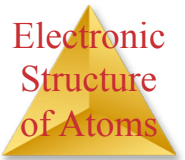


# Chapter 6

# Electronic Structure of Atoms



# Historical perspective

## Atomic spectra

- Bunsen, Kirchhoff, 1860
  - 1st spectroscope
  - 1st line spectrum
- Lockyer, 1868
  - He in solar system
- Balmer, 1885
  - H line spectrum

## Quantum theory

- Plank, 1900
  - Black body radiation
- Einstein, 1905
  - Photoelectric effect
- Bohr, 1913
  - Applied to atom structure

## Atomic structure

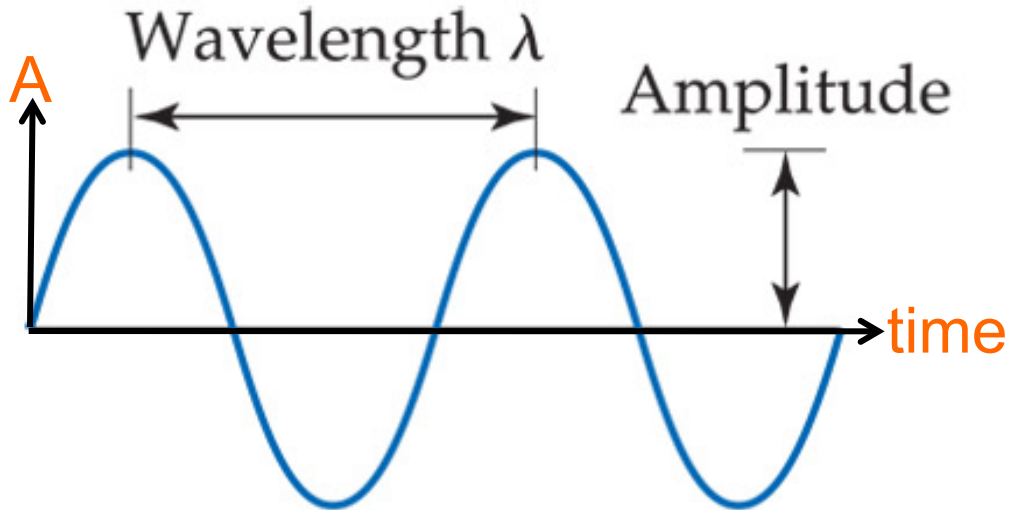
- Dalton, 1803
  - atomic nature
- Faraday, 1834
  - Electricity & Mag.
- Thompson, 1807
  - electrons e/m
- Millikan, 1911
  - oil drop
- Rutherford, 1911
  - gold foil/nucleus



# *Electro-magnetic radiation (light)*

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

# Waves

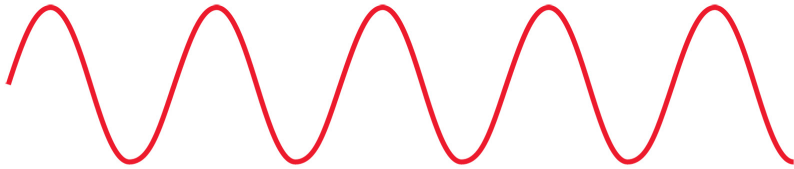


- Wave: some sort of periodic function
  - something that periodically changes vs. time.
- wavelength ( $\lambda$ ): distance between equivalent points
- Amplitude: “height” of wave, maximum displacement of periodic function.

# Waves

**Higher frequency**

**shorter wavelength**

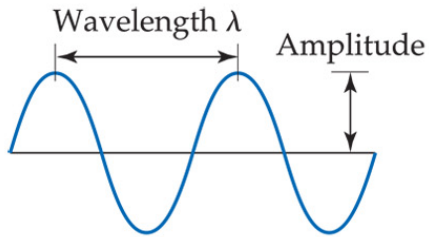


**lower frequency**

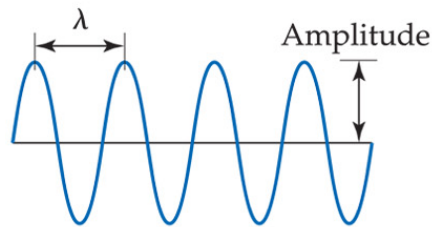
**longer wavelength**

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

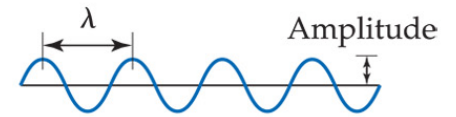
# Waves



(a) Two complete cycles of wavelength  $\lambda$



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



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

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$v = \text{wavelength} \times \text{frequency}$   
meters  $\times$  (1/sec) = m/sec

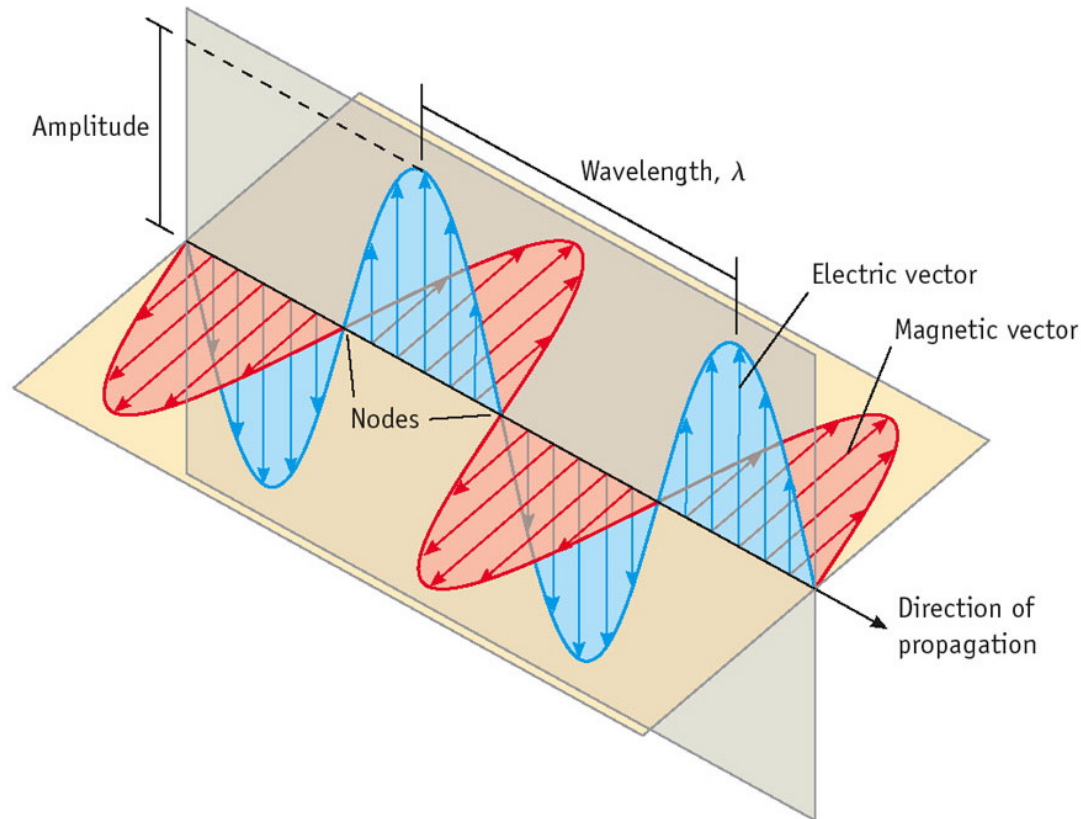
$$v = \lambda \nu$$

# Waves

## Major question:

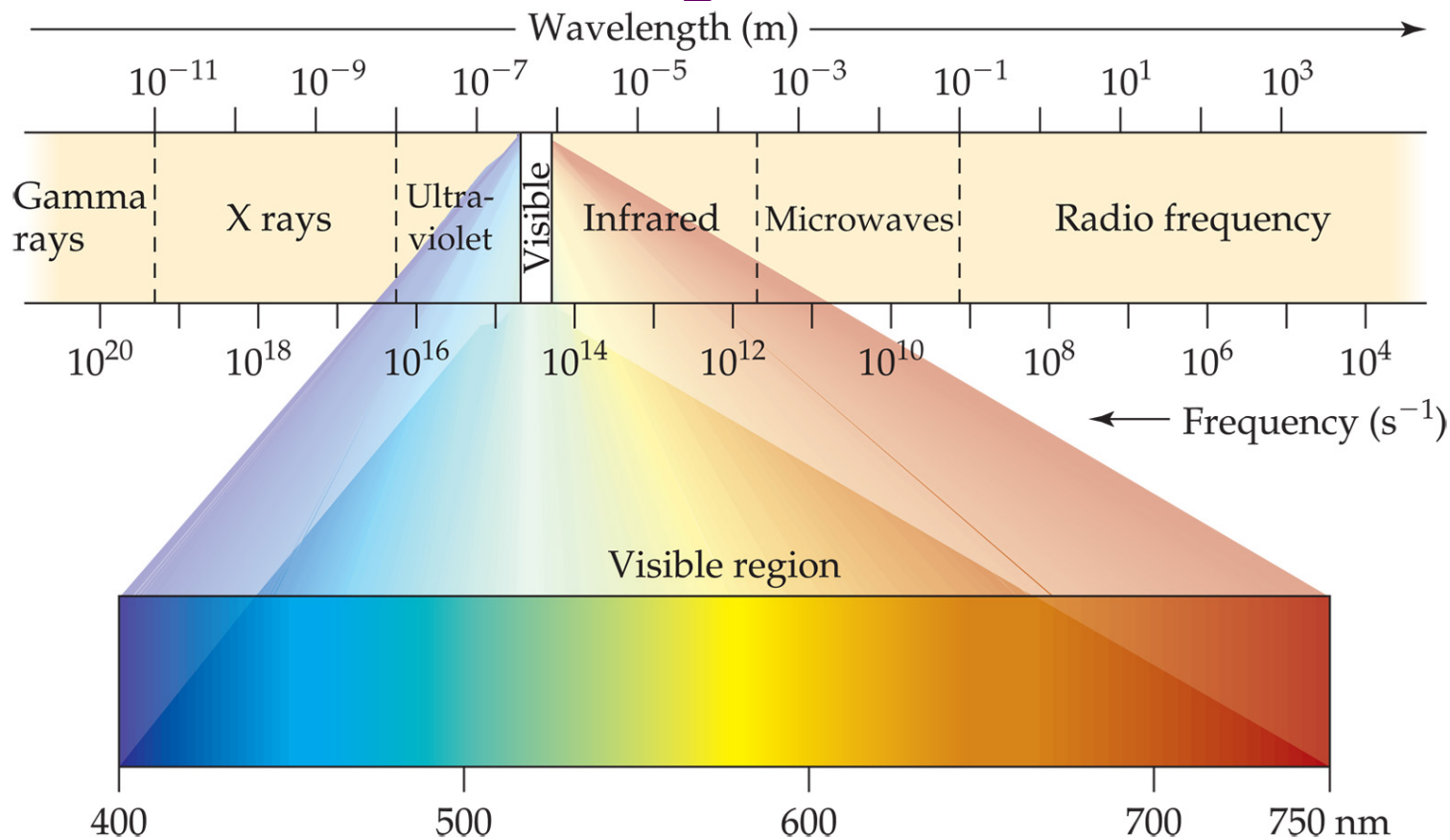
- What is waving?
- water wave:
  - water height (pressure)
- Sound wave:
  - air pressure
- Light?

# 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 ( $c$ ),  $3.00 \times 10^8$  m/s (in a vacuum).
- Therefore:  $c = \lambda \nu$

# Speed of light in other materials

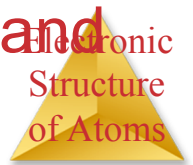
Index of refraction is:

$$n = c/v$$

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 19<sup>th</sup> century physics*

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

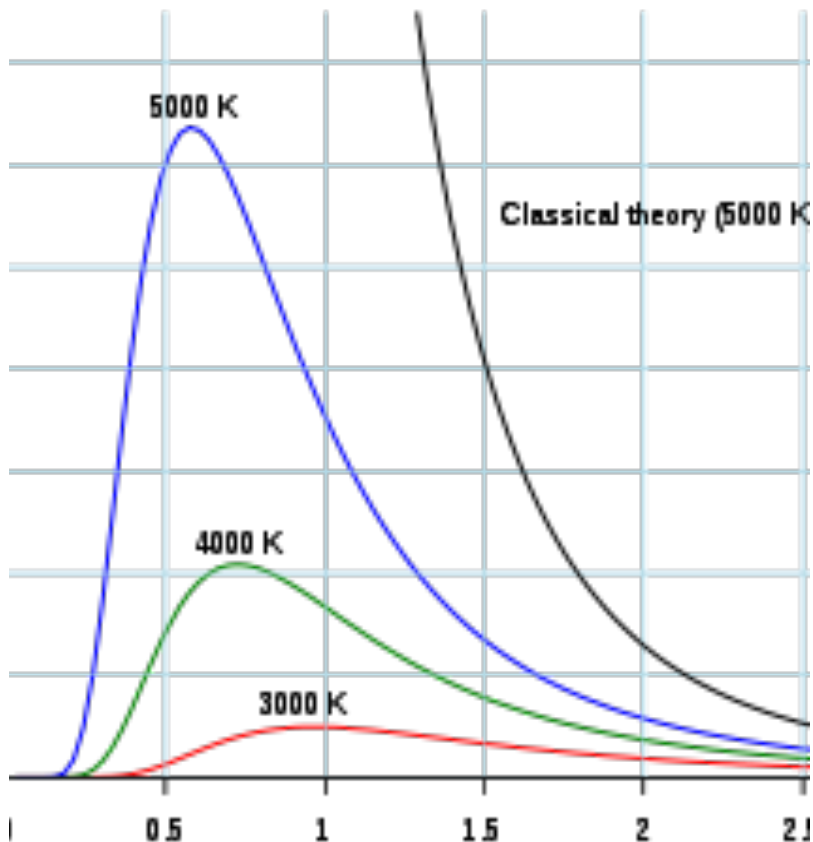
# The three mysteries of 19th century physics

## Mystery #1: Blackbody radiation



- Why does metal glow when heated?
- K.E. of electrons
- What light is given off?

# 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

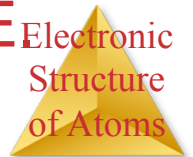


- Higher  $T$  leads to shorter wavelength of light
- More K.E., more  $E$
- Must be relationship between  $E$  and wavelength
- Plank concluded that energy is *quantized*. It comes in packets (like fruit snacks) and is proportional to frequency:

$$E = h\nu$$

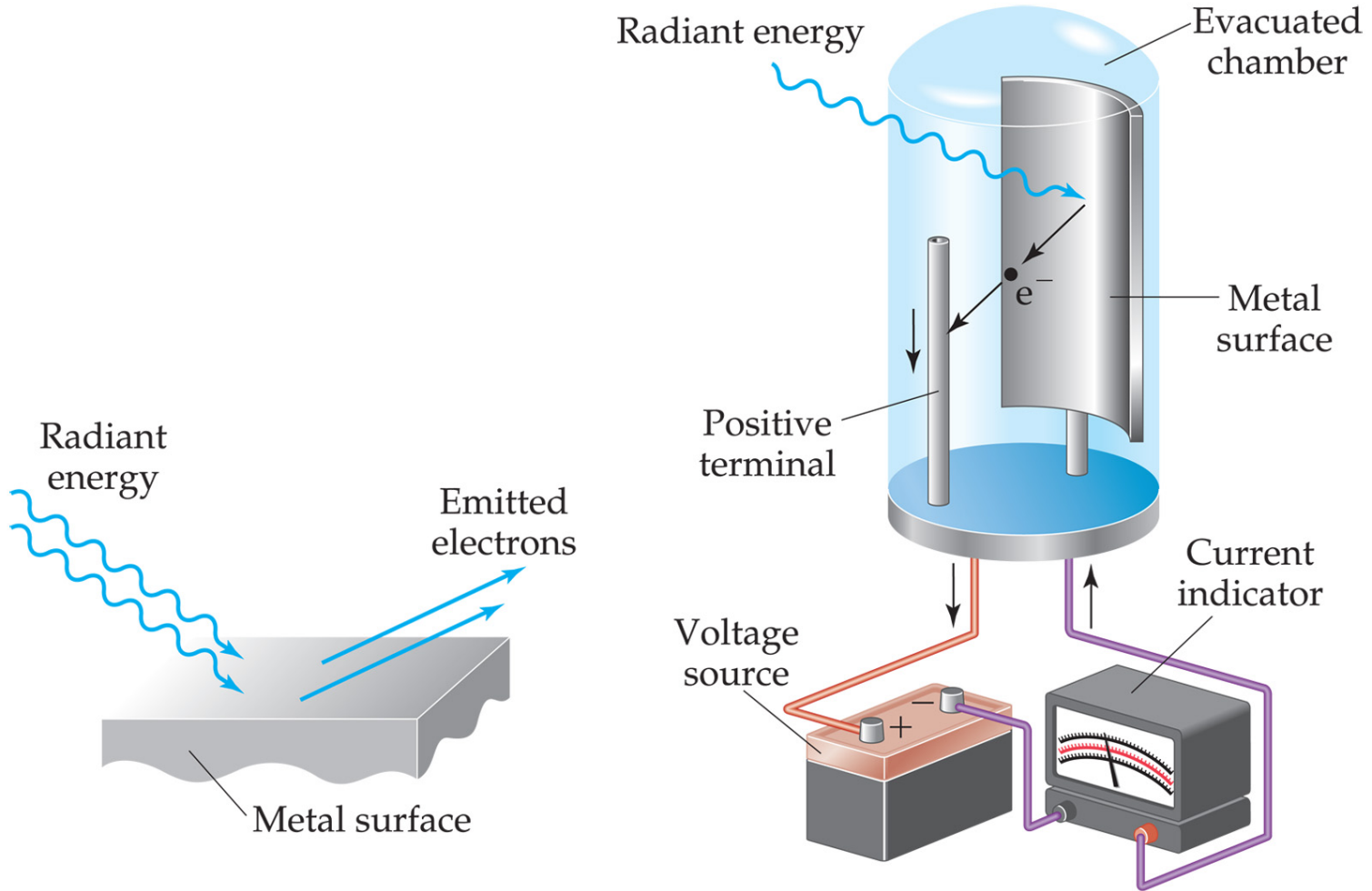
where  $h$  is Planck's constant,  $6.63 \times 10^{-34}$  J-s.

The minimum packet of  $E$



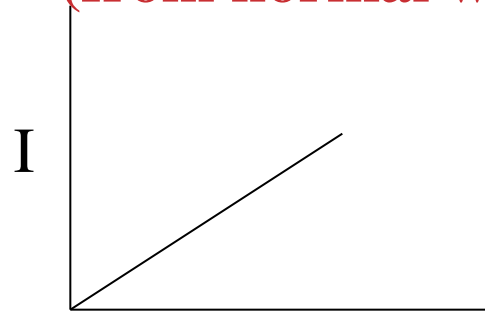
**What did Einstein get the Nobel Prize for?**

# Mystery #2: The Photo-electric effect



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

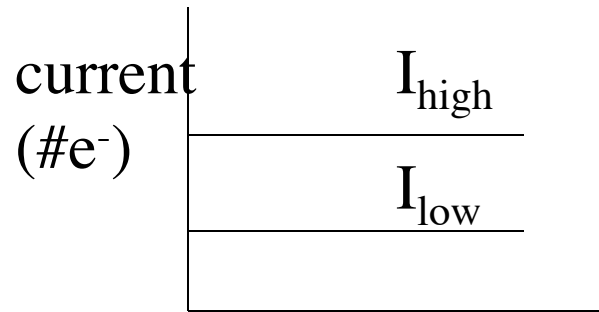
What might you expect  
(from normal waves)



K. E. e<sup>-</sup>

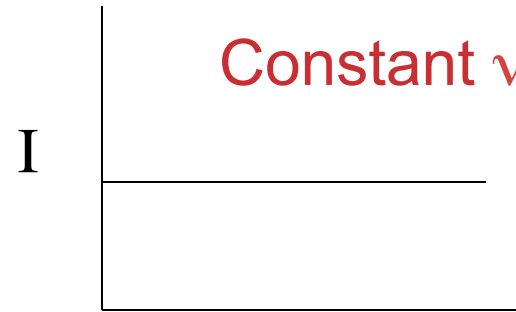


ν

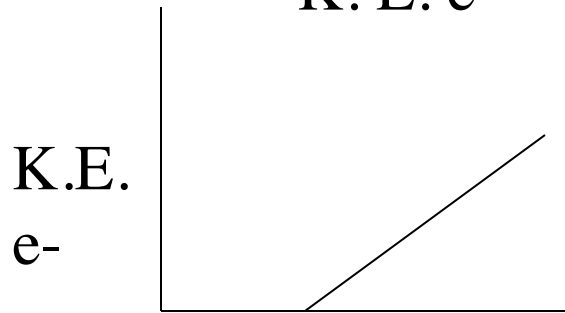


ν

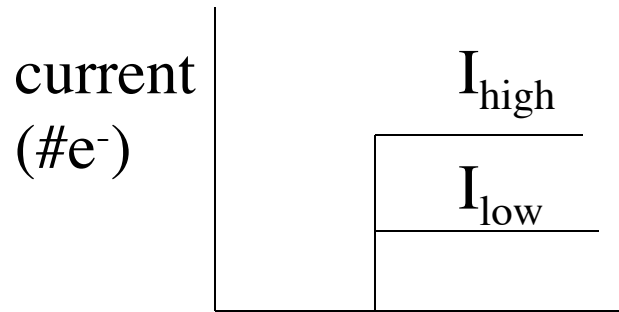
what do you see?



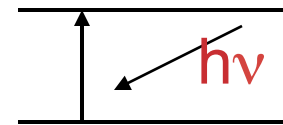
K. E. e<sup>-</sup>



ν



ν



*Einstein: Light is both a particle and a wave.*

e- K.E. "escape energy"

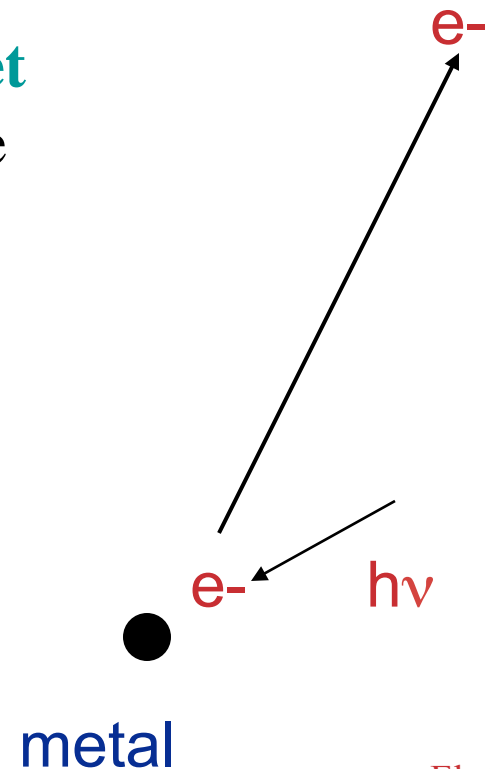
$$E_{\text{photon}} = 1/2mv^2 + h\nu_0 = E_{\text{electron}}$$

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).



# The Nature of Energy

- Energy,  $\lambda$ ,  $\nu$ , related:

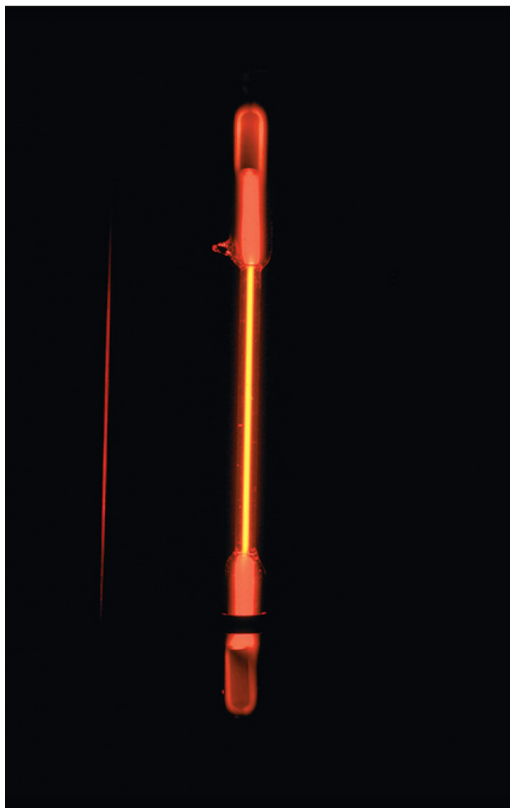
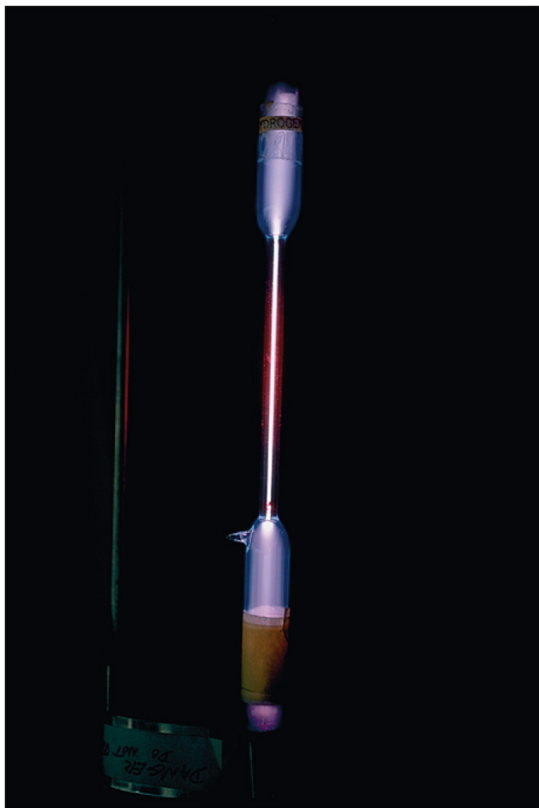
$$c = \lambda \nu$$

$$E = h\nu$$

**c = speed of light in vacuum,  
constant**



# Mystery number 3: element line spectrum



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.

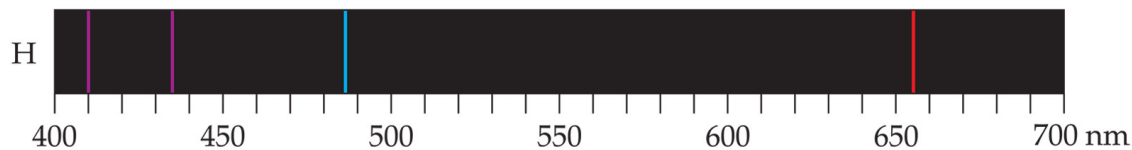
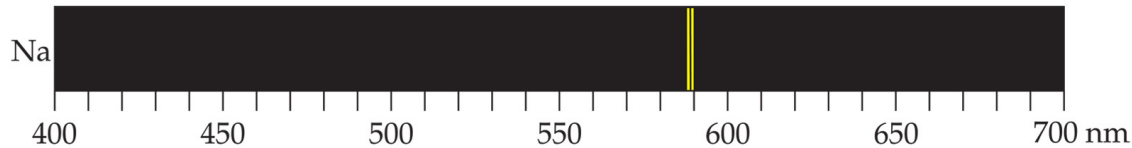
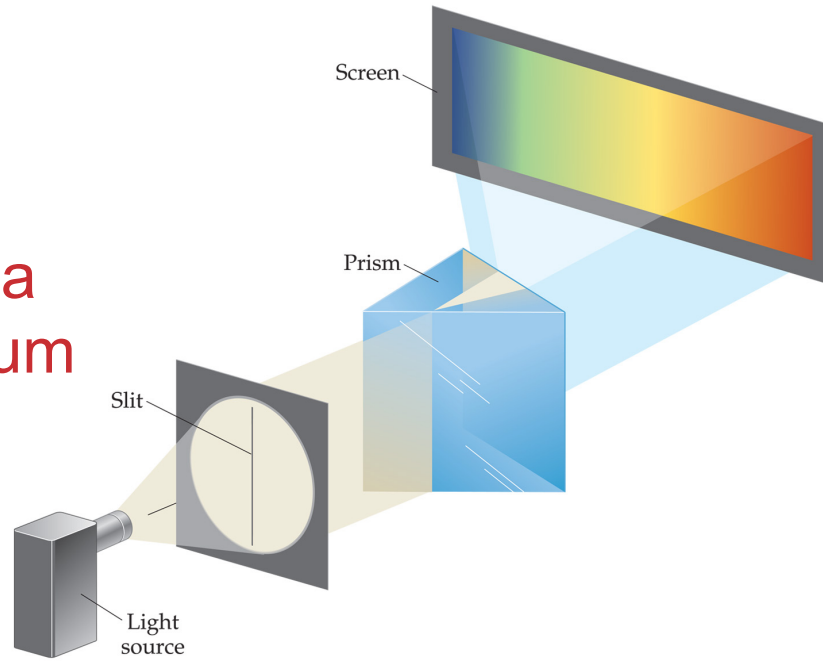
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Hydrogen

Neon

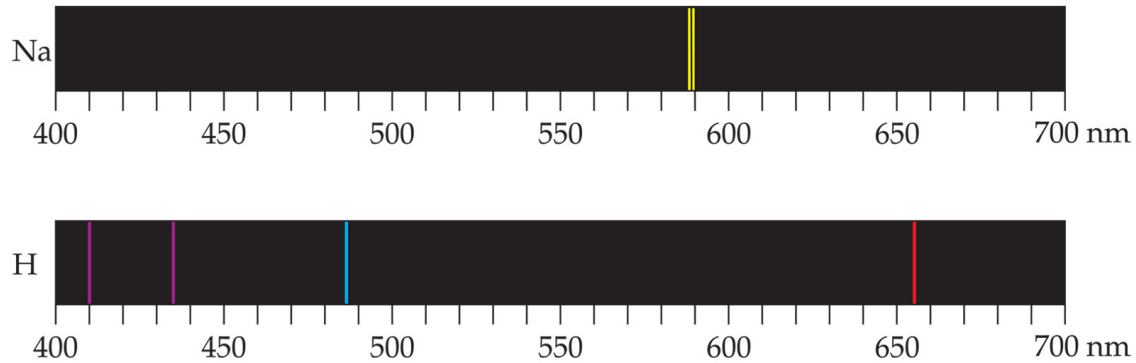
# The Nature of Light

White light shows a continuous spectrum



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

# Hydrogen Line spectra



Johann Balmer, School teacher

Just figured out that the lines fit a simple equation:

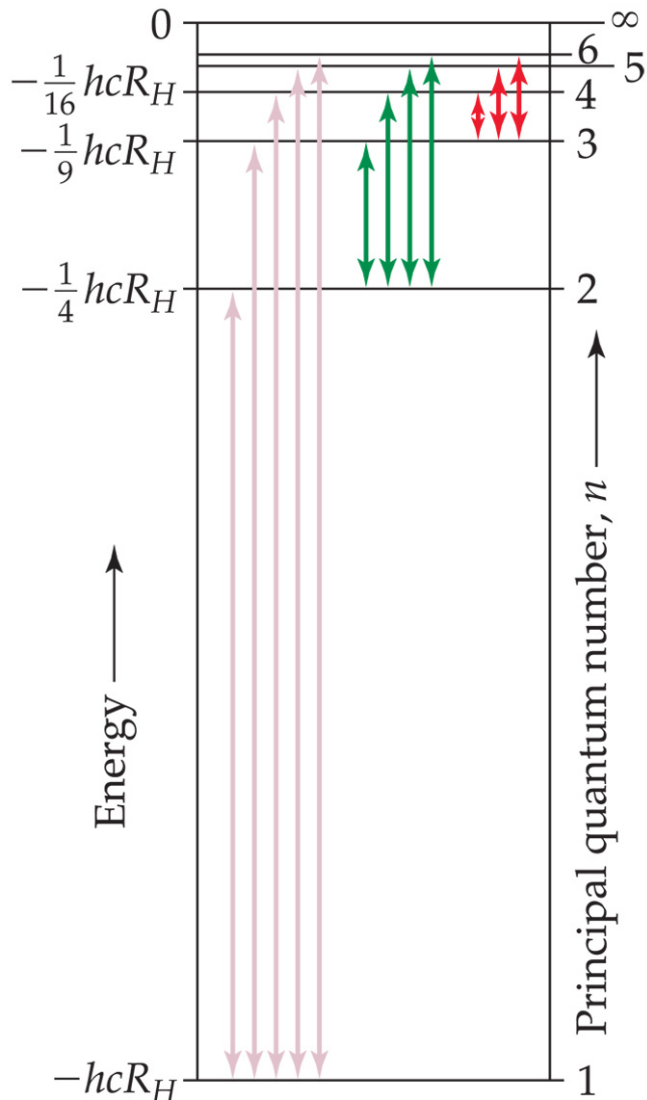
$$\frac{1}{\lambda} = (R_H) \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

$R_H$  = constant

$n_1$  and  $n_2$  are integers

But why?

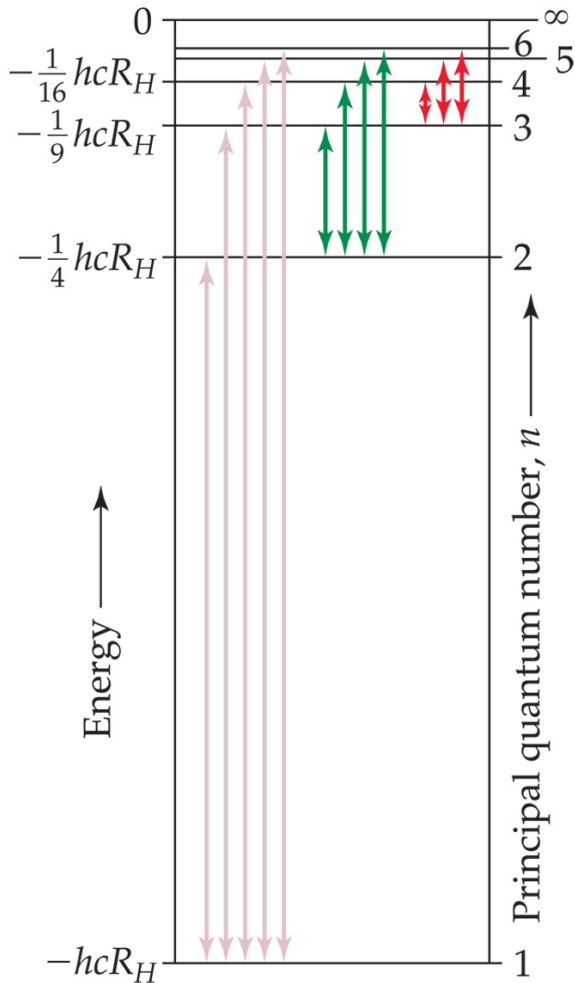
# The Nature of Energy



- 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).**

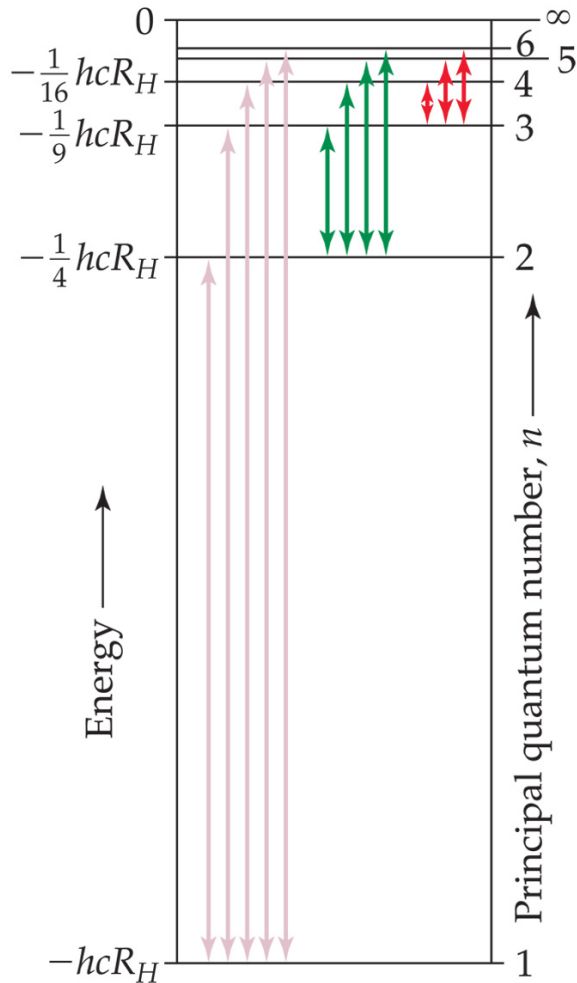
# The Nature of Energy



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

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

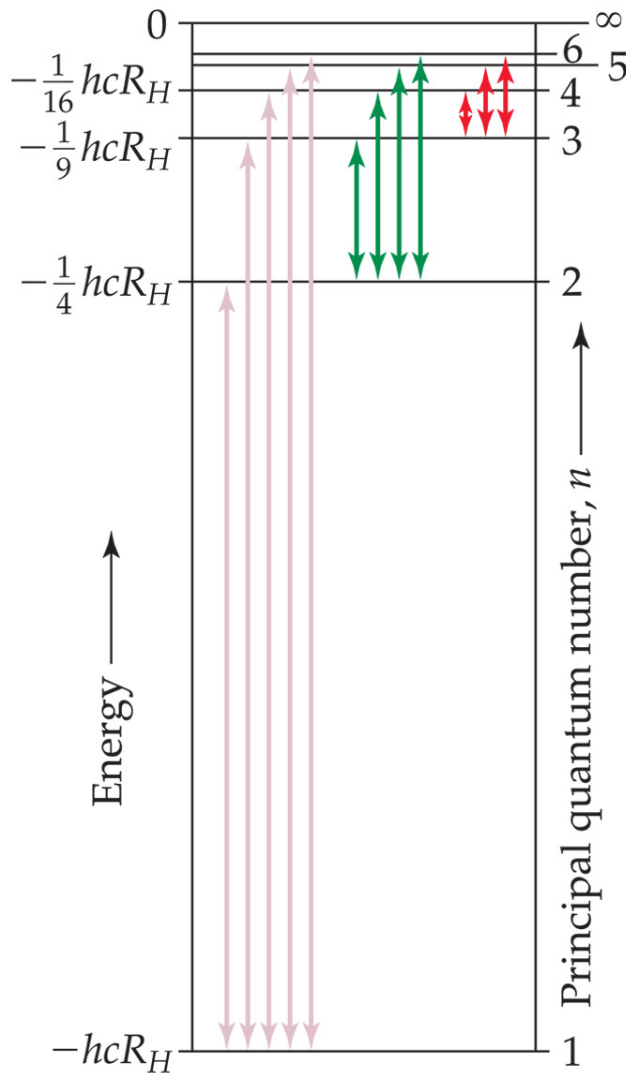
# The Nature of Energy



- 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

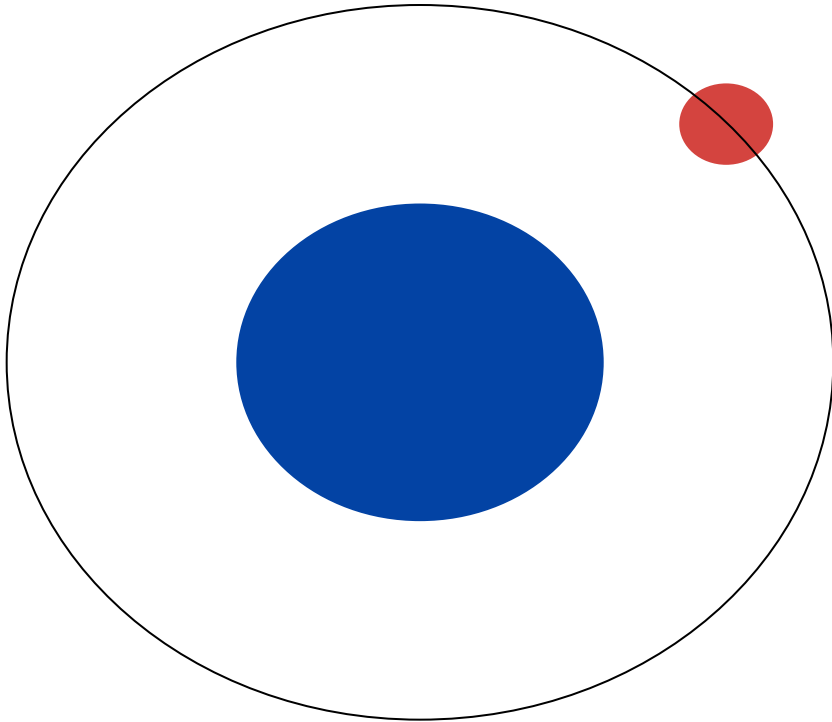


The energy absorbed or emitted from electron promotion or demotion can be calculated by the equation:

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

where  $R_H$  is the Rydberg constant,  $2.18 \times 10^{-18}$  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 nucleus like planets, Bohr could explain H, but no other elements.
- Too simple.

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

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

# The Wave Nature of Matter

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

$$\textit{Velocity} = \lambda \nu$$

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

like  $E=mc^2$

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

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

# Wave-like nature of matter

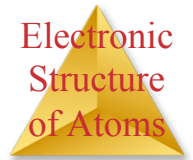
$$\lambda = \frac{h}{mv}$$

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

Example; What is  $\lambda$  for a 1 g ball?

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

wavelengths of everyday objects too small to measure.

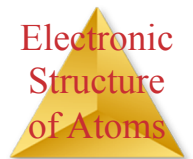


# Wave-like nature of matter

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

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

Wavelength of X-rays



# Electron microscopy

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

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# 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!

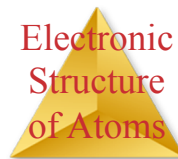
This is a result of the wave/particle duality of matter

# “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*

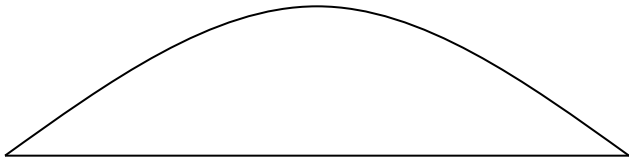


# Standing waves

$$l = (1/2)\lambda$$

$\nu_0$  = frequency

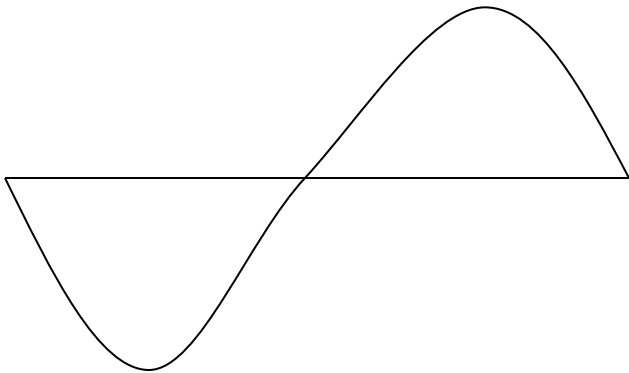
nodes = 2 (gotta have 2)



$$l = (2/2)\lambda = \lambda$$

$2\nu_0$  = frequency

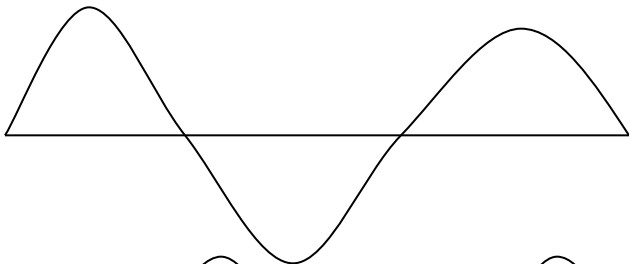
nodes = 3



$$l = (3/2)\lambda$$

$3\nu_0$  = frequency

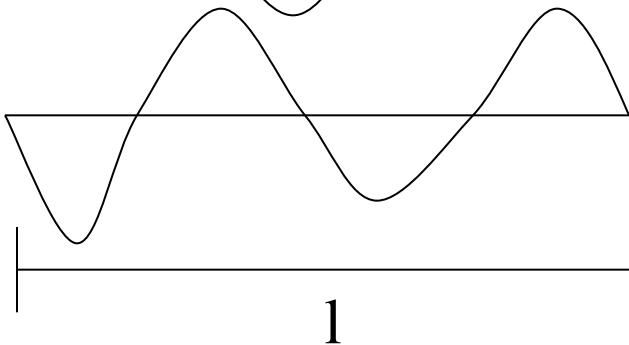
nodes = 4



$$l = (4/2)\lambda = 2\lambda$$

$4\nu_0$  = frequency

nodes = 5



Allowed  $\nu$  and  $\lambda$   
*quantized.*

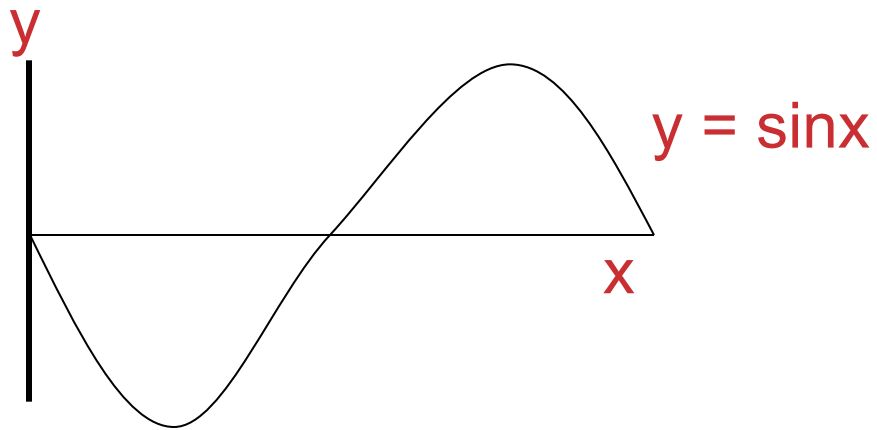
$l = (n/2)\lambda$ ,  $n$  is  
quantum #  
frequency =  $n\nu_0$

# Quantum mechanics

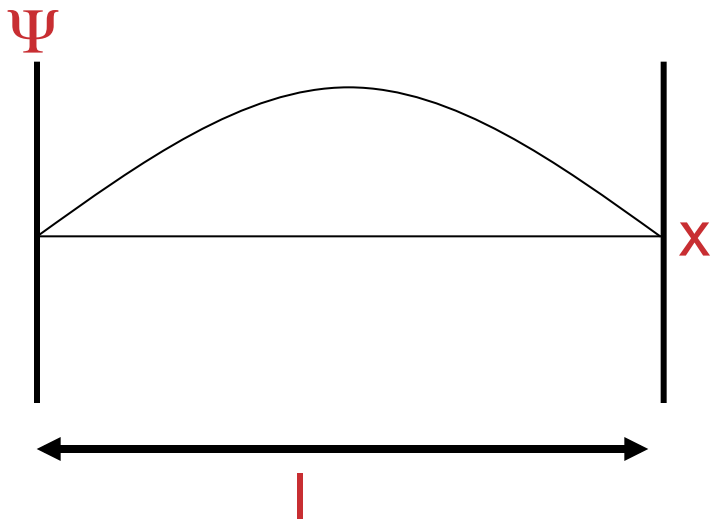
- Each electron can be explained using a standing wave equation (wavefunction)
- Quantized frequency corresponds to quantized Energy (Debroglie, Plank, etc.)
- Integer values are critical to this description: *quantum numbers*.

# Quantum mechanics

Examples of wave equations



Propagating wave



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

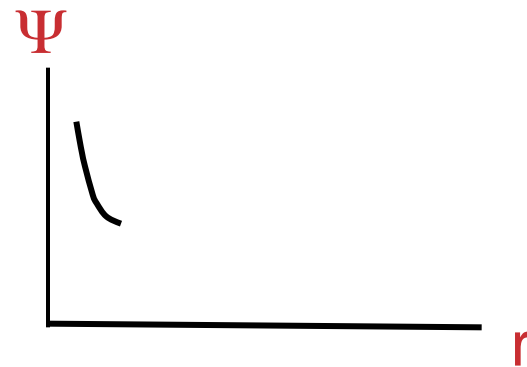
Standing wave

$l = 1/2\lambda$   
 $\nu_0 = \text{frequency}$   
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  $\Psi$ ? What is waving?

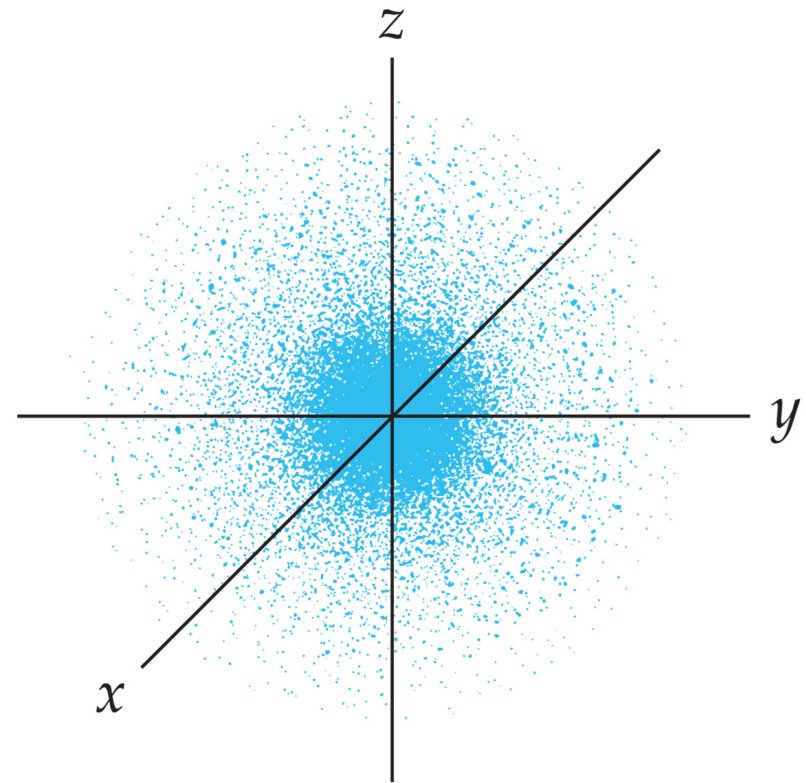
Born (1926):  $\Psi^2 =$  probability/volume of finding the electron.

# Quantum Mechanics

Plot of  $\Psi^2$  for hydrogen atom.

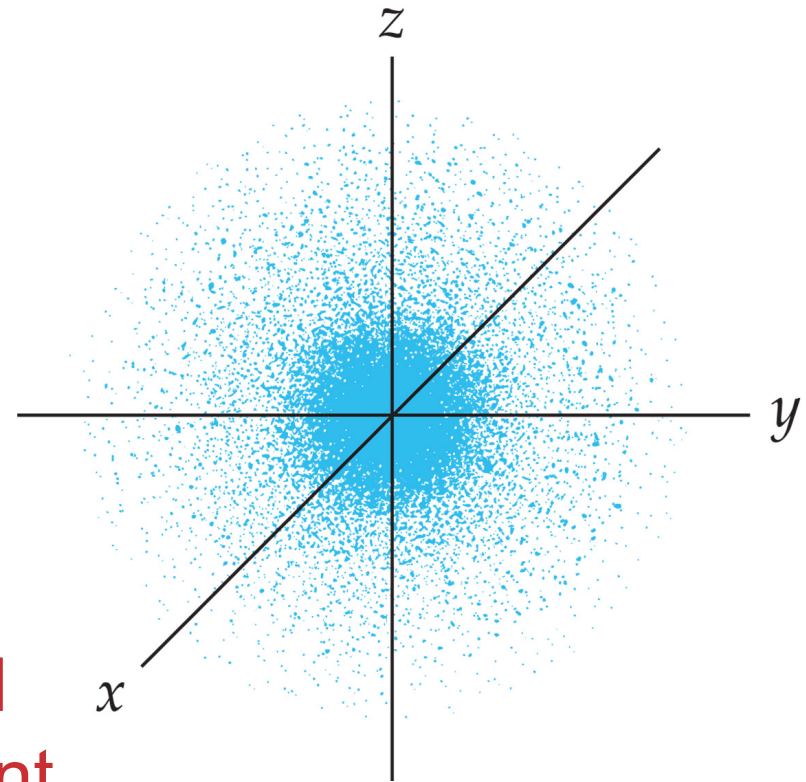
The closest thing we now have to a physical picture of an electron.

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* ( $\psi$ ).
- The square of the wave equation,  $\psi^2$ , 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, $l$

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

Value of $l$	0	1	2	3
Type of orbital	<i>s</i>	<i>p</i>	<i>d</i>	<i>f</i>

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	1s	0	1	1
2	0	2s	0	1	4
	1	2p	1, 0, -1	3	
3	0	3s	0	1	9
	1	3p	1, 0, -1	3	
	2	3d	2, 1, 0, -1, -2	5	
4	0	4s	0	1	16
	1	4p	1, 0, -1	3	
	2	4d	2, 1, 0, -1, -2	5	
	3	4f	3, 2, 1, 0, -1, -2, -3	7	

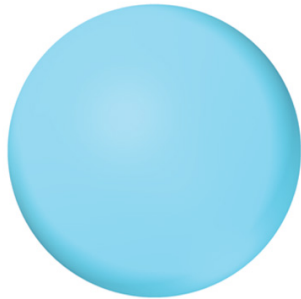
# s Orbitals



1s



2s

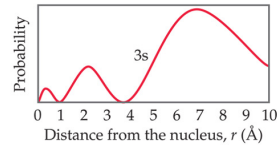
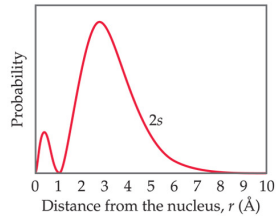
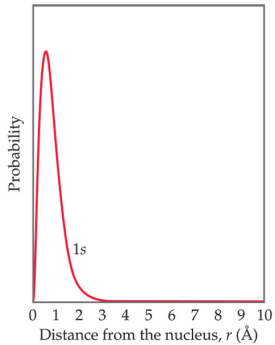


3s

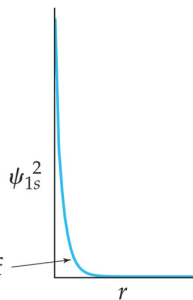
- Value of  $l = 0$ .
- Spherical in shape.
- Radius of sphere increases with increasing value of  $n$ .

# s Orbitals

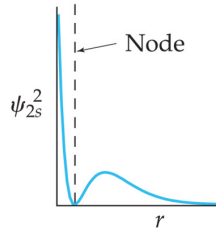
s orbitals possess  $n-1$  nodes, or regions where there is 0 probability of finding an electron.



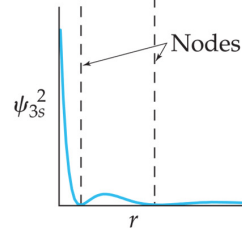
1s  
 $n = 1, l = 0$



2s  
 $n = 2, l = 0$

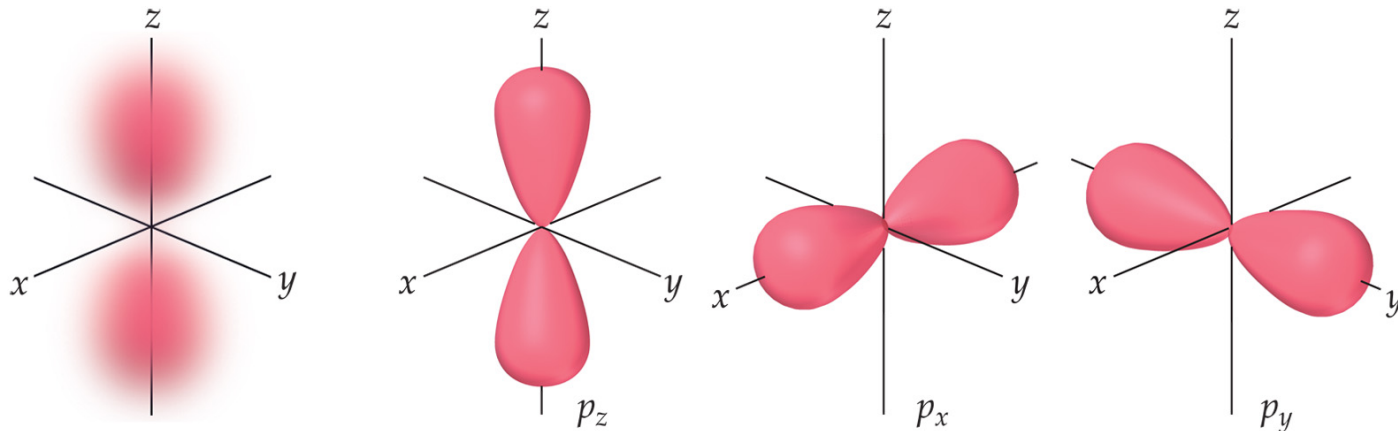


3s  
 $n = 3, l = 0$



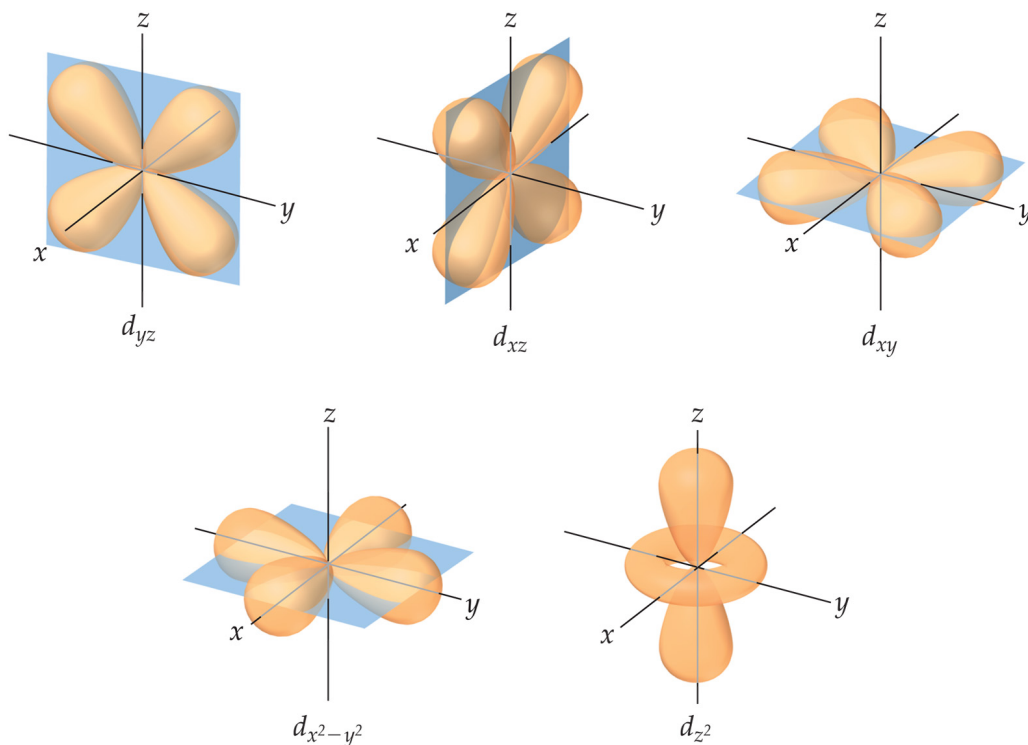
# $p$ Orbitals

- 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  $n$ .

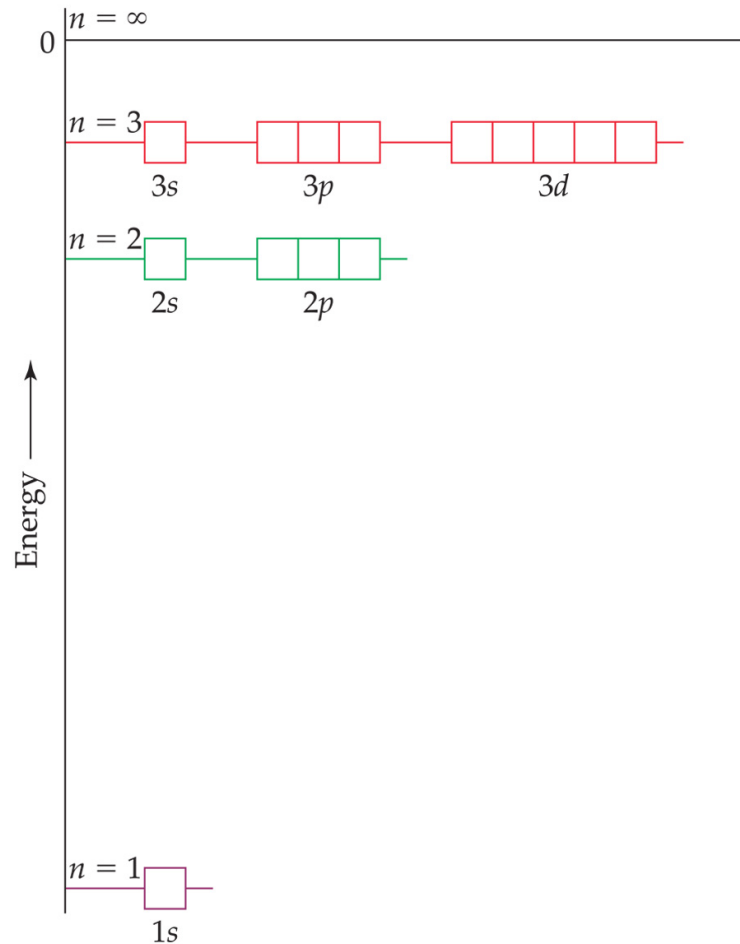
# Orbitals and nodes

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

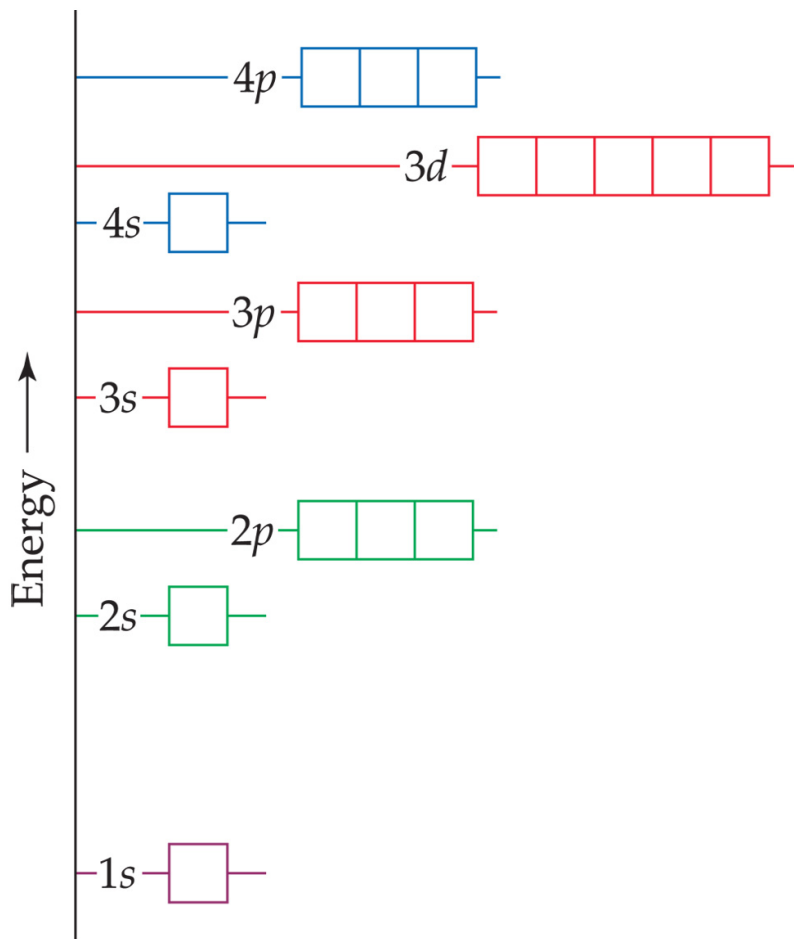
\*  $n$  = the shell, with  $n = 1$  the ground state or lowest possible energy shell. Thus  $n$  may have integral values from 1 - infinity.

# Energies of Orbitals

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

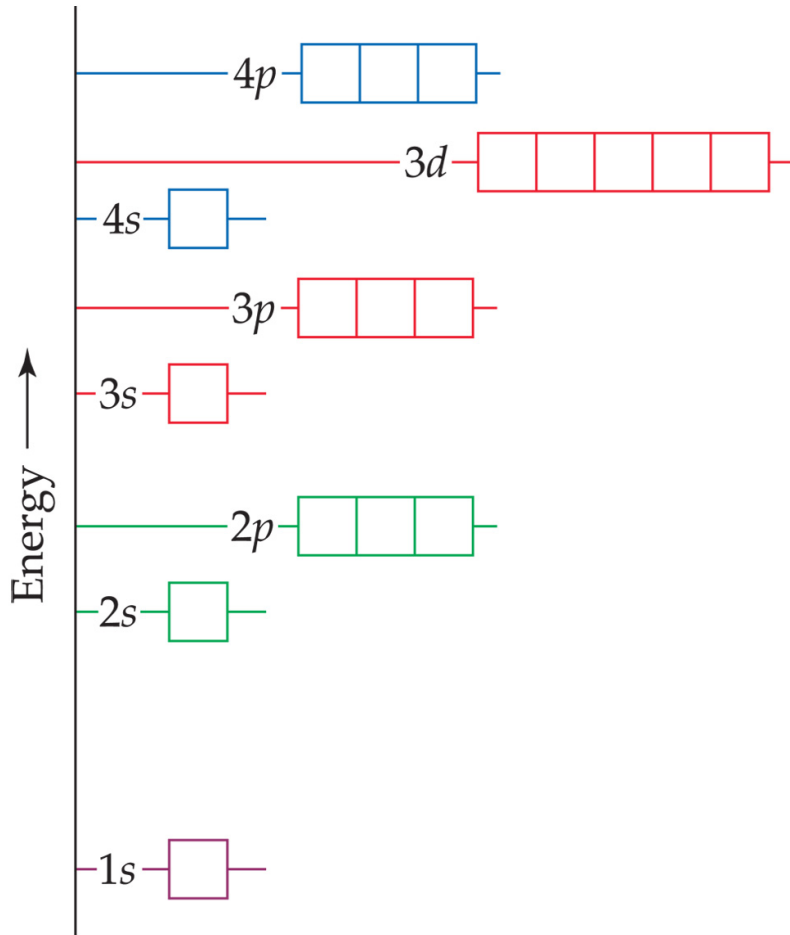


# Energies of Orbitals



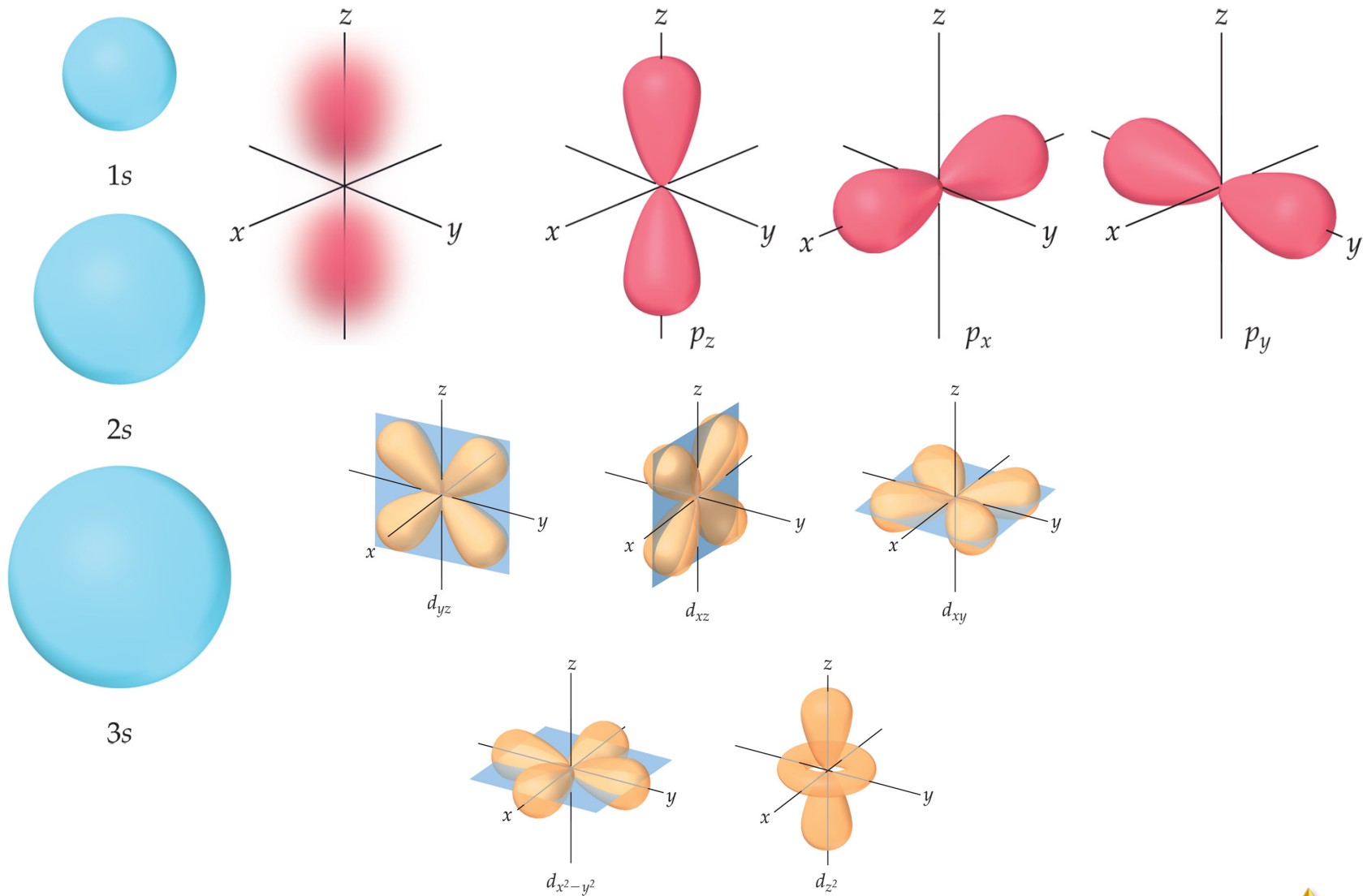
- As the number of electrons increases, though, so does the repulsion between them.
- Therefore, in many-electron atoms, orbitals on the same energy level are no longer degenerate.

# 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?



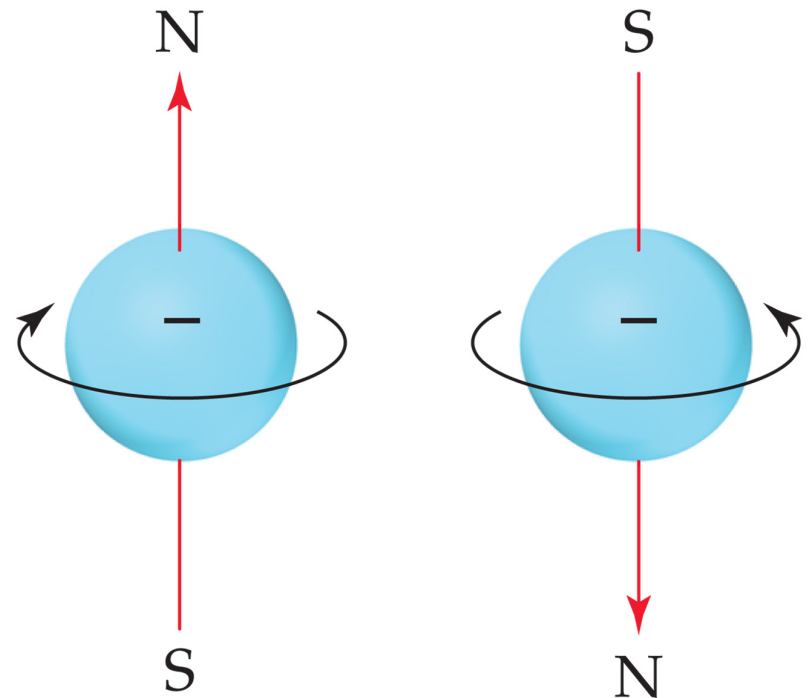
**The closer to the nucleus, the lower the energy**

# 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”

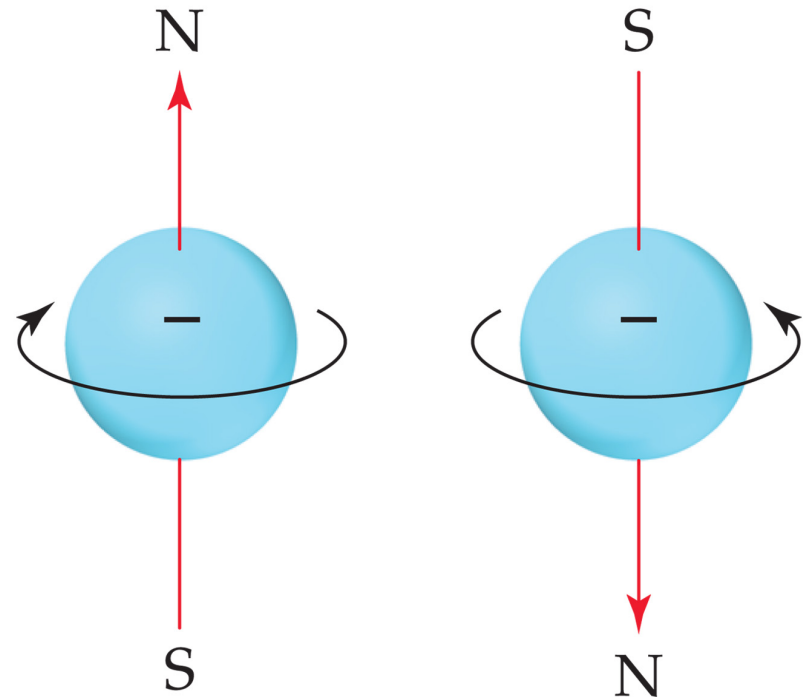
# Spin Quantum Number, $m_s$

- A fourth dimension required. Why?



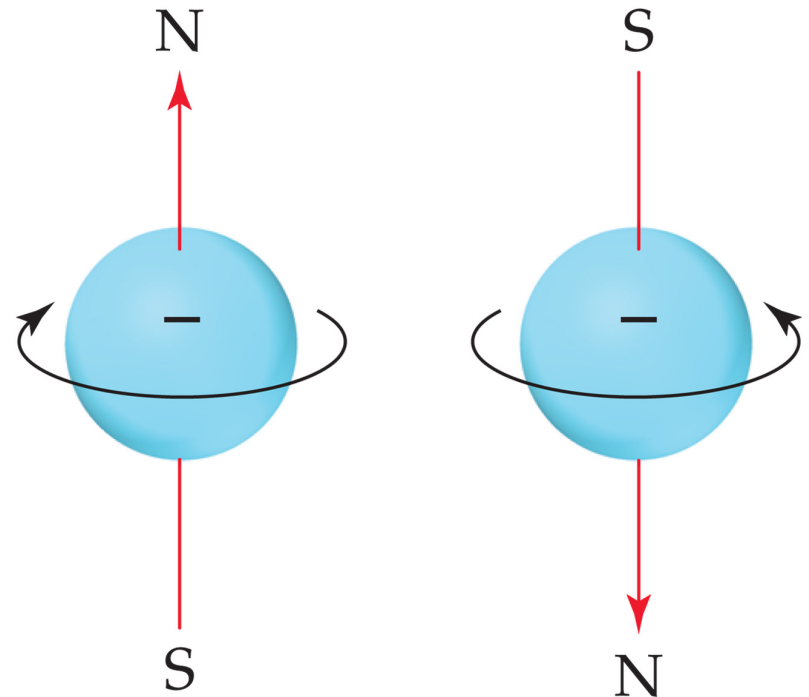
# Spin Quantum Number, $m_s$

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

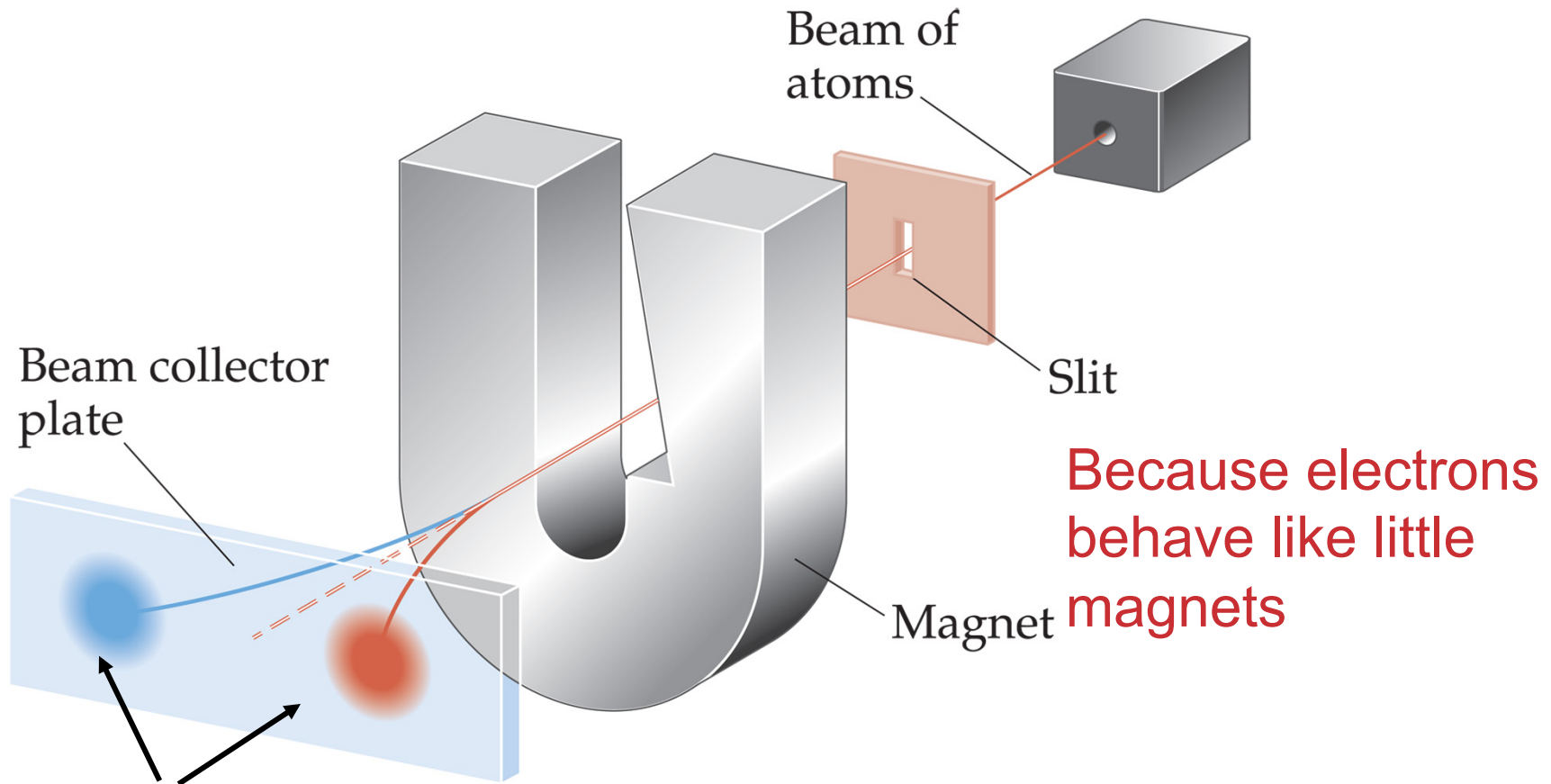


# Spin Quantum Number, $m_s$

- 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**

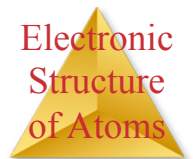


# Why do we call it “spin”



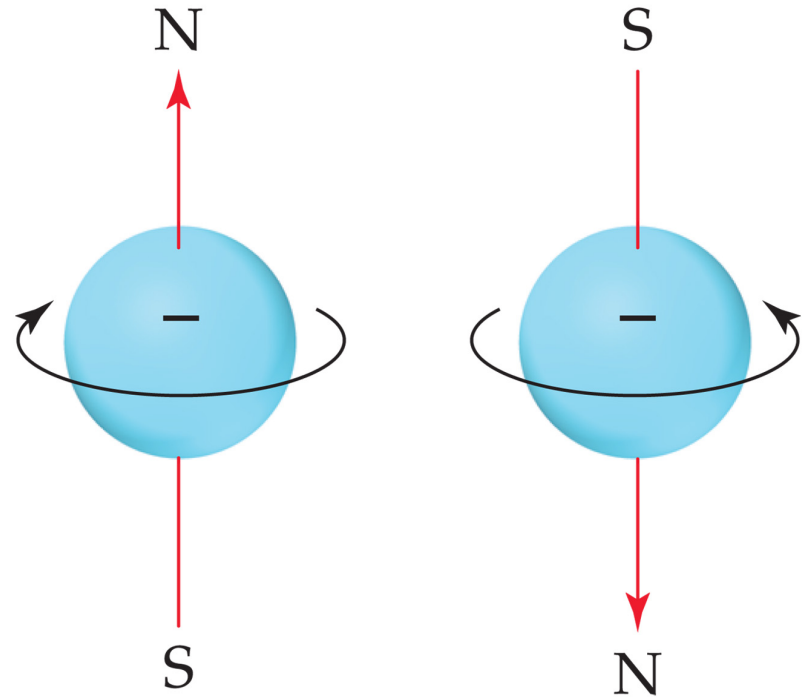
Note: apparently only two values for the magnetic field

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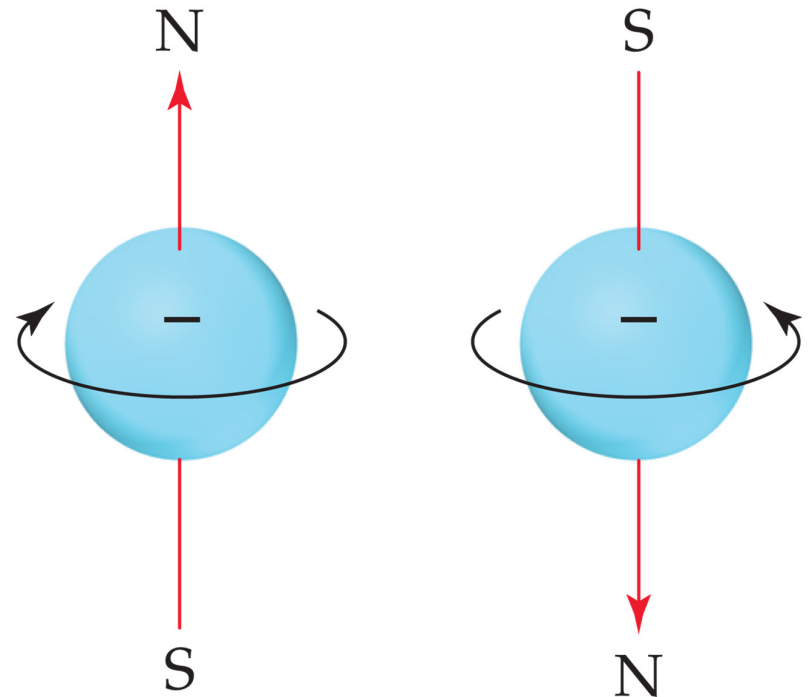
# Why do we call it “spin”

- 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<sup>5</sup>

# Electron Configurations



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

# Electron Configurations



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

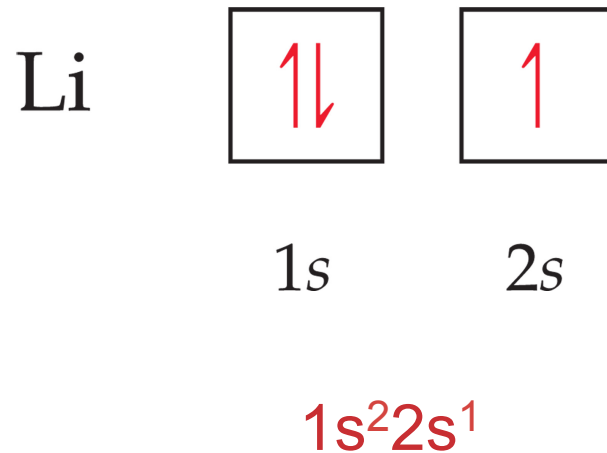
# Electron Configurations



- 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.

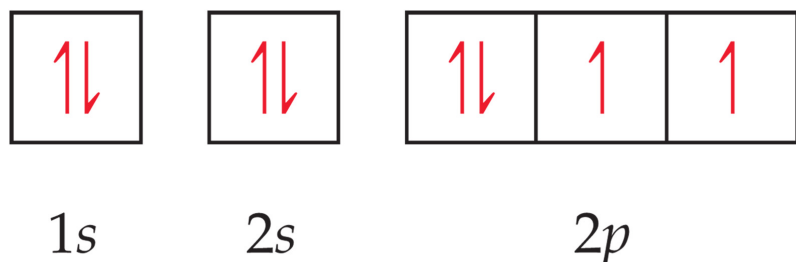
# Orbital Diagrams

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

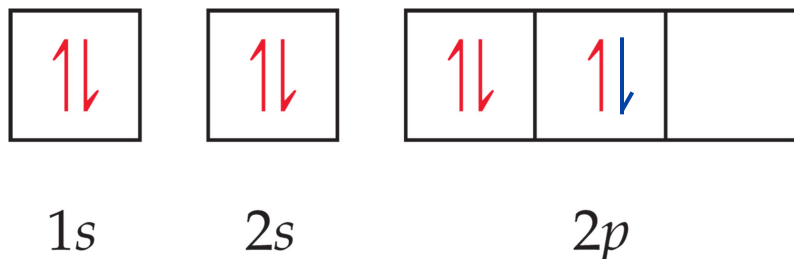


# Hund's Rule

(of maximum multiplicity)



NOT:

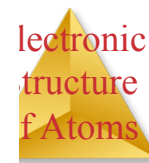


“For degenerate orbitals, the lowest energy is attained when the number of electrons with the same spin is maximized.”

# Electron configurations

**TABLE 6.3** Electron Configurations of Several Lighter Elements

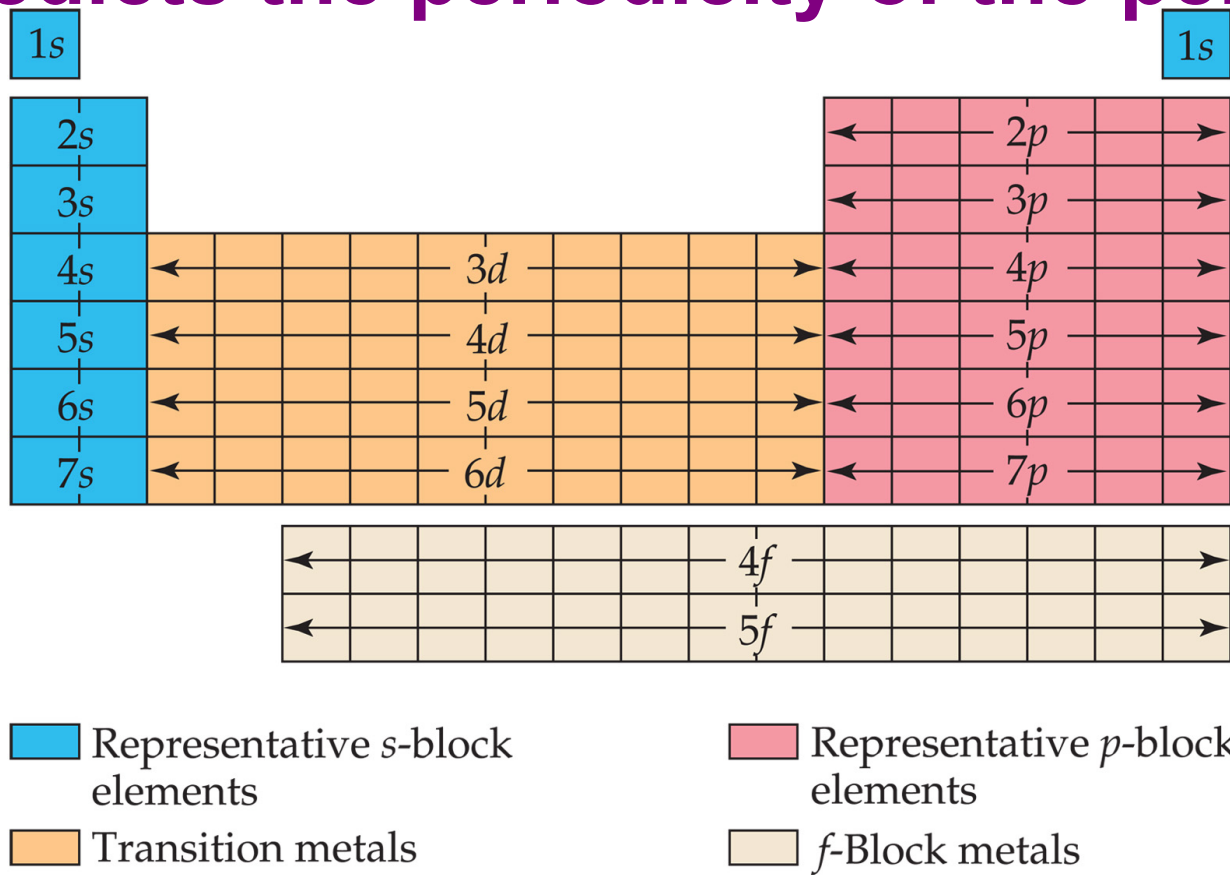
Element	Total Electrons	Orbital Diagram						Electron Configuration
		1s	2s	2p			3s	
Li	3	$\uparrow\downarrow$	$\uparrow$					$1s^2 2s^1$
Be	4	$\uparrow\downarrow$	$\uparrow\downarrow$					$1s^2 2s^2$
B	5	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$				$1s^2 2s^2 2p^1$
C	6	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$	$\uparrow$			$1s^2 2s^2 2p^2$
N	7	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$	$\uparrow$	$\uparrow$		$1s^2 2s^2 2p^3$
Ne	10	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$		$1s^2 2s^2 2p^6$
Na	11	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow$	$1s^2 2s^2 2p^6 3s^1$



# 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!!

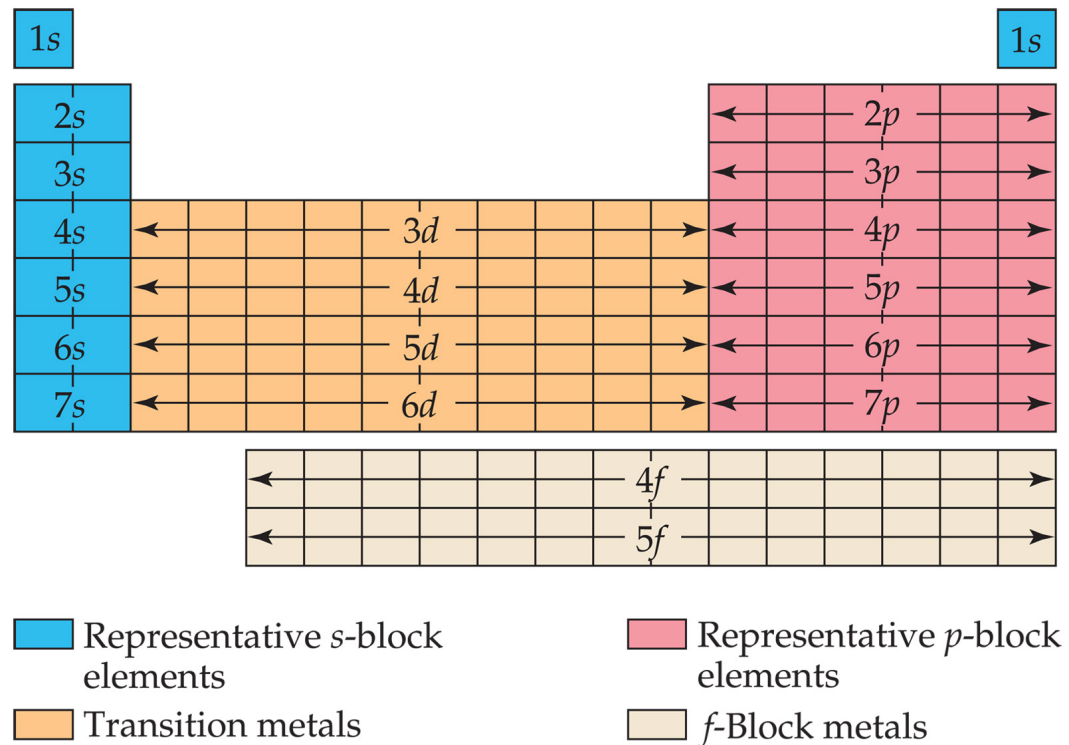


- We fill orbitals in increasing order of energy.
- Different blocks on the periodic table, then correspond to different types of orbitals.

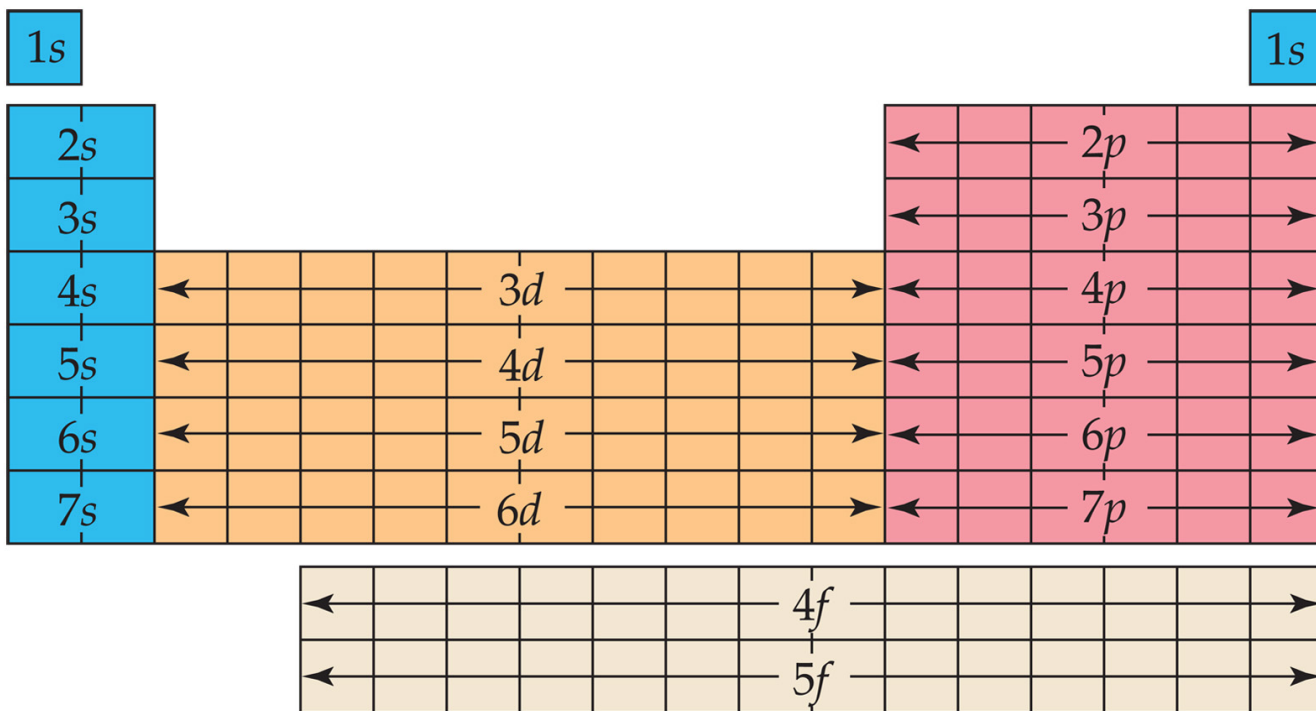
# Why do we accept this wacko stuff?

## 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



 Representative s-block elements

 Transition metals

 Representative p-block elements

 f-Block metals

- **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] $2s^2 2p^1$
Al	[Ne] $3s^2 3p^1$
Ga	[Ar] $3d^{10} 4s^2 4p^1$
In	[Kr] $4d^{10} 5s^2 5p^1$
Tl	[Xe] $4f^{14} 5d^{10} 6s^2 6p^1$

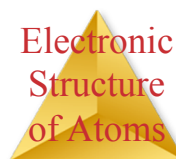
---

# Electron configurations of the elements

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

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

Metals
  Metalloids
  Nonmetals









# What's on the exam?

- Chapter 4
- Chapter 5 (everything)
- Chapter 6 (everything)

# Chapter 4.

Solution stoichiometry

Strong vs. weak electrolytes

Know strong electrolytes

strong acids

soluble salts

precipitation reactions

ionic equation

net ionic equation

Neutralization reactions

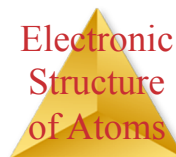
gas forming reactions



Molarity

Dilution

Titration



# Chapter 4.

Solution stoichiometry

Molarity

Dilution

Titration

Oxidation reduction

assigning oxidation numbers

who is oxidizing and reducing?

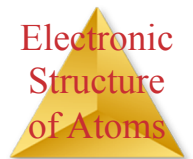
activity series

# Chapter 5

- Heat and work
- $E = q + w$
- $H$  is  $q$  at constant  $P$
- Hess' s law problems
- Heat of formation problems
- Calorimetry problems

# Chapter 6

- History of Light
- Electromagnetic radiation order
- Blackbody radiation
- Photo electric effect
- $E = h\nu$
- Quantum numbers
  - What are they for?
  - Why are there 3 or 4?
  - What does each stand for?
  - What is their relationship ( $l = 0, 1, 2 \dots n-1$ )
  - Electron configurations using the periodic table



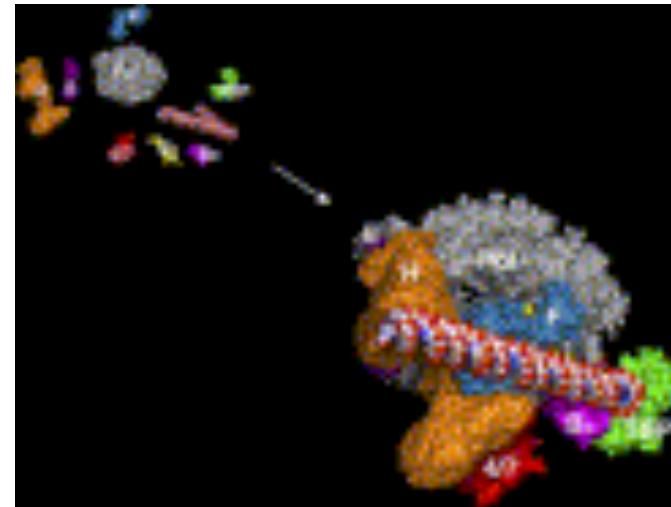
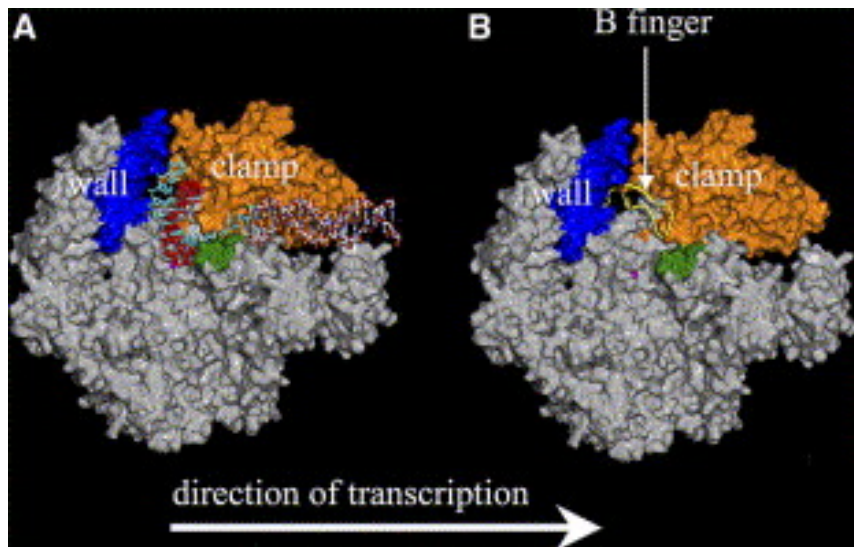
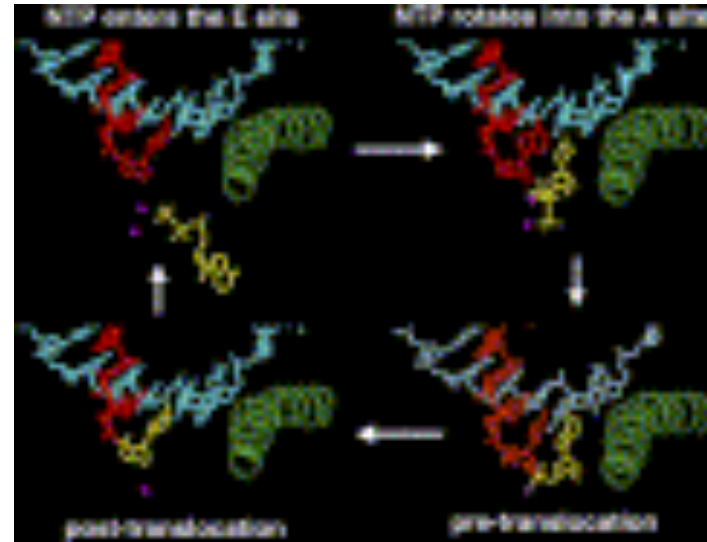


# The Nobel Prize in Chemistry 2008

Roger Kornberg

X-ray crystallography

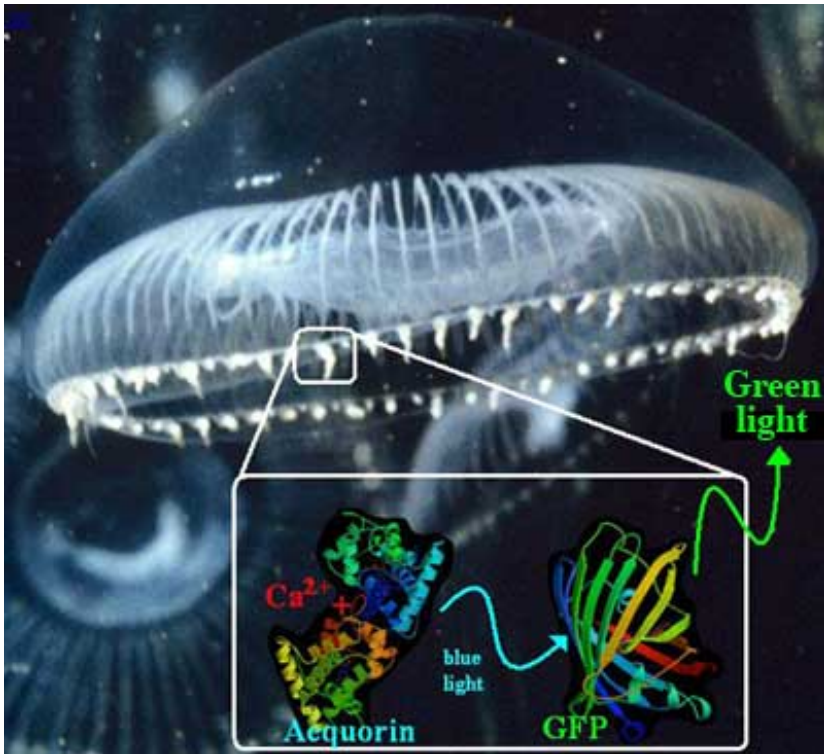
How does RNA Pol II  
decode DNA into RNA?



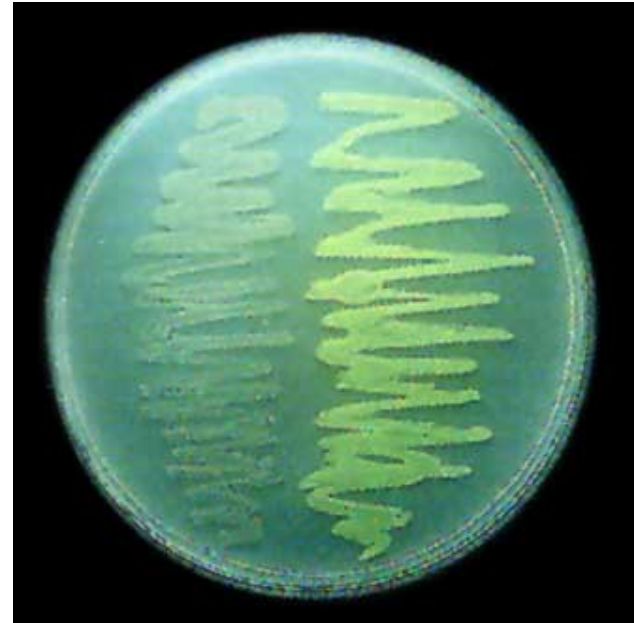
Electronic  
Structure  
of Atoms

# Nobel Prize in Chemistry

- Green Fluorescent protein

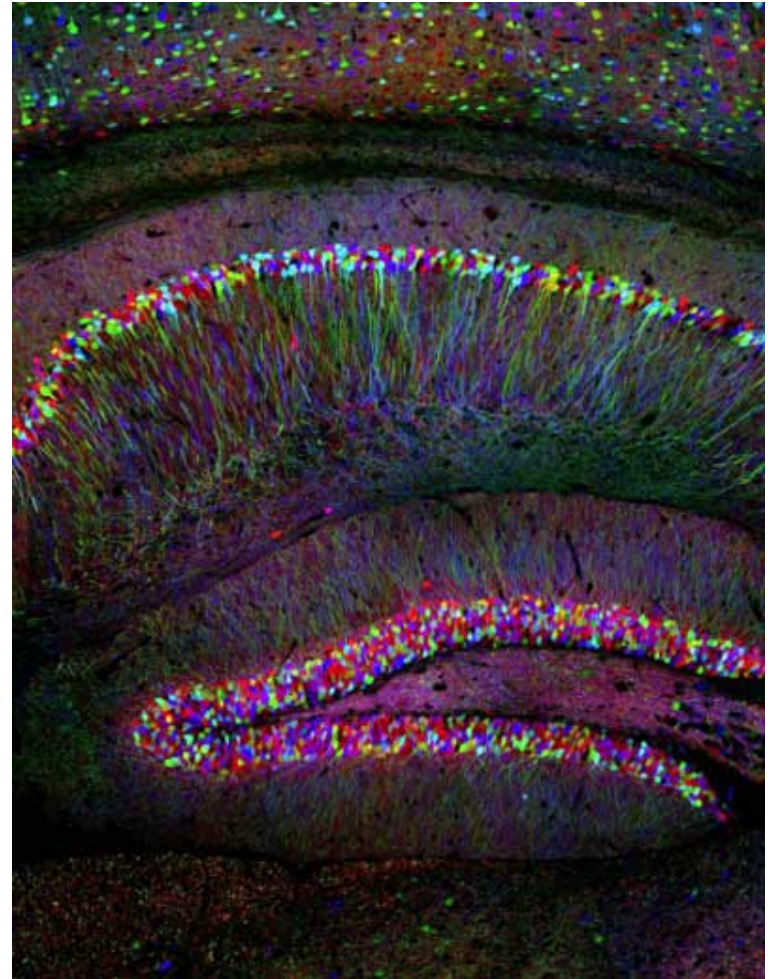


Osamu Shimonura



Marty Chalfie

GFP



Roger Tsien

Cerebral cortex  
Tsien/Chalfie  
Lichtman/Sanes

