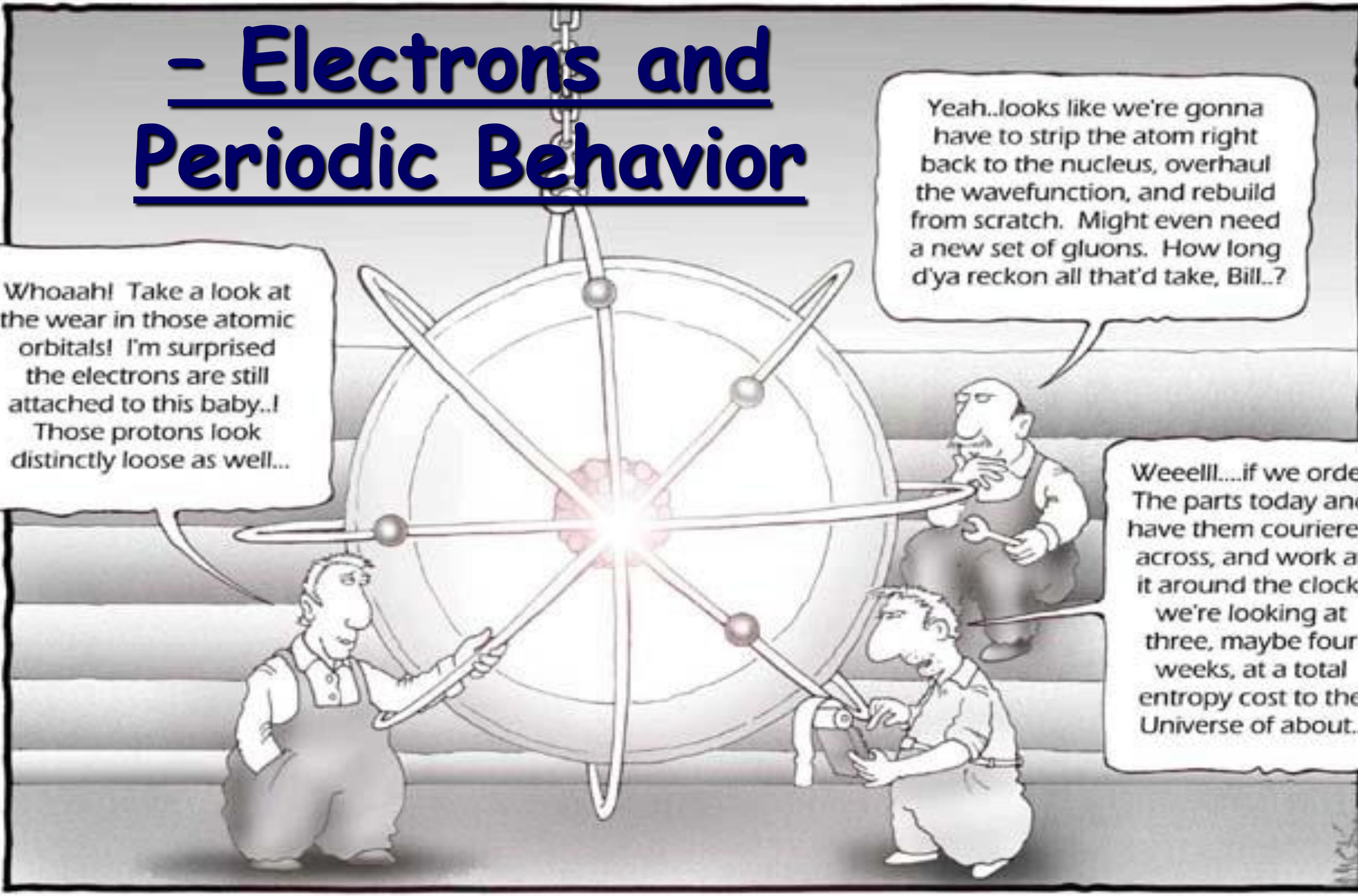
– Electrons and Periodic Behavior



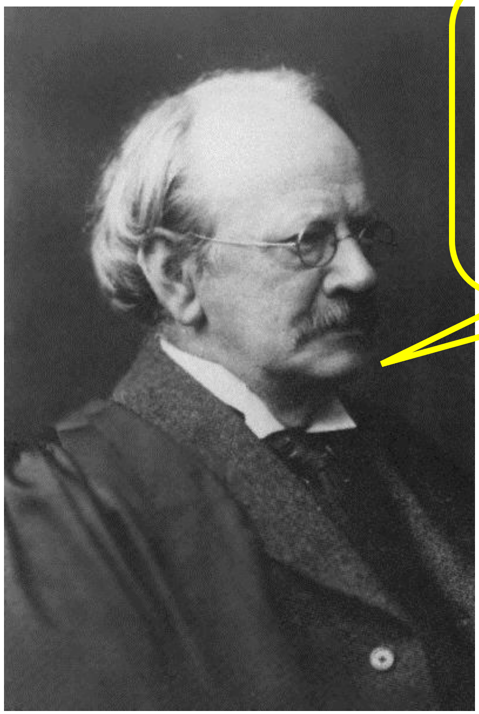
Quantum mechanics.

Cartoon courtesy of NearingZero.net

Wave-Particle Duality

JJ Thomson won the Nobel prize for describing the electron as a particle.

His son, George Thomson won the Nobel prize for describing the wave-like nature of the electron.



The electron is a particle!

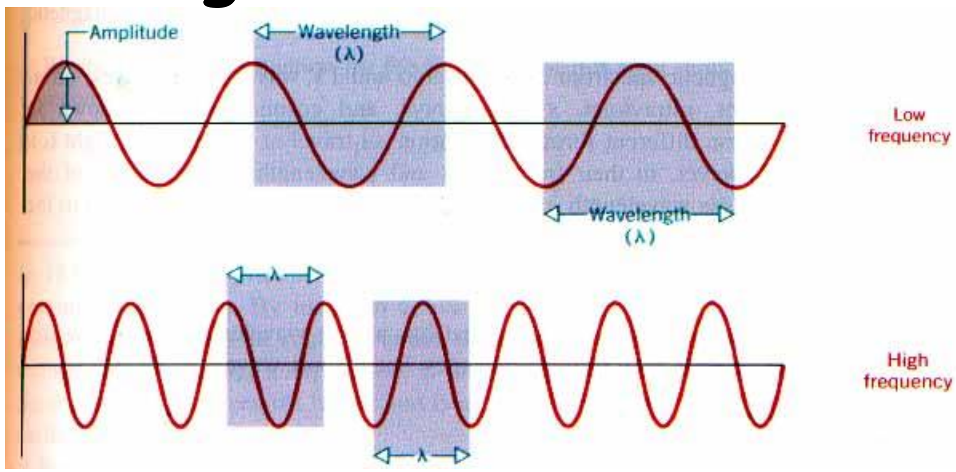
The electron is an energy wave!

The Wave-like Electron

Louis deBroglie

The electron propagates through space as an energy wave. To understand the atom, one must understand the behavior of electromagnetic waves.

Electromagnetic radiation propagates through space as a wave moving at the speed of light.



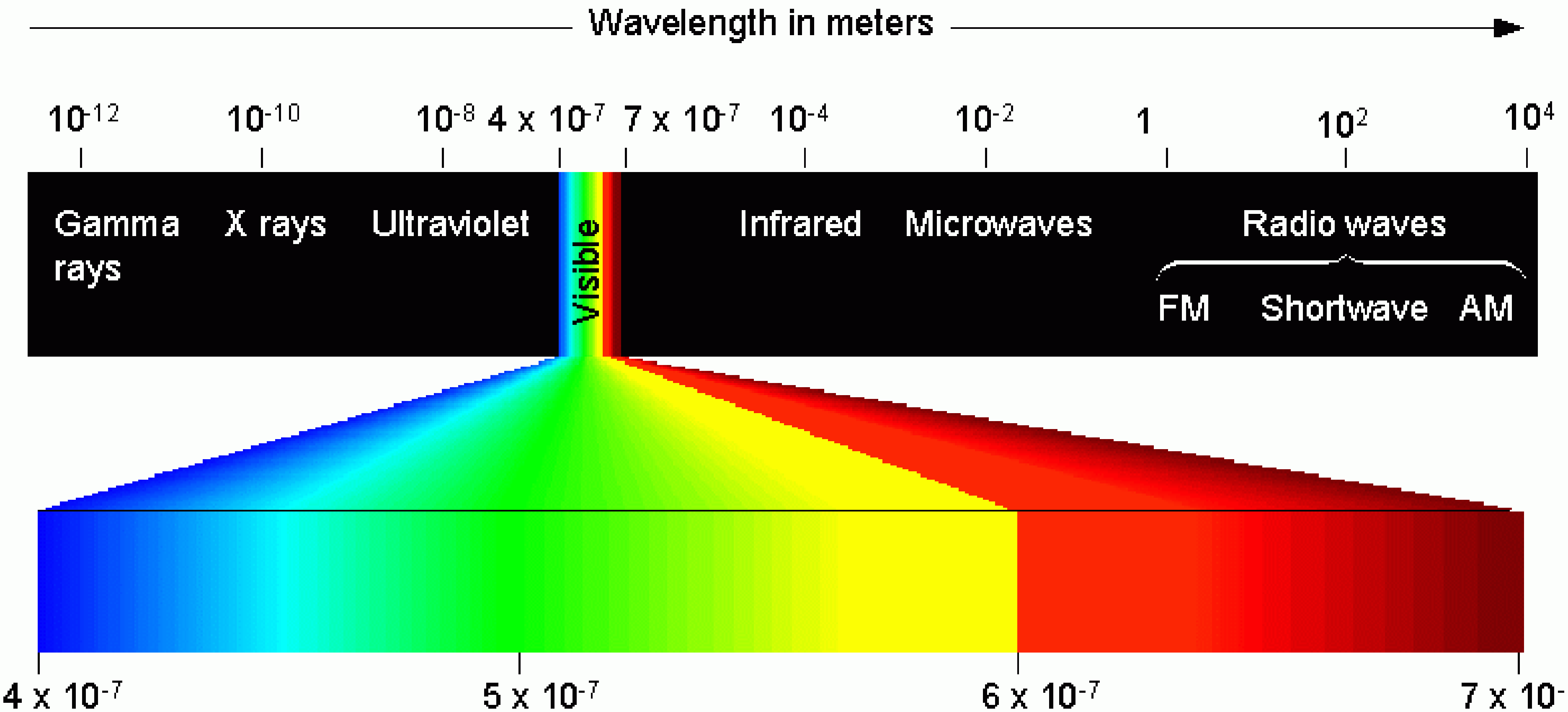
$$c = \nu\lambda$$

C = speed of light, a constant (3.00×10^8 m/s)

ν = frequency, in units of hertz (hz, sec^{-1})

λ = wavelength, in meters

Types of electromagnetic radiation:



Max Planck

- figured out that when a solid substance is heated, it gives off energy in "chunks"
- later called quantums of energy
  - *quantum means fixed amount*
- noticed that different substances released different "chunks" of energy

The energy (E) of electromagnetic radiation is directly proportional to the frequency (ν) of the radiation.

$$E = h\nu$$

E = Energy, in units of Joules ($kg \cdot m^2/s^2$)

h = Planck's constant (6.626 x 10-34 $J \cdot s$)

ν = frequency, in units of hertz (hz, sec^{-1})

Long Wavelength

=

Low Frequency

=

Low ENERGY

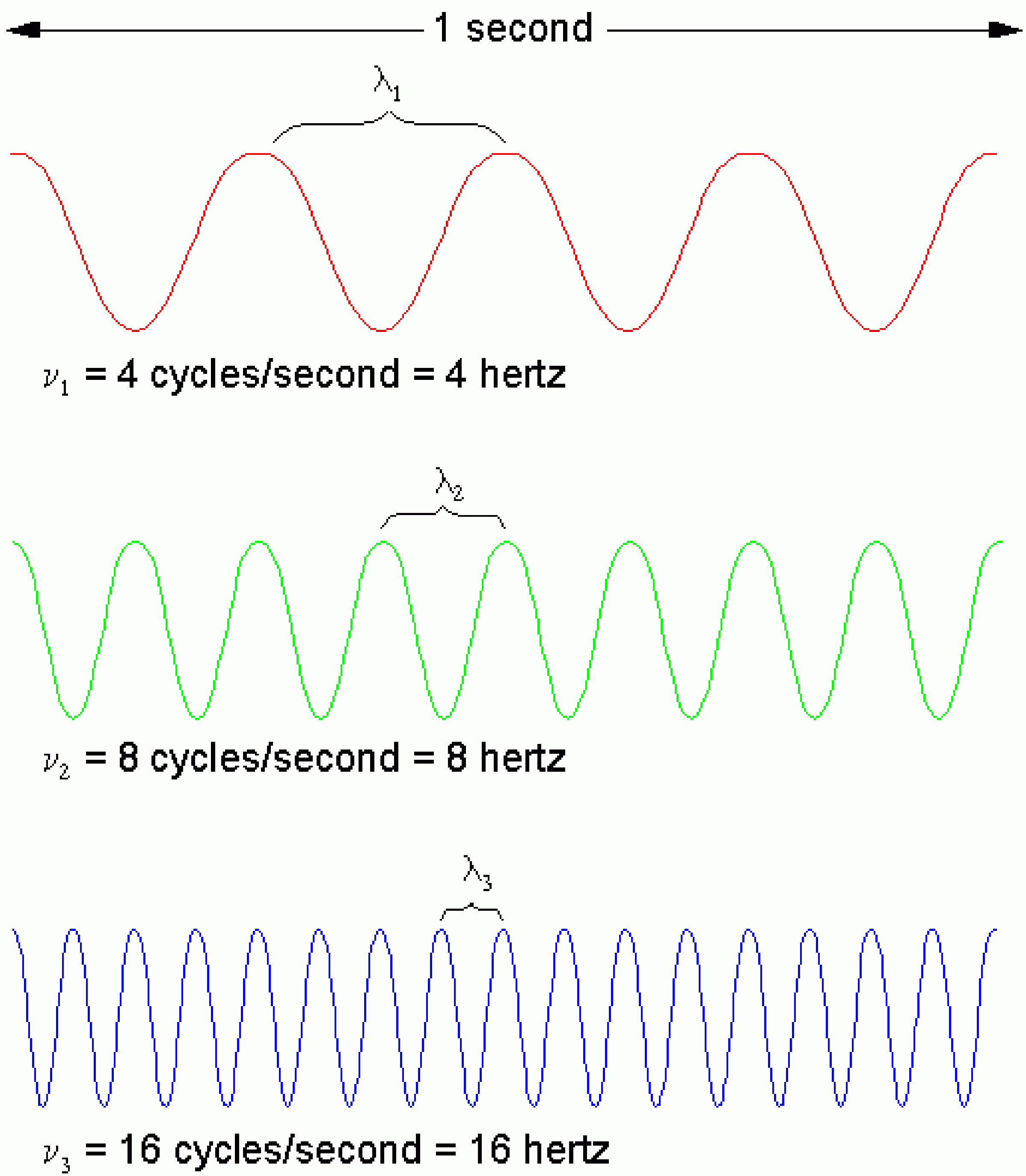
Short Wavelength

=

High Frequency

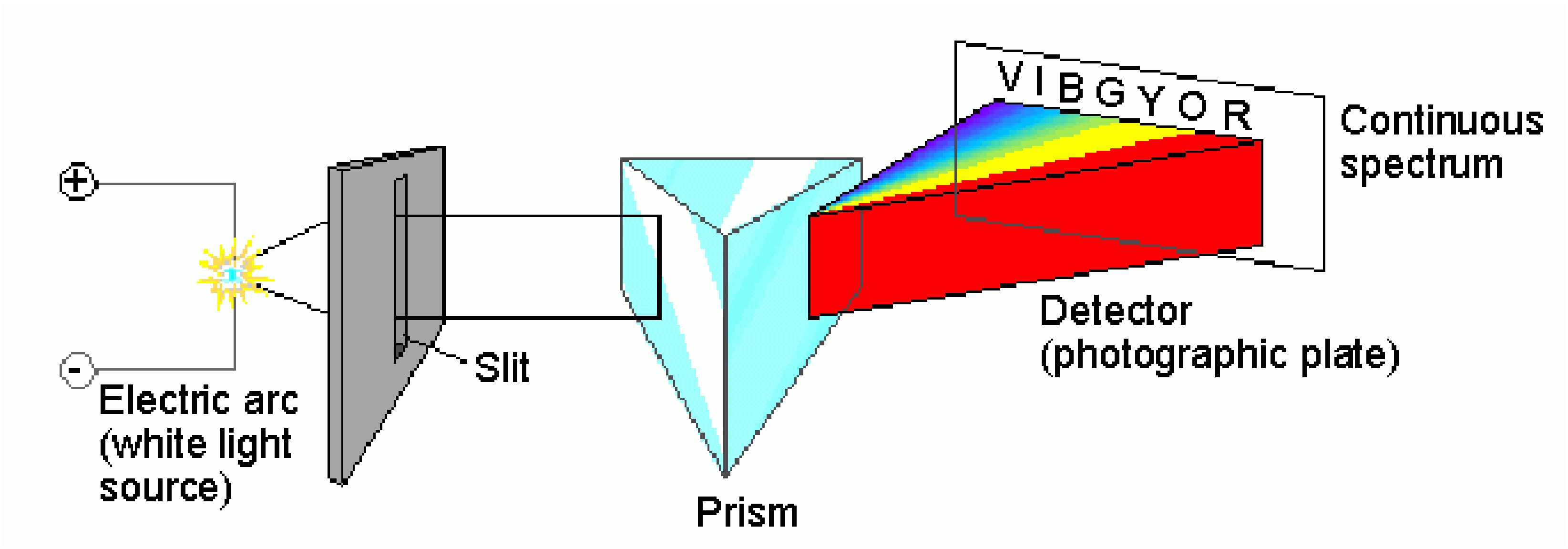
=

High ENERGY



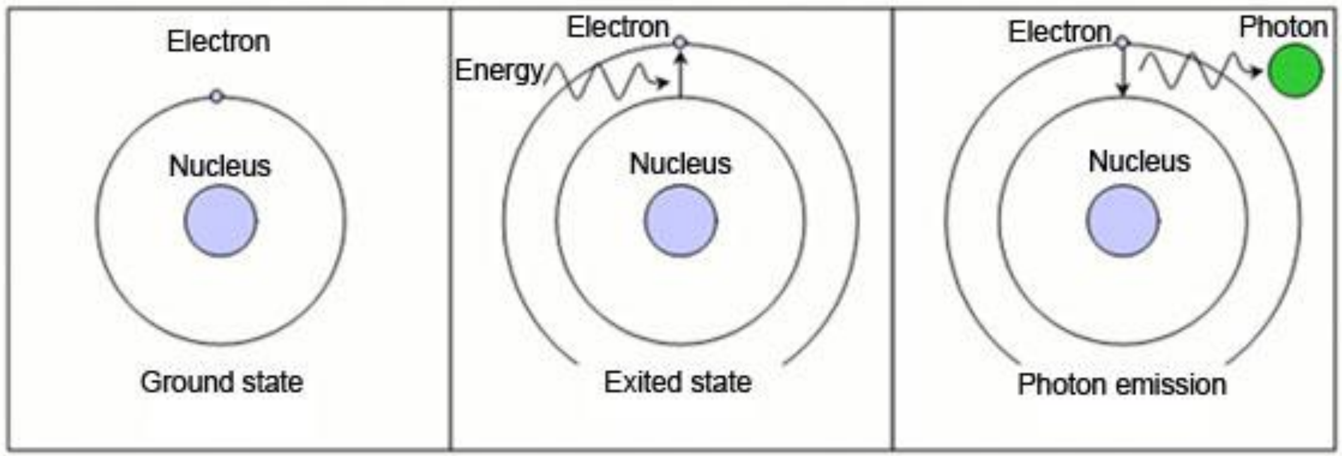
Spectroscopic analysis of the visible spectrum…

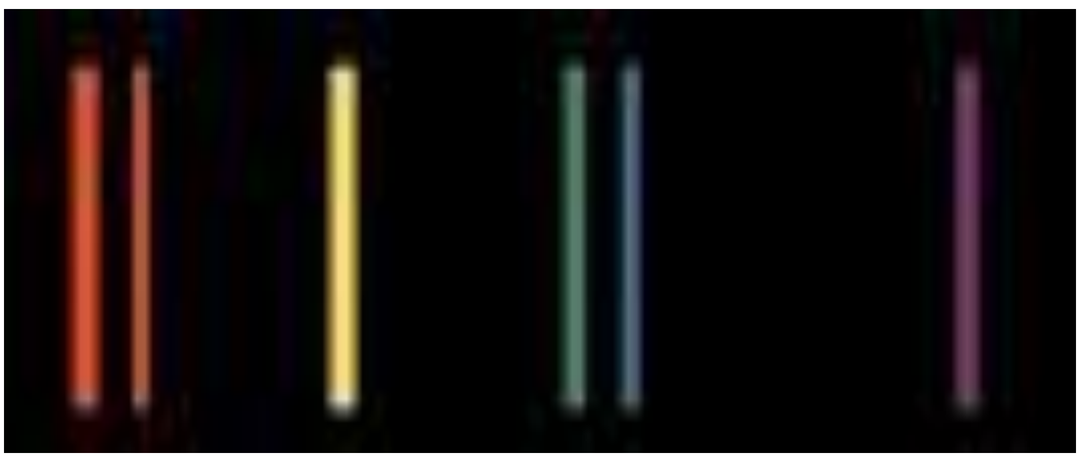
produces all of the colors in a continuous spectrum



How does matter produce light?

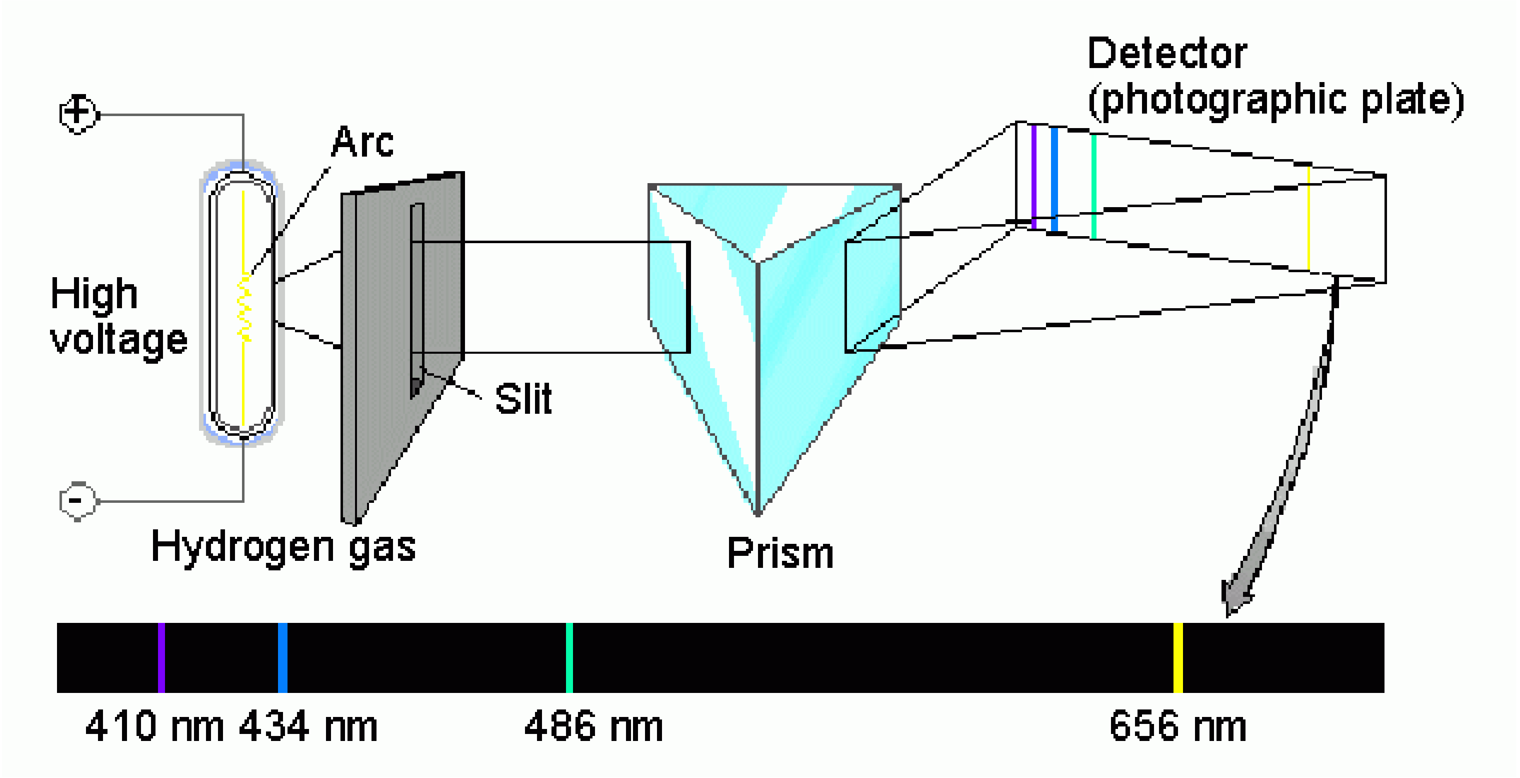
- When an electron in the **ground state** (lowest energy level that is natural) is promoted to an **excited (**higher) **state** it is temporary!!! The electron will fall back down to the ground state releasing light! (when viewed through a spectroscope- line emission spectra- gives a fingerprint of the atom)


Electron
Nucleus
Ground state
Electron
Energy
Nucleus
Exited state
Electron
Photon
Nucleus
Photon emission



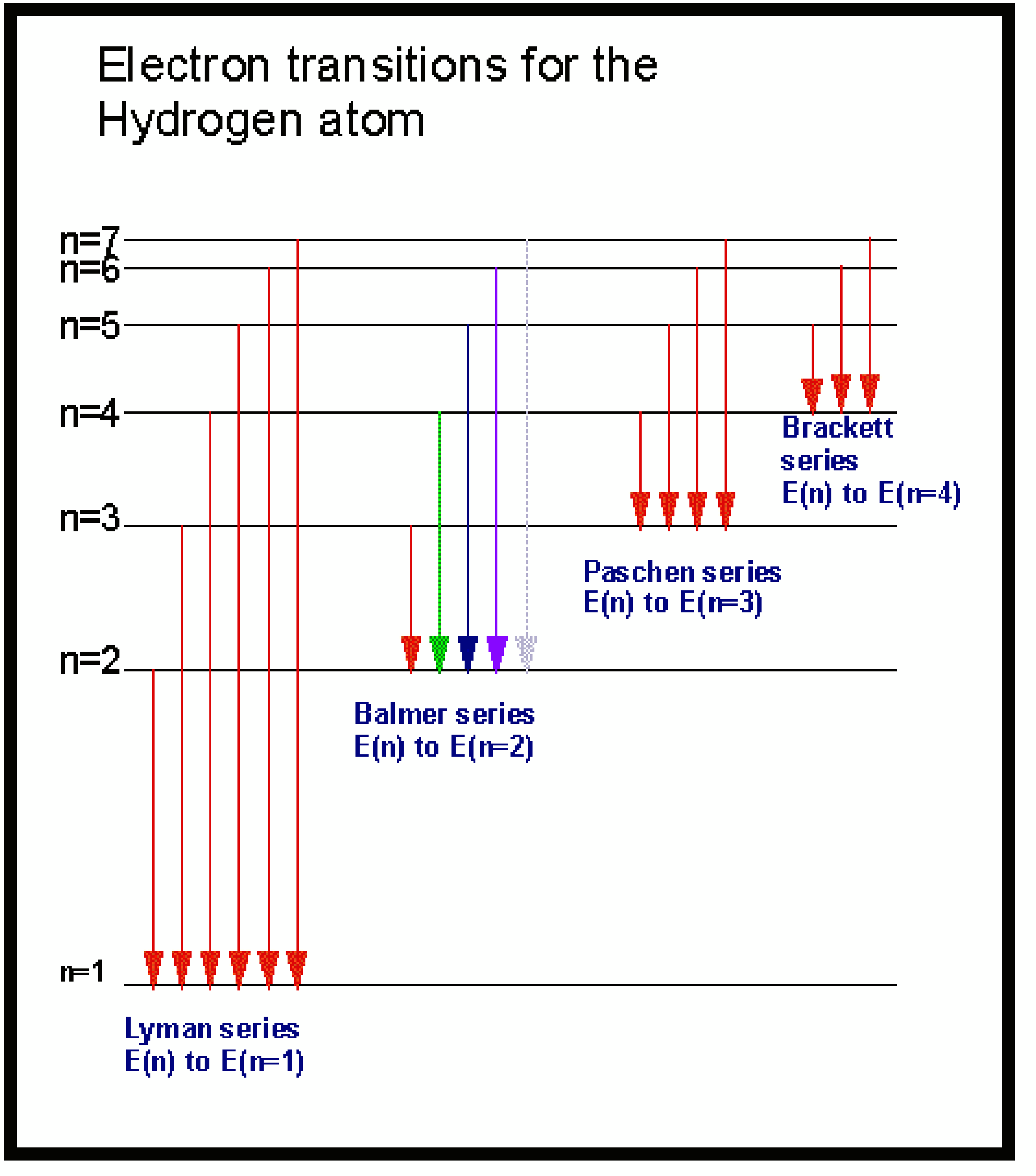
Spectroscopic analysis of the hydrogen spectrum…

produces a "bright line" spectrum

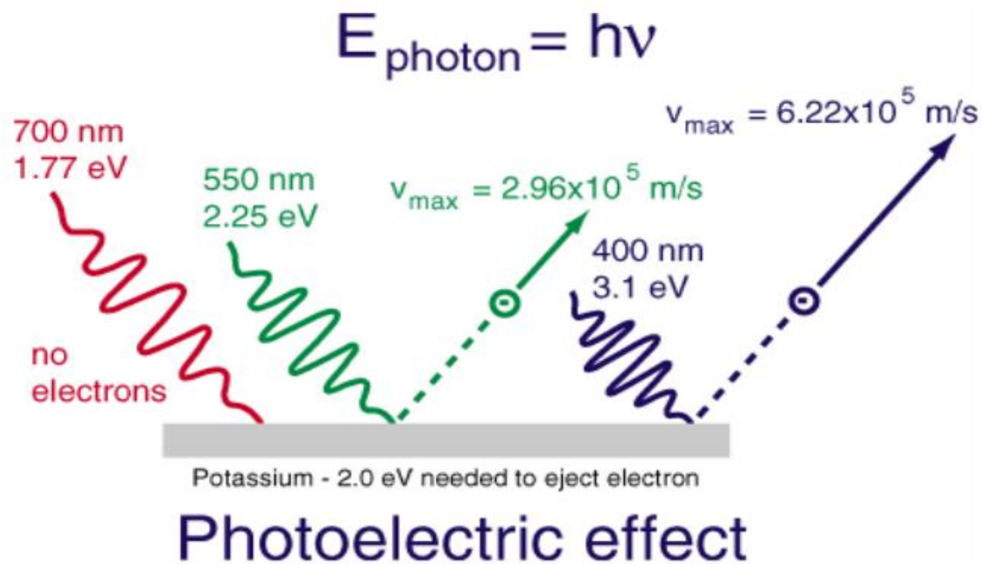


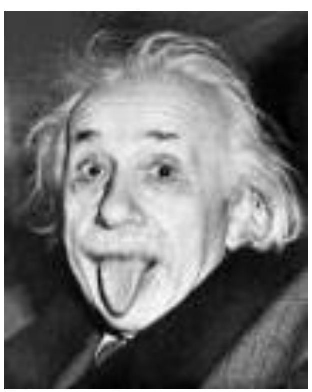
Electron transitions involve jumps of definite amounts of energy.

This produces bands of light with definite wavelengths.



How does light behave as a particle?





- Photoelectric effect-

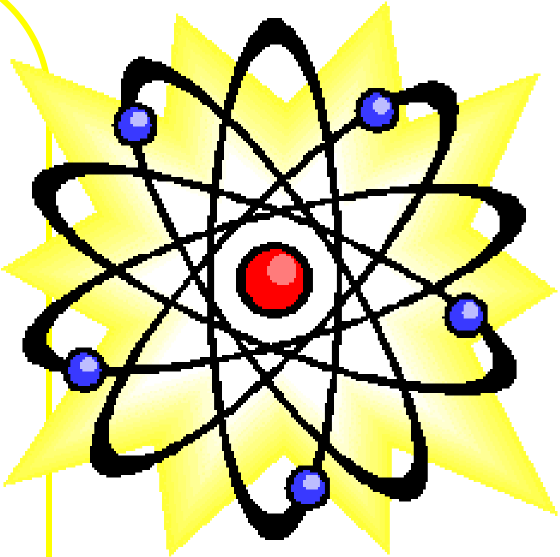
  when light of a particular frequency hits the surface of a metal an electron is ejected off the surface!

The Bohr Model of the Atom

Neils Bohr

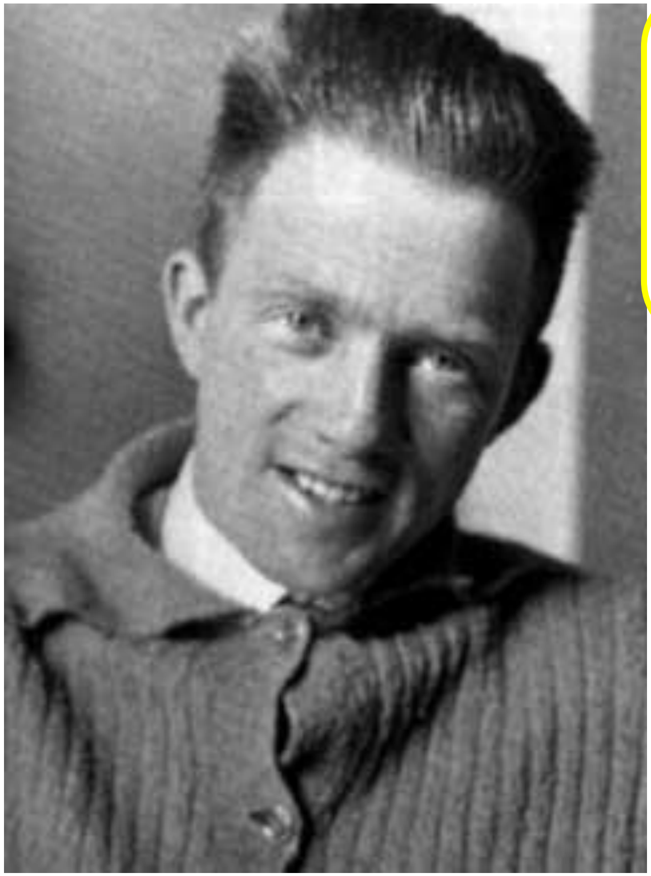
I pictured electrons orbiting the nucleus much like planets orbiting the sun.

But I was wrong!
They're more like bees around a hive



WRONG!!!

Heisenberg Uncertainty Principle



Werner Heisenberg

"One cannot simultaneously determine both the position and momentum of an electron."

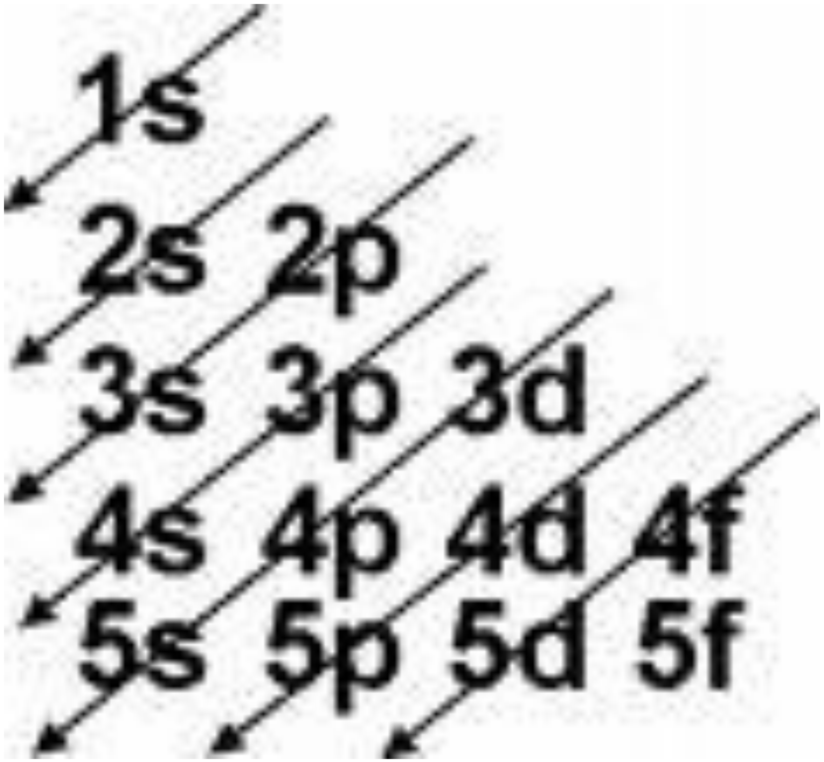
You can find out where the electron is, but not where it is going.

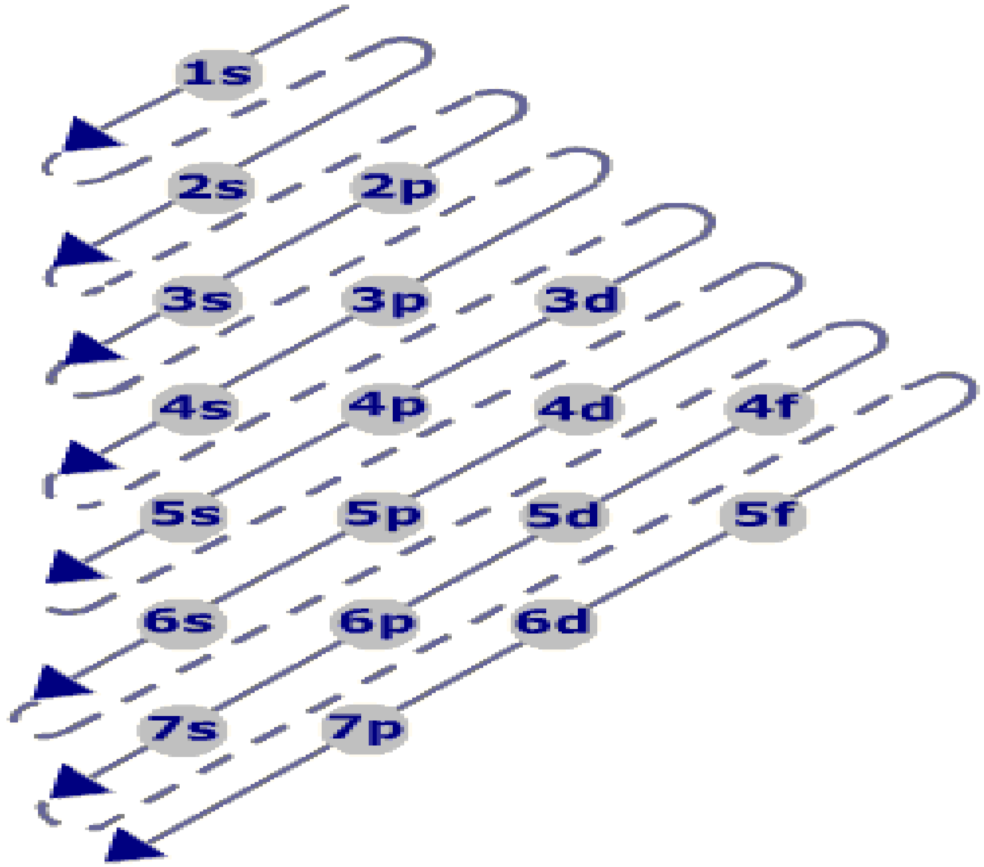
You can find out where the electron is going, but not where it is!

Electron Configuration

Simplest way to write which energy levels and sublevels are filled within the atom (How many e^- and where you can find them)

How do electrons fill in an atom?
The Diagonal Rule





Sublevel	# of electrons held
S	2
P	6
D	10
F	14

How to fill

- 1 Find total # of electrons
- 2 Write subshells in order of diagonal rule
- 3. Fill in subshells till all electrons are used
- 4. Last subshell may be partially filled.

Standard Notation of Fluorine

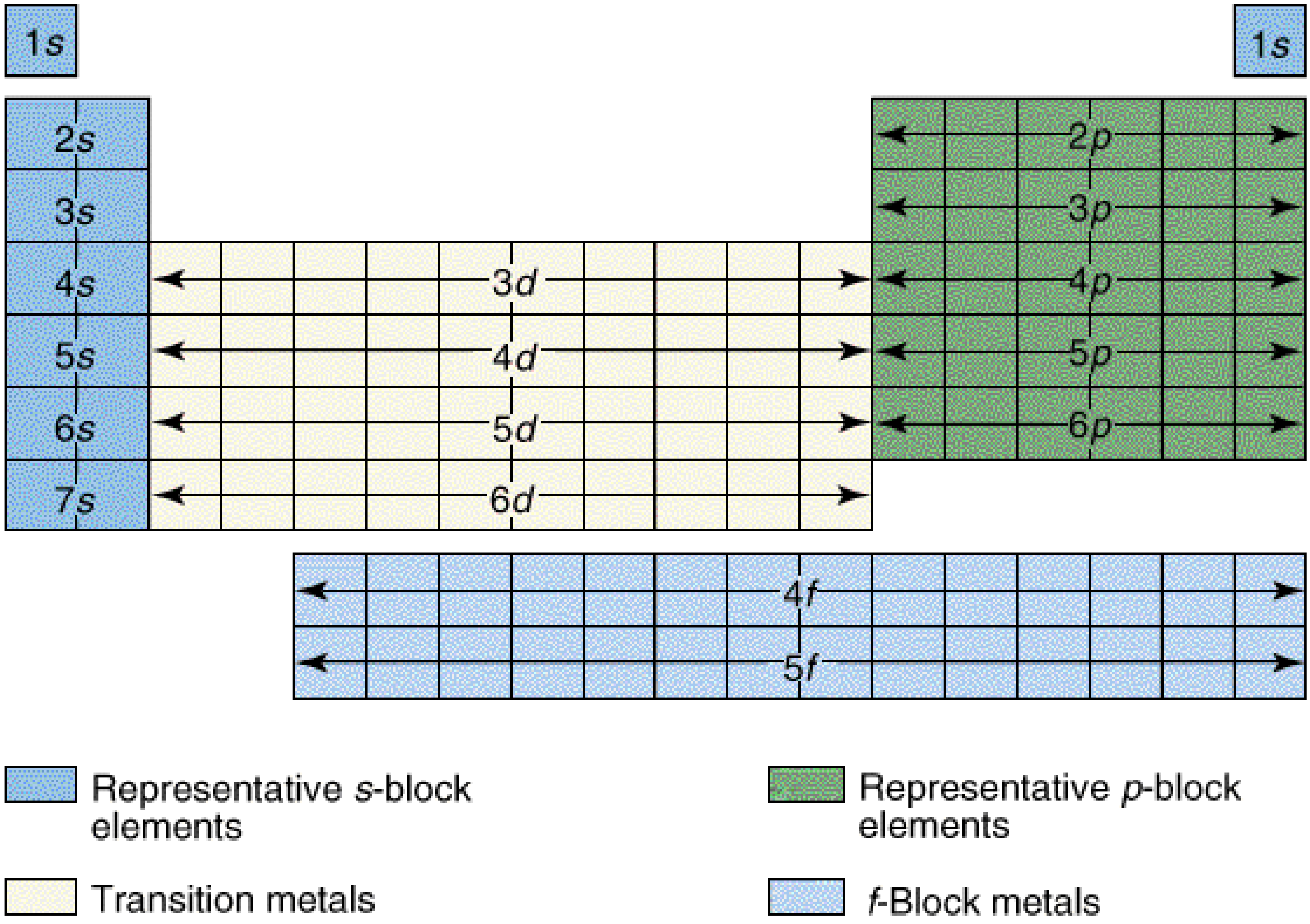
Number of electrons in the sub level 2,2,5

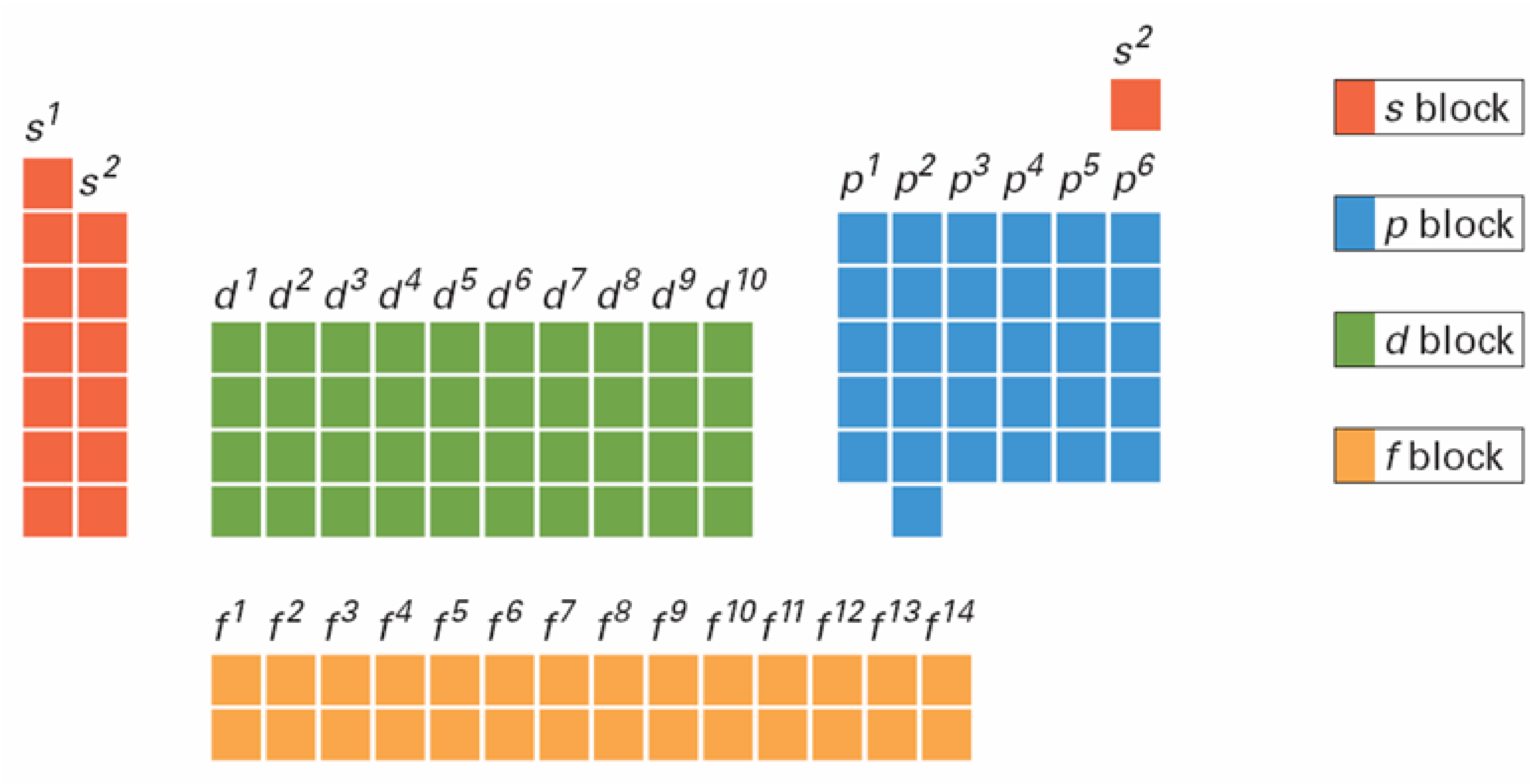
$$1s^2\ 2s^2\ 2p^5$$

Sublevels

Main Energy Level Numbers **1, 2, 2**

Blocks in the Periodic Table



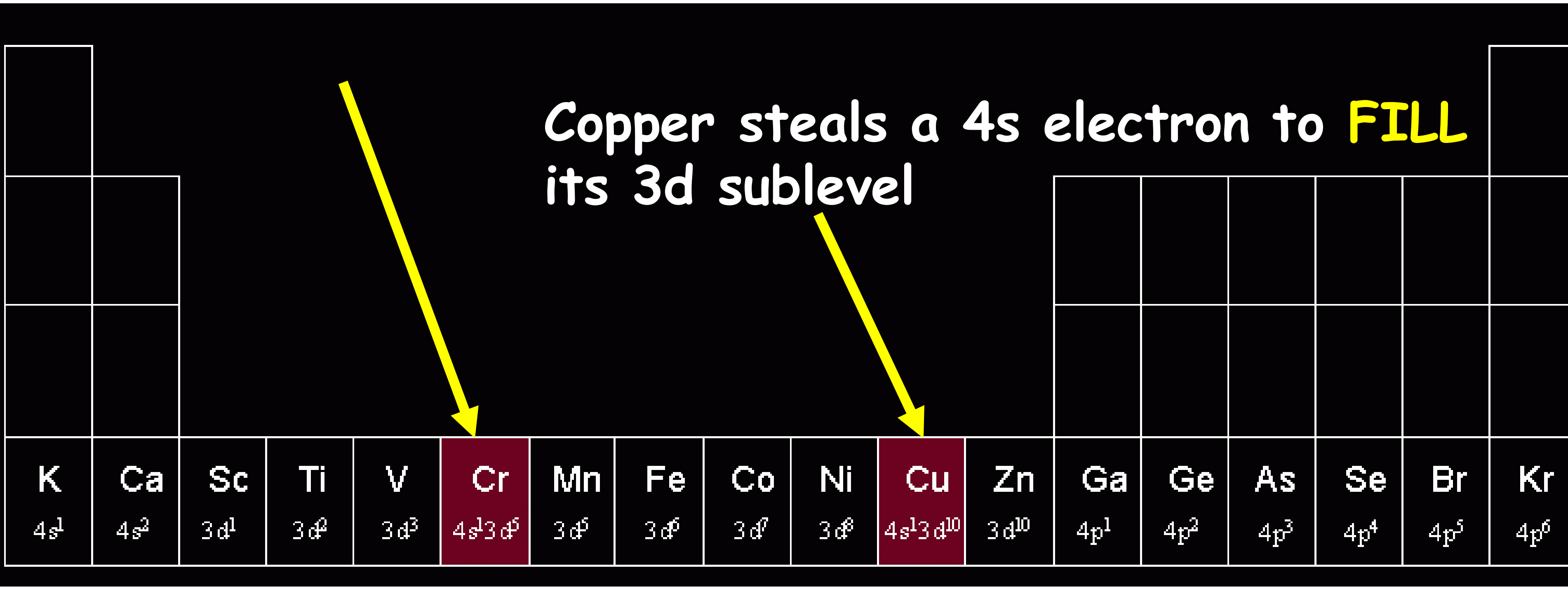


s1
s2
s2
d1 d2 d3 d4 d5 d6 d7 d8 d9 d10
p1 p2 p3 p4 p5 p6
f1 f2 f3 f4 f5 f6 f7 f8 f9 f10 f11 f12 f13 f14
s block
p block
d block
f block

Irregular conformations of Cr and Cu

THE DEVIANT Ds

Chromium steals a 4s electron to half fill its 3d sublevel (more stable)



Group or Family ↓

Period →

Alkaline earth metals

Halogens

Noble gases

Transition metals

Alkali metals

1 1A	2 2A	3	4	5	6	7	8	9	10	11	12	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
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
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
55 Cs	56 Ba	57 La*	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
87 Fr	88 Ra	89 Ac†	104 Unq	105 Unp	106 Unh	107 Uns	108 Uno	109 Une	110 Uun	111 Uuu							

*Lanthanides	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	71 Lu
† Actinides	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	103 Lr

Noble Gas Electron Configuration

- Even shorter way to write how electrons are distributed in atom
- Shows all electrons in shells common to a Noble Gas as the symbol in a bracket [symbol] then list leftover electrons in their energy levels and shells

- Neon $1s^2 2s^2 2p^6$
- Aluminum $1s^2 2s^2 2p^6 3s^2 3p^3$
- As you can see Al has the same distribution of inner electrons as neon therefore it can be written as:
- $[Ne]3s^2 3p^3$

Steps for Noble Gas Configuration

1 Find element on periodic table.

2 Find number of electrons

3 Find Group 8 element from period above target element

4 Write group 8 element symbol in [brackets]

5 Subtract noble gases electrons from initial elements

6 Start filling from S subshell of initial elements period # til all electrons are placed

Element	Configuration notation	Orbital notation	Noble gas notation
Lithium	$1s^22s^1$	1s: ↑↓ 2s: ↑ 2p: _ _ _	$[He]2s^1$
Beryllium	$1s^22s^2$	1s: ↓ 2s: ↑↓ 2p: _ _ _	$[He]2s^2$
Boron	$1s^22s^2p^1$	1s: ↑↓ 2s: ↑↓ 2p: ↑ _ _	$[He]2s^2p^1$
Carbon	$1s^22s^2p^2$	1s: ↑↓ 2s: ↑↓ 2p: ↑ ↑ _	$[He]2s^2p^2$
Nitrogen	$1s^22s^2p^3$	1s: ↑↓ 2s: ↑↓ 2p: ↑ ↑ ↑	$[He]2s^2p^3$
Oxygen	$1s^22s^2p^4$	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑ ↑	$[He]2s^2p^4$
Fluorine	$1s^22s^2p^5$	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑	$[He]2s^2p^5$
Neon	$1s^22s^2p^6$	1s: ↑↓ 2s: ↑↓ 2p: ↑↓ ↑↓ ↑↓	$[He]2s^2p^6$

Orbital Notation or Diagrams

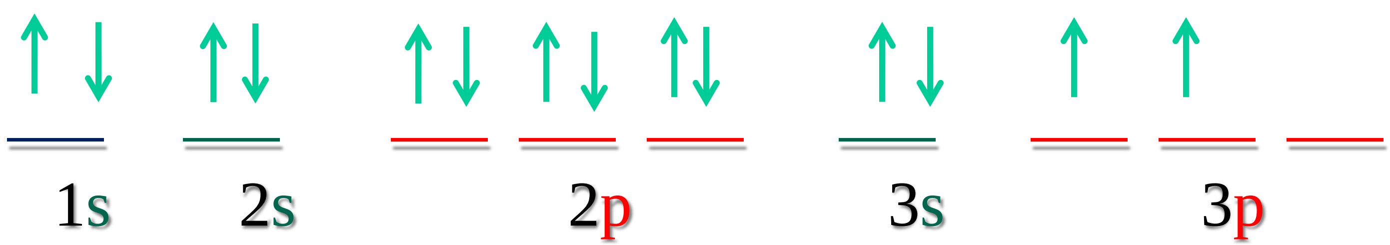
Simply use horizontal or vertical lines and arrows instead of exponents to represent the electrons

1 arrow = 1electron **Each line holds 2 electrons**

# of lines for S P D F must hold same number of electrons as in longhand electron configuration

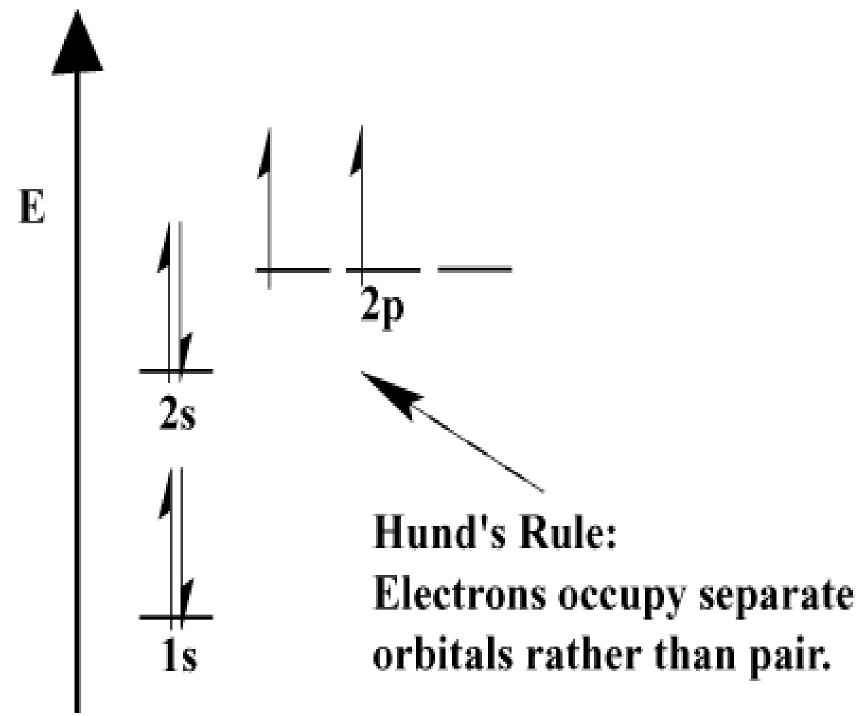
S = 2e^- so 1 line P= 6e^- so 3 lines d=10e^- so 5 lines

f= 14e^- so 7 lines



Rules for electron filling:

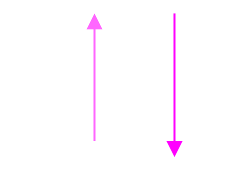
- Aufbaus Rule- must fill the lowest energy level available first!
- Hunds Rule -orbitals of equal energy sublevel will be occupied by one electron before a second one may enter (pairing up)
  - **No one can have seconds until everyone from same row has gone through once!!**



Pauli Exclusion Principle

Wolfgang
Pauli

Two electrons occupying the same orbital must have opposite spins- this eliminates 2 electrons within the same atom having the same QNS



How do electrons fill in an atom?
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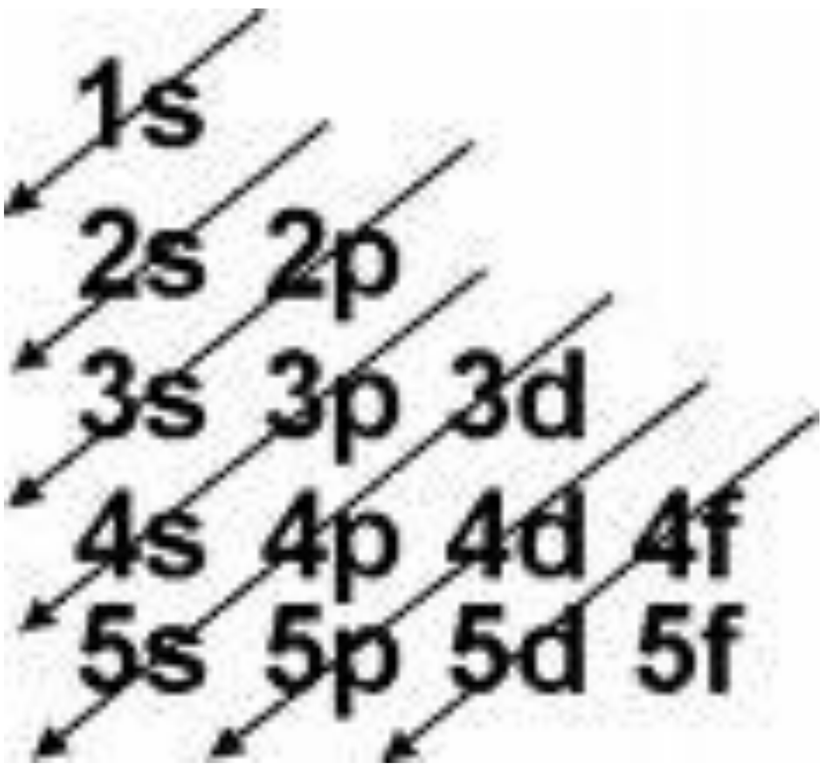
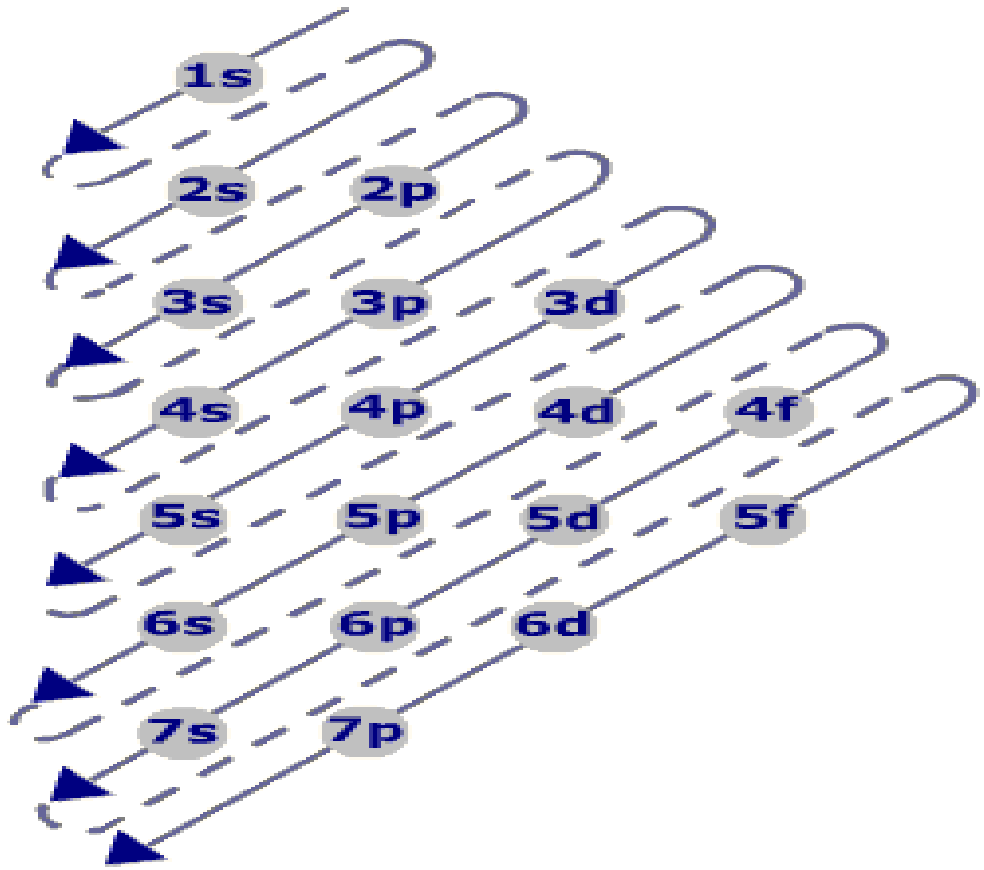
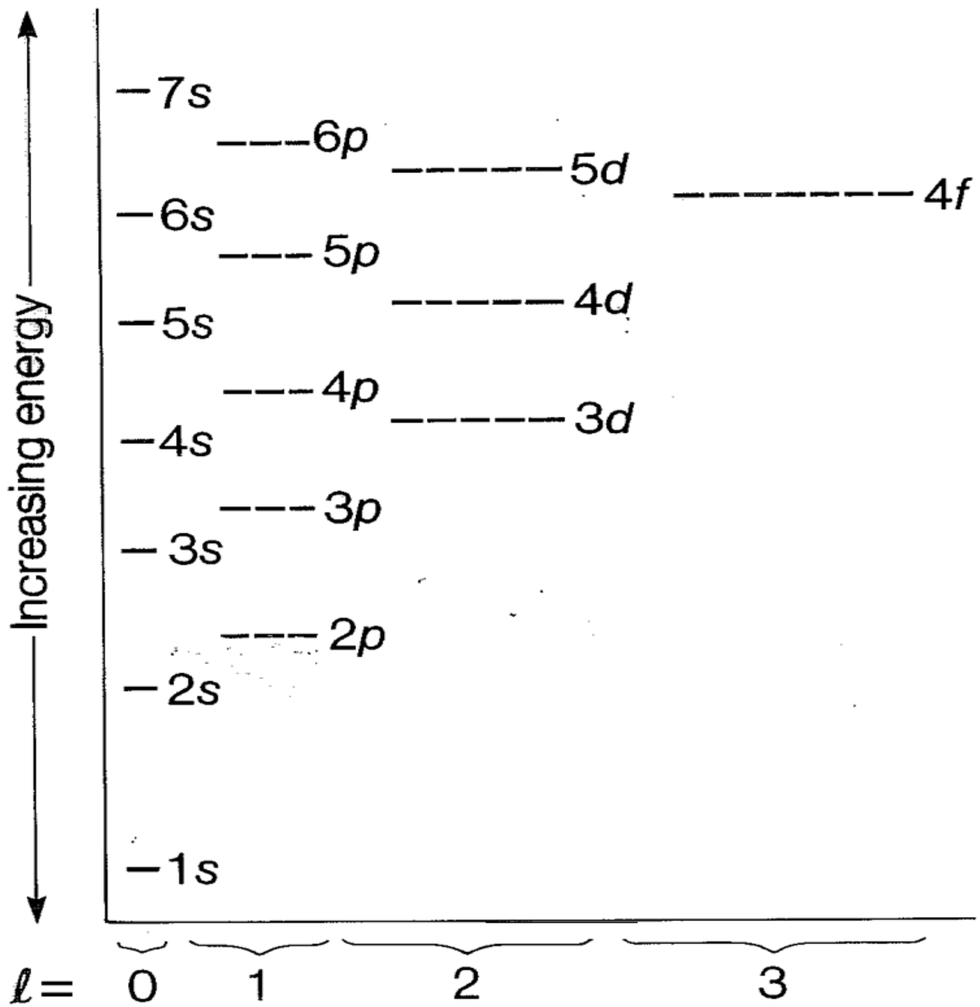




Things you must accept to do orbital Diagrams (energy diagrams)

- Energy builds further away from the nucleus
- Each line represents an orbital
- Each orbital can hold only two electrons
- We as a group will decide to place positive spin arrows in first..this is arbitrary NOT A RULE Just so all our QNS are the same
- Each electron is represented by an arrow
- In an orbital the two electrons must point in different directions
- Remember from the diagonal rule 4s fills before 3d breaking Aufbau's Rule

Orbital Diagrams



Energy Levels, Sublevels, Electrons

Energy Level (n)	Sublevels in main energy level (n sublevels)	Number of orbitals per sublevel	Number of Electrons per sublevel	Number of electrons per main energy level ($2n^2$)
1	s	1	2	2
2	s p	1 3	2 6	8
3	s p d	1 3 5	2 6 10	18
4	s p d f	1 3 5 7	2 6 10 14	32

Quantum Mechanical Model of the Atom

Mathematical laws can identify the regions outside of the nucleus where electrons are most likely to be found.

These laws are beyond the scope of this class…VERY SIMPLY PUT… each electron has four numbers to describe the probability of finding the electron within the atom…

Schrodinger Wave Equation

Erwin Schrodinger

$$-\frac{h^2}{8\pi^2 m}\frac{d^2\psi}{dx^2}+V\psi=E\psi$$

Equation for probability of a single electron being found along a single axis (x-axis)

What are they?

- Characterization of the orbital that an electron occupies
- Describes the following:
  - distance from the nucleus
  - shape
  - position with respect to the 3 dimensional axis
  - direction of spin of the electron

4 Quantum Numbers

n= Principle Quantum Number

distance from the nucleus (Denotes Size)

values 1-7 (note 7 periods on the P.T.)

L = Sublevel

shape of the cloud

values 0→3 (0 = s, 1 = p, 2 = d, 3 = f)

m = magnetic

orientation about the axis values –3 →3

S= spin

Direction of movement within orbital

+ ½ or – ½

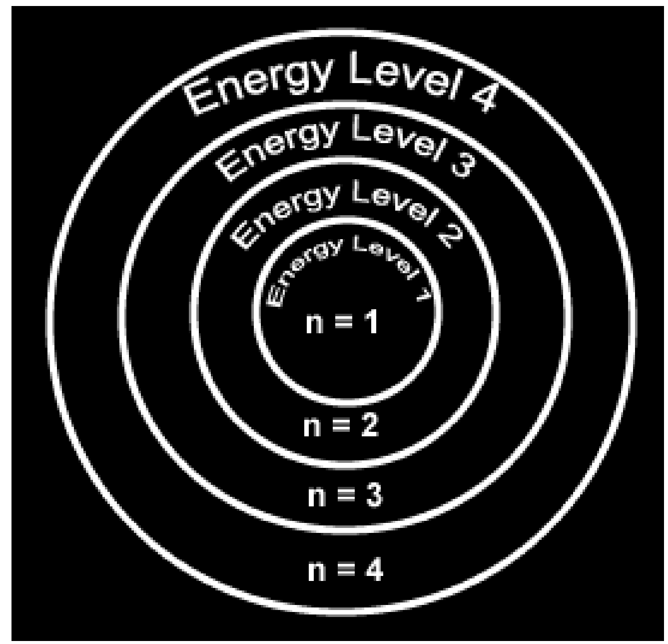
Electron Energy Level (Shell)

Generally symbolized by n, it denotes the probable distance of the electron from the nucleus.

n is numerically represented by positive integers. (1-7)

Number of electrons that can fit in a shell:

$$2n^2$$



Angular Momentum

- Abbreviated by ℓ
- Describes the shape of an orbital by using a number and letter designation.
- Determined by all values between
- (n-1) through 0

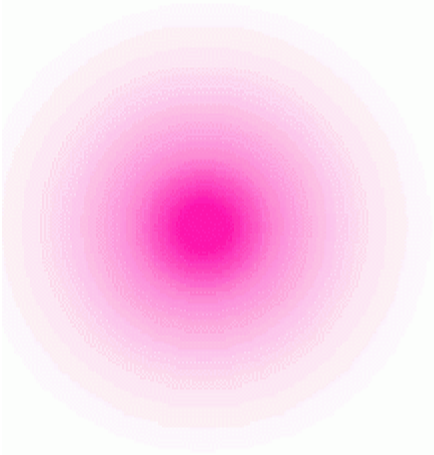
ℓ	Letter
0	s
1	p
2	d
3	f

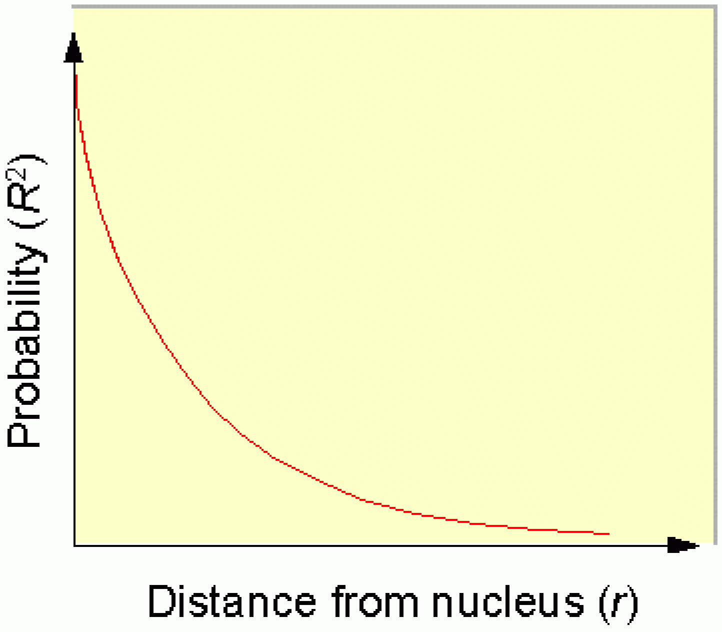
There are predicted shapes out to ℓ=6

Which would be a i orbital but don't

exsist in nature.

An orbital is a region within an energy level where there is a probability of finding an electron. This is a probability diagram for the s orbital in the <u>first energy level…</u>

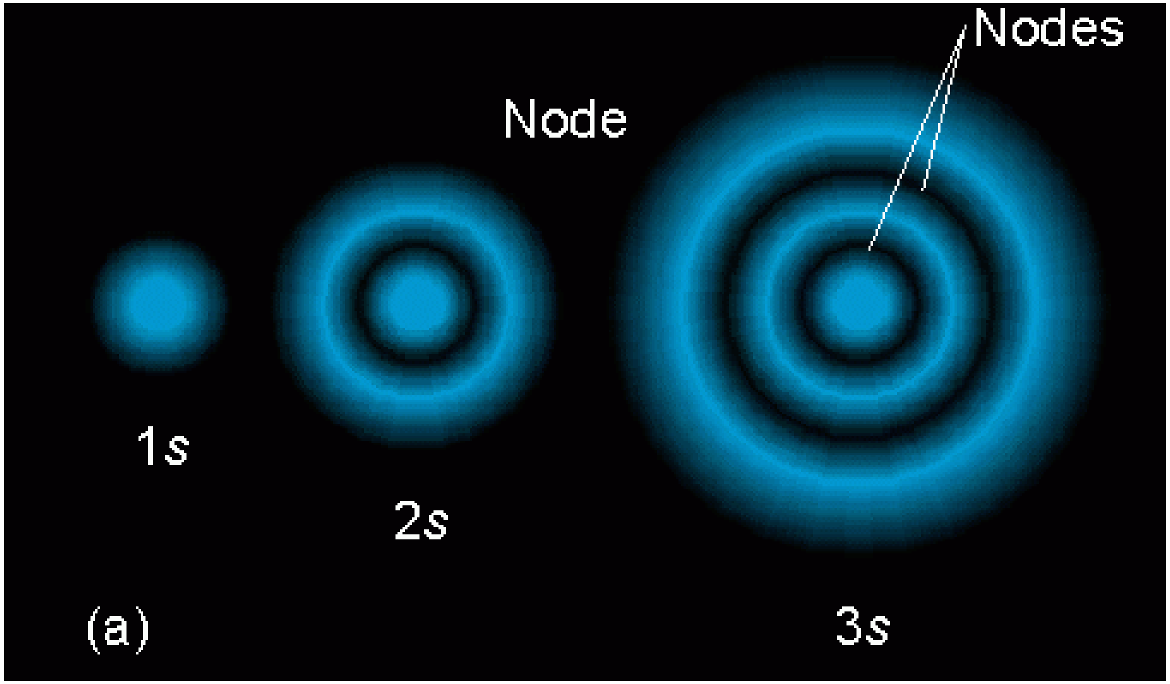




<u>Orbital shapes</u> are defined as the surface that contains <u>90%</u> of the total electron probability.

Sizes of *s* orbitals

Orbitals of the same shape (s, for instance) grow larger as n increases…

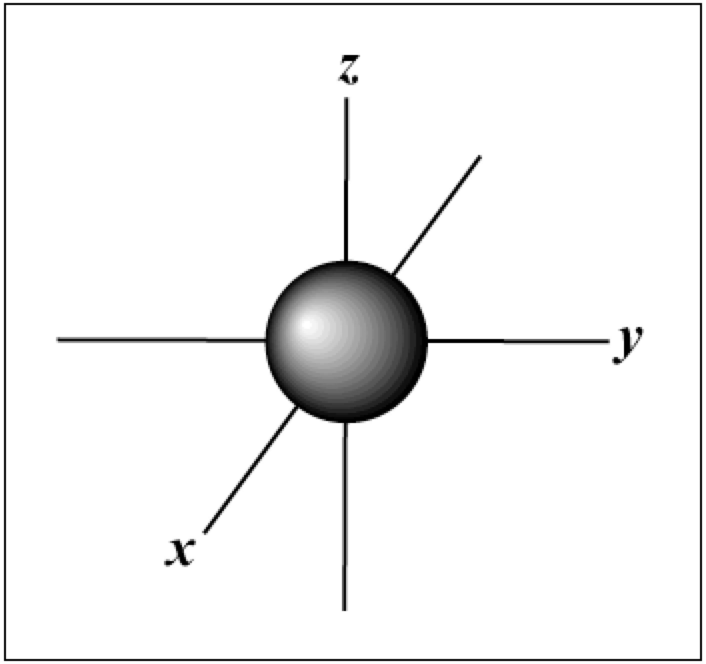


Nodes are regions of low probability within an orbital.

Shapes

- s orbital is spherical
- p orbital looks like a dumbell
- d orbitals look like 2 dumbells
- f orbitals look like flowers

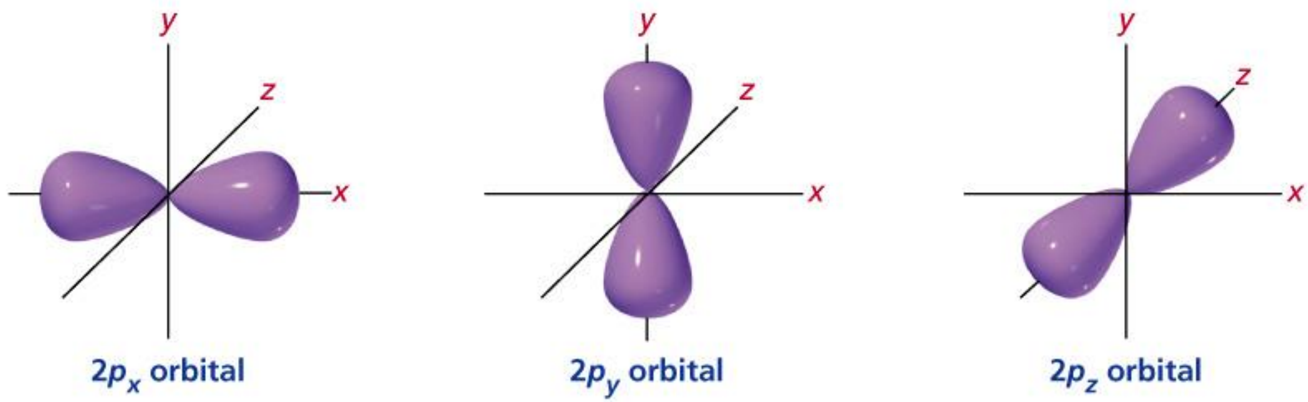
The *s* orbital has a spherical shape centered around the origin of the three axes in space.

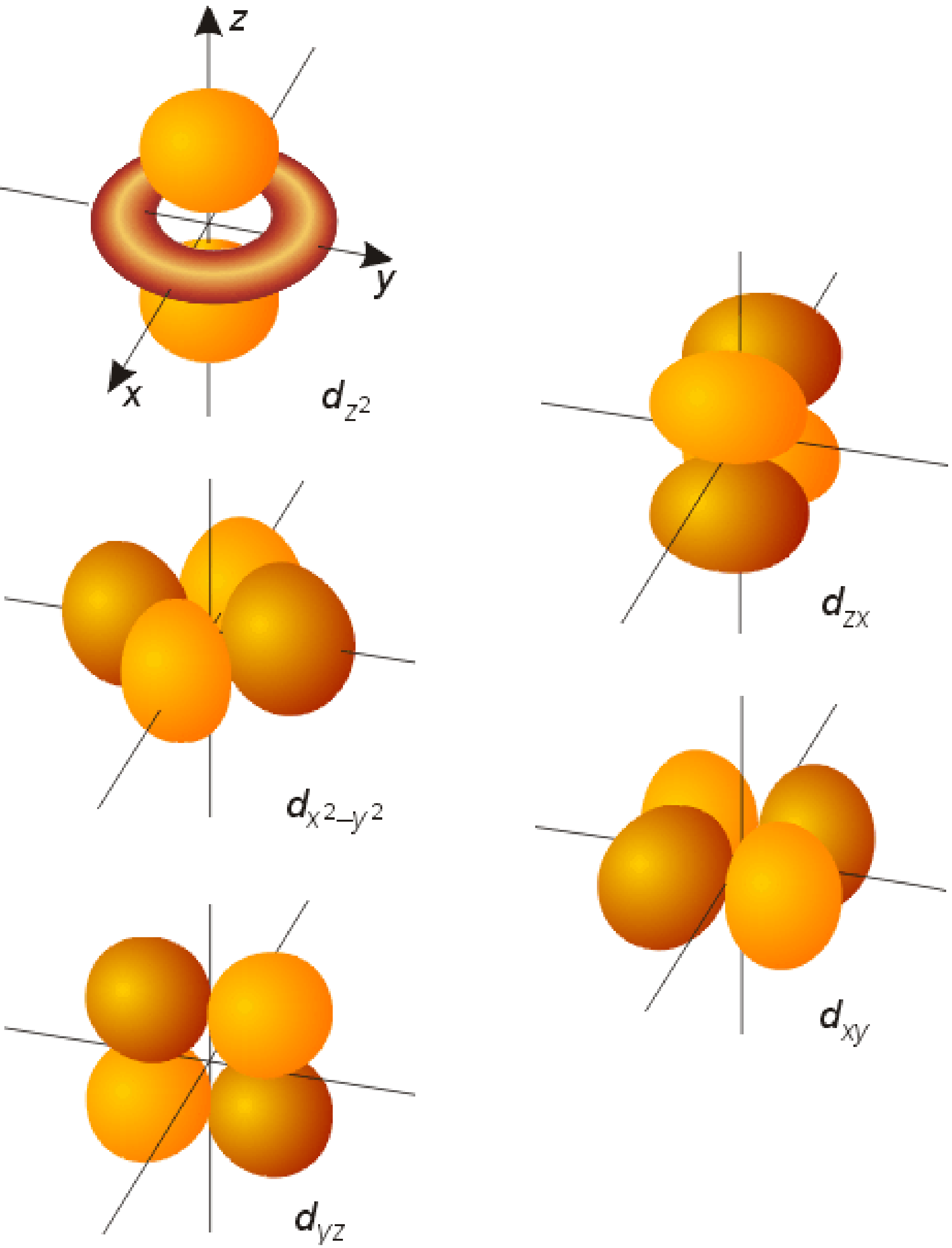


s orbital shape

P orbital shape

There are three dumbbell-shaped *p* orbitals in each energy level above n = 1, each assigned to its own axis (x, y and z) in space.

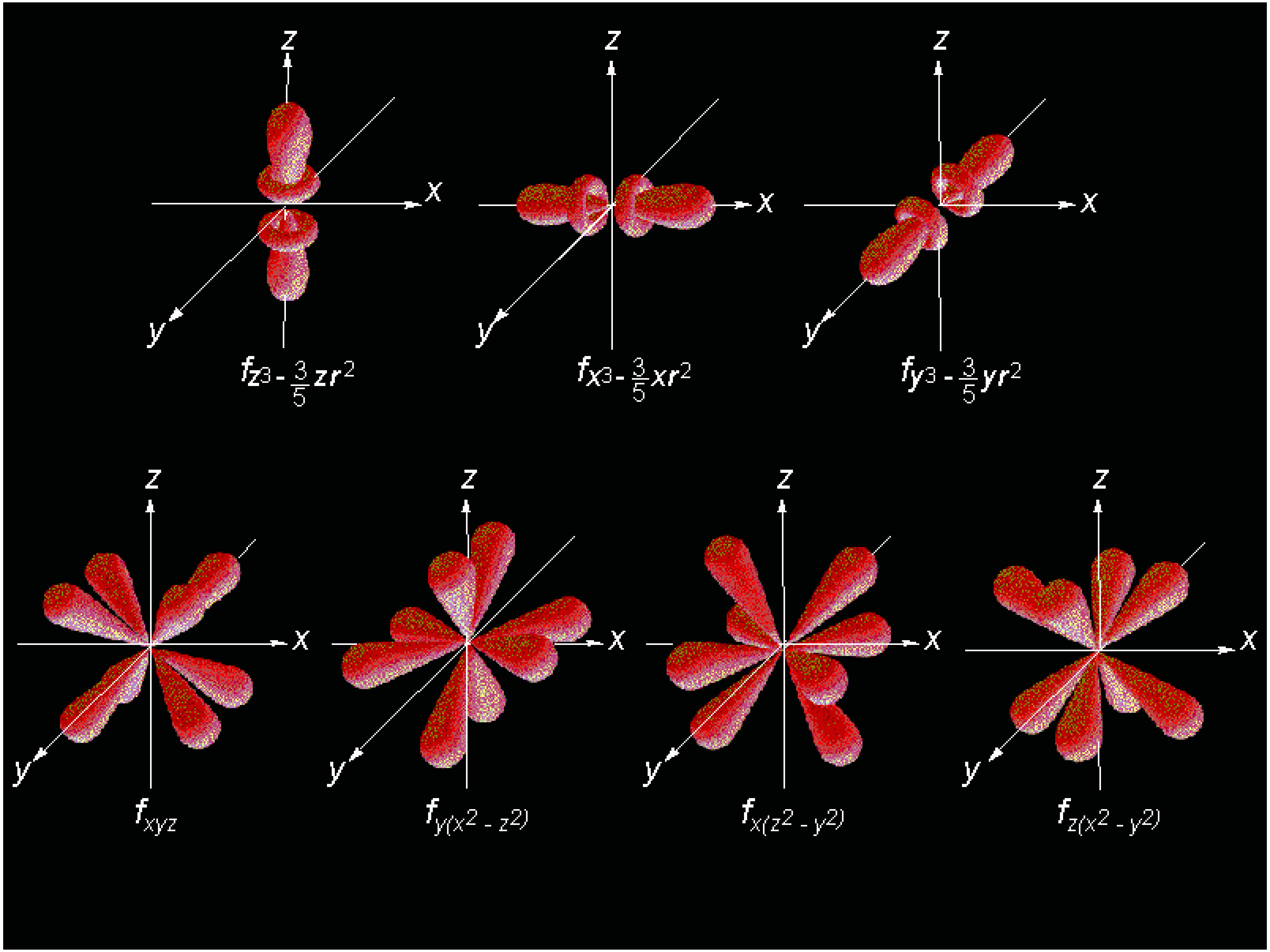




Things get a bit more shapes complicated with the five *d* orbitals that are found in the *d* sublevels beginning with n = 3. To remember the shapes, think of "double dumbells"

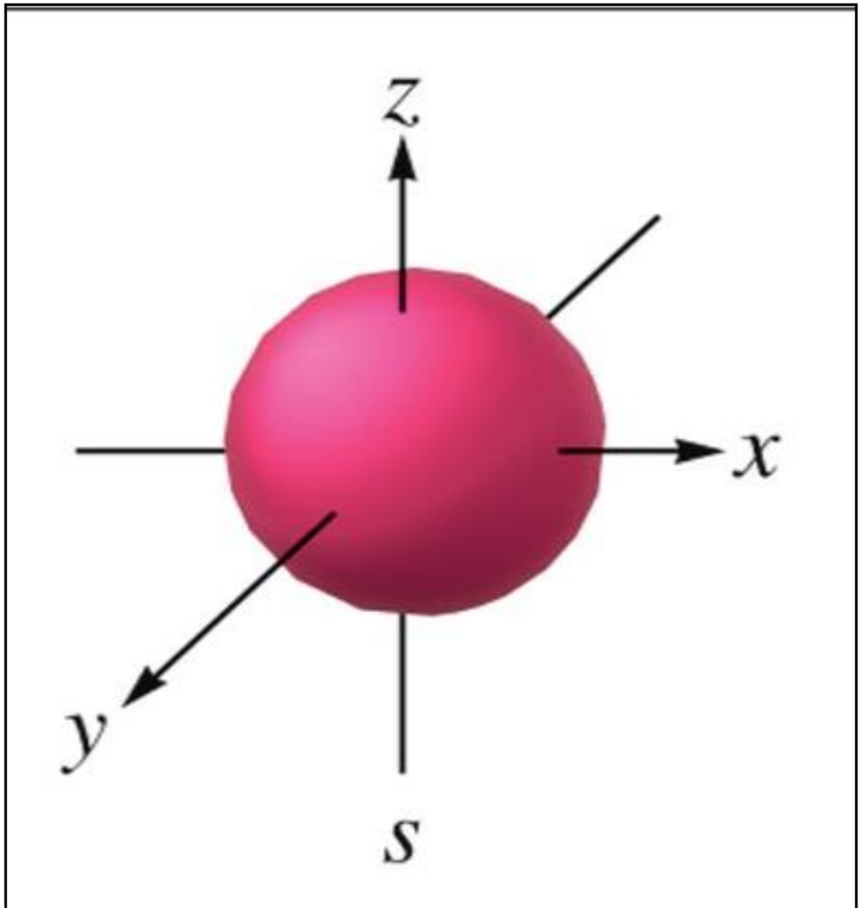
...and a "dumbell with a donut"!

Shape of f orbitals

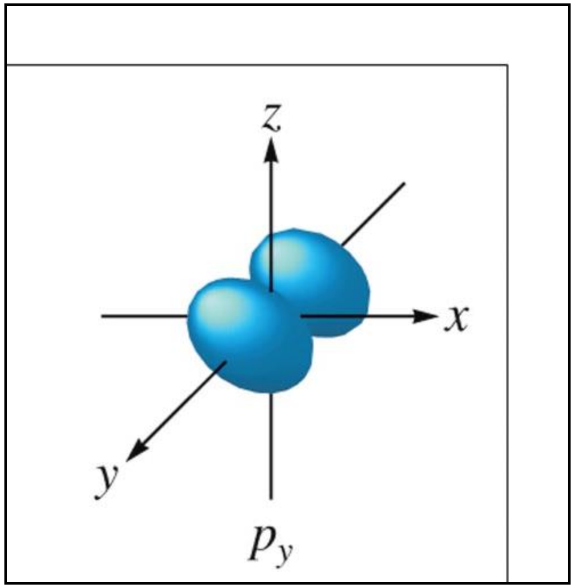


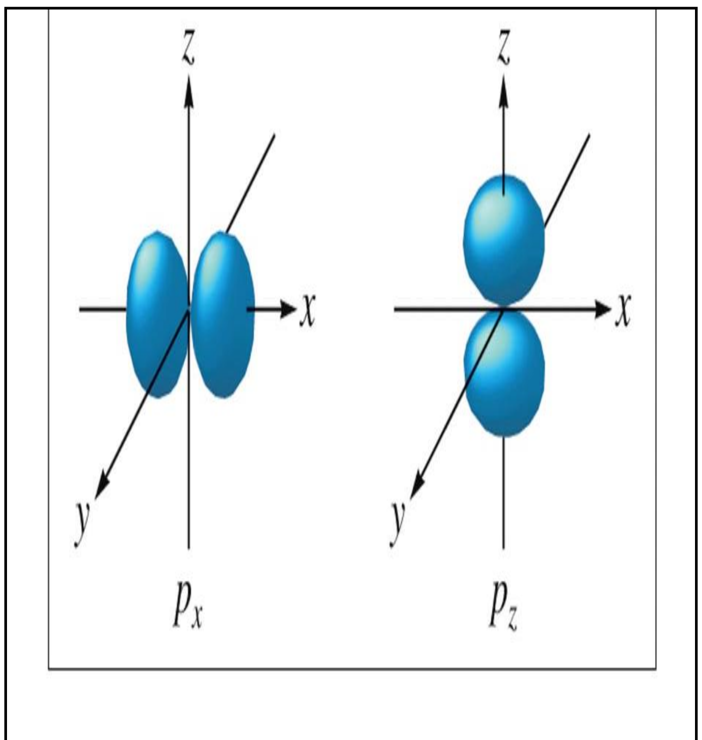
Magnetic Quantum Number

- Abbreviated by m_l or simply m
- Represents the 3 dimensional orientation of an orbital (how your orbital is pointed)
- Assumes all values from $-\ell \ldots 0 \ldots +\ell$



- If n=1
- Then ℓ=0
- So m_l=0
  - Think of spinning a basketball on your finger…does it really change where it is in 3D??





- For a p orbital…
  - n=2
  - ℓ=1
  - m_l= -1, 0, +1
    - The p orbital can exist in different planes in 3D

Spin Quantum Number

- Abbreviated by m_s or s
- Uses +1/2 or -1/2
- Indicates the direction of spin of an electron
- Electrons in the same orbital MUST have opposite spins!