Grade 12 Chemistry

Structure and Properties of Matter

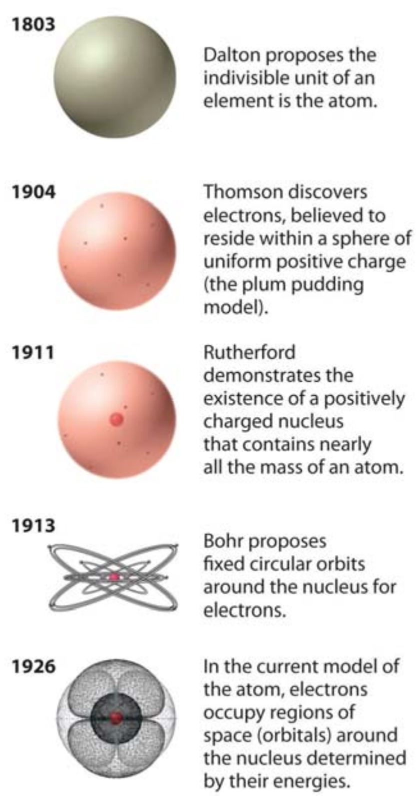
Class 5

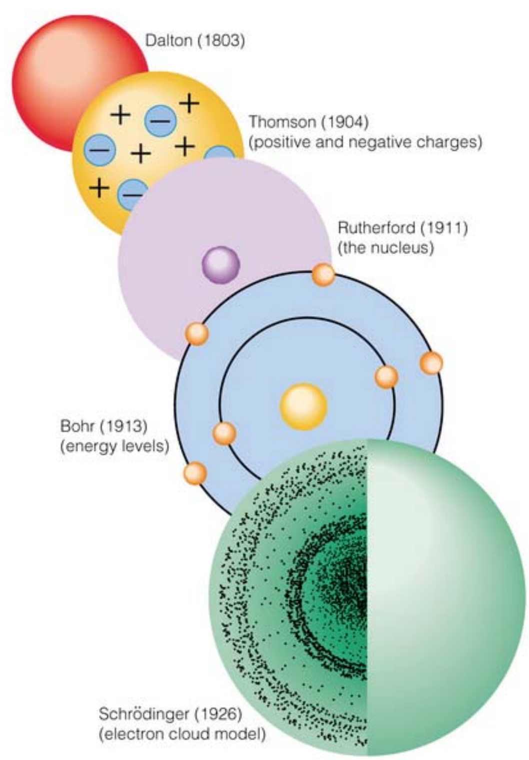
Overall Expectations

- Assess the benefits to society and evaluate the environmental impact of products and technologies that apply principles related to the structure and properties of matter
- Investigate the molecular shapes and physical properties of various types of matter
- Demonstrate an understanding of atomic structure and chemical bonding, and how they relate to the physical properties of ionic, molecular, covalent network, and metallic substances

How much do you remember?

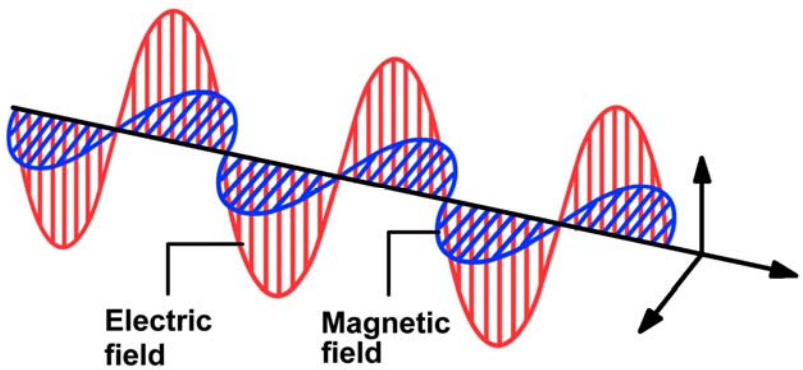
1. Who organized the periodic table?
2. How did he arrange the elements?
3. Who came up with the atomic theory?
4. What did Rutherford's gold foil experiment conclude?
5. Who discovered the neutron?
6. Who came up with the idea of electron orbits?

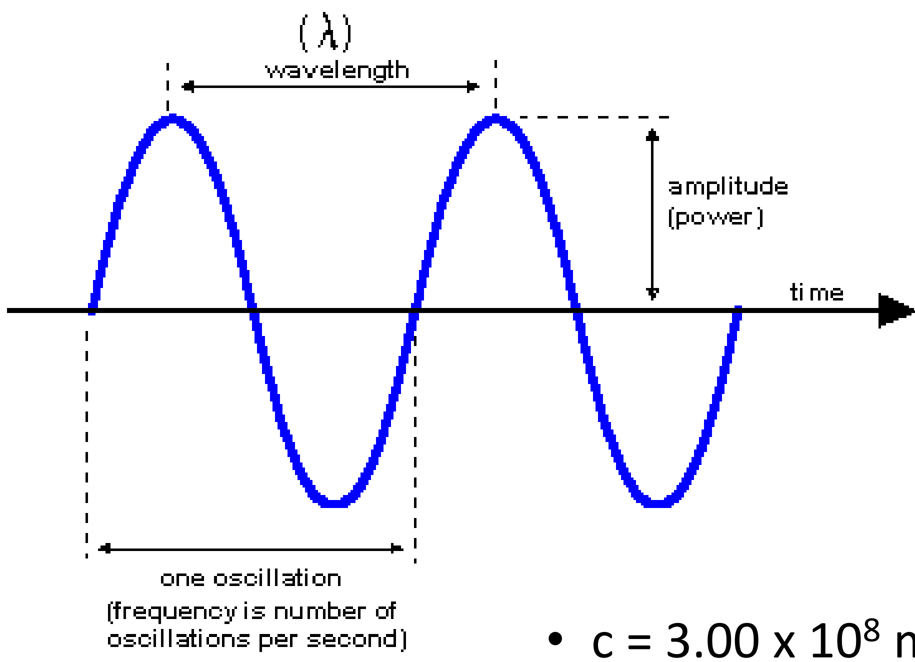




Light

- Light is electromagnetic radiation
- Classical theory states that light is an electromagnetic wave with no mass or specific position in space



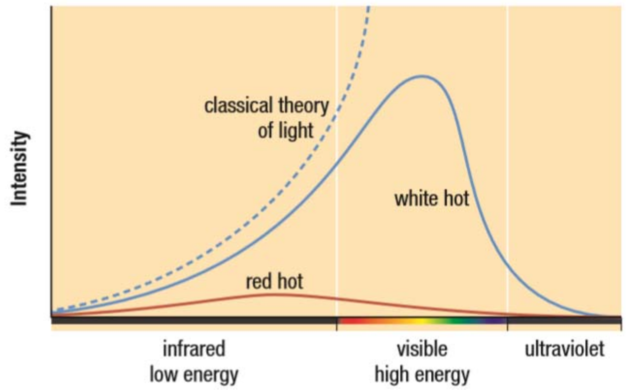


Speed of light $c = \lambda f$

- $c = 3.00 \times 10^8$ m/s
- λ = wavelength (m)
- f = frequency (1/s)

The Beginning of Quantum

- In 1900, Max Planck studied blackbody radiation
- Classical theory predicted energy curve should rise continuously with temperature



- Planck concluded that matter can gain or lose energy in whole-number multiples
- Energy is quantized

$$E = nhf$$

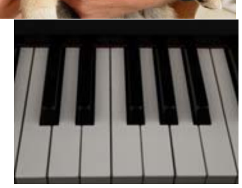
Where:

E = energy

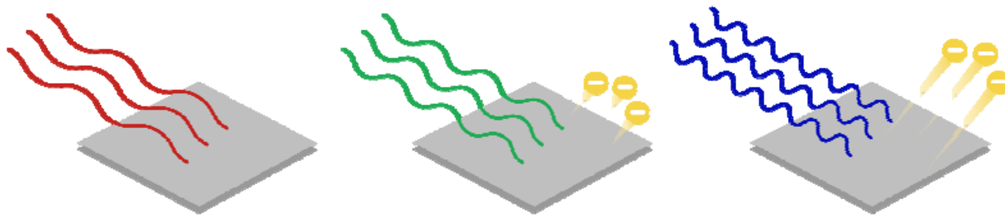
n = integer (1, 2, 3, …)

h = Planck's constant 6.63×10^{-34} J•s

f = frequency of the radiation (1/s)



Photoelectric Effect



- In 1905, Einstein noticed that electrons are ejected from the surface of certain metals exposed to light at a threshold frequency
- Below threshold frequency, no electrons were ejected; non-continuous
- Einstein suggested that a beam of light is actually a stream of particles called **photons** (Wave-Particle Duality)

- Photons – particles of light
- Each photon possesses energy given by the equation

$$E = h\frac{c}{\lambda}$$

Where:

E = energy

h = Planck's constant 6.63×10^{-34} J•s

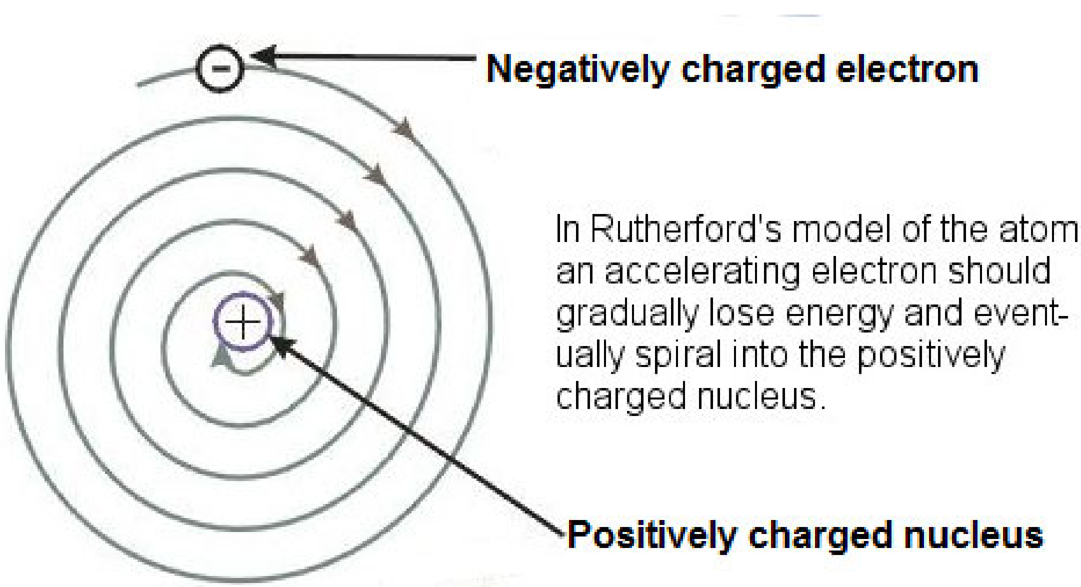
c = 3.00×10^{8} m/s

λ = wavelength (m)

Conclusion: Light behaves as a wave AND a particle

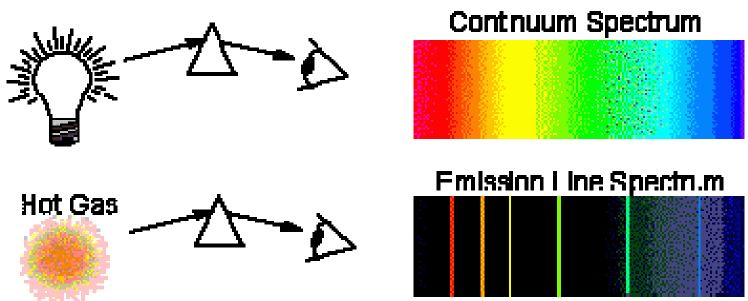
What about electrons?

The Atomic Model



Emission Spectrum

- Two Types:
  - **Continuous Spectrum** – when white light passes through a prism
  - **Line Spectrum** – when a sample of gas absorbs energy and emits energy at particular wavelengths



Line Spectrum

- Atoms give off light when heated or energetically excited
- An excited atom does not give off a continuous distribution of all wavelengths but rather a series of discrete lines – **a line spectrum**
- Each element has its own spectral “signature”

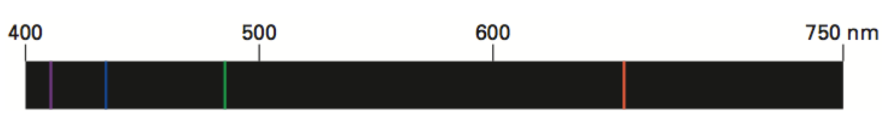
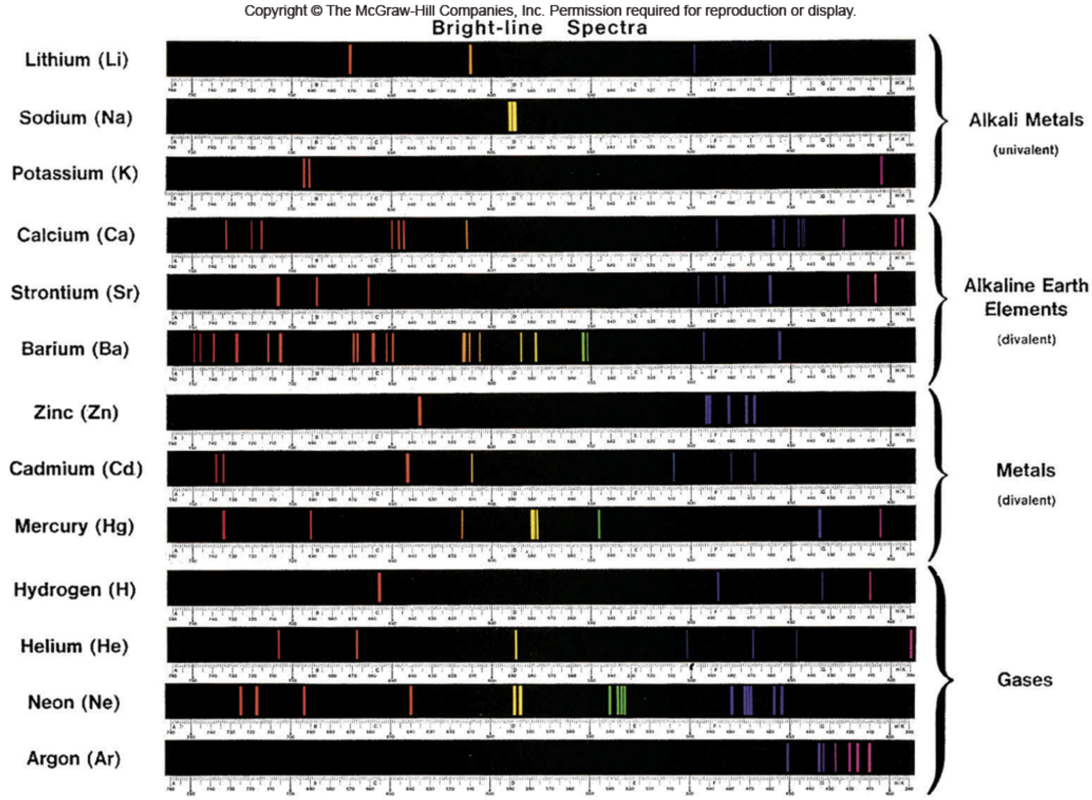
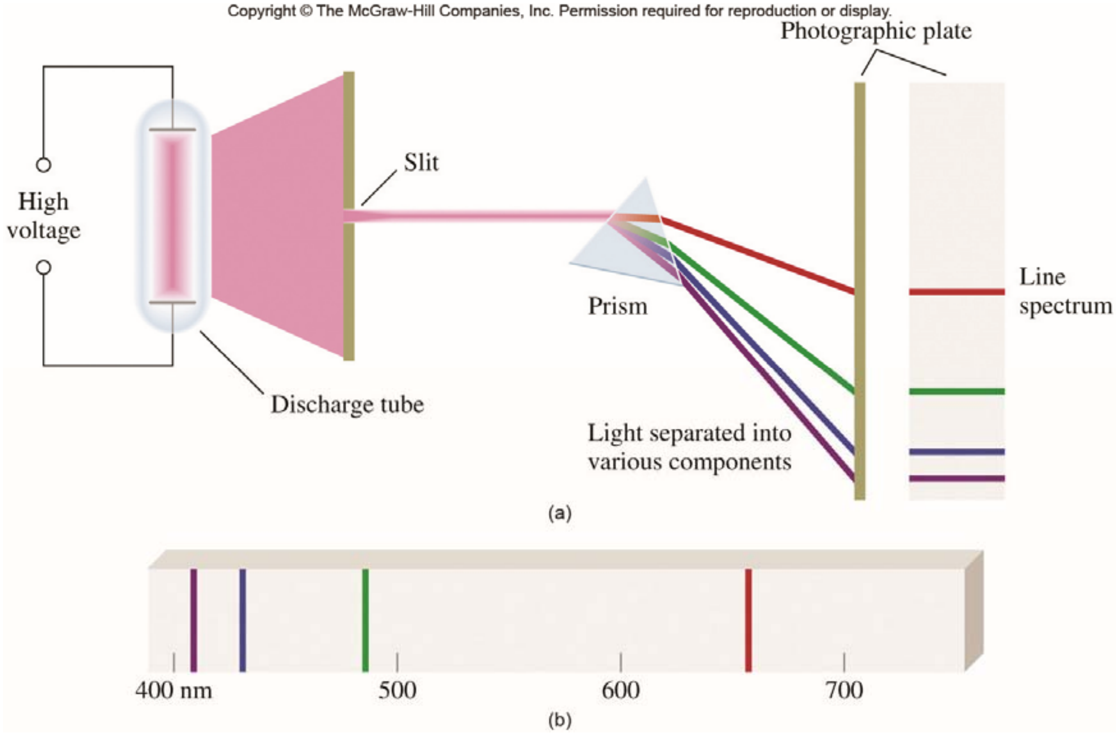


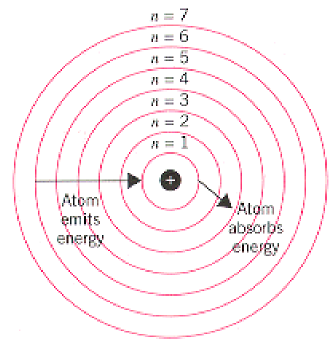
Figure 3.7 The discrete, coloured lines of this spectrum are characteristic of hydrogen atoms. No other atoms display this pattern of coloured lines.

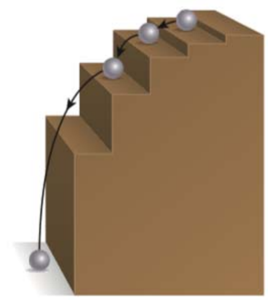

Bright-line Spectra
Lithium (Li)
Sodium (Na)
Potassium (K)
Calcium (Ca)
Strontium (Sr)
Barium (Ba)
Zinc (Zn)
Cadmium (Cd)
Mercury (Hg)
Hydrogen (H)
Helium (He)
Neon (Ne)
Argon (Ar)
Alkali Metals
(univalent)
Alkaline Earth Elements
(divalent)
Metals
(divalent)
Gases

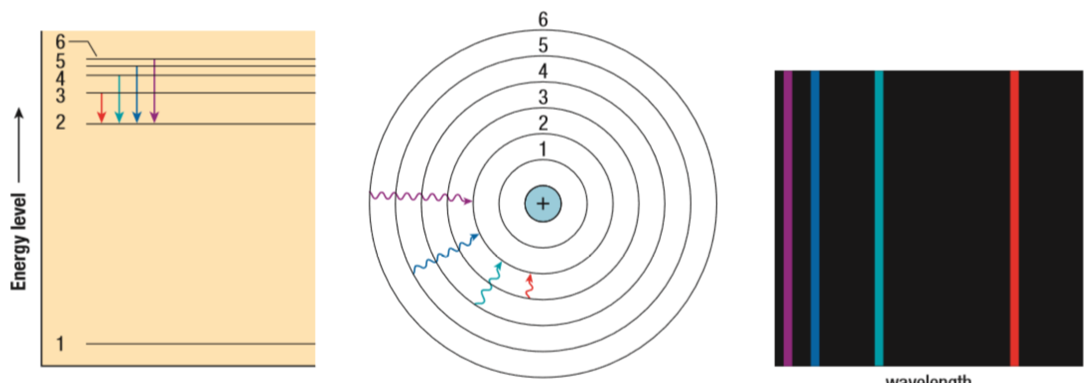

Photographic plate
Slit
High voltage
Prism
Line spectrum
Discharge tube
Light separated into various components
(a)
400 nm
500
600
700
(b)

The Bohr Model of the Atom

- The atom has only specific, allowable energy levels called stationary states. Each stationary state corresponds to the atom’s electrons occupying fixed, circular orbits around the nucleus
- While in its stationary state, atoms do not emit energy

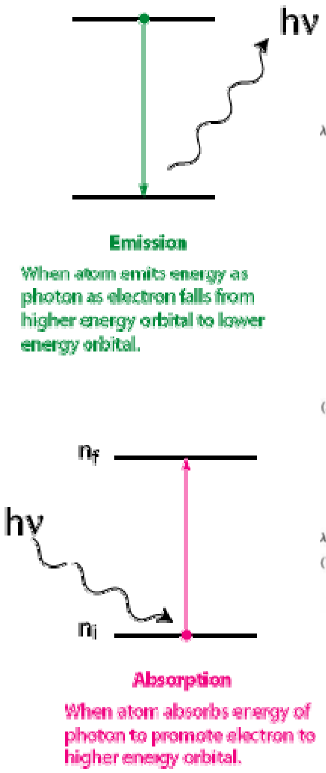


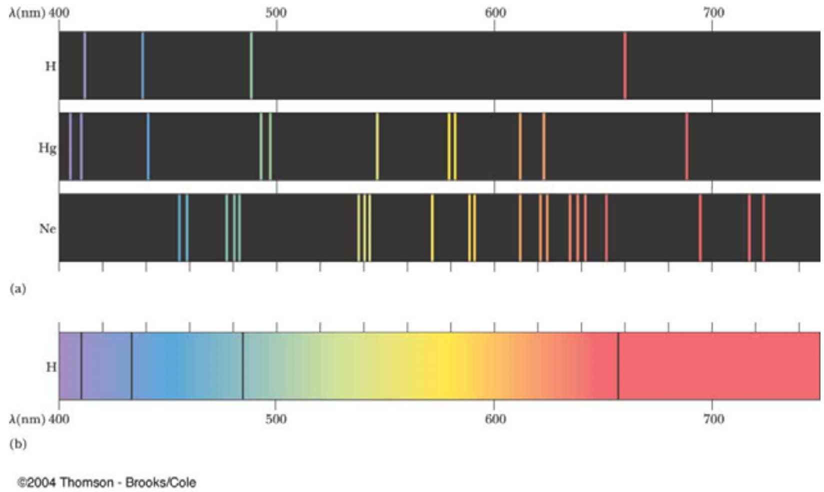




(a) (b) (c)

- An atom changes stationary states by emitting or absorbing a specific quantity of energy exactly equal to the difference in energy between the two stationary states

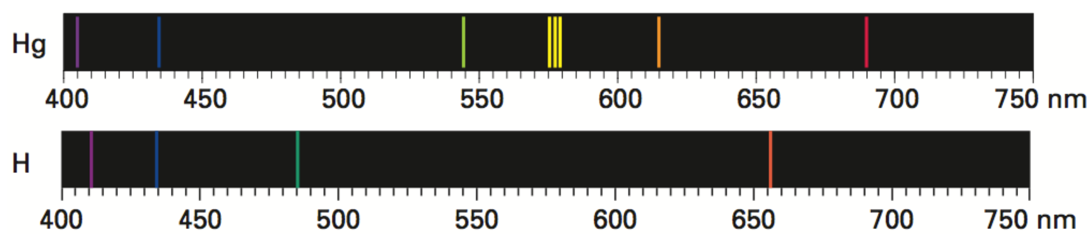




Conclusion: Electrons only exist at certain energy levels

Limitations of Bohr's Model

- This model fails for atoms with more than one electron; his model only worked for H and other ions with only one electron like He^{+} Bohr could not explain the emission spectra of atoms with two or more electrons



The Dual Nature of the Electron

- In the 1920s, Louis de Broglie proposed that matter has properties of waves and that electrons are standing stationary waves

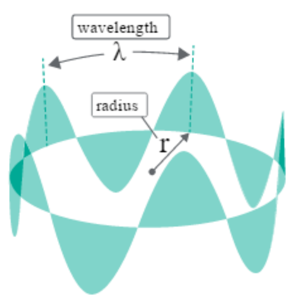
$$\lambda = \frac{h}{mv}$$

Where:
h = Plank's constant (6.63×10^{-34} J•s)
m = mass
v = velocity

Checkpoint

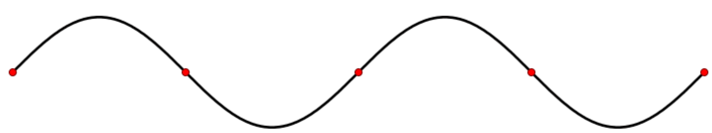
EXAMPLE 7.5

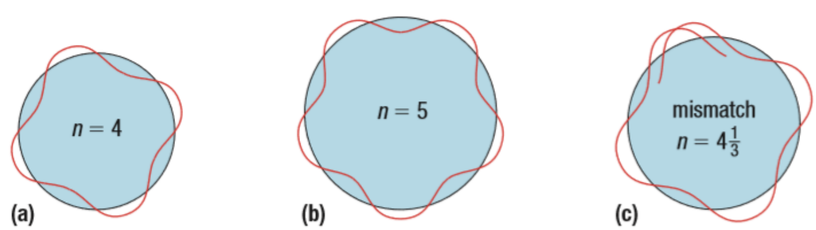
Calculate the wavelength of the "particle" in the following two cases: (a) The fastest serve in tennis is about 150 miles per hour, or 68 m/s. Calculate the wavelength associated with a 6.0×10^{-2}-kg tennis ball traveling at this speed. (b) Calculate the wavelength associated with an electron (9.1094×10^{-31} kg) moving at 68 m/s.



Standing Wave

Destructive Interference





- Only certain wavelengths are allowed for any standing wave
- Vibrations are quantized

Schrodinger Equation

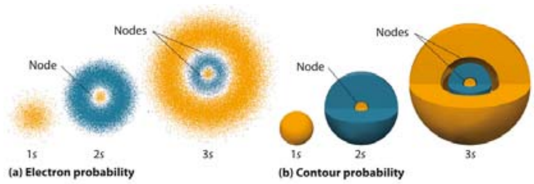
- 1926, Erwin Schrodinger used mathematics and statistics to combine deBroglie's idea of matter waves and Einstein's idea of quantized energy particles

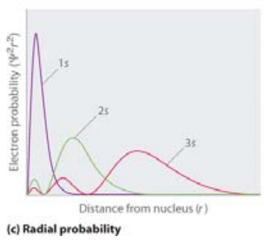
$$H\psi = E\psi$$

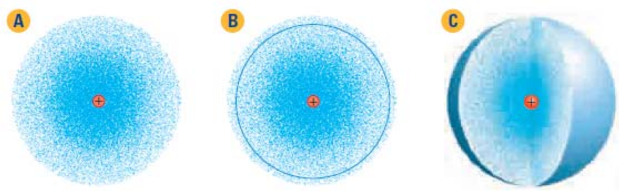
- ψ^2 was more useful – represents the probability of finding an electron in a given region within the atom

$$H(t)|\psi(t)\rangle = i\hbar\frac{\partial}{\partial t}|\psi(t)\rangle$$

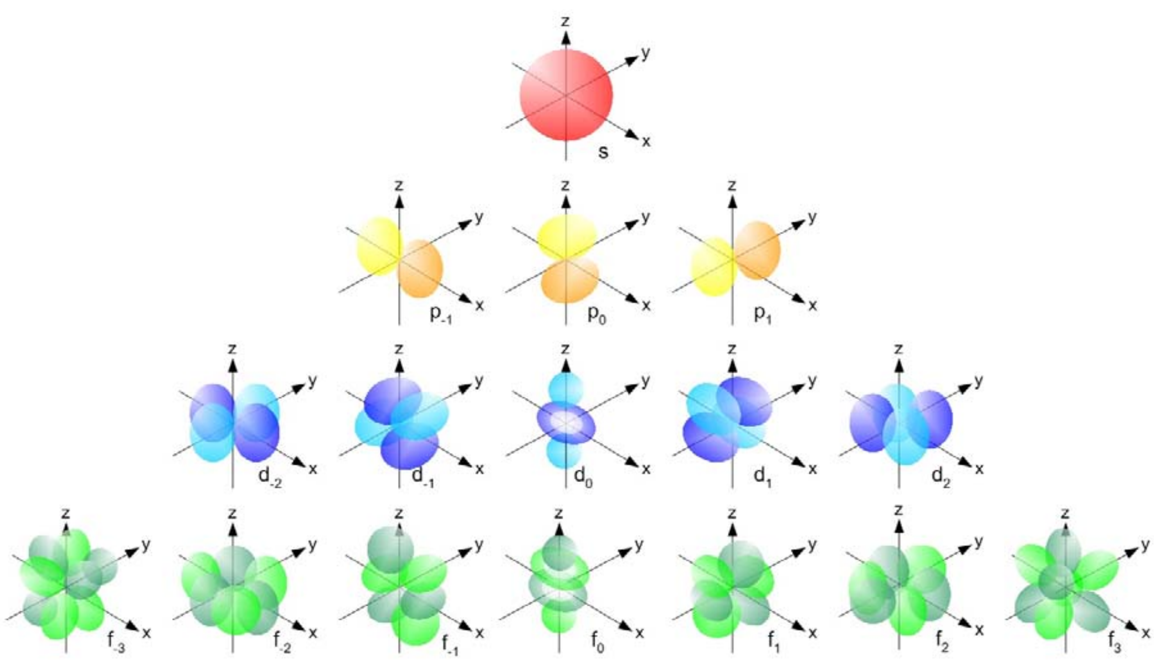
Electron Probability Density Graphs



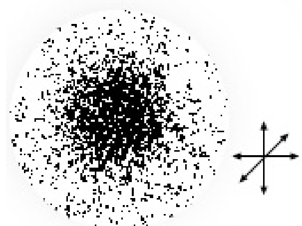




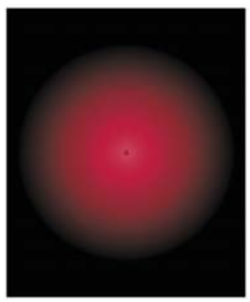
- Help chemists visualize the space in which electrons are most likely to be found around atoms
- Indicates where there is a high probability (90%) of finding electrons



Heisenberg's Uncertainty Principle



The Electron Density Plot for the Hydrogen Atom. The nucleus is in the center.



- It is impossible to know both the **momentum (*Δp*)** and the **position (*Δx*)** of a particle with certainty

$$\Delta x \Delta p \geq \frac{h}{4\pi}$$

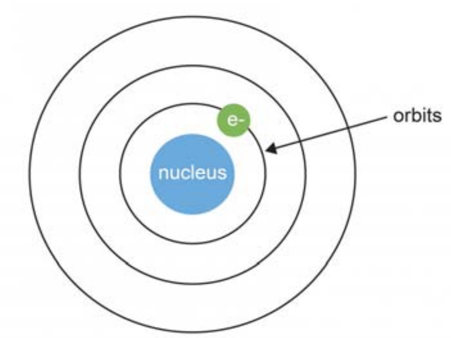
Where: h = 6.63 x 10^{-34} J•s

Summary: Quantum Model

- Describes atoms as having defined quantities of energy based on the wave-like properties of electrons
- Each electron surrounding a nucleus is described by a set of quantum numbers that describes where the electron would spend *most* of its time
- "Orbital" – refers to the three-dimensional probability distribution graphs NOT Bohr's orbitals

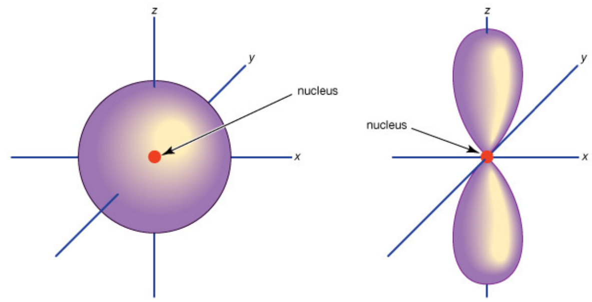
Orbits vs. Orbitals

ORBIT



- 2-dimensional
- Distance from nucleus is fixed
- Path is circular

ORBITAL

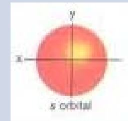
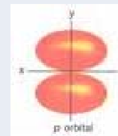
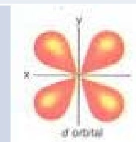
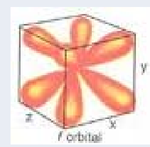


- 3-dimensional
- Distance from nucleus varies
- No set path

Quantum Numbers

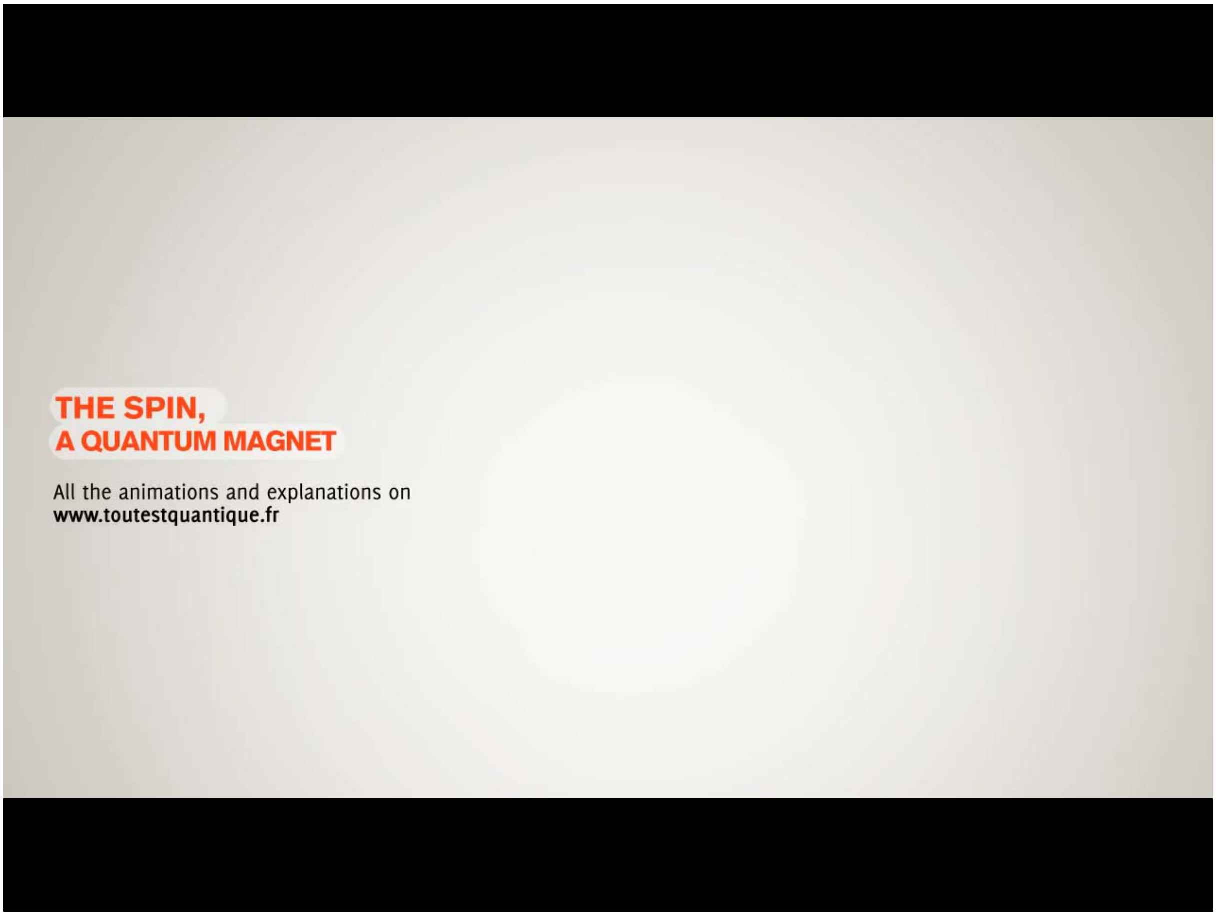
Orbitals have a variety of shapes:

1. Principal Quantum (n) – orbital size and energy level
   - $n = 1, 2, 3 \ldots \infty$
   - The max number of electrons at a given energy level is $2n^2$
2. The Angular-Momentum Quantum Number (l) – orbital shape or subshell
   - $l = 0$ to n-1 (ex: if n=3, l=0, 1, 2)

l = ?	Letter	Max # of electrons	Shape
0	s	2	
1	p	6	
2	d	10	
3	f	14	

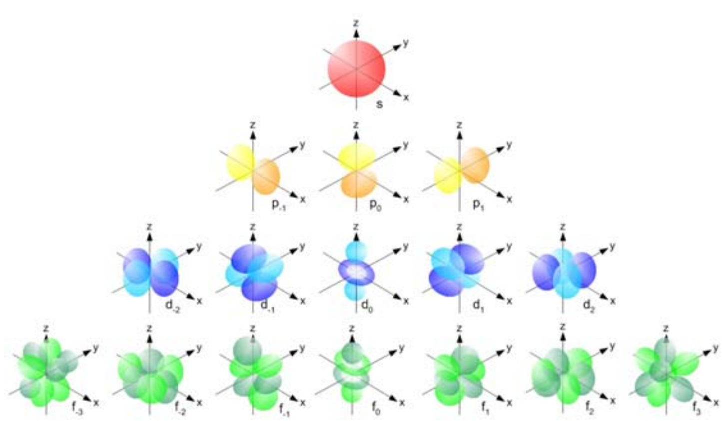
Quantum Numbers

3. Magnetic Quantum Number (m_l) – orbital orientation
   – m_l= -l to +l (ex: if l=2, m_l= -2, -1, 0, 1, 2)
4. Spin Quantum Number (m_s) – electron spin direction
   – m_s = ±½ (+½ = $\uparrow$ and -½ = $\downarrow$)
   – m_s defines the rotational direction of each of the two electrons in a given orbital
   – Independent of the other three quantum numbers



Summarizing the Four Quantum Numbers for Electrons in Atoms

Quantum Number Name	Symbol	Allowed Values	Property
principal	n	positive integers (1, 2, 3, etc.)	orbital size and energy
orbital-shape	l	integers from 0 to $(n - 1)$	orbital shape
magnetic	m_l	integers from $-l$ to $+l$	orbital orientation
spin	m_s	$+\frac{1}{2}$ or $-\frac{1}{2}$	electron spin direction



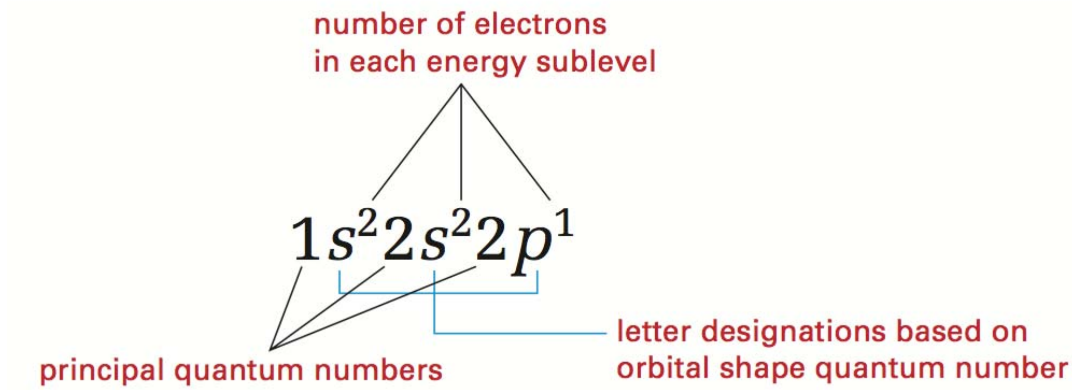
Checkpoint

If $n=3$, what are the allowed values for l and m_l and what is the total number of orbitals in this energy level?

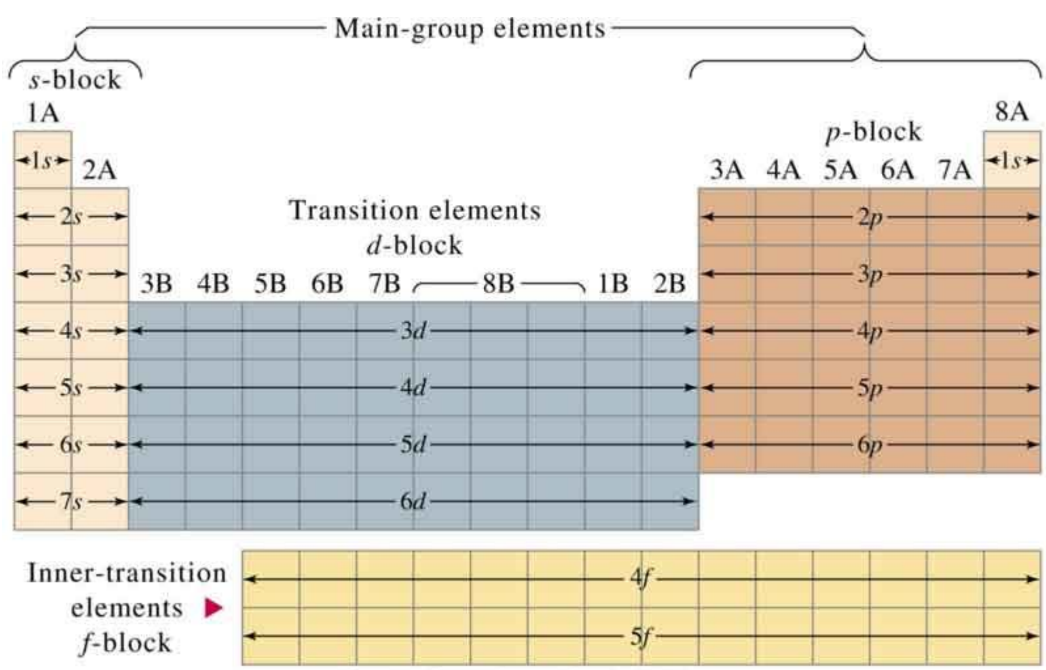
- Hydrogen has only 1 electron
  - n=1, l=0 (s), m_l=0, m_s= +½ ($1s^1$)
- Helium has 2 electrons
  - n=1, l=0 (s), m_l=0, m_s= +½
  - n=1, l=0 (s), m_l=0, m_s= -½ ($1s^2$)
- Lithium has 3 electrons
  - n=1, l=0 (s), m_l=0, m_s= +½
  - n=1, l=0 (s), m_l=0, m_s= -½
  - n=2, l=0 (s), m_l=0, m_s= +½ ($1s^2 2s^1$)
- Predict the quantum number of Beryllium's electrons

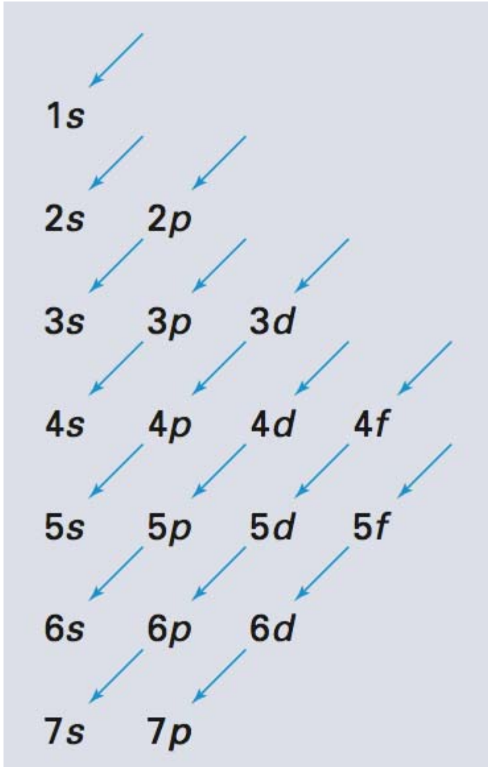
Electron Configuration

- A shorthand notation that shows the number and arrangement of electrons in its orbitals



Electron Configurations and the Periodic Table





- For main group elements, the last number of the group number is the same as the number of valence electrons
- The *n* value of the highest occupied energy level is the period number

Checkpoint

Write the electronic configuration for the following:

a) Boron

b) Silicon

c) Bismuth

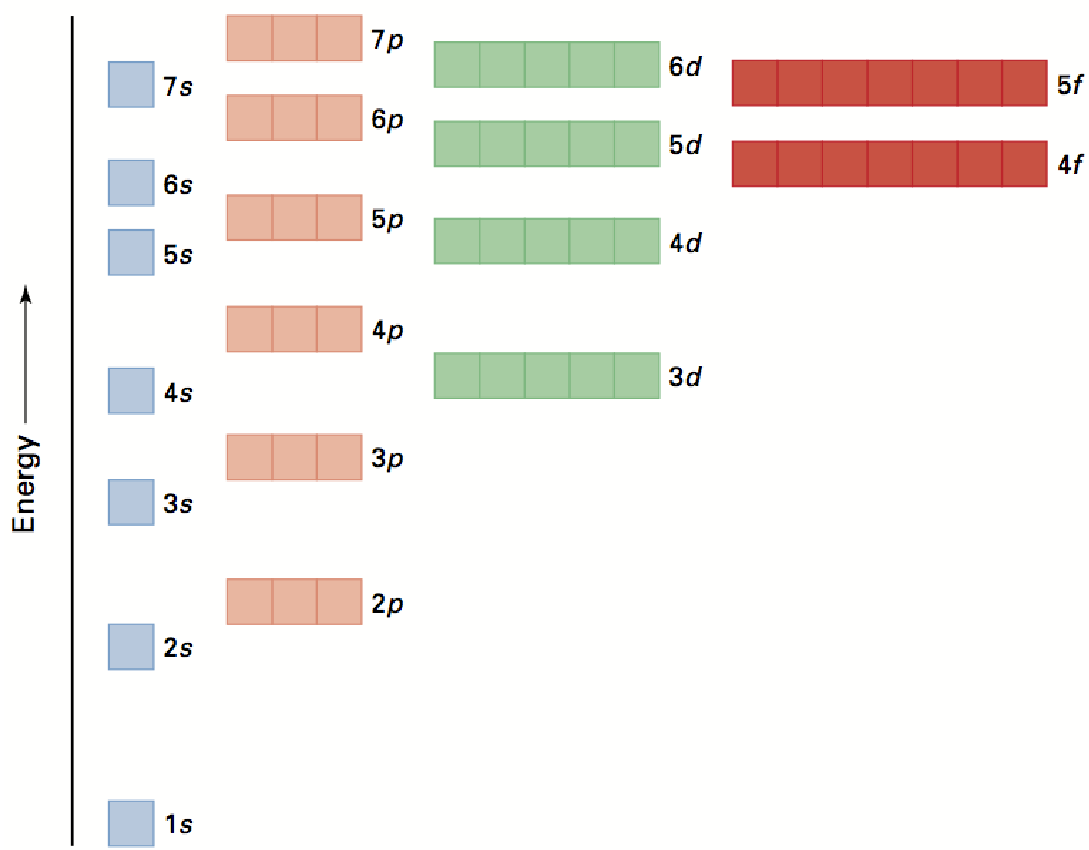
Orbital Filling Diagrams

- How do we apply quantum numbers to electrons?
  1. **Aufbau Principle** – Electrons occupy the lowest energy orbital available
  2. **Pauli Exclusion Principle** – Only two electrons of opposite spin can occupy an orbital; no two electrons in an atom can have the same four quantum numbers
  3. **Hund's Rule** – Electrons in the same subshell occupy available orbitals singly before pairing up

Orbital Filling Diagrams

1. Draw the order of subshells.
2. How many electrons are in the atom?
3. Write out the electronic configuration in the order of the subshells.
   – s → 2 electrons
   – p → 6 electrons
   – d → 10 electrons
   – f → 14 electrons

4. Draw boxes or lines to represent each subshell
   - s → 1 box/line
   - p → 3 boxes/lines
   - d → 5 boxes/lines
   - f → 7 boxes/lines
5. Fill in the electrons using up and down arrows, filling one subshell at a time. Single arrows point up by convention.



Checkpoint

- Give the electronic configuration and draw an orbital-filling diagram of:
  a) Nitrogen (Z=7)
  b) Aluminum (Z=13)
  c) Arsenic (Z=33)