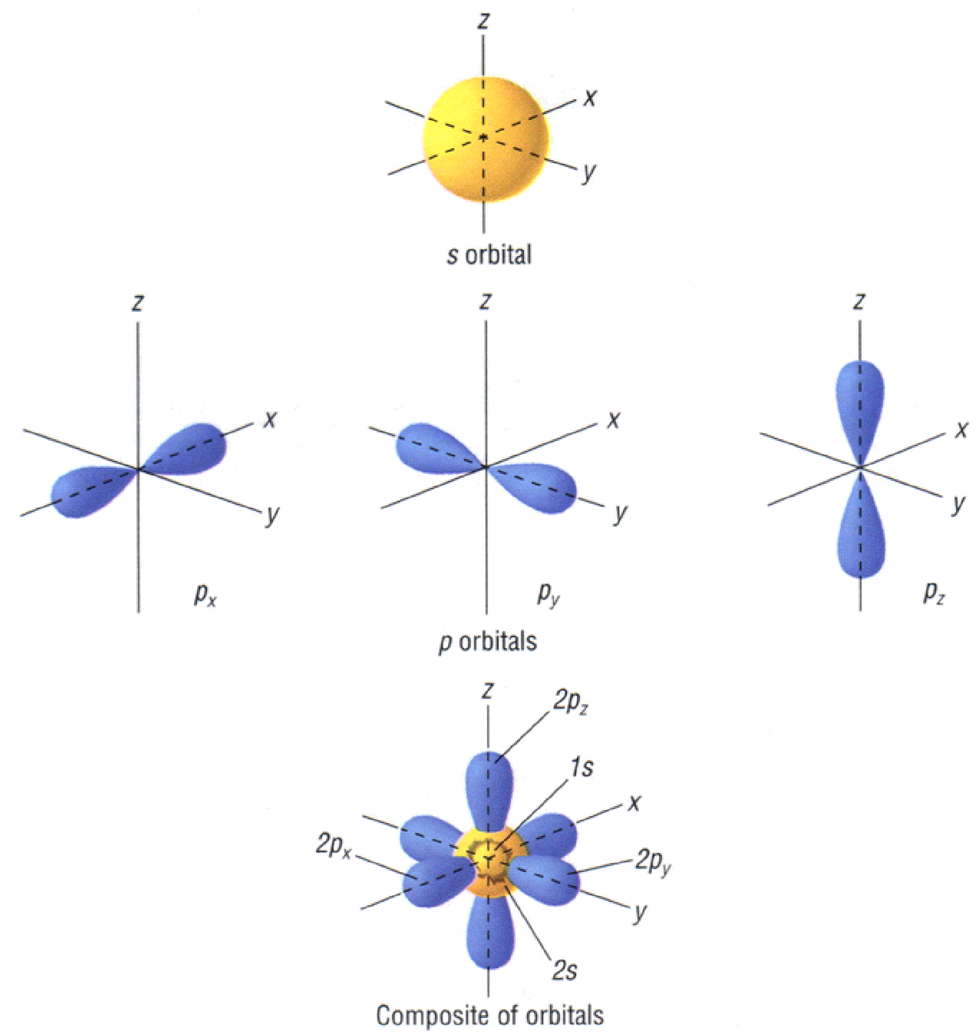
p **orbitals**

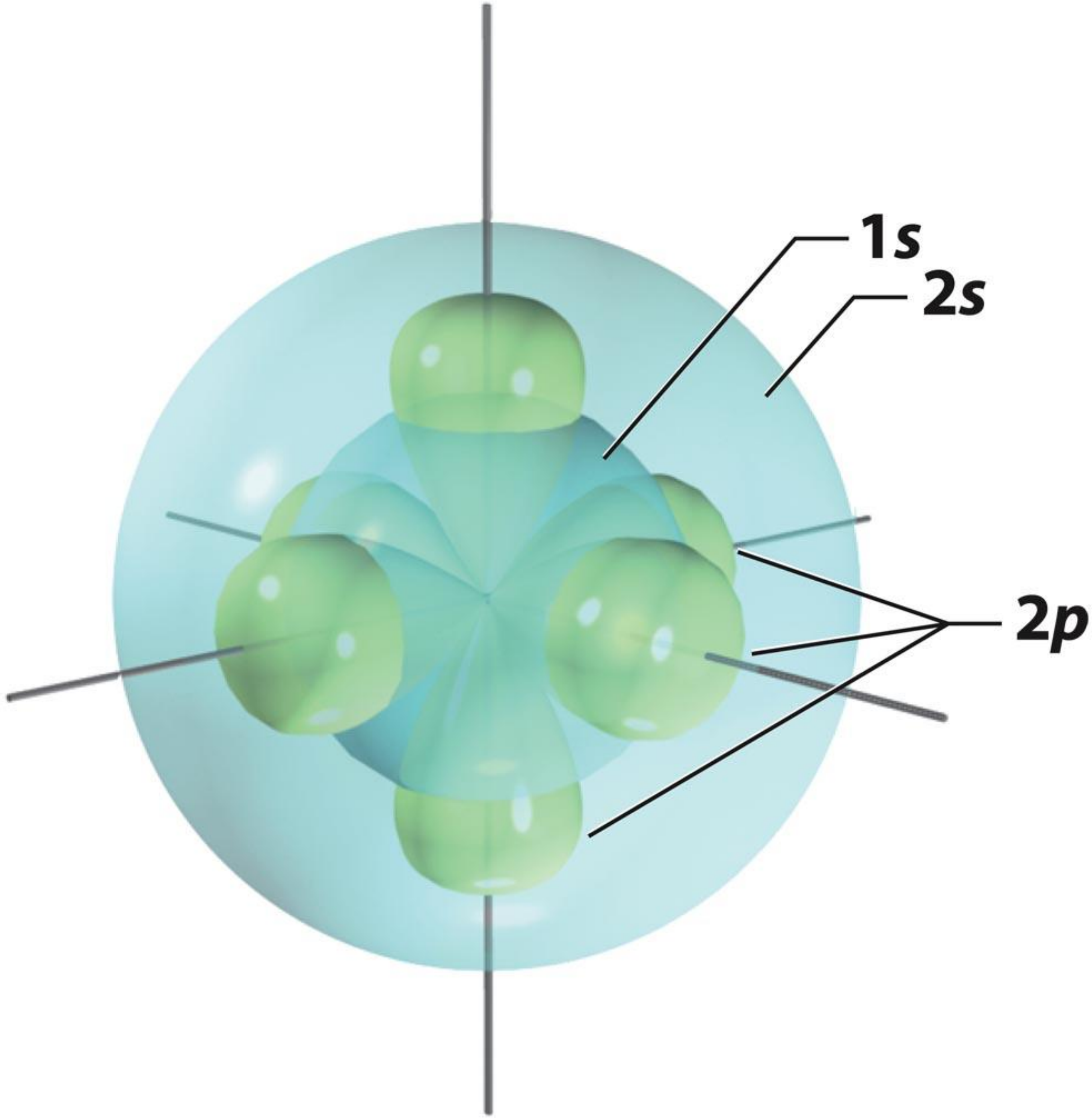


d **orbitals**

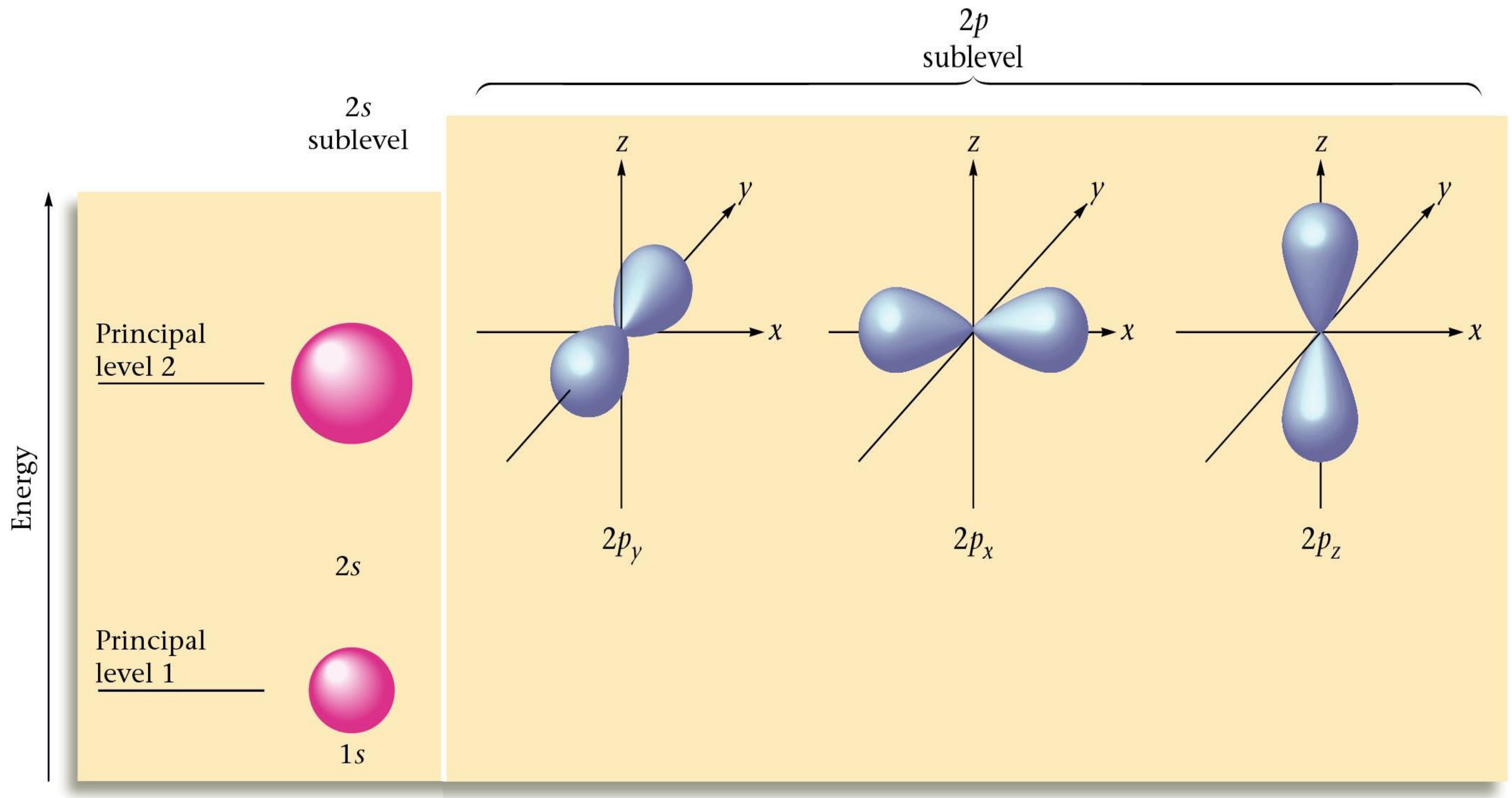


Atomic Orbitals




1s
2s
2p

Principal Energy Levels 1 and 2



Homework

1. Describe the shape of s and porbitals.
2. Define orbital.
3. How does the size of the orbital change as n increases?

Quantum Numbers
(used to represent different energy states)

Wolfgang Pau

- Pauli Exclusion Principle
  - No two electrons in an atom can have the same 4 quantum numbers.

1. Principal # (n) → energy level
2. Secondary # (l) → sublevel *(s,p,d,f)*, shape
3. Magnetic # (m_l) → tilt or orientation of orbital
4. Spin # (m_s) → electron spin

Courtesy Christy Johannesson www.nisd.net/communicationsarts/pages/chem

l = 0 (s)
l = 1 (p)
l = 2 (d)
l = 3 (f)

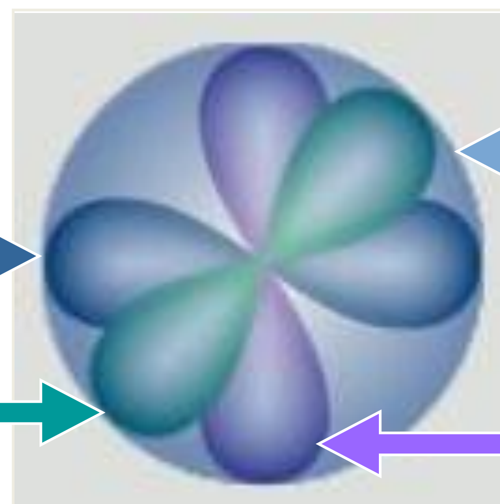
Quantum Numbers

n = 1
l = 0

1s

n = 2
l = 0

n = 2
l = 1



2s

$2p_x$

$2p_y$

$2p_z$

n = 2
l = 1

n = 2
l = 1

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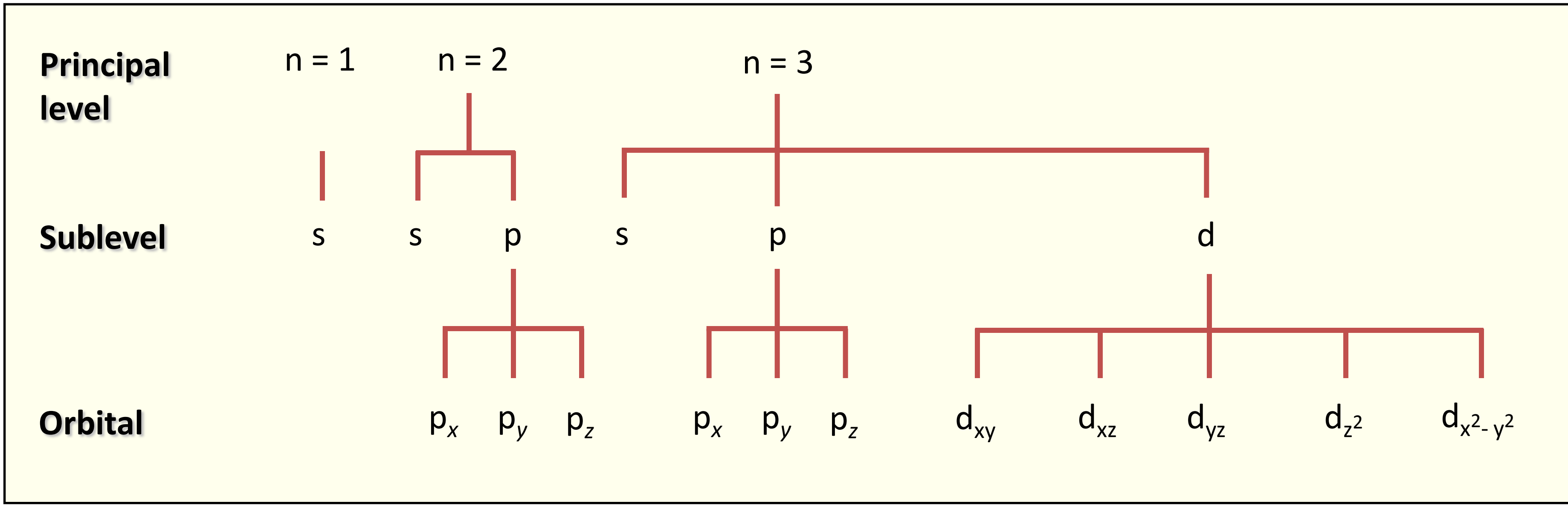
Quantum Numbers

n (principal quantum #) $n = 1,2,3\ldots$	l (secondary quantum #) $l = 0,1\ldots n-1$	m_l (magnetic quantum #) $m_l = -l$ to $+l$	m_s (spin quantum #) $m_s = ½, -½$

Quantum Numbers



- n = # of sublevels per level
- # of orbitals per sublevel: **1** (s), **3** (p), **5** (d), **7** (f)

(think magnetic quantum #)

Courtesy Christy Johannesson www.nisd.net/communicationsarts/pages/chem

Maximum Capacities of Subshells and Principal Shells

n	1	2		3			4				...*n*
l	0	0	1	0	1	2	0	1	2	3	
Subshell designation	*s*	*s*	*p*	*s*	*p*	*d*	*s*	*p*	*d*	*f*	
Orbitals in subshell	1	1	3	1	3	5	1	3	5	7	
Subshell capacity	2	2	6	2	6	10	2	6	10	14	
Principal shell capacity	2	8		18			32				...$2n^2$

Hill, Petrucci, General Chemistry *An Integrated Approach* 1999, page 320

Practice

1. Decide whether each set of quantum #s is allowed.

a) $n = 1 \;\; l = 0 \;\; m_l = 0 \;\; m_s = ½$

b) $n = 2 \;\; l = 1 \;\; m_l = -1 \;\; m_s = -½$

c) $n = 3 \;\; l = 3 \;\; m_l = 2 \;\; m_s = ½$

d) $n = 0 \;\; l = 0 \;\; m_l = 0 \;\; m_s = ½$

2. Give the complete set of quantum #s for a lone e^- in the third orbital of a 2p sublevel.

Homework

- p. 159 #1-9,11,12