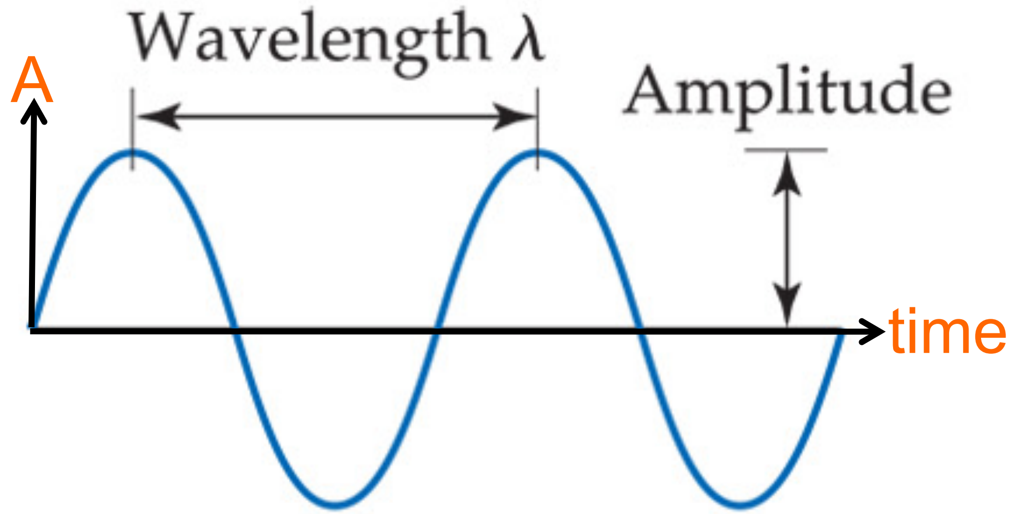
Electro-magnetic radiation (light)

- The nature of light
  - light is a wave
- The nature of waves
  - What is a wave?
  - What is waving?

Waves

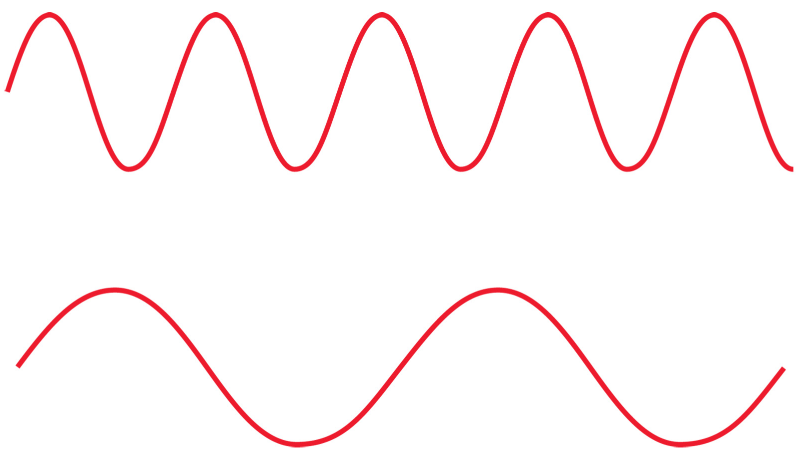


- Wave: some sort of periodic function
  - something that periodicaly changes vs. time.
- wavelength (λ): distance between equivalent points
- Amplitude: "height" of wave, maximum displacement of periodic function.

Waves

Higher frequency

shorter wavelength

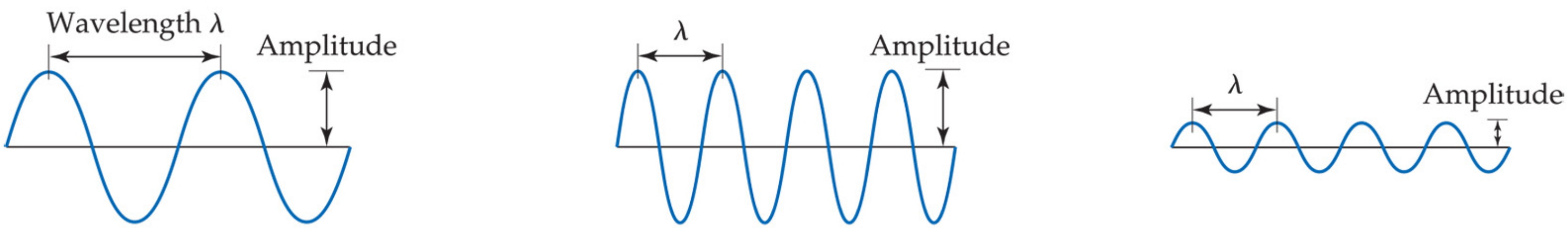


lower frequency

longer wavelength

- The number of waves passing a given point per unit of time is the frequency (ν).
- For waves traveling at the same velocity, the longer the wavelength, the smaller the frequency.

Waves



(a) Two complete cycles of wavelength λ

(b) Wavelength half of that in (a); frequency twice as great as in (a)

(c) Same frequency as (b), smaller amplitude

v = wavelength x frequency

meters x (1/sec) = m/sec

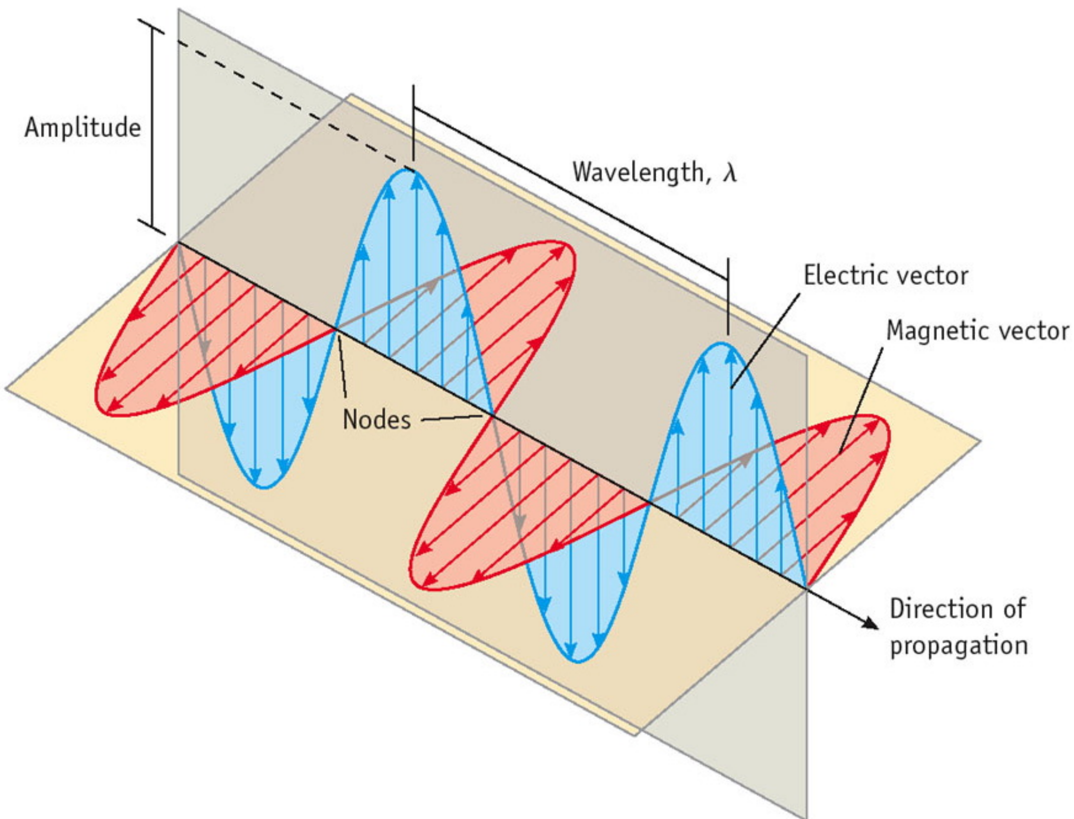
$v = \lambda\nu$

Waves

Major question:

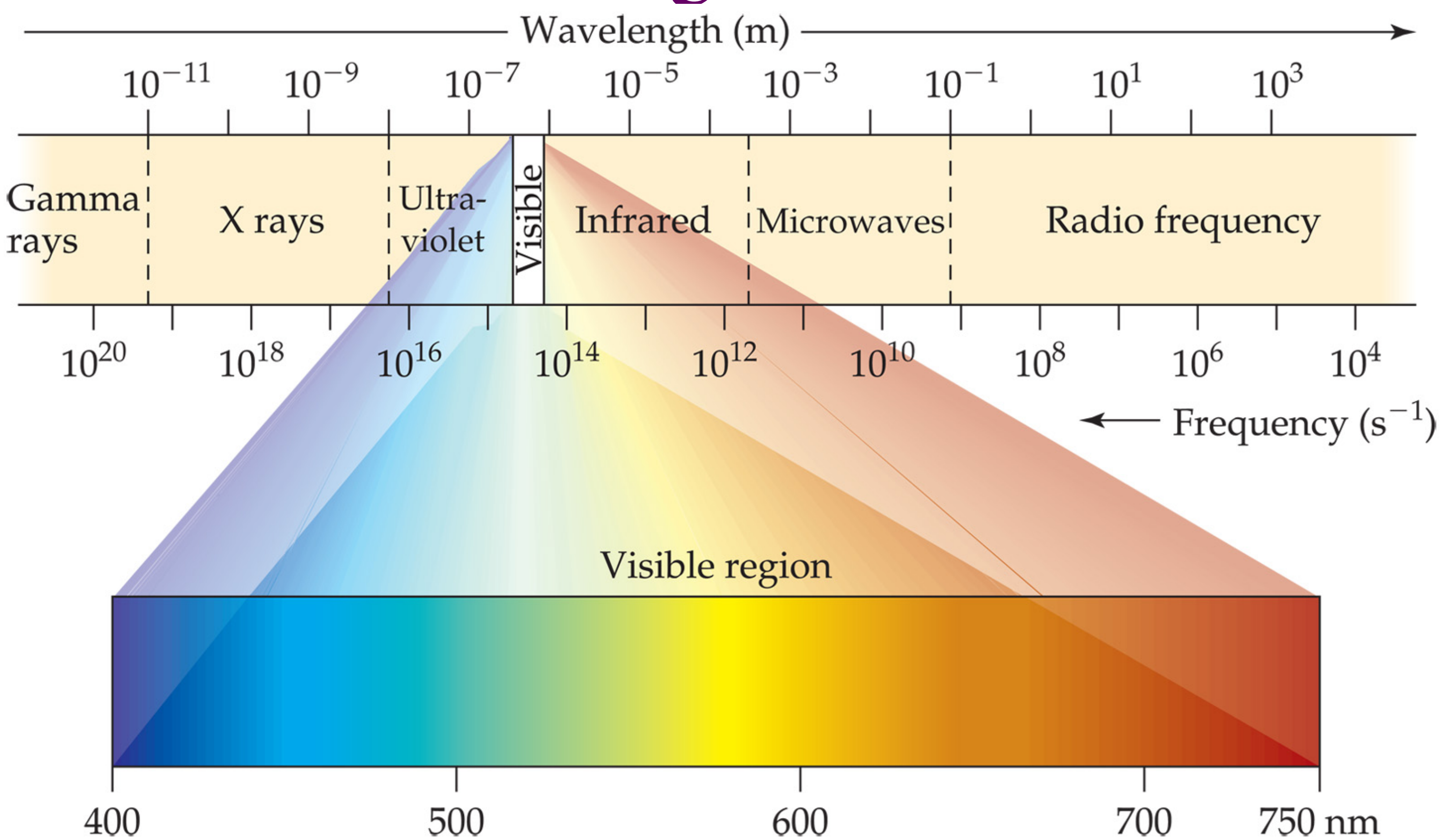
- What is waving?
- water wave:
  - water height(pressure)
- Sound wave:
  - air pressure
- Light?

Light waves.



- What is waving? Electric field, and perpendicular magnetic field.
- Faraday thought this, Maxwell proved it.

Electromagnetic Radiation



- **All electromagnetic radiation travels the speed of light (*c*), 3.00 × 10^8 m/s (in a vacuum).**
- **Therefore:** $c = \lambda \nu$

Speed of light in other materials

Index of refraction is:

$n = c/v$

The index of refraction of some common materials are given below.

material	n	material	n
Vacuum	1	Crown Glass	1.52
Air	1.0003	Salt	1.54
Water	1.33	Asphalt	1.635
Ethyl Alcohol	1.36	Heavy Flint Glass	1.65
Fused Quartz	1.4585	Diamond	2.42
Whale Oil	1.460	Lead	2.6

Values of n come from the CRC Handbook of Chemistry and Physics

The major issue of late 19th century physics

- What is light?
- Light and energy?
- How does light interact with matter?

History of the atom

Atomic spectra

Bunsen, Kirchhoff, 1860
- 1st spectroscope
- 1st line spectrum

Lockyer, 1868
- He in solar system

Balmer, 1885
- H line spectrum
- Rydberg 1890
  - generalized line spectra, but no explanatio!

Quantum theory

- Plank,1900
  - ➢ Black body radiation
- Einstein, 1905
  - ➢ Photoelectric effect
- Bohr, 1913
  - ➢ Application of atomic theory to the electronic spectra of hydrogen

Atomic structure

- Dalton, 1803
  - ➢ atomic nature
- Faraday, 1834
  - ➢ Electricity & Mag.
- Thompson, 1807
  - ➢ electrons e/m
- Millikan, 1911
  - ➢ oil drop
- Rutherford, 1911
  - ➢ gold foil/nucleus

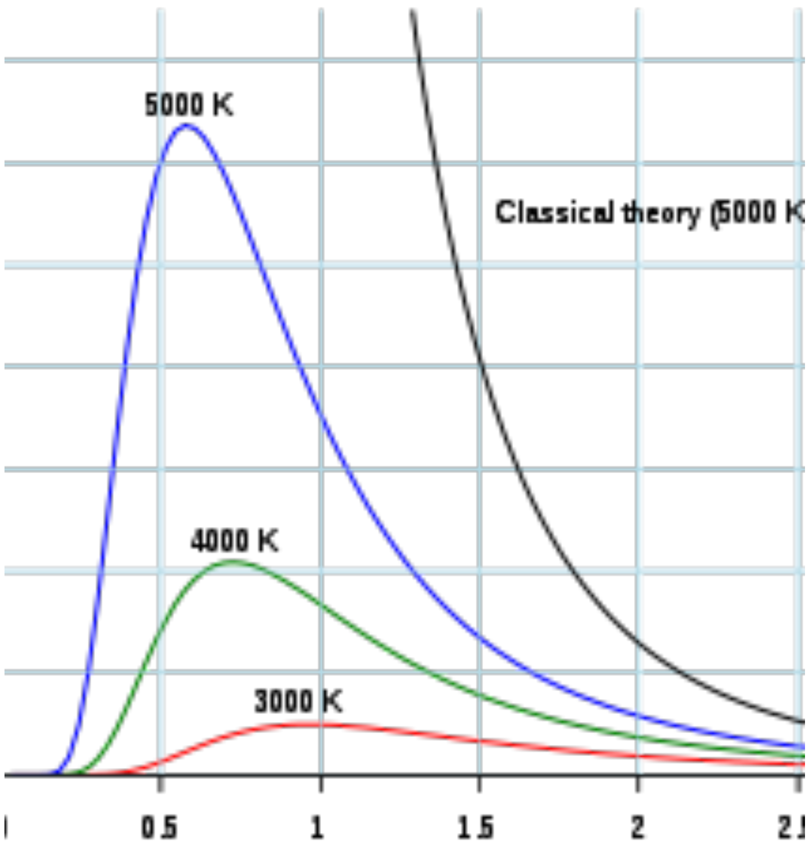
History of the atom continued

- Bohr, 1913
  - Application of atomic theory to the electronic spectra of hydrogen
- Pauli 1923
- De Broglie, 1924
  - Wave properties of particles
- Born, Heisenberg, Dirac, 1925
  - Quantum mechanics
- Schrodinger, 1926
  - Wave mechanics
- Heisenberg, 1927
  - Uncertainty principle
- Heitler, London, 1927
  - Valence Bond theory
  - Pauling, 1928,1930
    - hybridization, resonance
- Mulliken, 1928
  - Molecular orbital theory

Discrete atoms

← **Chemical bonding**

Black Body Radiation



Spectral output of a black body.

Black shows that predicted from classical electricity & magnetism

Colored curves are what you actually get.

Light is emitted when atoms vibrate (or oscillate), but they can only oscillate with an energy given by:

- **$E = nh\nu$**

Mystery 1: Black body radiation

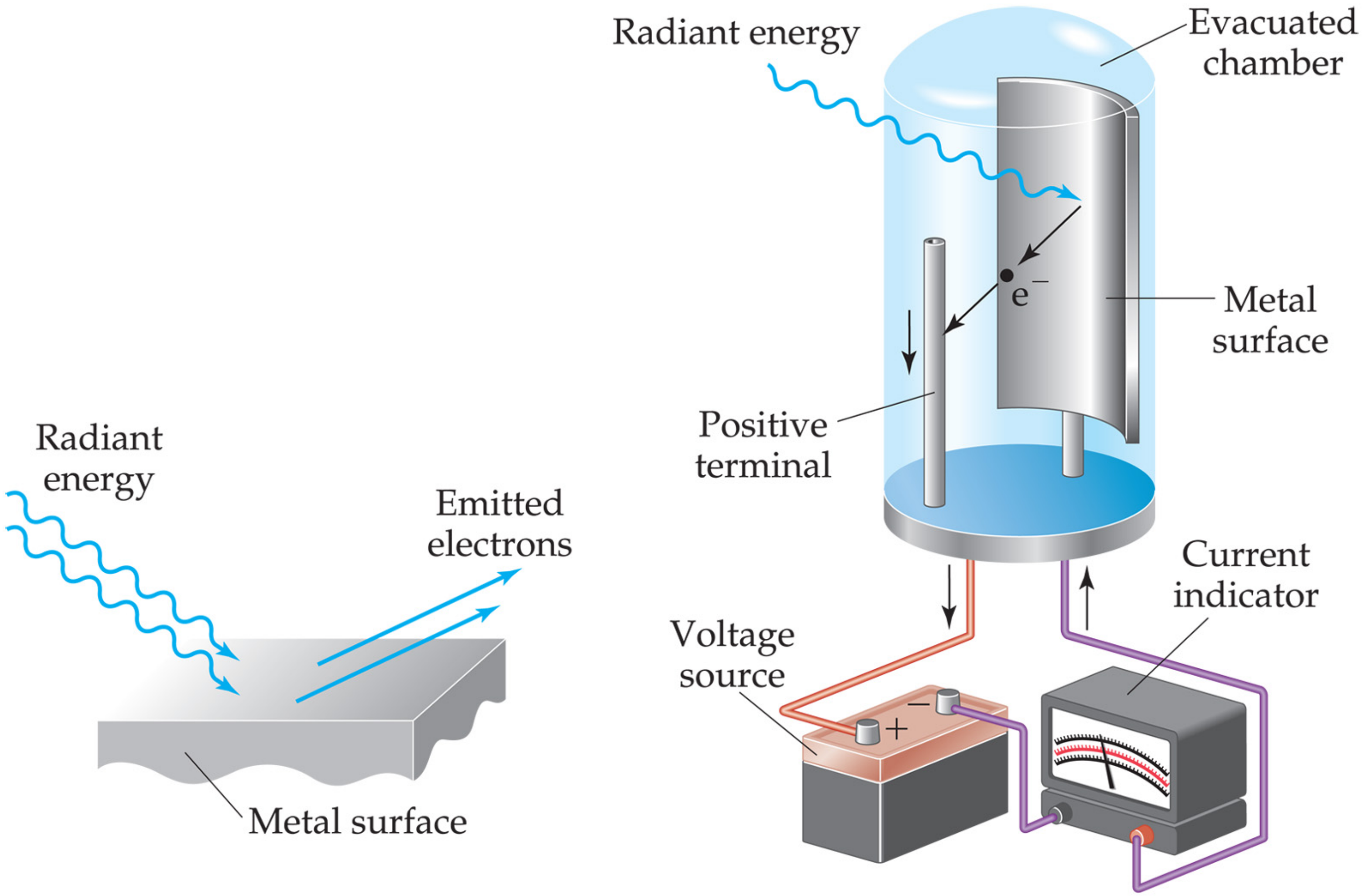
- Higher T leads to shorter wavelength of light
- More K.E., more E
- Must be relationship between E and wavelength
- Plank concluded that energy is *quantized. It comes in packets (like fruit snacks) and is* proportional to frequency:

  $E = h\nu$

  where h is Planck's constant, 6.63×10^{-34} J-s. The minimum packet of E.

What did Einstein get the Nobel Prize for?

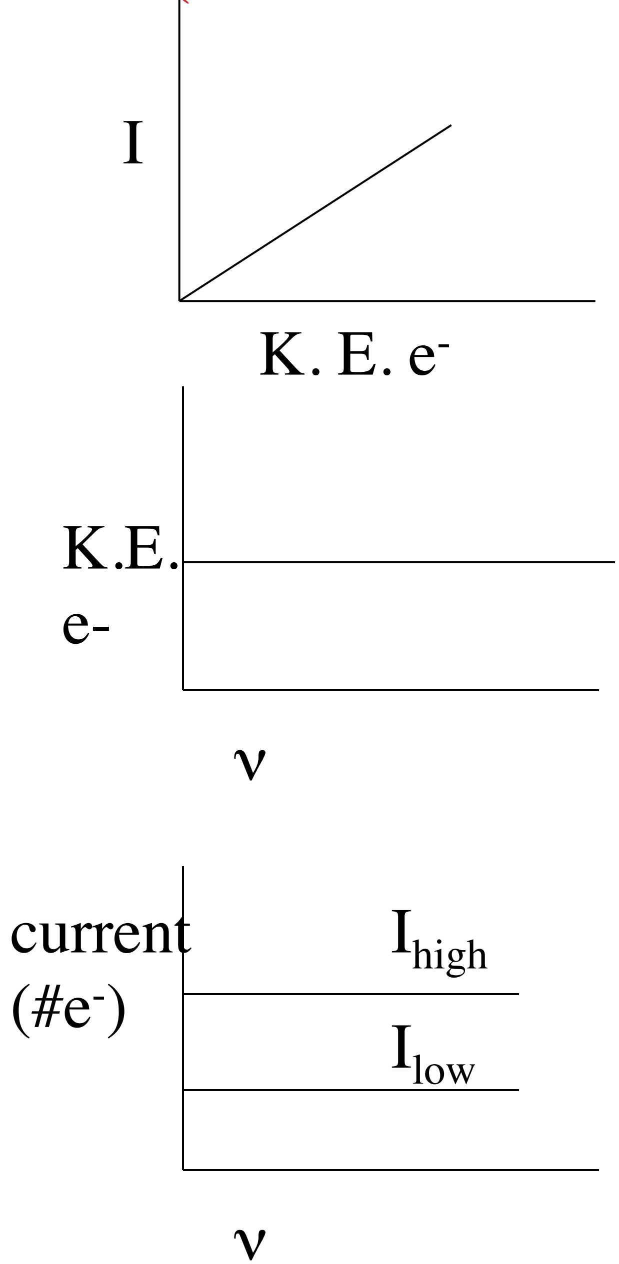
Mystery #2: The Photo-electric effect

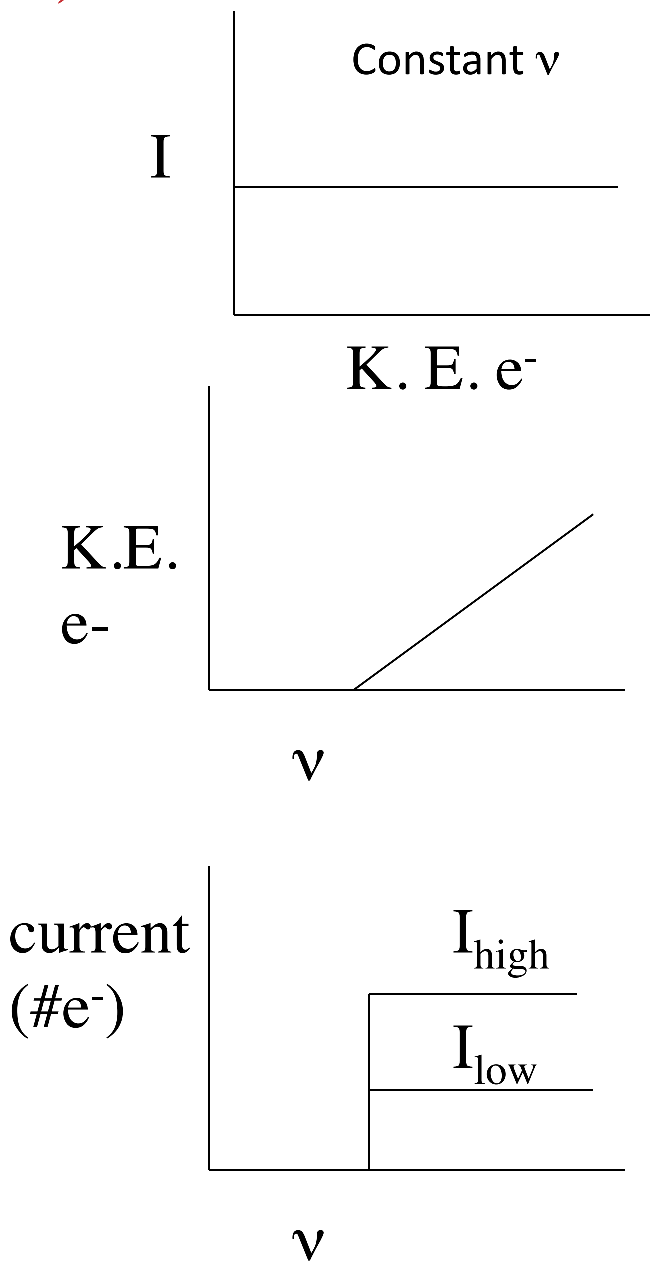


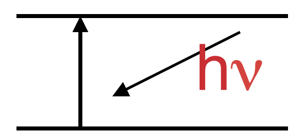
Note, this is what a photo cell does
Turn light into work (current)

What might you expect (from normal waves)

what do you see?





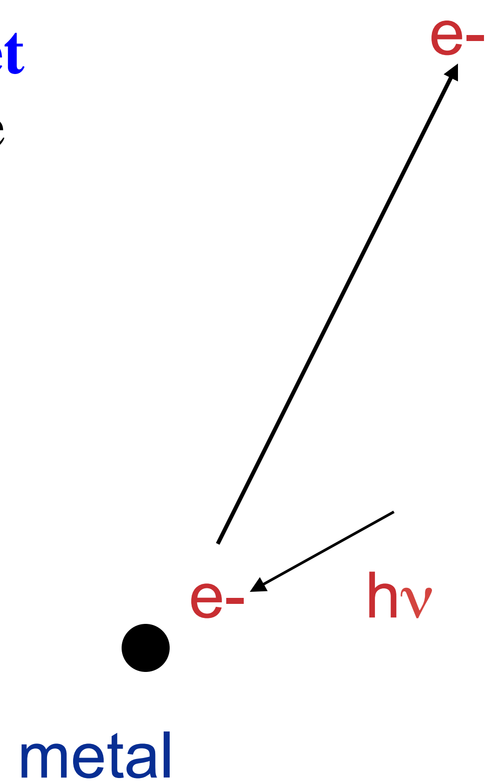


Einstein: Light is both a particle and a wave.

e- K.E.   "escape energy"

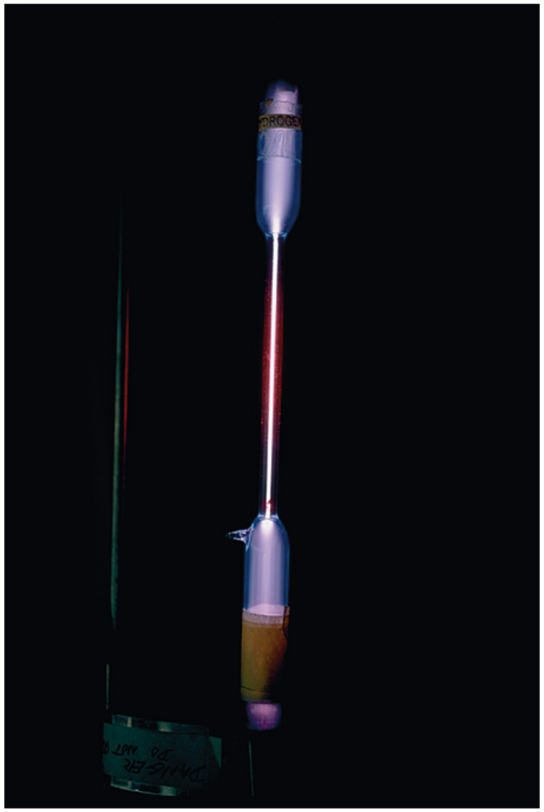
$$E_{photon} = 1/2mv^2 + h\nu_o = E_{electron}$$

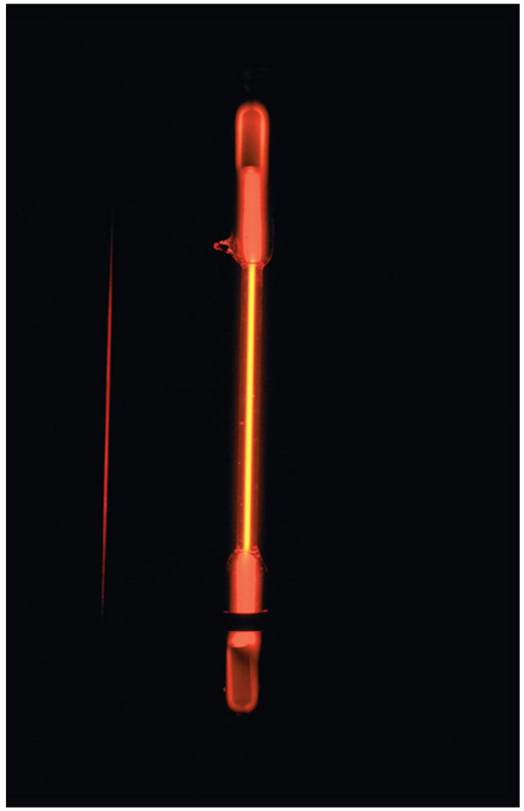
light comes in **packets of energy**. Each **packet** runs into **one electron**. Each packet must have enough E to break electron loose from metal. The rest of the energy goes into kinetic energy.



Frequency tells us the E of each packet.
I tells us how many **packets/second** we get.
More packets, more current (more electrons knocked off).

Mystery number 3: element line spectrum





Hydrogen

Neon

Gas discharge tube (full of some elemental gas)

Gives off specific frequencies of light only.

Different elements give off different colors.

i.e. different energies.

The Nature of Energy

- Energy, λ, ν, related:

$$c = \lambda\nu$$

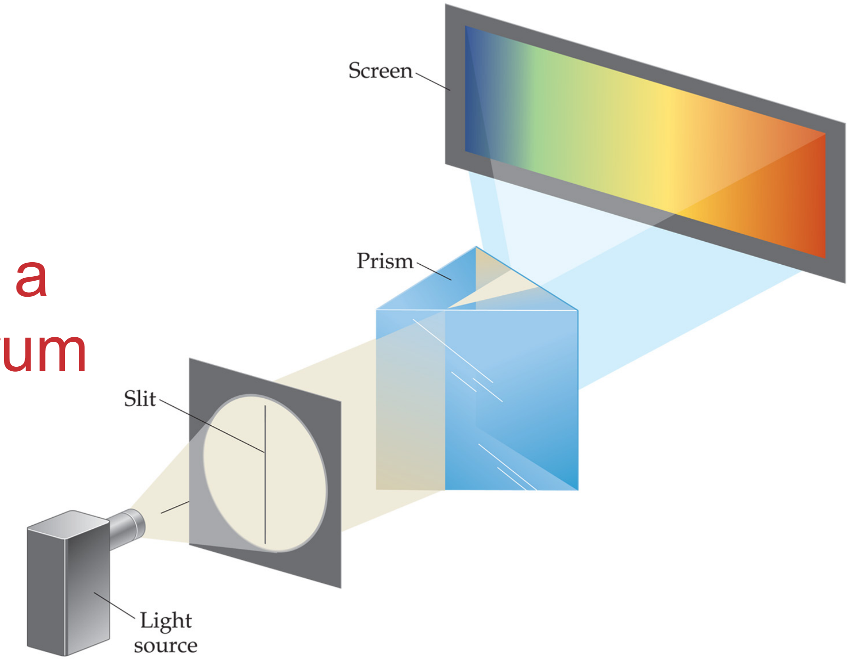
$$E = h\nu$$

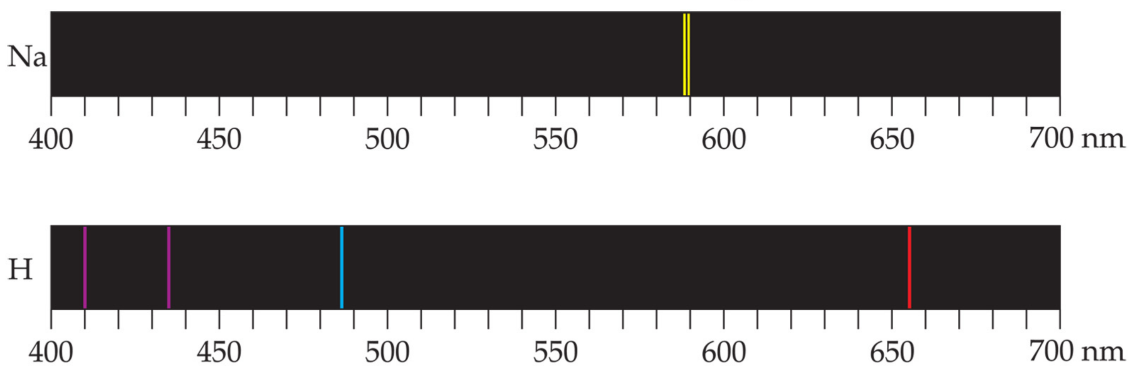
c= speed of light in vacuum,

constant

The Nature of Light

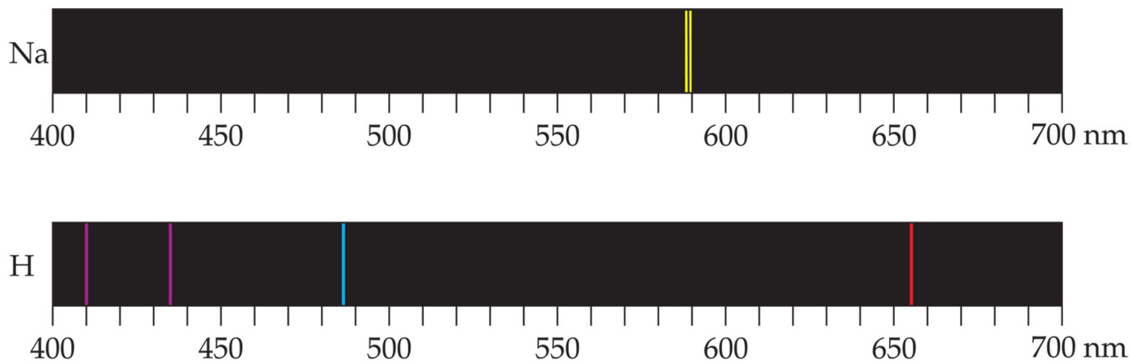
White light shows a continuous spectrum





- A line spectrum of discrete wavelengths is observed from an element

Hydrogen Line spectra



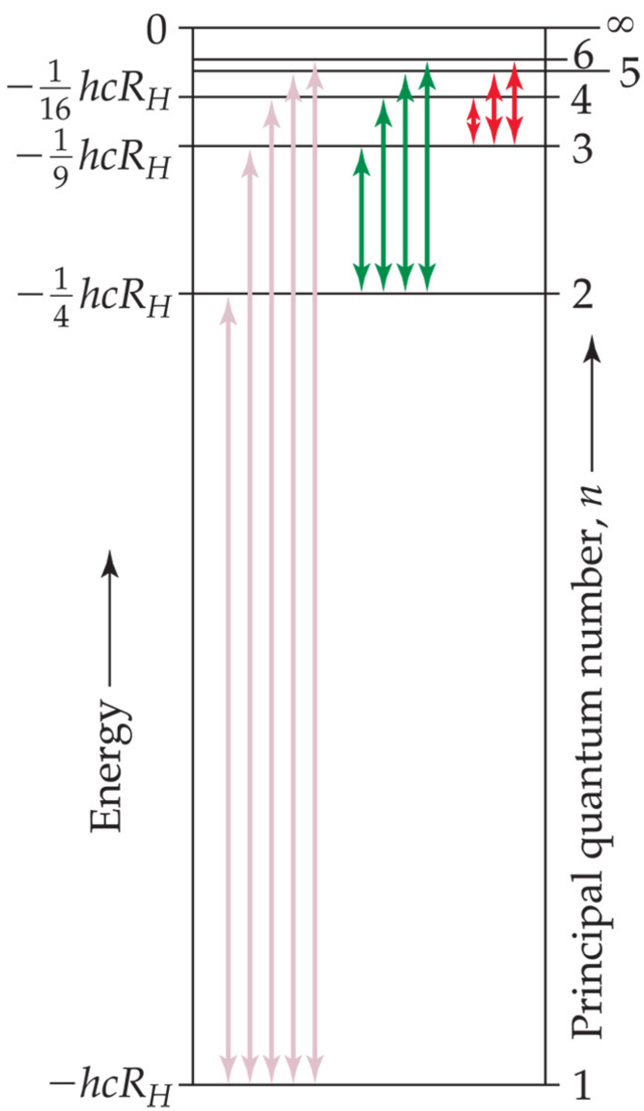
Johann Balmer, School teacher
figured out that the lines fit a simple equation:

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

R_H=constant
n_1 and n_2 are integers

But why?

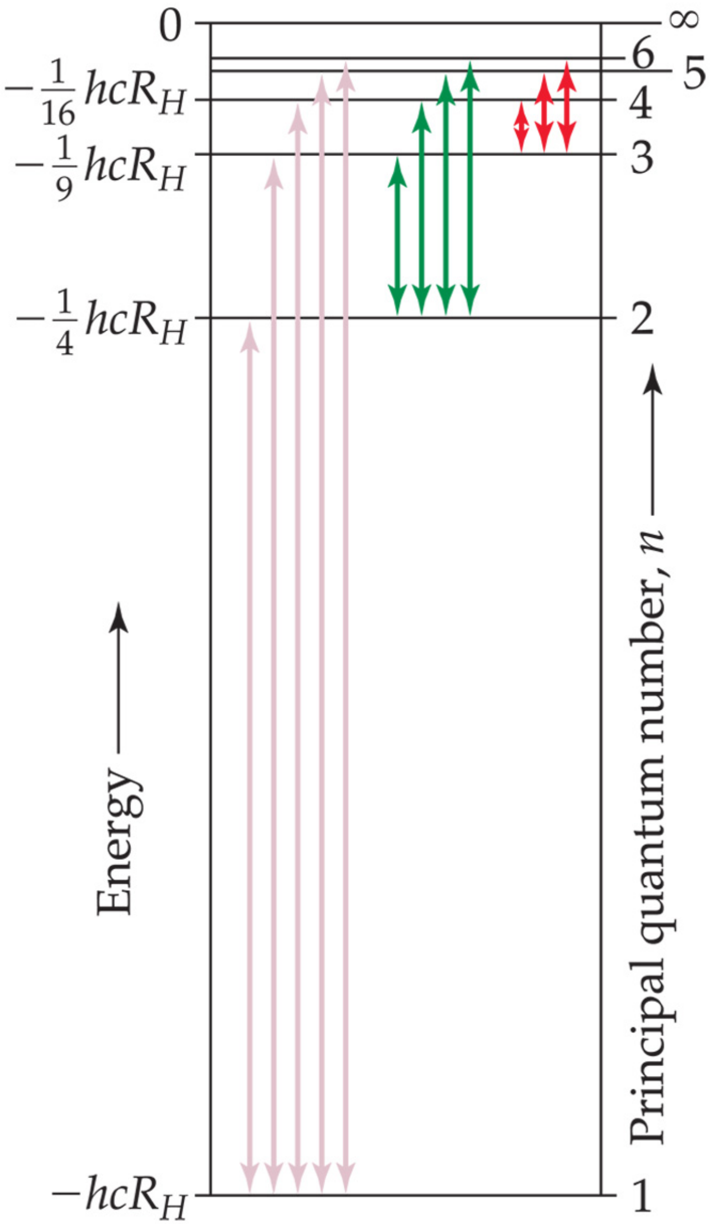
The Nature of Energy



- Niels Bohr adopted Planck's assumption and explained these phenomena in this way:

  1. **Electrons in an atom can only occupy certain orbits (corresponding to certain energies).**

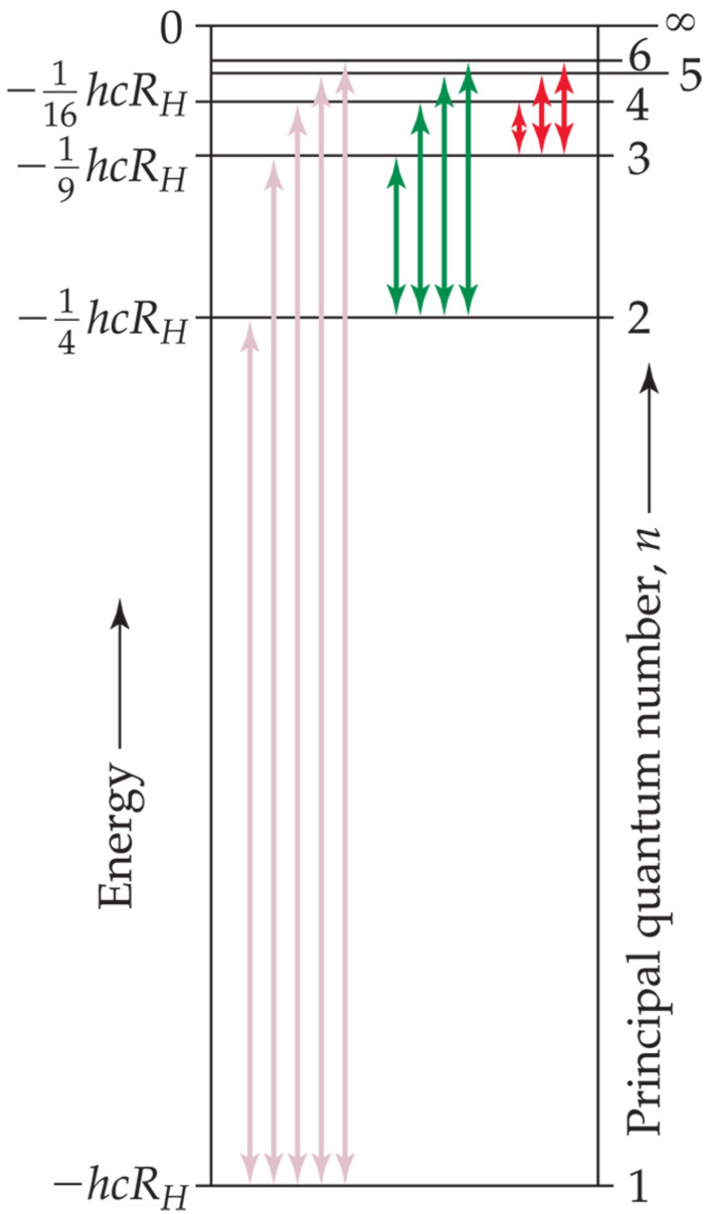
The Nature of Energy



- Niels Bohr adopted Planck's assumption and explained these phenomena in this way:

2. Electrons in permitted orbits have specific, "allowed" energies;

The Nature of Energy

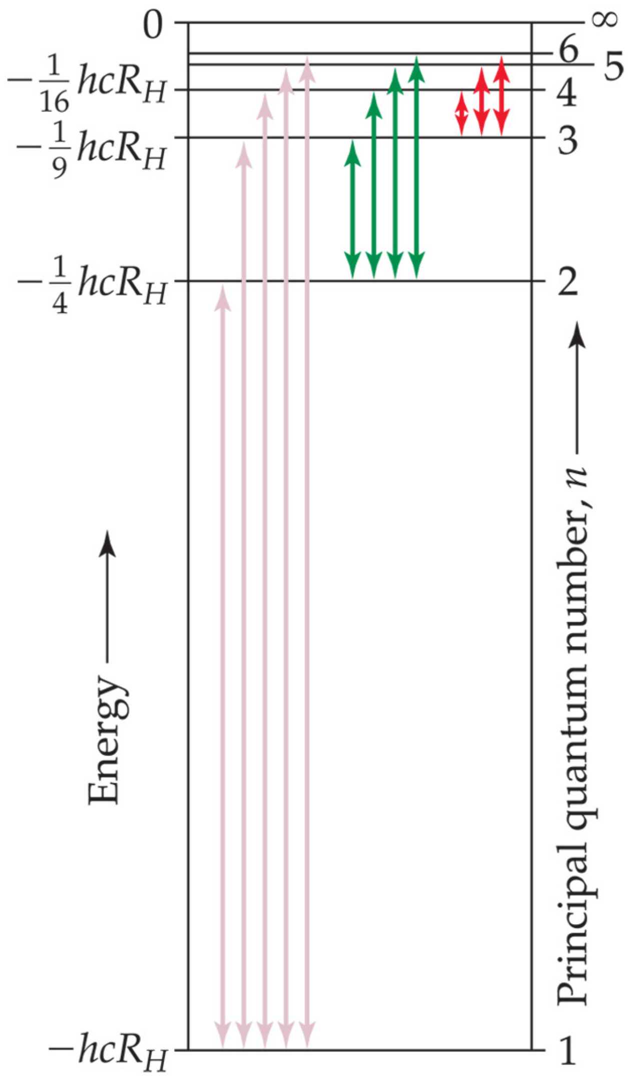


- Niels Bohr adopted Planck's assumption and explained these phenomena in this way:

3. Energy is only absorbed or emitted in such a way as to move an electron from one "allowed" energy state to another; the energy is defined by

$$E = h\nu$$

The Nature of Energy

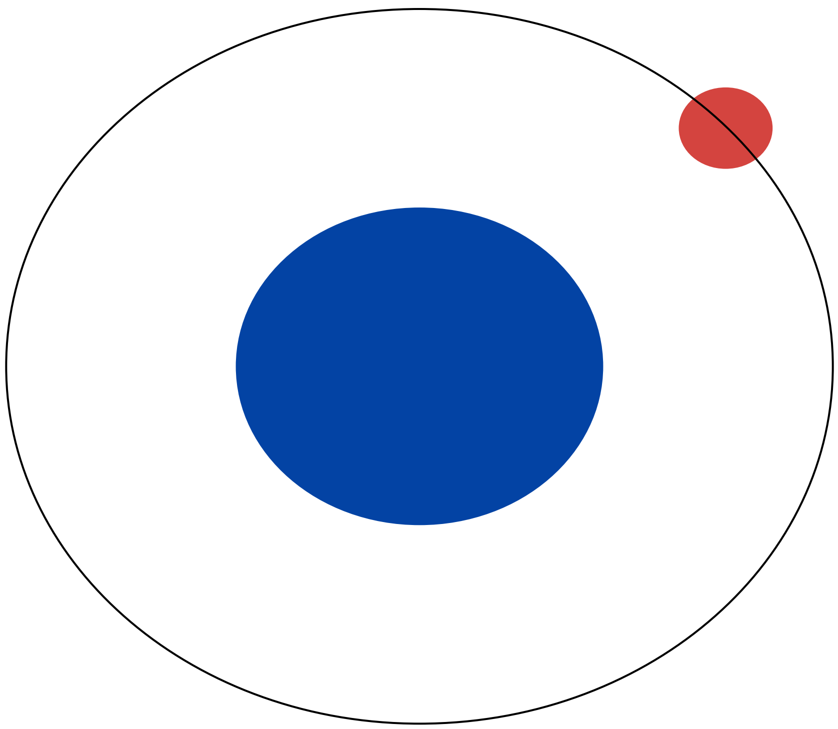


The energy absorbed or emitted from electron promotion or demotion can be calculated by the equation:

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

where R_H is the Rydberg constant, 2.18×10^{-18} J, and n_i and n_f are integers, the initial and final energy levels of the electron.

Bohr.



- Using a model that had electrons orbiting the nuceus like planets, Bohr could explain H, but no other elements.
- Too simple.

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

$R_E = 1/2 m_e c^2 \alpha^2$

The Wave Nature of Matter

- Louis de Broglie: if light can be a particle, maybe matter can be wave-like.

$$Velocity = \lambda \nu$$

like $E=mc^2$

$$\nu = \frac{velocity}{\lambda}$$

$$E = m(velocity)^2 = h\nu = h\frac{velocity}{\lambda}$$

$$\lambda = \frac{h}{m(velocity)}$$

Wave-like nature of matter

$$\lambda = \frac{h}{mv}$$

However, the higher the mass, the smaller the wavelength & $h = 6.63 \times 10^{-34}$ J-s, a really small number.

Example; What is λ for a 1 g ball?

$$\lambda = \frac{6.63 \times 10^{-34} kgm^2/s}{.001kg(1m/s)} = 6.63 \times 10^{-31} \text{ m}$$

wavelengths of everyday objects too small to measure.

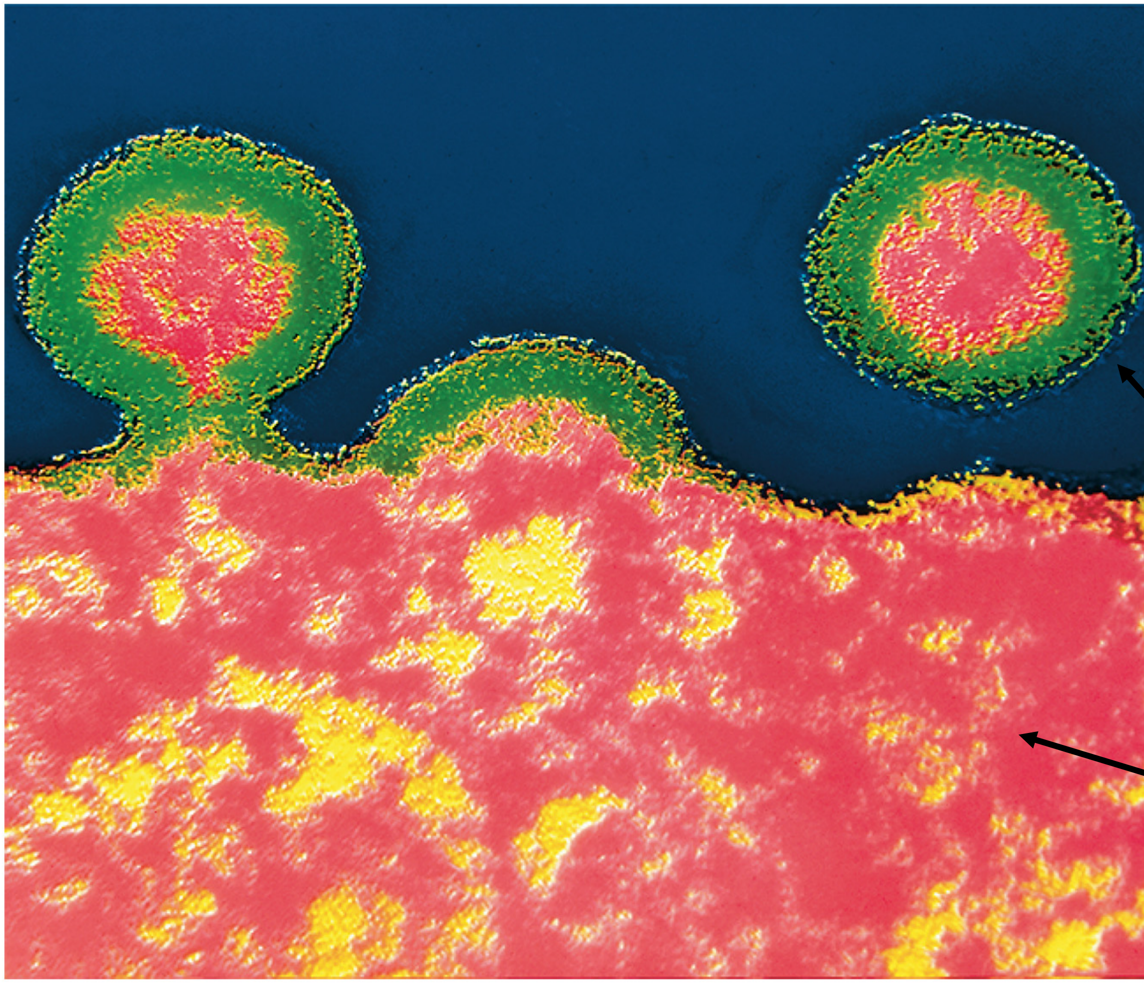
Wave-like nature of matter

- What about an electron? $v = 6 \times 10^6$ m/s:
- $m = 9.1 \times 10^{-28}$ g.

$$\lambda = \frac{6.63\text{x}10^{-34}\text{kgm}^2/\text{s}}{9.1 \times 10^{-28} (6 \times 10^6 \text{ m/s})} = 1.22 \times 10^{-10} \text{ m} = .122 \text{ nm}$$

Wavelength of X-rays

Electron microscopy



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Because electron wavelengths are very small, you can use them to look at very small things.

HIV virus
100 nm, (light microscope limit 400 nm)

T-lymphocyte

The Uncertainty Principle

- Heisenberg showed that the more precisely the momentum of a particle is known, the less precisely is its position known:

$$(\Delta x)\ (\Delta mv) \geq \frac{h}{4\pi}$$

- our uncertainty of the whereabouts of an electron can be greater than the size of the atom!

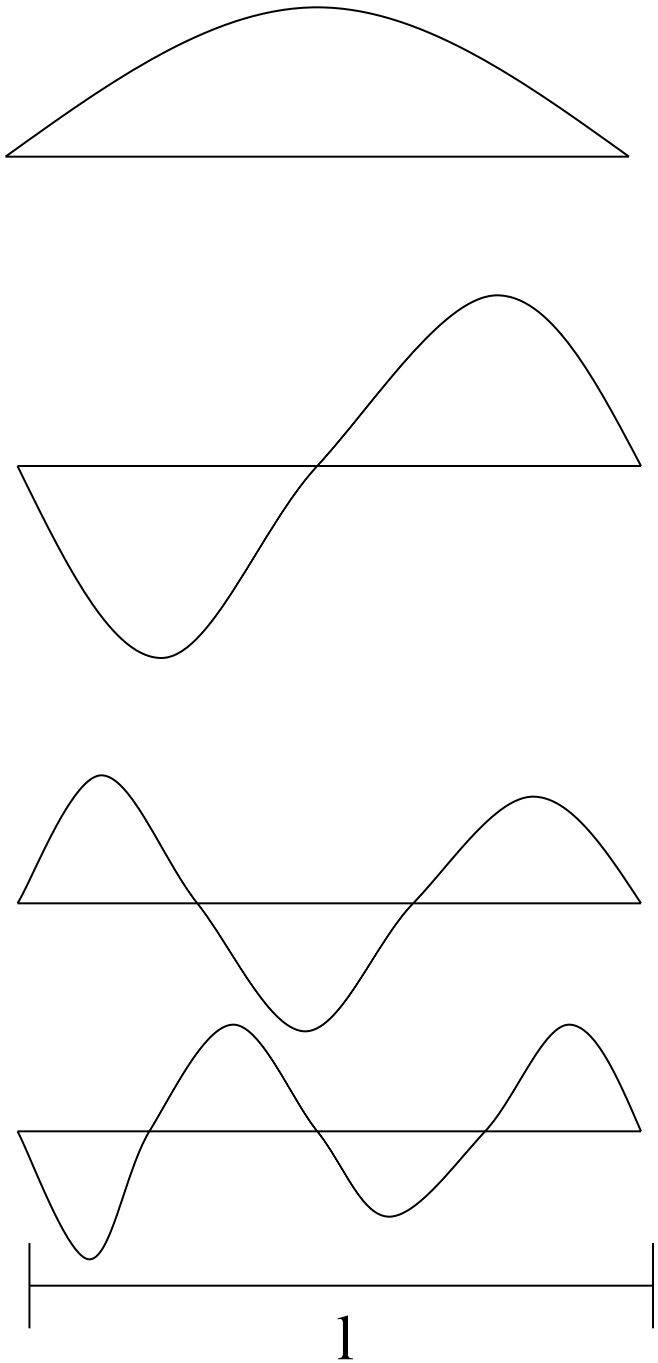
This is a result of the wave/particle duality of matter

"The clues"

- 1. Plank: E of light is quantized & depends on frequency
- 2. Einstein/photo-electric effect: Light behaves like a particle when it interacts with matter
- 3. Emission spectra/Bohr: Potential E. of electrons are quantized in an atom
- 4. Debroglie: wave/particle duality of electrons (matter).
- 5. Standing waves: are quantized inherently

Born/Schroedinger/Jordan: use standing wave analogy to explain electron P.E. in atoms. *Quantum Mechanics*

Standing waves



$l=(1/2)\lambda$
ν_O=frequency
nodes = 2 (gotta have 2)

$l = (2/2)\lambda = \lambda$
$2\nu_O$=frequency
nodes = 3

$l=(3/2)\lambda$
$3\nu_O$=frequency
nodes = 4

$l=(4/2)\lambda=2\lambda$
$4\nu_O$=frequency
nodes = 5

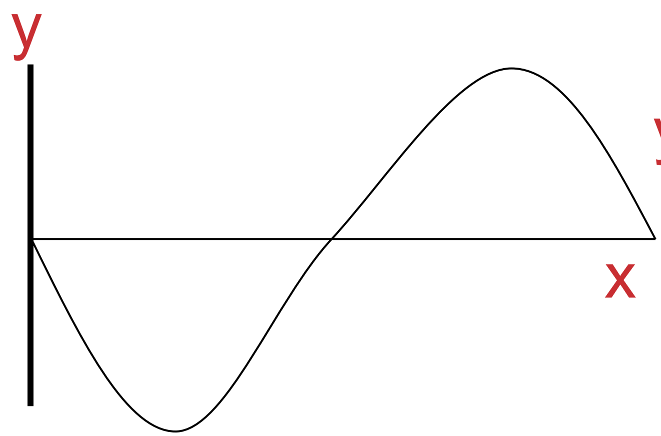
Allowed ν and λ *quantized*.
$l = (n/2)\lambda$, n is quantum #
frequency = $n\nu_O$

Quantum mechanics

- Each electron can be explained using a standing wave equation (wavefunction)
- Quantized frequency corresponds to quantized Energy (Debroglie, Plank, etc.)
- Integer values are critical to this description: *quantum numbers.*

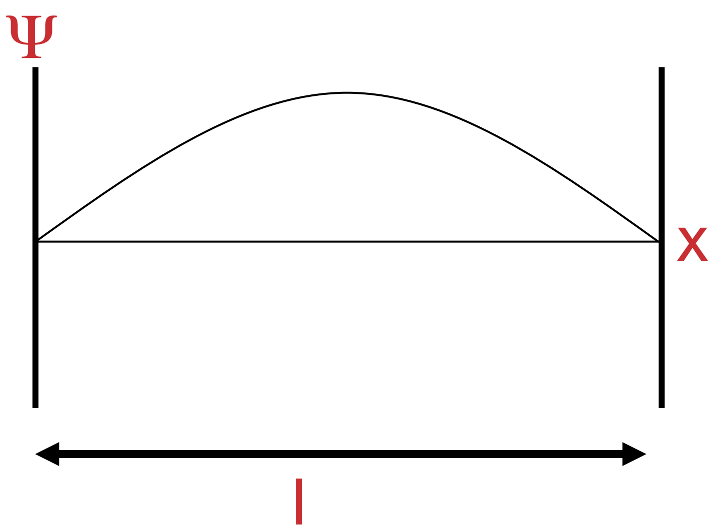
Quantum mechanics

Examples of wave equations



y = sinx

Propagating wave



$$\Psi = \sqrt{\frac{2}{l}} \sin \frac{\pi x}{l}$$

Standing wave

$l = 1/2\lambda$

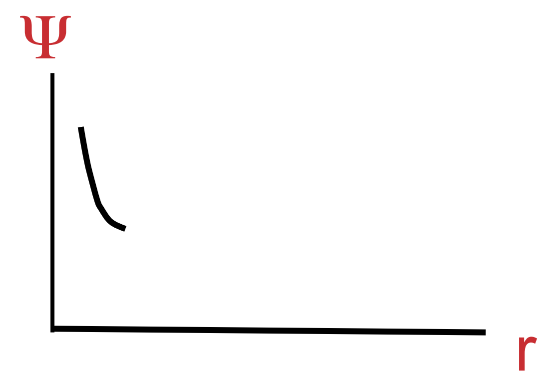
ν_O=frequency

nodes = 2

Quantum mechanics

- Using math we do NOT want to deal with, you can do the same thing for an electron in hydrogen:

$$\Psi = \frac{1}{\sqrt{\pi}} e^{-r}$$



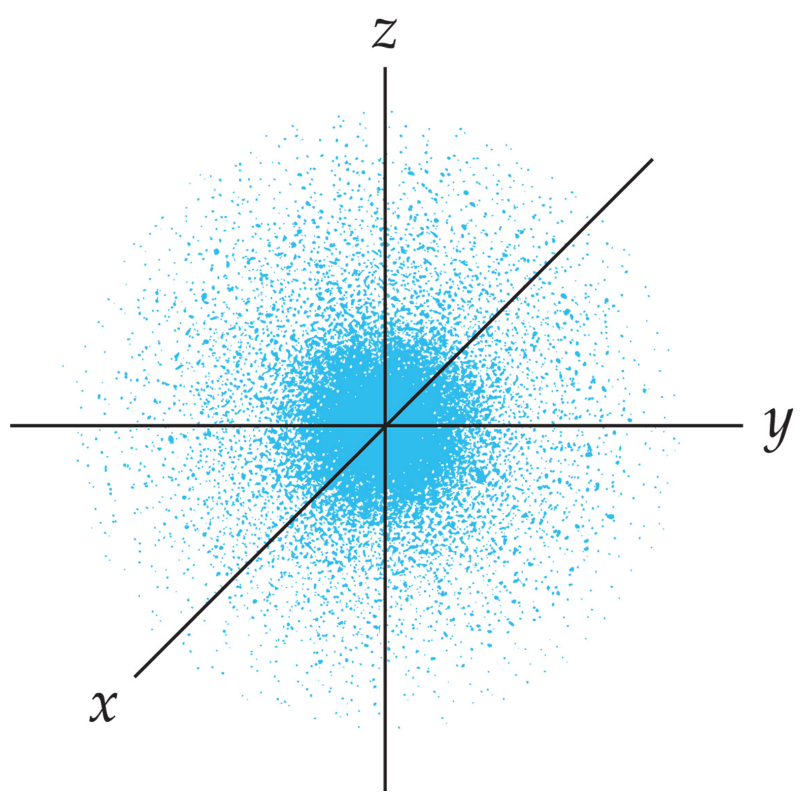
But what, physically is Ψ? What is waving?
Born (1926): Ψ^2= probability/volume of finding the electron.

Quantum Mechanics

Plot of Ψ^2 for hydrogen atom.

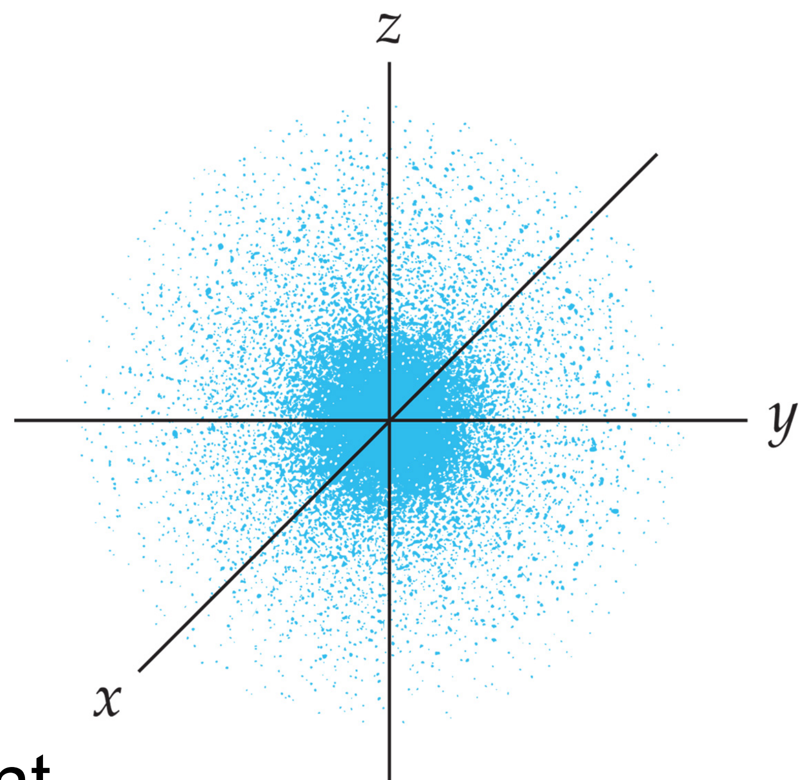
The closest thing we now have to a physical picture of an electron.

90% contour, will find electron in blue stuff 90% of the time.



Quantum Mechanics

- The wave equation is designated with a lower case Greek *psi* (ψ).
- The square of the wave equation, ψ^2, gives a probability density map of where an electron has a certain statistical likelihood of being at any given instant in time.



Quantum Numbers

- Solving the wave equation gives a set of wave functions, or orbitals, and their corresponding energies.
- Each orbital describes a spatial distribution of electron density.
- An orbital is described by a set of three quantum numbers (integers)
- Why three?

Quantum numbers

- 3 dimensions.

- Need three quantum numbers to define a given wavefunction.
- Another name for wavefunction: Orbital (because of Bohr).

Principal Quantum Number, *n*

- The principal quantum number, *n*, describes the **energy level** on which the orbital resides.
- Largest E difference is between E levels
- The values of *n* are integers ≥ 0.
- 1, 2, 3,...n.

Azimuthal Quantum Number, *l*

- defines **shape** of the orbital.
- Allowed values of *l* are integers ranging from 0 to $n - 1$.
- We use letter designations to communicate the different values of *l* and, therefore, the shapes and types of orbitals.

Azimuthal Quantum Number, *l*

$l = 0, 1 \ldots, n-1$

Value of *l*	0	1	2	3
Type of orbital	*s*	*p*	*d*	*f*

So each of these letters corresponds to a shape of orbital.

Magnetic Quantum Number, m_l

- Describes the **three-dimensional orientation** of the orbital.
- Values are integers ranging from -*l* to *l*:

$$-l \leq m_l \leq l.$$

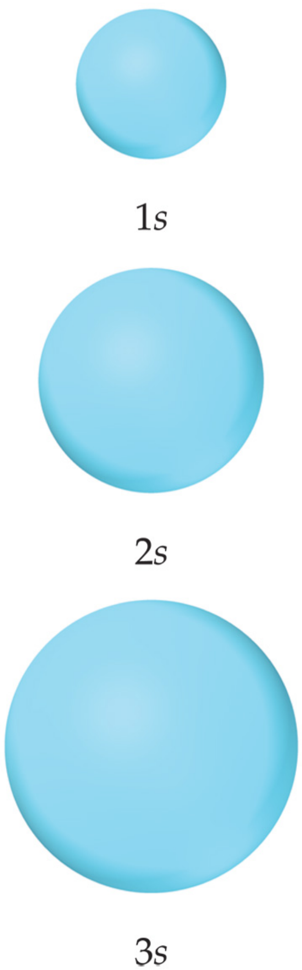
- Therefore, on any given energy level, there can be up to:
- 1 *s* *(l=0)* orbital (m_l=0),
- 3 *p* *(l=1)* orbitals, (m_l=-1,0,1)
- 5 *d* *(l=2)* orbitals, (m_l=-2,-1,0,1,2)
- 7 *f* *(l=3)* orbitals, (m_l=-3,-2,-1,0,1,2,3)

Magnetic Quantum Number, m_l

- Orbitals with the same value of *n* form a shell.
- Different orbital types within a shell are subshells (s, p, d, f).

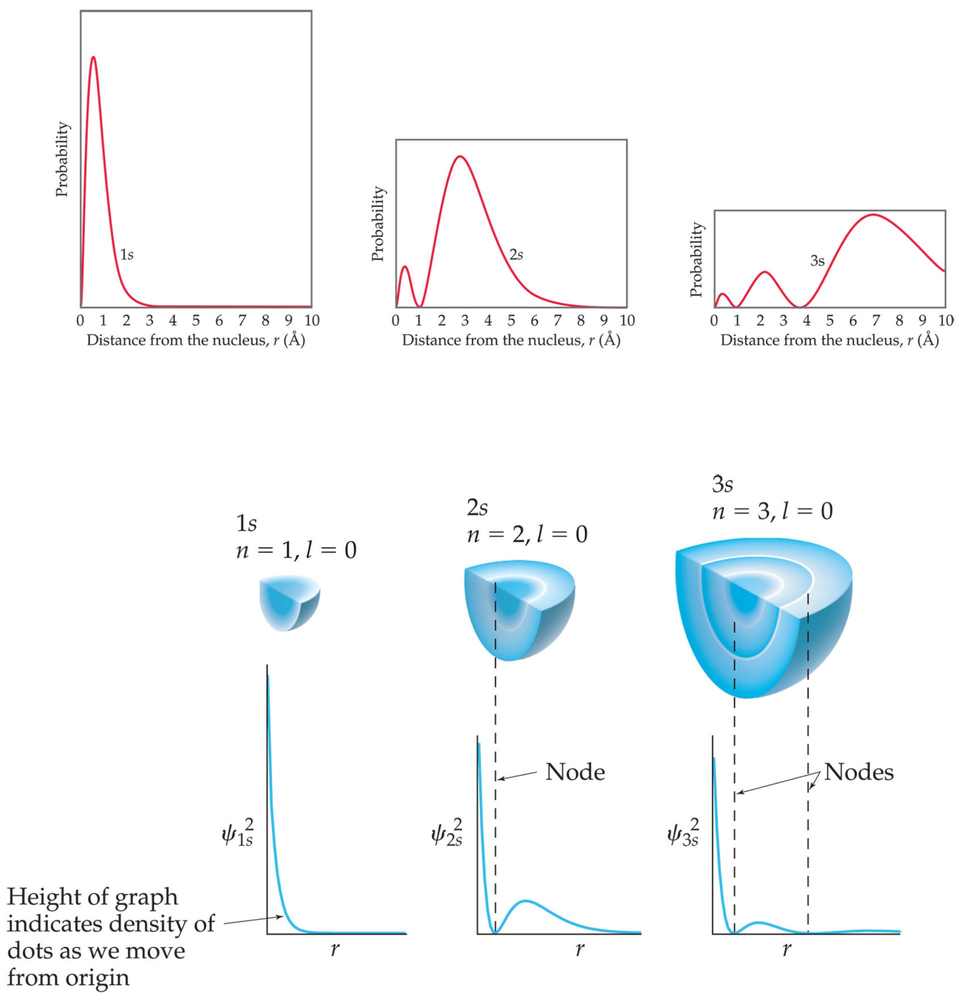
n	Possible Values of *l*	Subshell Designation	Possible Values of m_l	Number of Orbitals in Subshell	Total Number of Orbitals in Shell
1	0	1*s*	0	1	1
2	0	2*s*	0	1	
	1	2*p*	1, 0, −1	3	4
3	0	3*s*	0	1	
	1	3*p*	1, 0, −1	3	
	2	3*d*	2, 1, 0, −1, −2	5	9
4	0	4*s*	0	1	
	1	4*p*	1, 0, −1	3	
	2	4*d*	2, 1, 0, −1, −2	5	
	3	4*f*	3, 2, 1, 0, −1, −2, −3	7	16

s Orbitals



- Value of *l* = 0.
- Spherical in shape.
- Radius of sphere increases with increasing value of *n*.

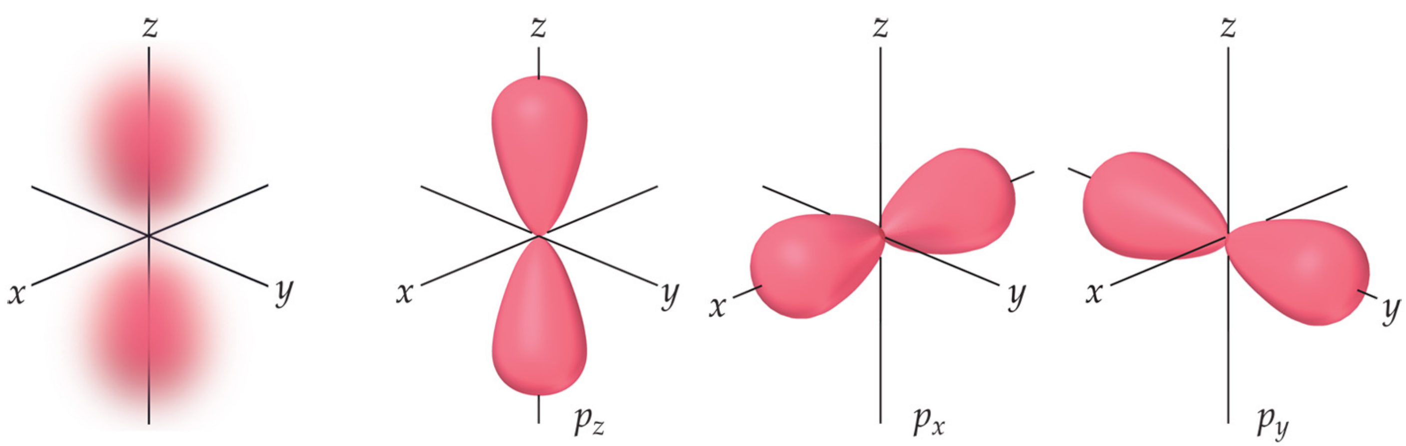
s Orbitals



s orbitals possess *n*−1 nodes, or regions where there is 0 probability of finding an electron.

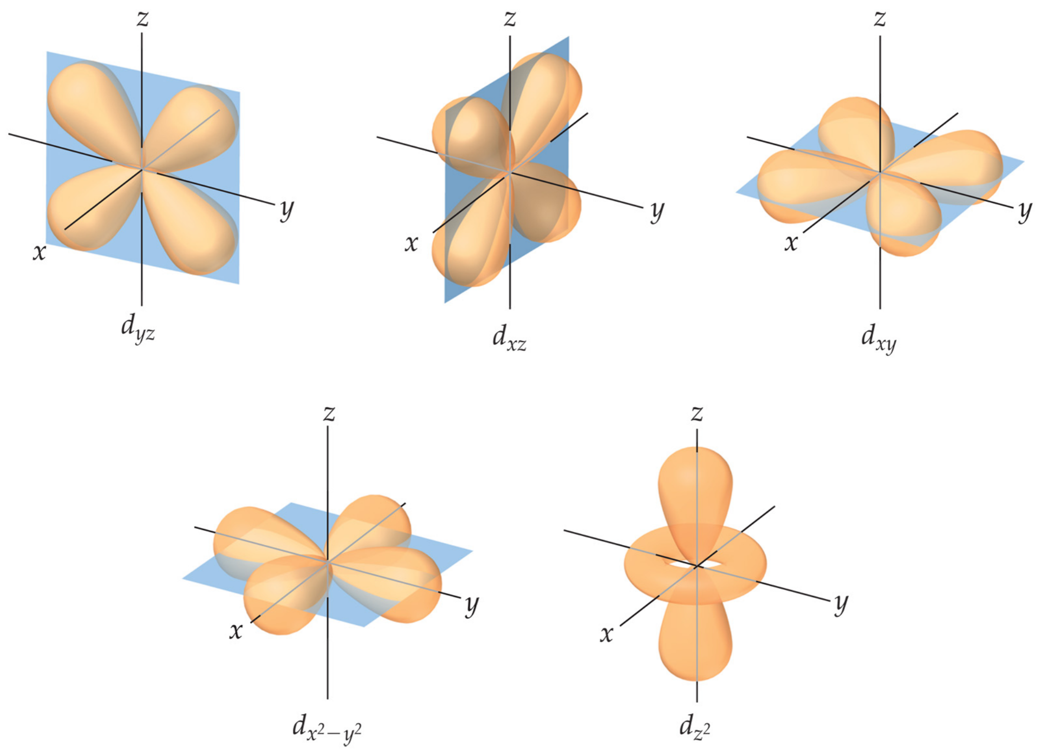
p Orbitals

- Value of *l* = 1.
- Have two lobes with a nodal plane between them.



Note: always 3 p orbitals for a given n

d Orbitals



- Value of *l* is 2.
- 2 nodal planes
- Four of the five orbitals have 4 lobes; the other resembles a *p* orbital with a doughnut around the center.

Note: always 5 d orbitals for a given n.

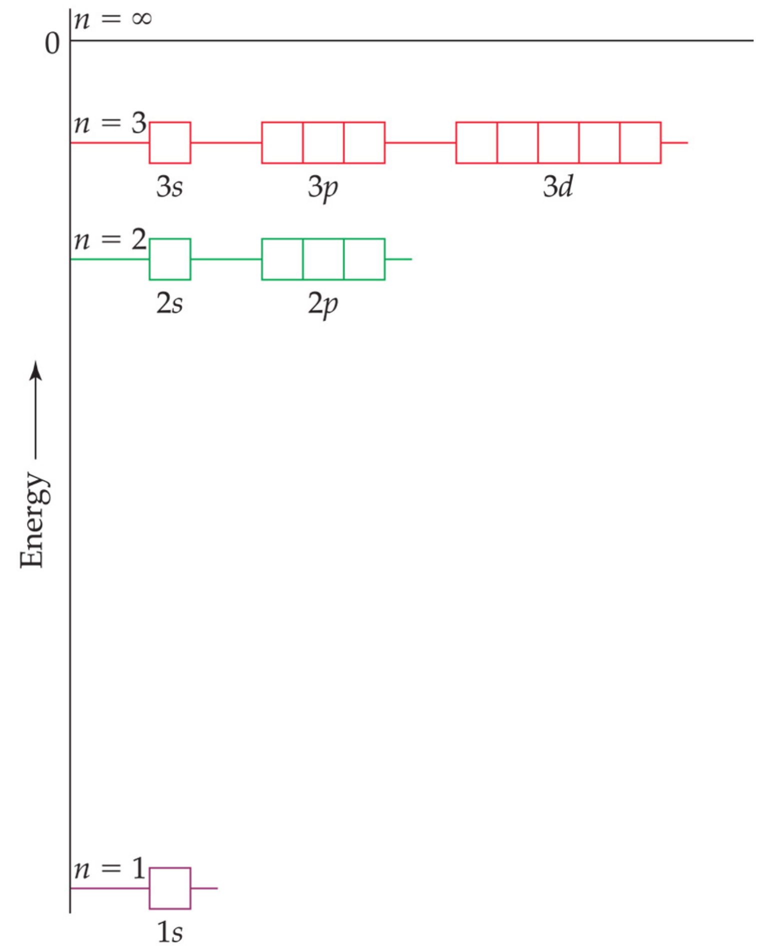
Orbitals and nodes

Orbital	Symmetry	Node geometry	Spherical nodes/shell*	Orbitals/set
s	spherical	spherical	n-1	1
p	cylindrical around x, y, or z axis	1 planar remainder spherical	n - 1	3
d	complex	2 planar surfaces diagonal to Cartesian axis; remainder spherical	n - 2	5
f	complex	complex	n - 3	7

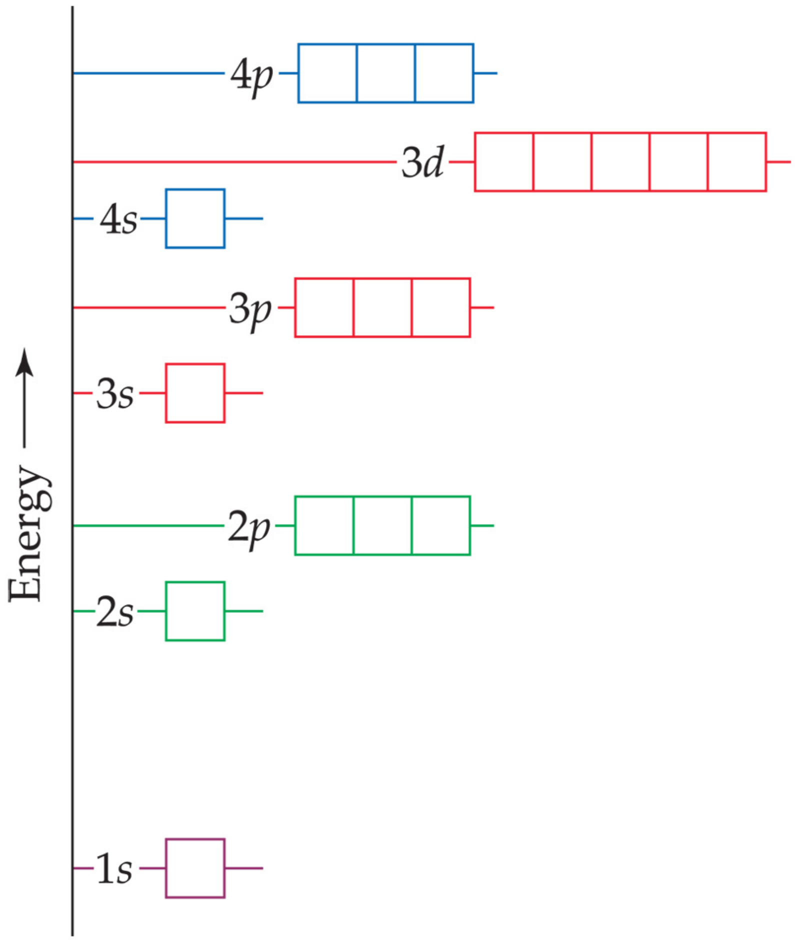
* n = the shell, with n = 1 the ground state or lowest possible energy shell. Thus n may have integral values from 1 - infinity.

Energies of Orbitals

- For a one-electron hydrogen atom, orbitals on the same energy level have the same energy.
- That is, they are degenerate.

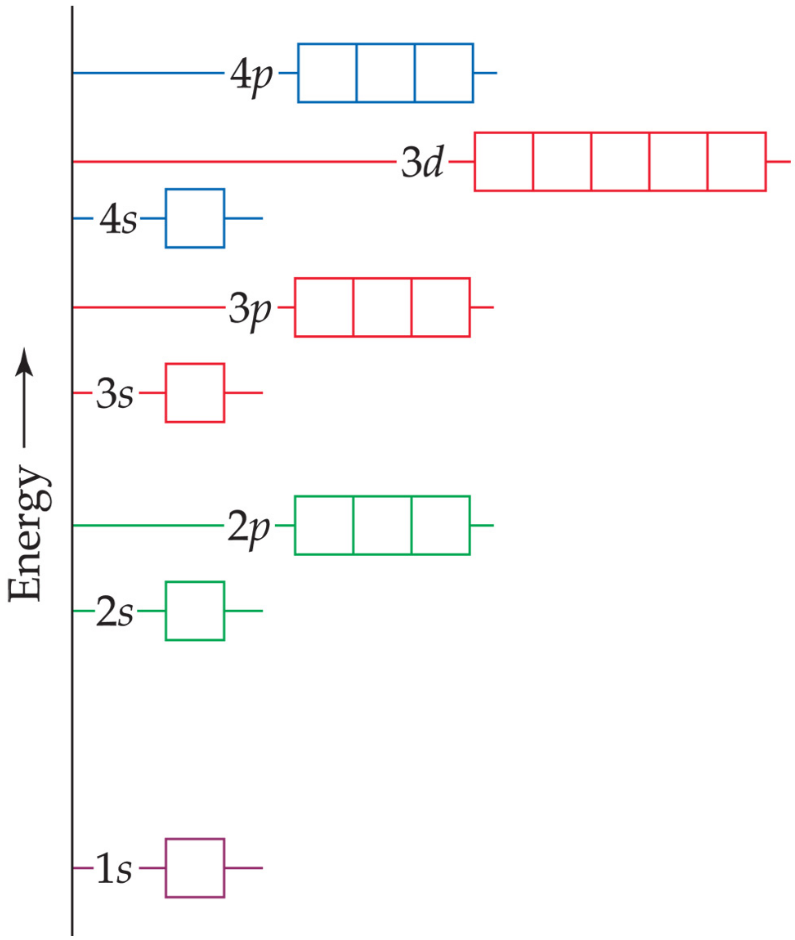


Energies of Orbitals



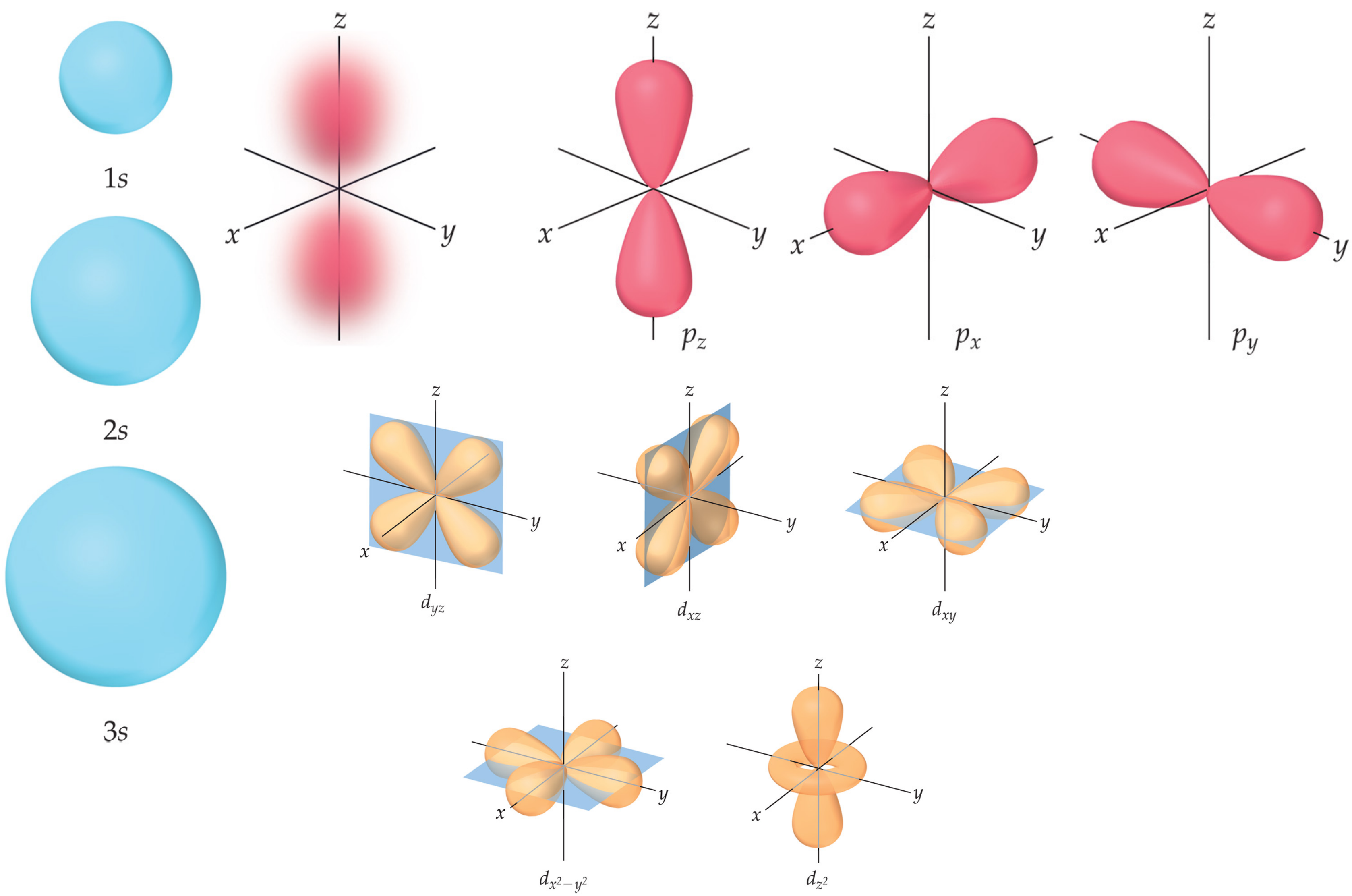
- As the number of electrons increases, though, so does the repulsion between them.
- Therefore, in many-electron atoms, orbitals on the same energy level are no longer degenerate.

Energies of Orbitals



- For a given energy level (n):
- Energy:
- s<p<d<f
- s lowest energy, where electrons go first
- Next p
- Then d

Why?



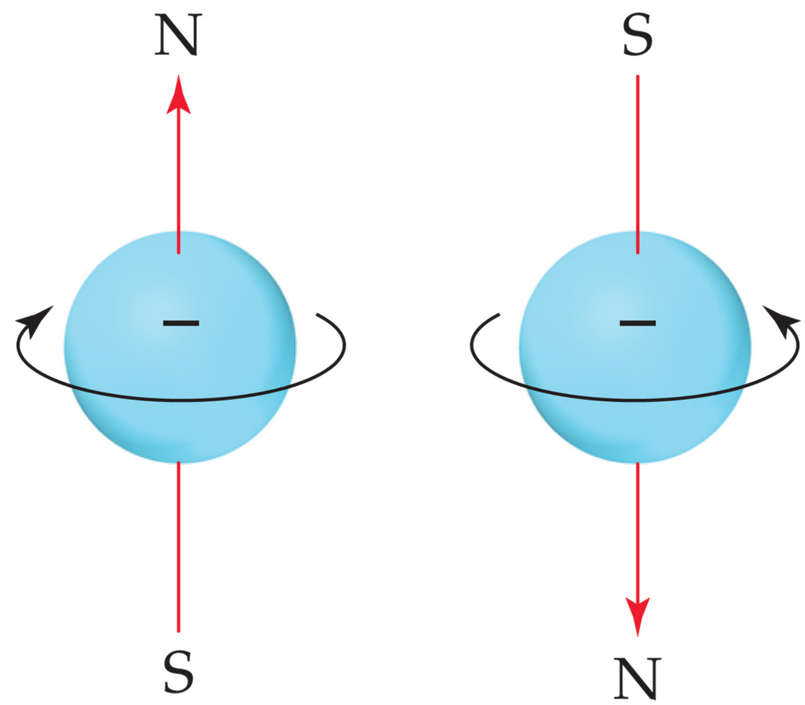
The closer to the nucleus, the lower the energy

The problem with quantum mechanics

- It's not hard to solve equations for the various wavefunctions if they are all alone (like H)
- The problem is what happens in the presence of other electrons
- The electron interaction problem
- Electron interaction so complex, exact solutions are only possible for H!
- Electron probabilities overlap a lot, must interact a lot, repulsion keeps them from ever "touching"

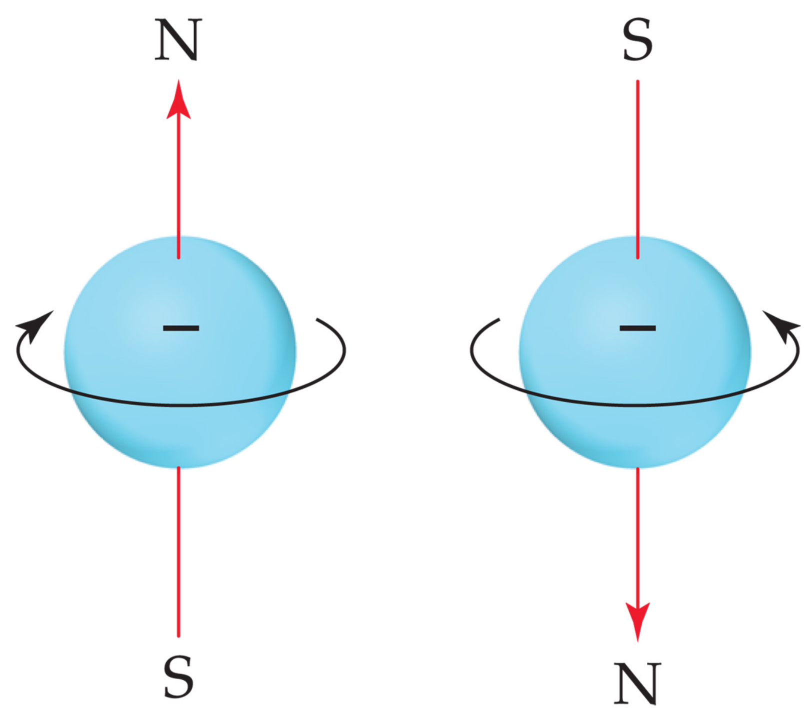
Spin Quantum Number, m_s

- A fourth dimension required. Why?



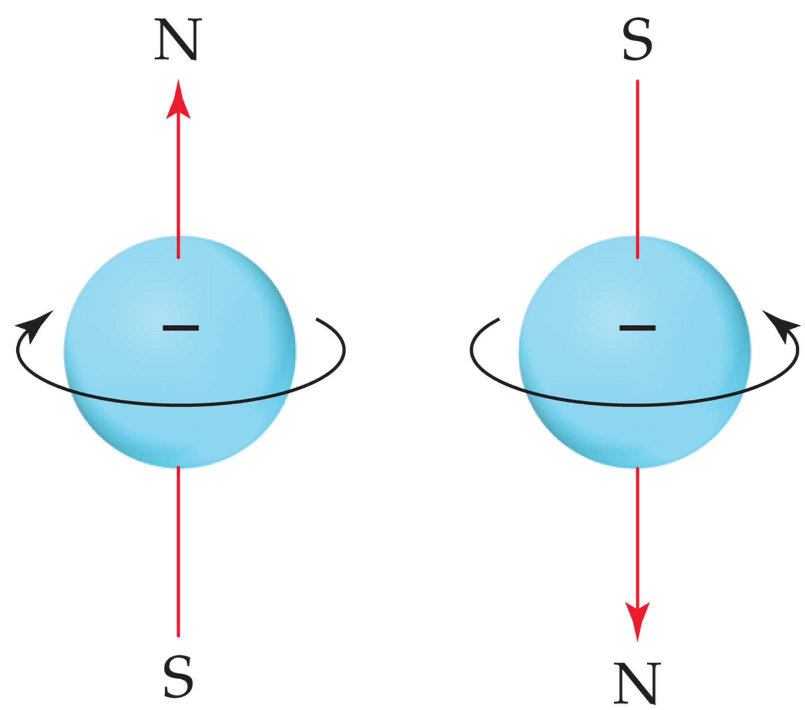
Spin Quantum Number, m_s

- A fourth dimension required. Why?
- Time. Adding time changes E
- Another integer (quantum number) needed.
- Time *dependent Schroedinger equation.*

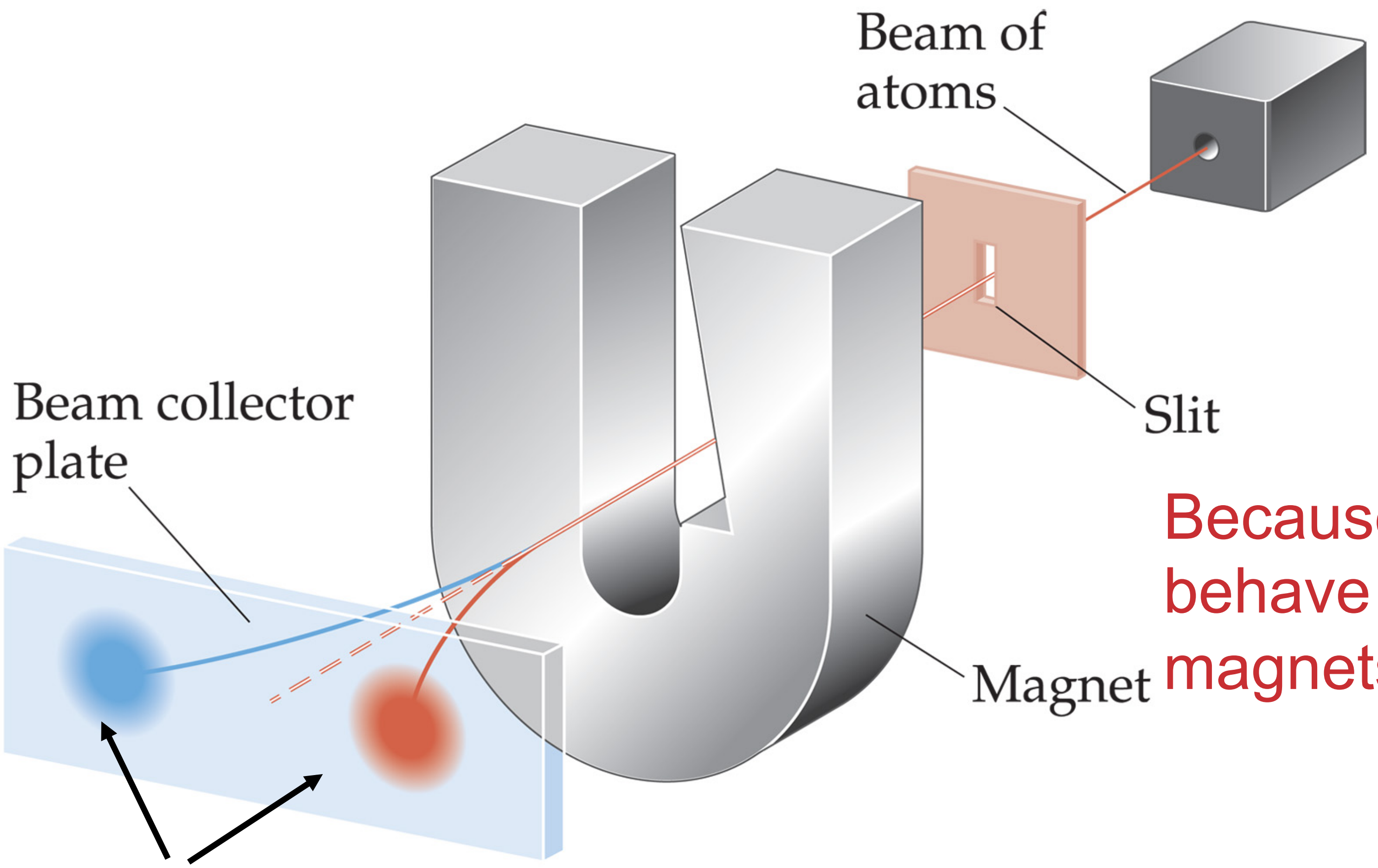


Spin Quantum Number, m_s

- This leads to a fourth quantum number, the spin quantum number m_s.
- The spin quantum number has only 2 values **+1/2 and -1/2**
- **Describes magnetic field vector of electron**



Why do we call it "spin"

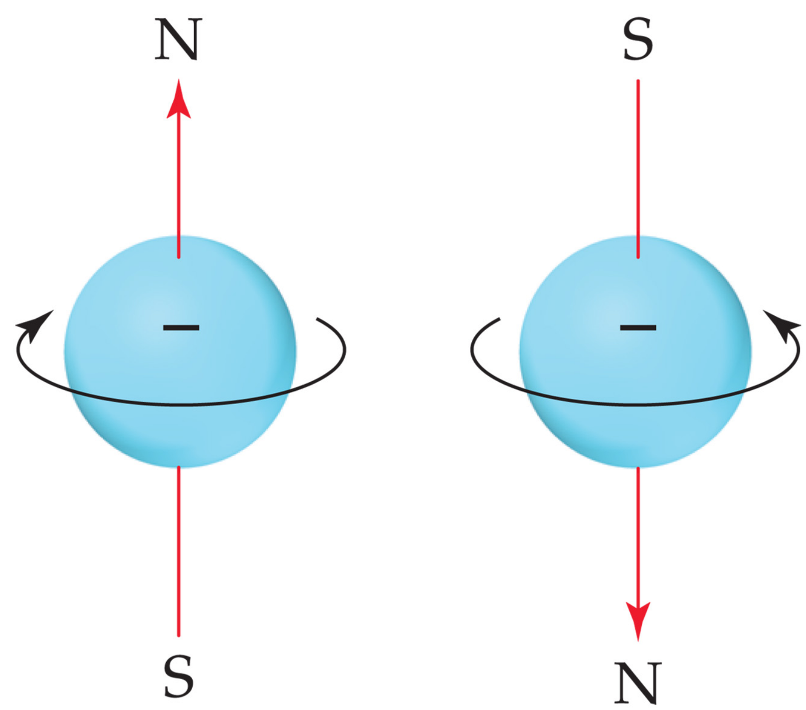


Because electrons behave like little magnets

Note: apparently only two values for the magnetic field

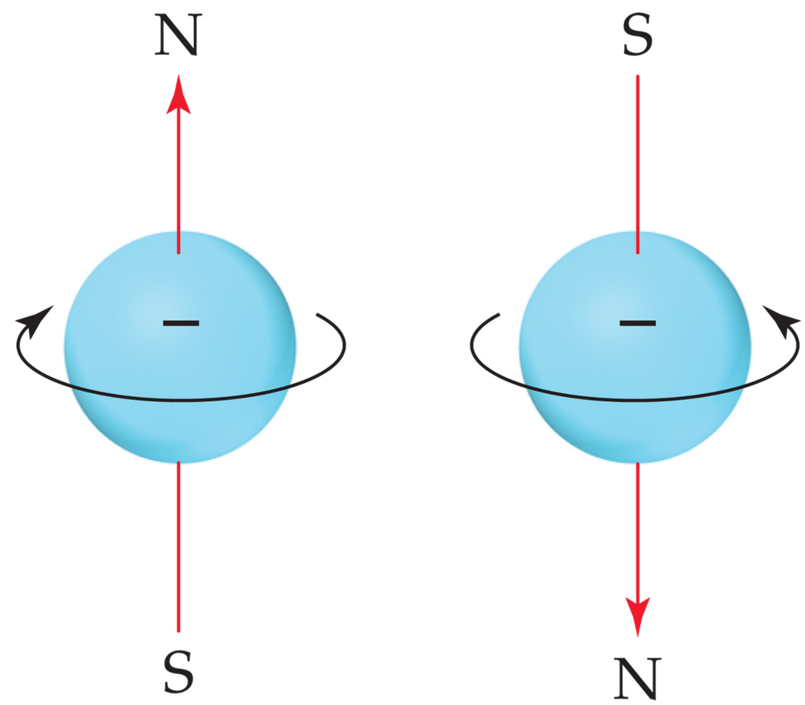
Why do we call it “spin”

- And charges that spin produce magnetic fields



Pauli Exclusion Principle

- No two electrons in the same atom can have exactly the same energy.
- For example, no two electrons in the same atom can have identical sets of quantum numbers.



Electron Configurations Every electron has a name

- Name of each electron unique
- Name consists of four numbers:
- n, l, m_l, m_s
- Example:
- Mr. George Herbert Walker Bush
- We must learn to name our electrons
- Unlike people, there is a lot in the "name" of an electron.

$4p^5$

Electron Configurations

$4p^5$

- Distribution of all electrons in an atom
- Consist of
  - Number denoting the energy level

Electron Configurations

$4p^5$

- Distribution of all electrons in an atom
- Consist of
  - Number denoting the energy level
  - Letter denoting the type of orbital

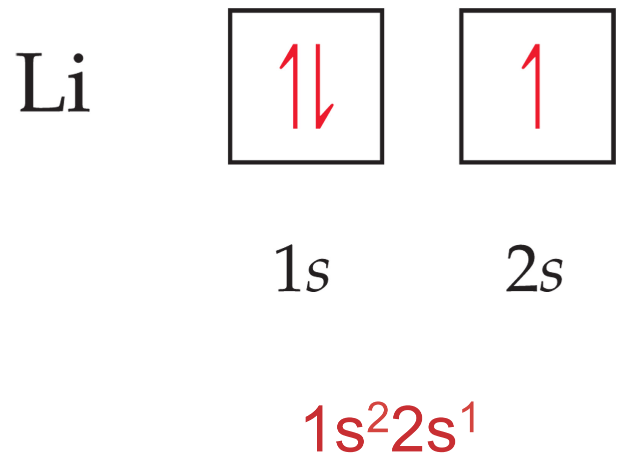
Electron Configurations

$4p^5$

- Distribution of all electrons in an atom.
- Consist of
  - Number denoting the energy level.
  - Letter denoting the type of orbital.
  - Superscript denoting the number of electrons in those orbitals.

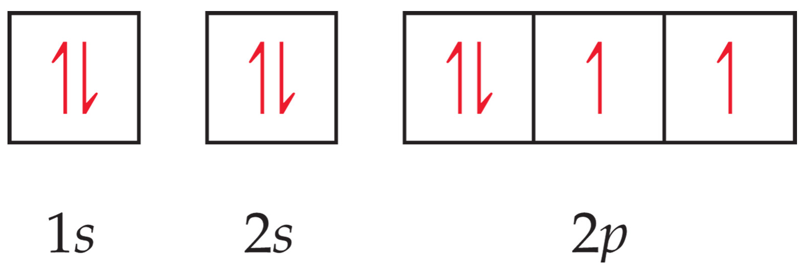
Orbital Diagrams

- Each box represents one orbital.
- Half-arrows represent the electrons.
- The direction of the arrow represents the spin of the electron.

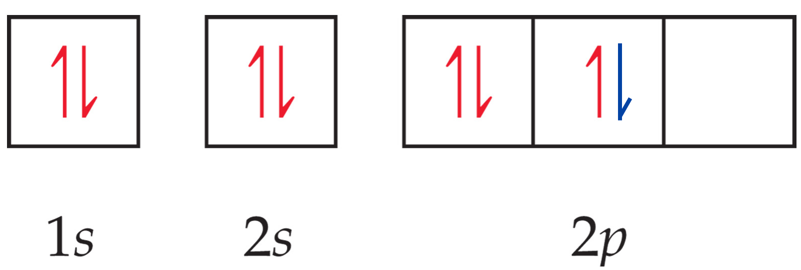


Hund's Rule

(of maximum multiplicity)



NOT:



"For degenerate orbitals, the lowest energy is attained when the number of electrons with the same spin is maximized."

Electron configurations

TABLE 6.3 Electron Configurations of Several Lighter Elements

Element	Total Electrons	Orbital Diagram 1s	2s	2p			3s	Electron Configuration
Li	3	↑↓	↑					$1s^22s^1$
Be	4	↑↓	↑↓					$1s^22s^2$
B	5	↑↓	↑↓	↑				$1s^22s^22p^1$
C	6	↑↓	↑↓	↑	↑			$1s^22s^22p^2$
N	7	↑↓	↑↓	↑	↑	↑		$1s^22s^22p^3$
Ne	10	↑↓	↑↓	↑↓	↑↓	↑↓		$1s^22s^22p^6$
Na	11	↑↓	↑↓	↑↓	↑↓	↑↓	↑	$1s^22s^22p^63s^1$

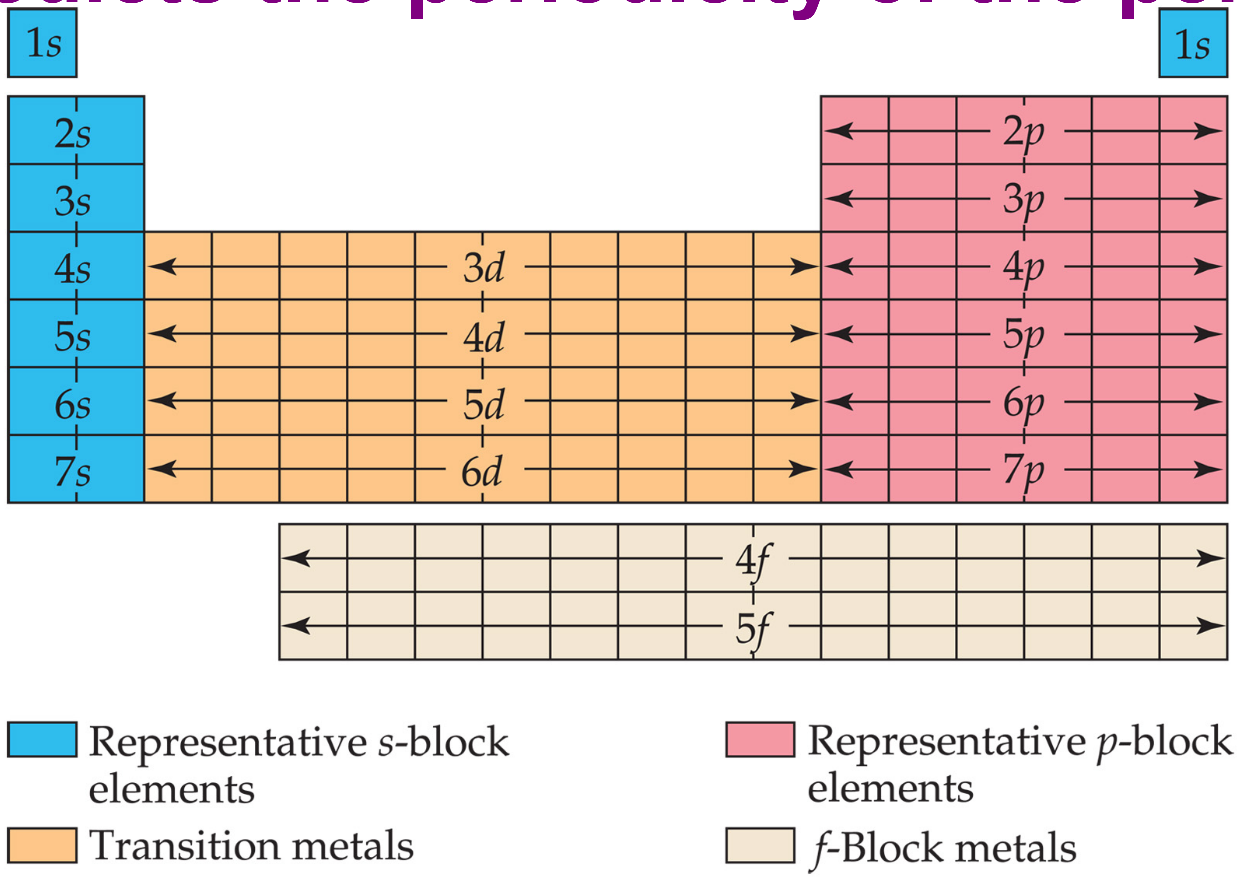
Why do we accept this wacko stuff?

- It must explain all the data
- It should predict things

- Q.M. is consistent with all our data (photoelectric effect, emission spectra of elements, dual wave/ particle weirdness, etc.
- One prediction: elements with similar electron configuration should have similar chemical properties

Why do we accept this wacko stuff?

It predicts the periodicity of the periodic table!!



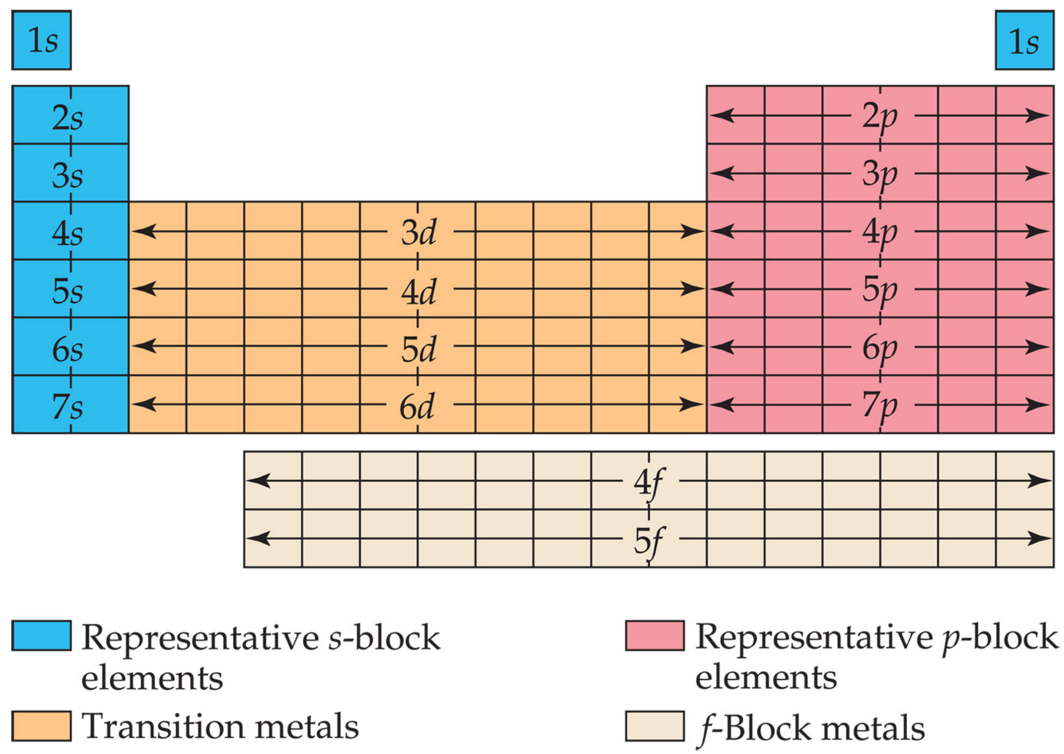
- **We fill orbitals in increasing order of energy.**
- **Different blocks on the periodic table, then correspond to different types of orbitals.**

Why do we accept this wacko stuff?

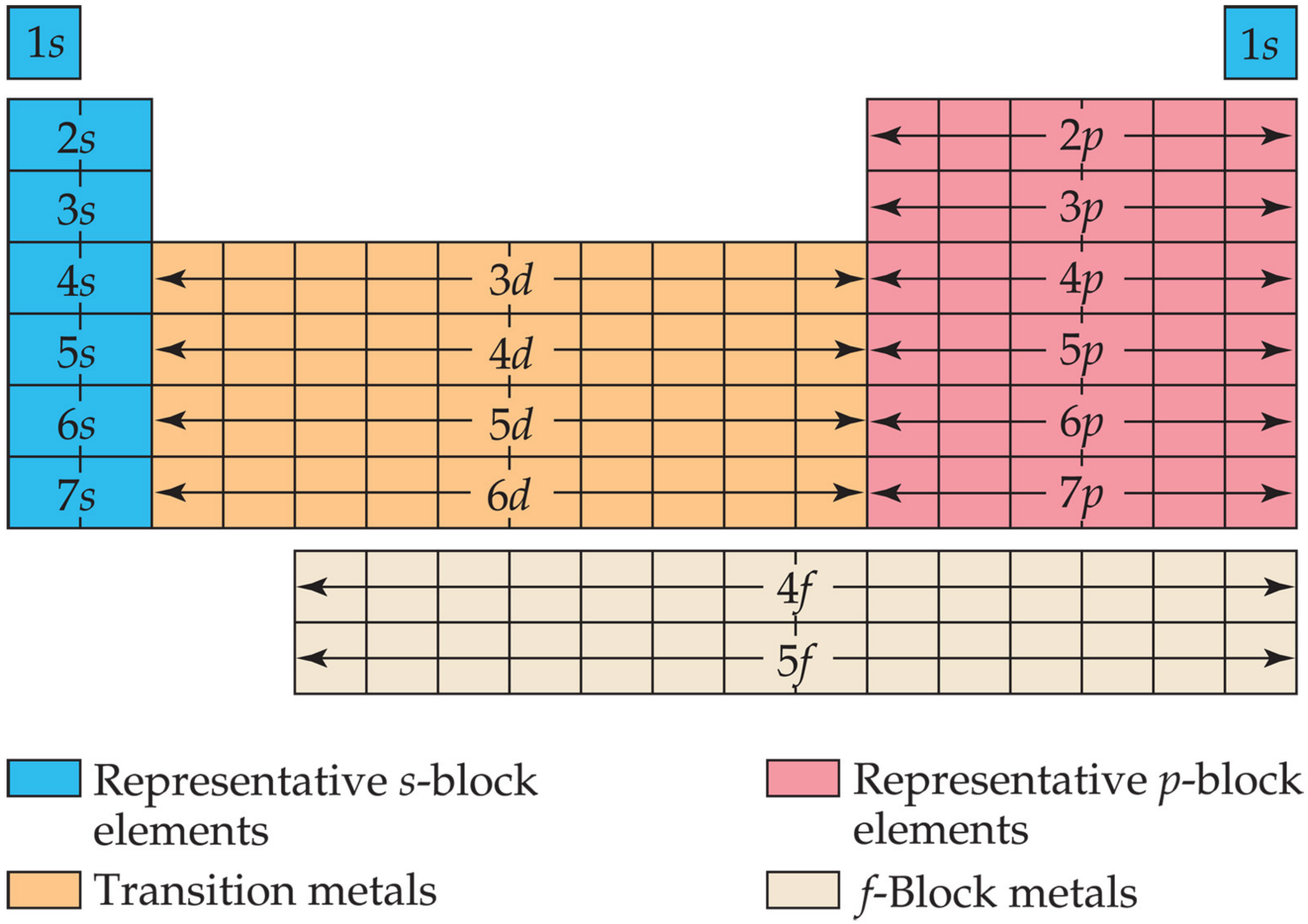
It predicts the periodicity of the periodic table!!

- **Remember: The periodic table was arranged the way it was based on chemical properties.**

- **Totally empirical, until now. Based only on observation.**



Periodic Table



- **Periodic table tells you about the last electron that went in!!!**
- **Periodic table also makes it easy to do electron configurations.**

Short cut for writing electron configurations

TABLE 6.4 Electron Configurations of the Group 2A and 3A Elements

Group 2A	
Be	$[He]2s^2$
Mg	$[Ne]3s^2$
Ca	$[Ar]4s^2$
Sr	$[Kr]5s^2$
Ba	$[Xe]6s^2$
Ra	$[Rn]7s^2$
Group 3A	
B	$[He]2s^22p^1$
Al	$[Ne]3s^23p^1$
Ga	$[Ar]3d^{10}4s^24p^1$
In	$[Kr]4d^{10}5s^25p^1$
Tl	$[Xe]4f^{14}5d^{10}6s^26p^1$

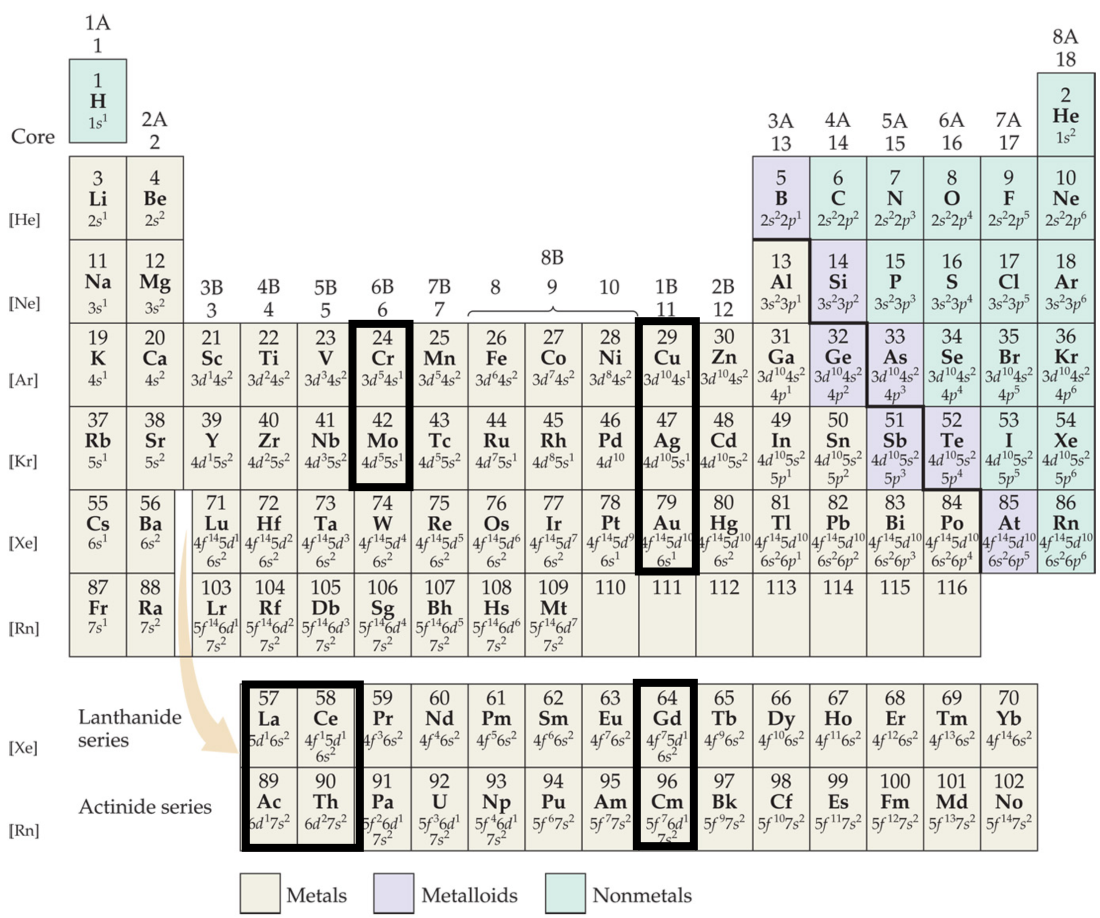
Electron configurations of the elements

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

Core	Series														
[Xe]	Lanthanide series	57 **La** $5d^1 6s^2$	58 **Ce** $4f^1 5d^1 6s^2$	59 **Pr** $4f^3 6s^2$	60 **Nd** $4f^4 6s^2$	61 **Pm** $4f^5 6s^2$	62 **Sm** $4f^6 6s^2$	63 **Eu** $4f^7 6s^2$	64 **Gd** $4f^7 5d^1 6s^2$	65 **Tb** $4f^9 6s^2$	66 **Dy** $4f^{10} 6s^2$	67 **Ho** $4f^{11} 6s^2$	68 **Er** $4f^{12} 6s^2$	69 **Tm** $4f^{13} 6s^2$	70 **Yb** $4f^{14} 6s^2$
[Rn]	Actinide series	89 **Ac** $6d^1 7s^2$	90 **Th** $6d^2 7s^2$	91 **Pa** $5f^2 6d^1 7s^2$	92 **U** $5f^3 6d^1 7s^2$	93 **Np** $5f^4 6d^1 7s^2$	94 **Pu** $5f^6 7s^2$	95 **Am** $5f^7 7s^2$	96 **Cm** $5f^7 6d^1 7s^2$	97 **Bk** $5f^9 7s^2$	98 **Cf** $5f^{10} 7s^2$	99 **Es** $5f^{11} 7s^2$	100 **Fm** $5f^{12} 7s^2$	101 **Md** $5f^{13} 7s^2$	102 **No** $5f^{14} 7s^2$

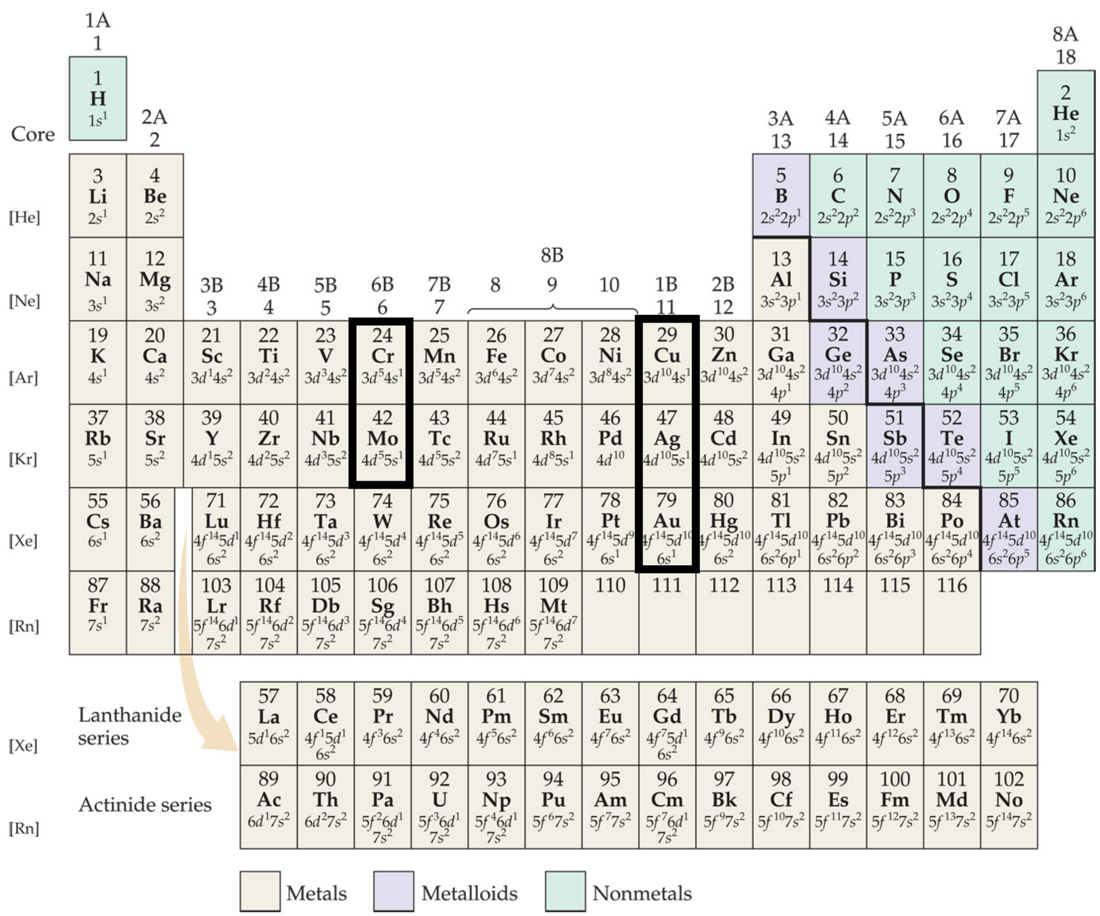
Metals | Metalloids | Nonmetals

Some Anomalies



Some irregularities occur when there are enough electrons to half-fill *s* and *d* orbitals on a given row.

Some Anomalies



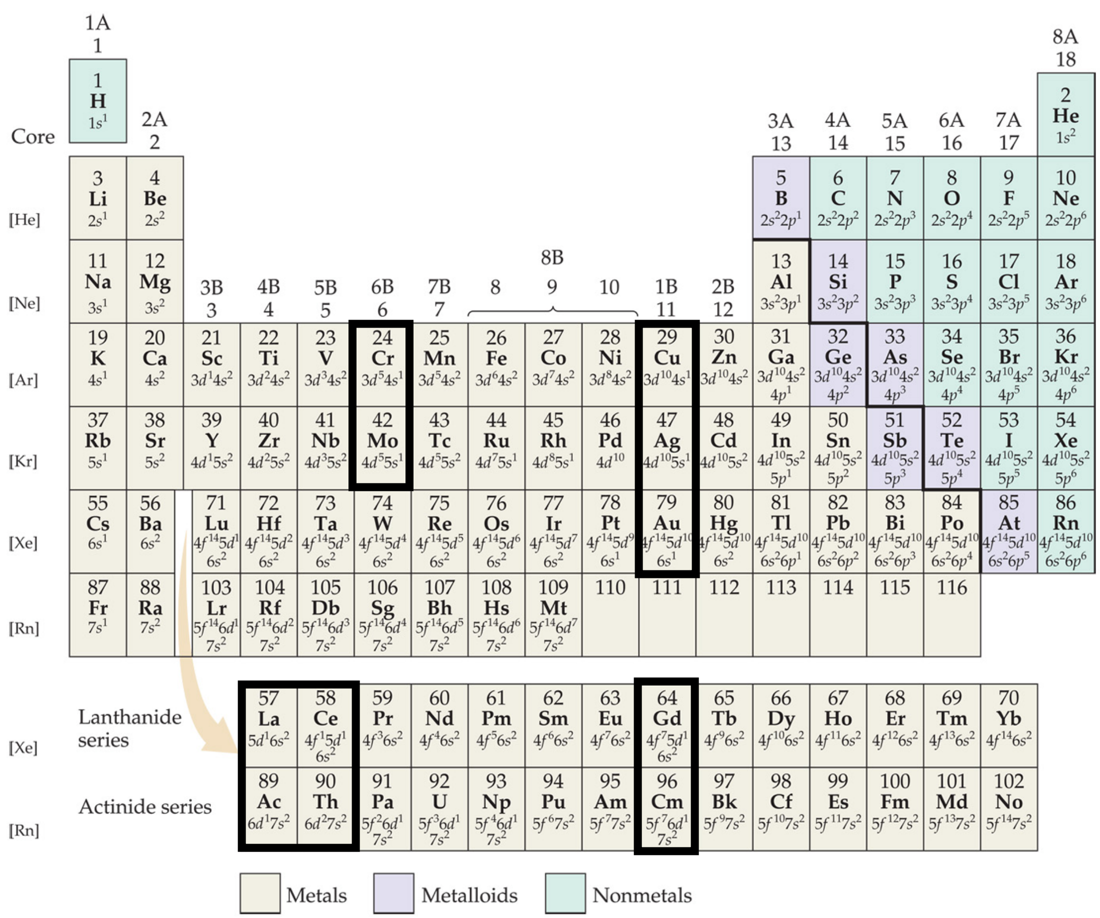
For instance, the electron configuration for Chromium, is

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

rather than the expected

[Ar] $4s^2\ 3d^4$.

Some Anomalies



- This occurs because the 4*s* and 3*d* orbitals are very close in energy.
- These anomalies occur in *f*-block atoms, as well.