

Chemistry 3A

Introductory General Chemistry

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Concepts

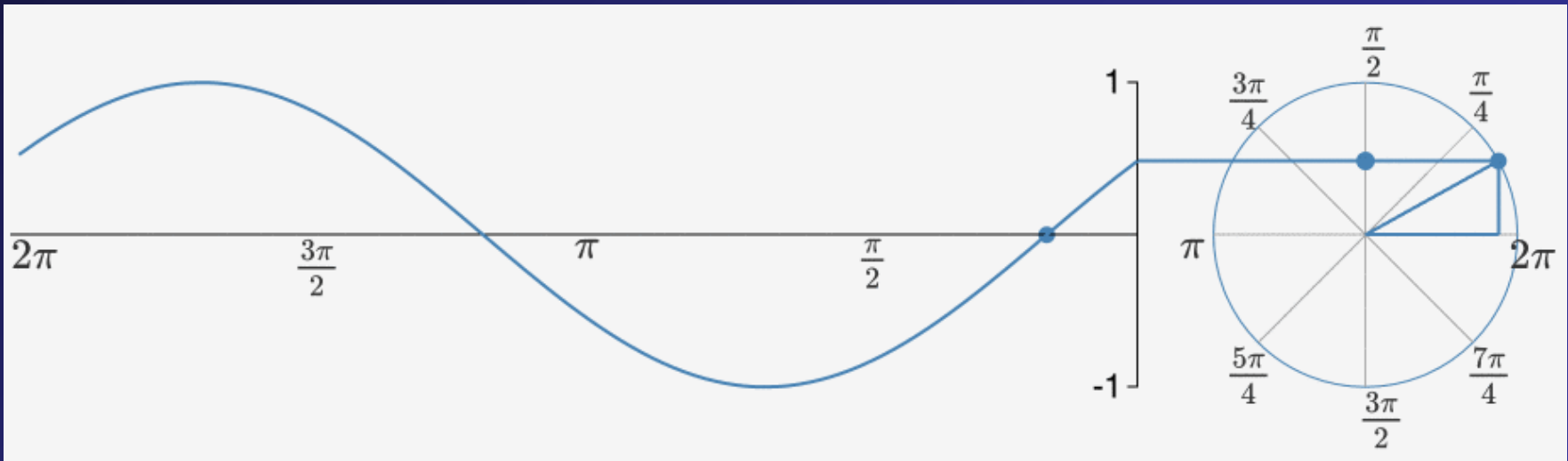
- Waves
- Observations on Atom Structure
- Electromagnetic Radiation & Spectrum
- Quantized Nature of Energy
- Energy levels of Electrons
- Quantum Mechanical Model of Atom
- Electron "Identity": Quantum Numbers
- Shells, Subshells
- Orbitals and their "Shapes"
- Populating (Filling) Electrons: Electron Configuration
- Related Trends in Periodic Table
- Atomic Radius, Ionization Energy, Electron Affinity

Waves

- Waves are a sinusoidal motion or movement

There are terms used to talk about waves

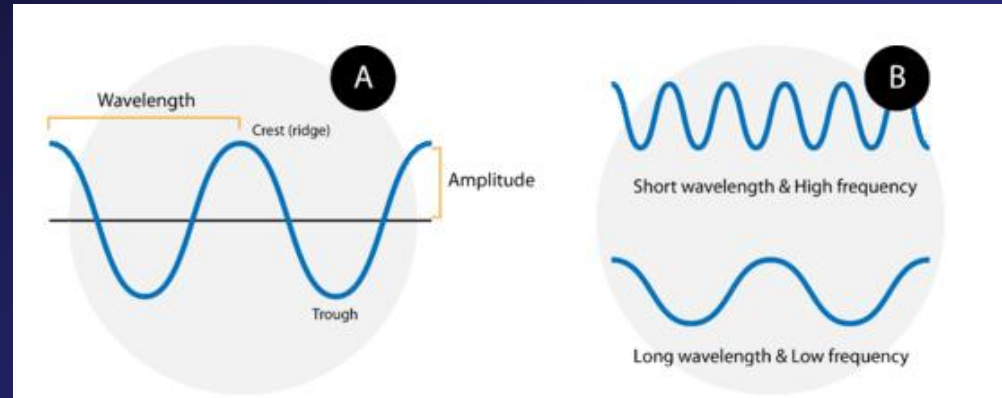
- **Crest** – the topmost point of a wave
- **Trough** – the bottommost point of a wave



Waves

Wavelength

- Distance between to corresponding points on adjacent waves
- Crest-to-crest or trough-to-trough distance
- Represented by Greek letter *lambda* (λ)
- Units: some form of meter (m, nm, cm, km)



Frequency

- The number of waves that pass a point in a certain period of time
- Represented by Greek letter *nu* (ν)
- Units: per unit time [waves per second (s^{-1})]
Also called Hertz (Hz)

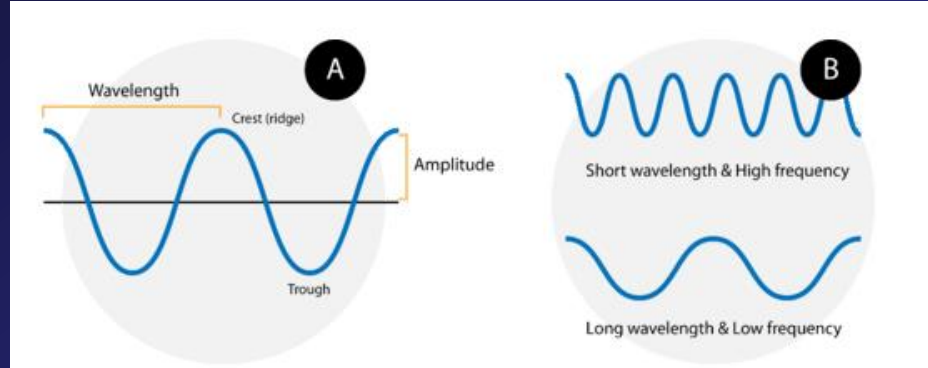
Waves

Amplitude

- maximum distance or displacement of a wave from its resting (zero) position

Also called equilibrium position

- represents height of wave's crest (or trough)
- indicator of the energy of a wave
higher amplitude → higher energy



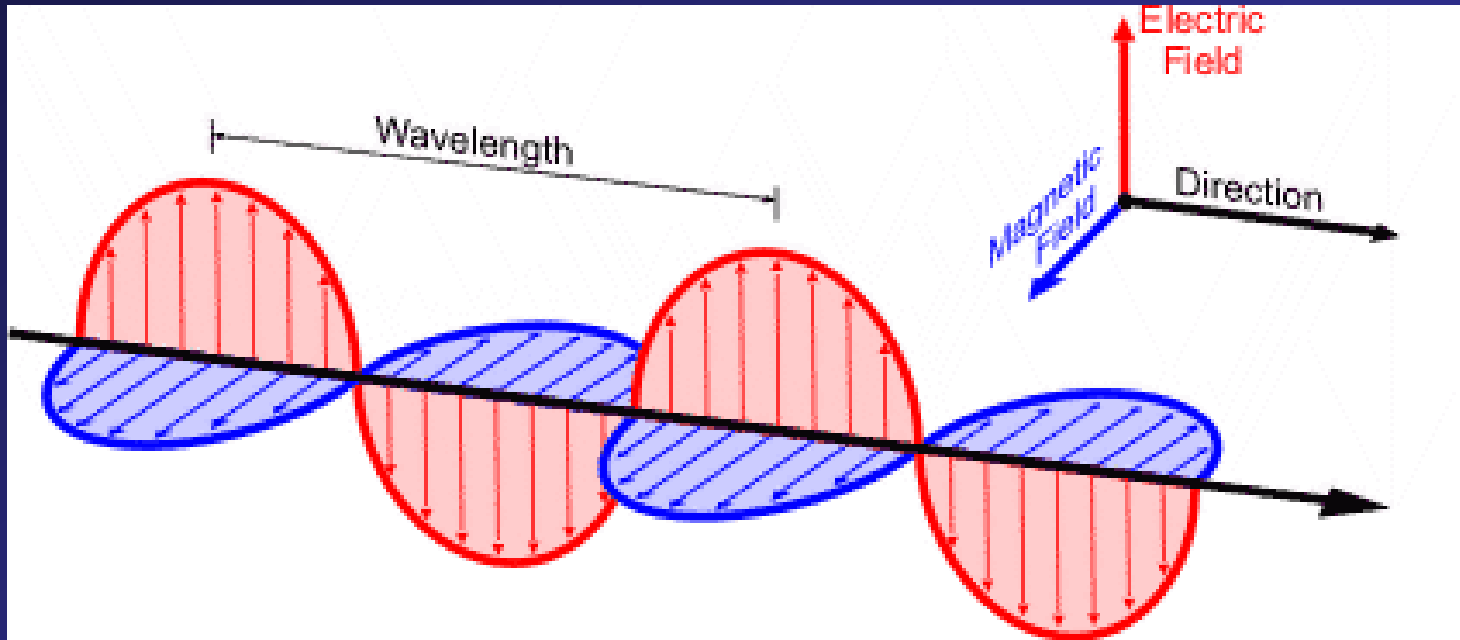
Wave speed

- distance a wave disturbance travels through a medium in a given amount of time
- Units: distance per unit time (m/s)
- indicated usually by v (velocity)
wave speed = wavelength (m) \times frequency (s^{-1})

$$v = \lambda \nu$$

Waves

- Electromagnetic radiation is a wave
- Actually a wave with **two** components perpendicular (at right angles, 90° to) each other
 - Electric field wave
 - Magnetic field wave



Waves

It is typical to talk about EM waves with respect to their **wavelength**

Frequency is required when you make energy calculations—*slide after next slide!*

In the visible light spectrum, wavelengths are from 400-700 nanometers (nm).

Color	Wavelength	Description
Red	~650 nm	Longest visible wavelength; warm and low-energy.
Green	~530 nm	Mid-spectrum; often associated with nature and balance.
Blue	~470 nm	Shorter wavelength; cooler and higher-energy.

Making A Calculation

What is the frequency of green light at 530 nm?

$$c = \lambda \nu$$

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

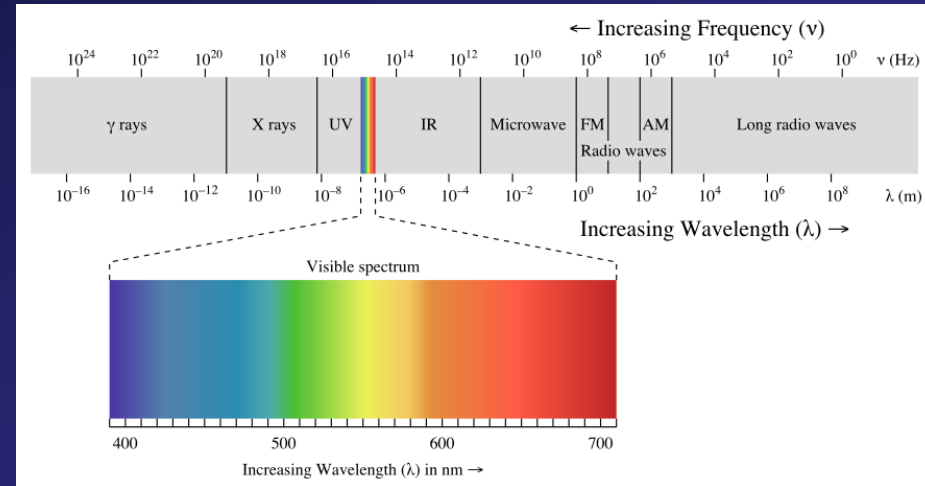
$$\nu = \frac{3.00 \times 10^8 \text{ m}}{\text{s}} \times \frac{1}{530 \text{ nm}} \times \frac{10^9 \text{ nm}}{1 \text{ m}} = 5.66 \times 10^{14} \text{ s}^{-1}$$

or $5.66 \times 10^{14} \text{ Hz}$

Algebra + conversion factors +
a natural constant (the speed of light)

Electromagnetic Spectrum

- Spans from radio waves to gamma rays
- Our eyes detect a narrow part of the EM spectrum called “visible light”



EM radiation has ENERGY that is calculated using the famous Planck's equation

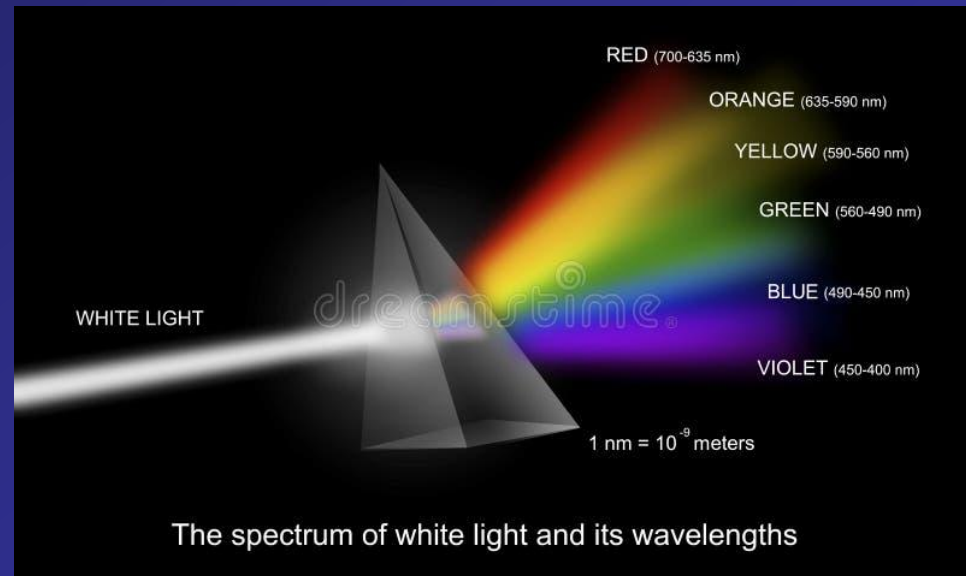
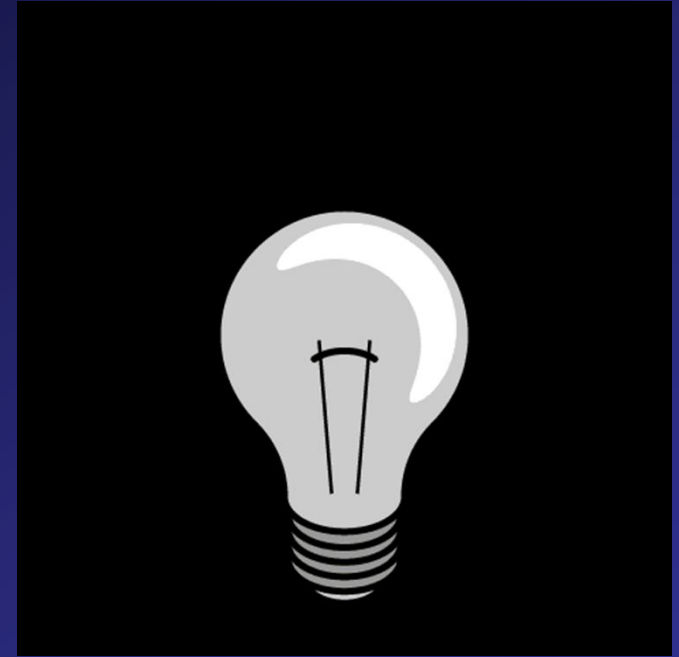
$$E = h\nu$$

EM radiation that is “ionizing” has frequencies in the UV and higher (X-rays, γ rays)

The Light Bulb

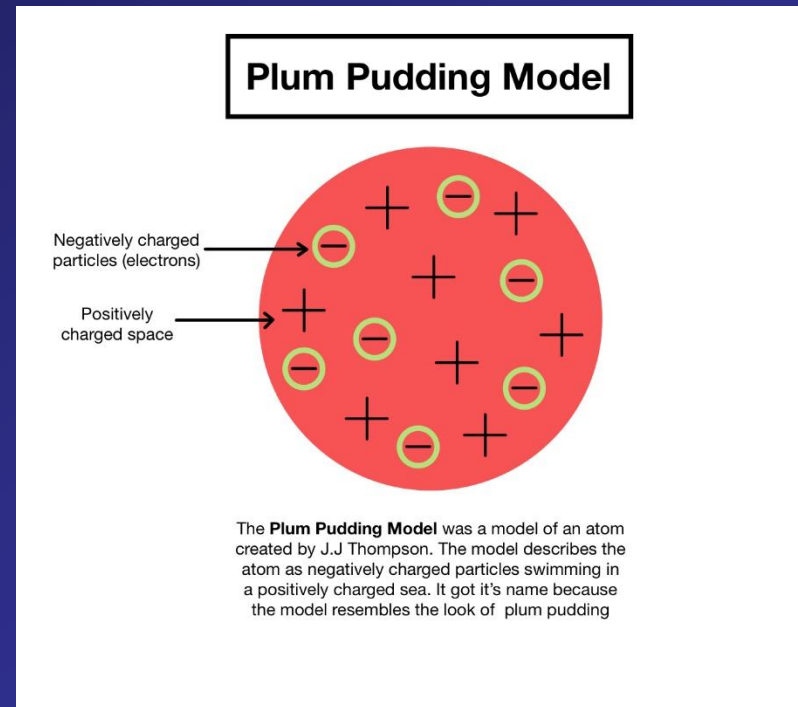
- An electric current passes through the tungsten filament
Element symbol W
- The filament resists the flow, heating to $\sim 3000\text{ }^{\circ}\text{C}$
- At this high temperature, it emits **blackbody radiation**
- The spectrum spans visible wavelengths, so the glow appears white light

⚡ This is not from tungsten's atomic emission spectrum



Historical Observations

- 1897 JJ Thomson discovers the electron is a negatively charged particle
- Proposes “plum pudding” model of atom
- Electrons and protons in a mix (cloud)
- No nucleus



Historical Observations

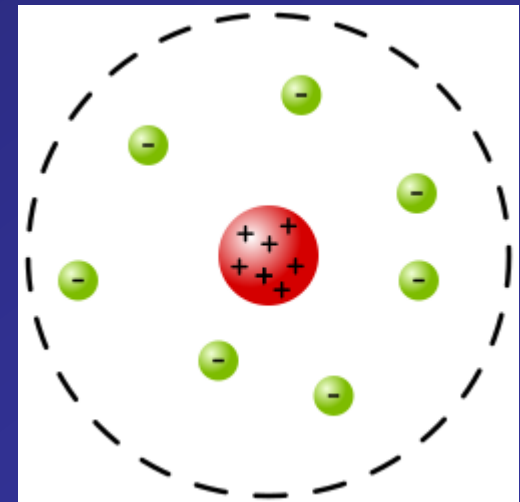
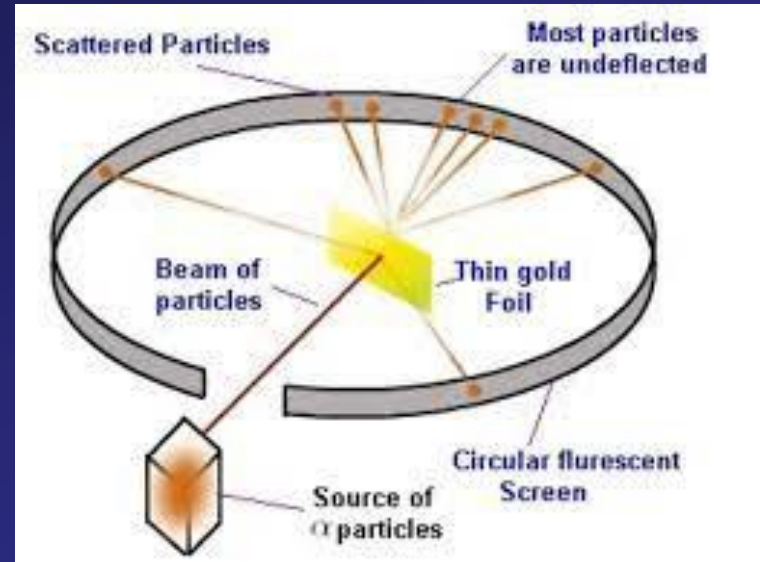
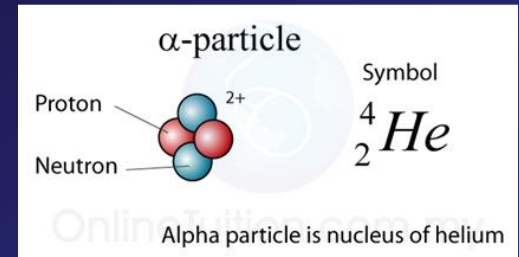
1909 Rutherford aimed alpha particles at thin gold foil and he sees that they pass through the foil and some are deflected

Alpha particles are positively charged helium atoms (${}^4\text{He}$)

Important findings

- Gold foil atoms must have space in them or the alpha particles would not pass through
- The deflections would have to be positive charges in the gold foil, since positive charges repel

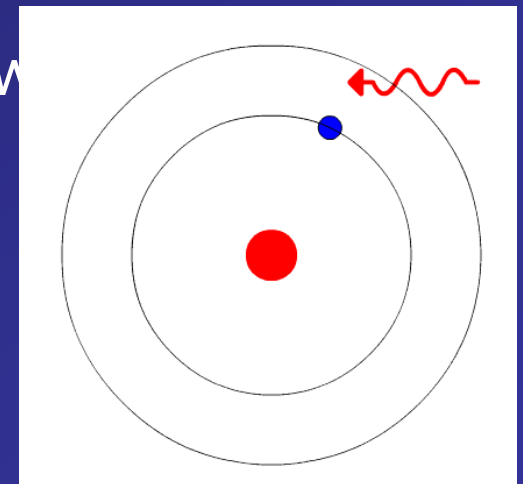
Rutherford proposes a positively charged "nucleus" with electrons outside of it (did not describe orbits)



Atoms & Orbiting Electrons

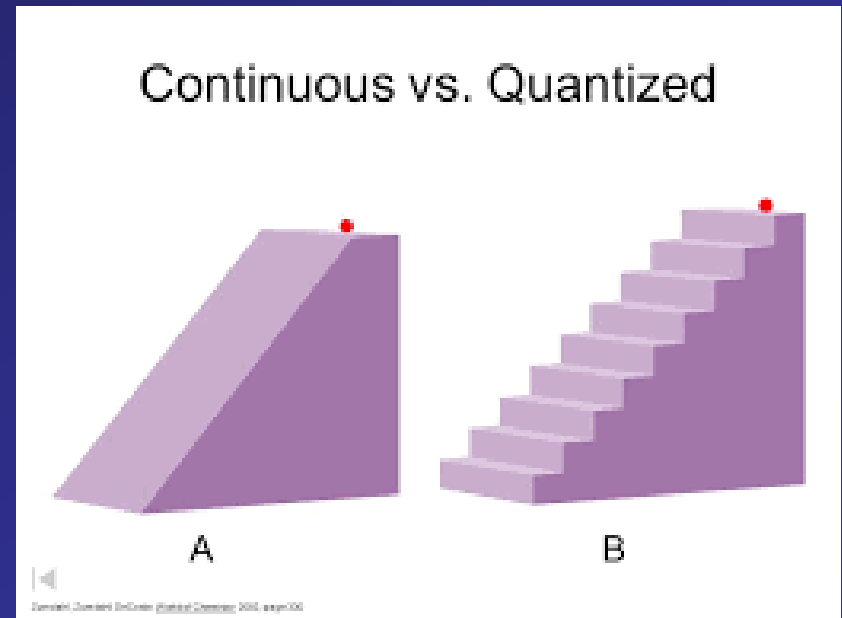
The Bohr Model

- Electrons orbit the nucleus in fixed energy levels
- Electrons can transition between energy levels by absorbing or emitting a photon as a quantum of energy
- Transition between a ground state and excited state
- Explained the hydrogen spectrum with mathematical precision.



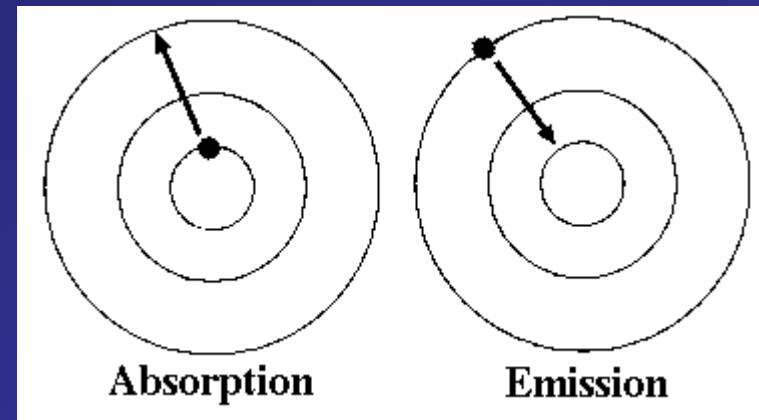
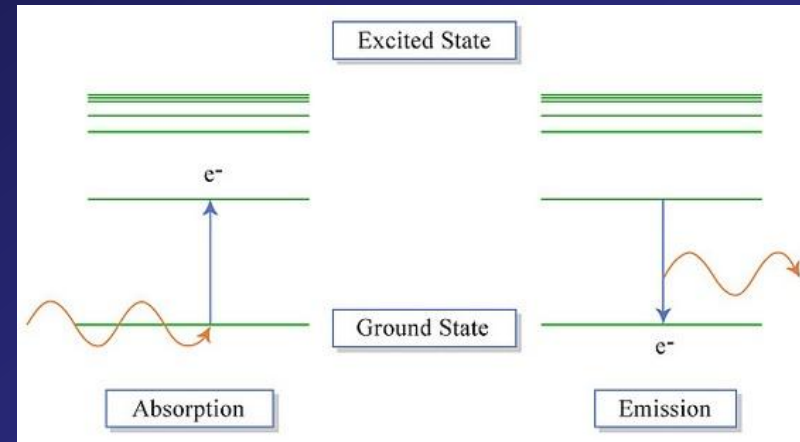
Quantized vs Continuous

- Energy level transitions in atoms are **quantized**
- Electrons cannot absorb just any amount of energy
- They can only move between specific energy levels with exact energy differences
- If incoming photon does not have required amount of energy, no transition occurs



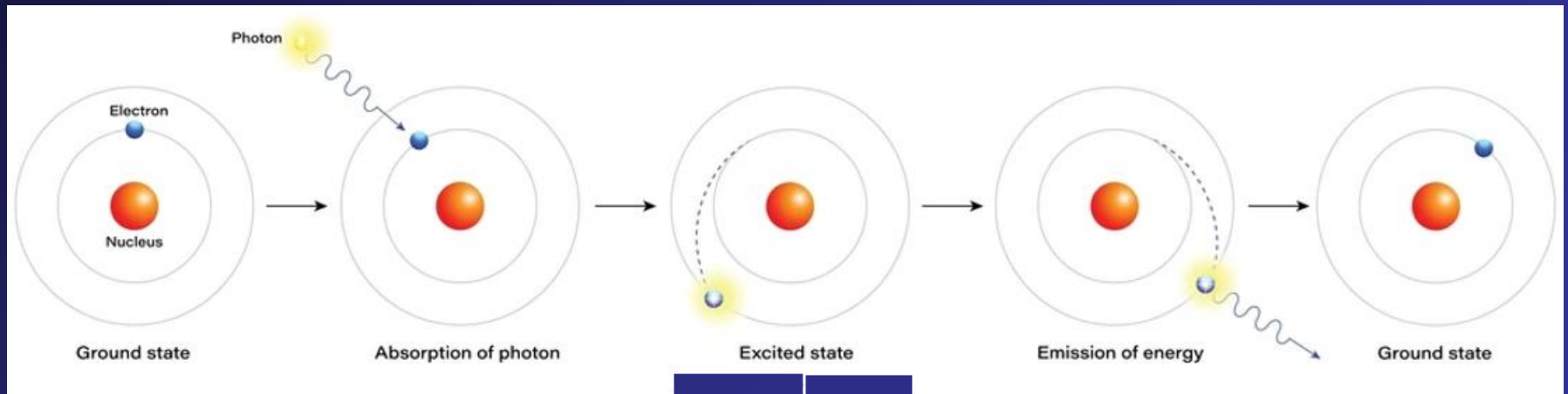
Electrons & Energy Levels

- An electron in an atom moves in an orbit about the nucleus of the atom
- Its usual orbit is in a **ground state**, its lowest energy state
- When given energy through the "**absorption**" of a **photon** (EM radiation), it jumps to another a higher energy level, an orbit further away from the nucleus. It is in an **excited state**.
- When it loses its energy back to its ground state, it **emits** a photon.



Electrons and EM Radiation

- This cycle of absorbing photon in the **ground state** and being excited (the **excited state**) to an energy level and then relaxing or losing that energy back to the ground state with **emission** of a photon shown below



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Emission Spectrum

These electron energy level transitions explain the **emission spectra** observations

- Hydrogen atom has only a few energy transitions
- Iron atom has numerous

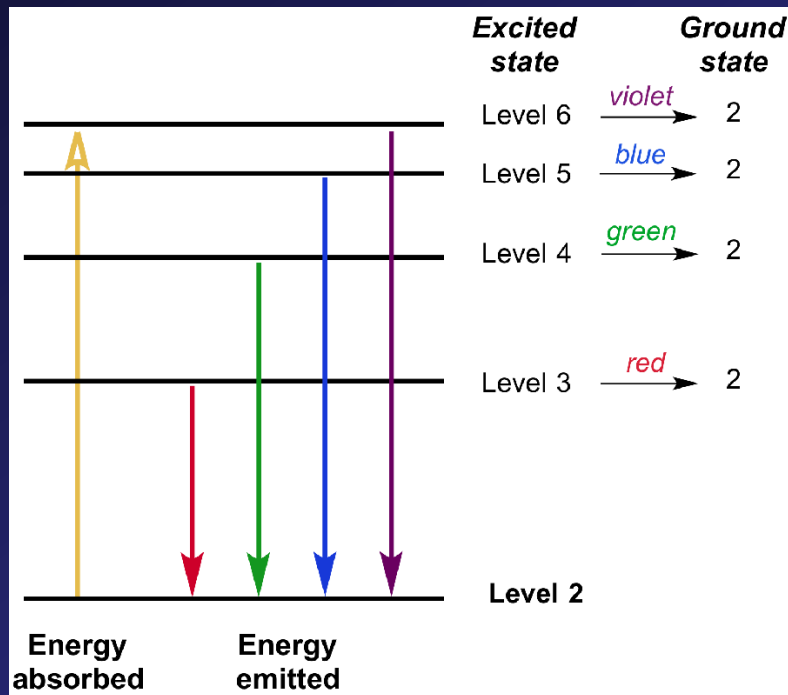


Figure 3.2.1.2: Atomic Emission Spectrum of Hydrogen.

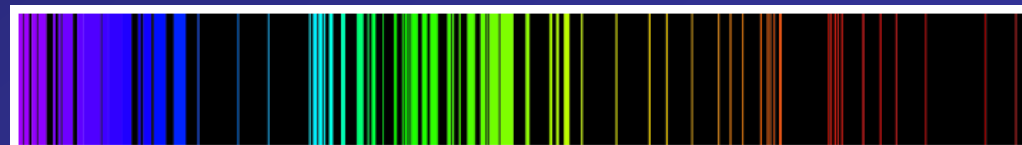
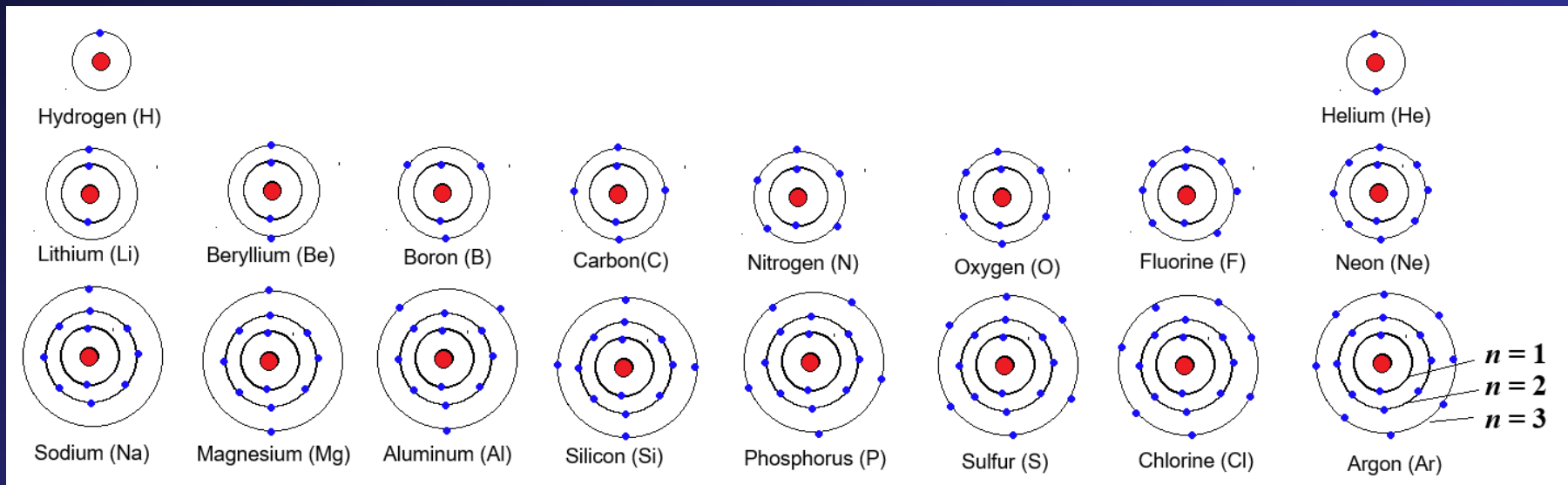


Figure 3.2.1.3: Atomic Emission Spectrum of Iron.

Quantum Mechanics Model

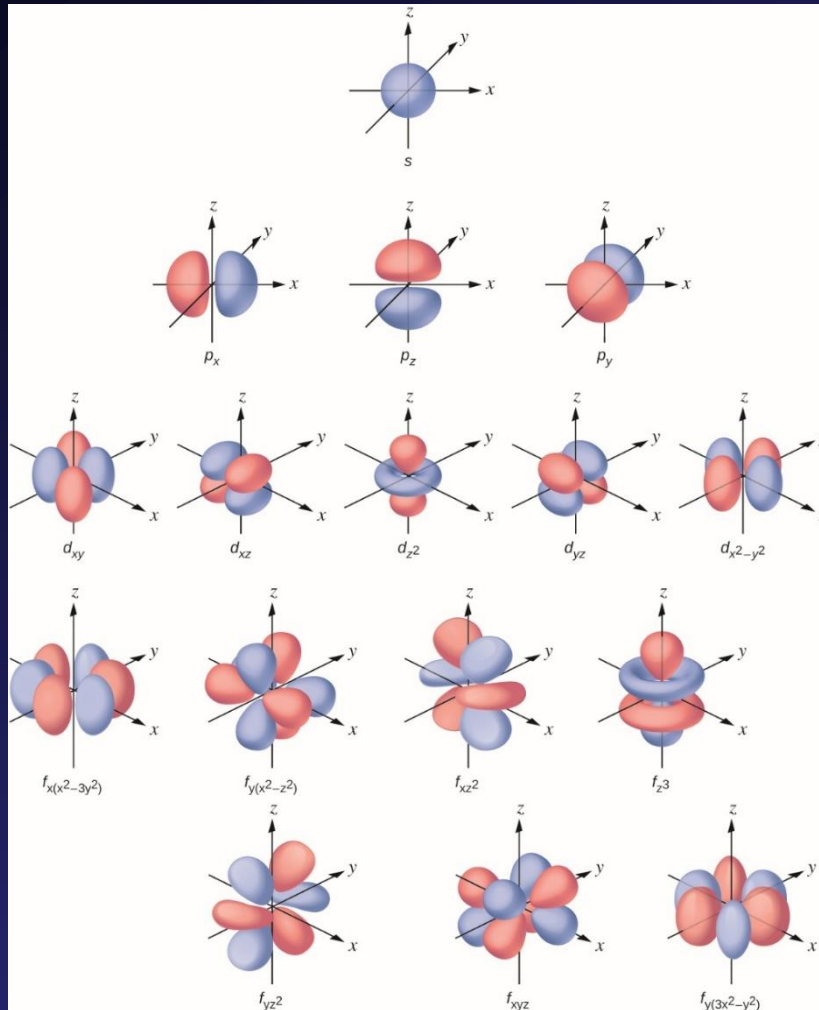
- Following Bohr's explanations, a model for how electrons exist
- Electrons have an identity: quantum numbers
- There are 4 quantum numbers for each electron



Quantum Numbers: Meaning

- Principal Quantum Number (n)
 - describes main energy level or shell the electron is in
 - the bigger the number (1,2,3,...), the farther the electron is from the nucleus and the higher its energy
- Angular Momentum Quantum Number (l)
 - describes shape of the orbital, or the region where the electron is likely to be found
 - the shapes are given letter names (come from spectroscopy):
 s (sphere), p (dumbbell), d (cloverleaf), f (double cloverleaf)
- Magnetic Quantum Number (m_l)
 - Describes the orientation of the orbital in 3D space
 - For a dumbbell-shaped p-orbital, this number tells you if it's on the x, y, or z-axis
- Spin Quantum Number (m_s)
 - Describes an electron's intrinsic "spin"
 - There are only two possible values: spin up (+1/2) or spin down (-1/2)
 - Every orbital can hold a maximum of two electrons
 - They must have opposite spins

Orbitals and Electron Configurations



Orbital shapes are actually mathematical calculations describing the space where an electron can be found

Probabilities functions

Depending on the value of n (the shell), there is possible

- only one s orbital
- three p orbitals
- five d orbitals
- seven f orbitals

Each orbital can have only two electrons

Orbitals and Electron Configurations

Table 3.3.1: Atomic shell and subshell structure with the number of electrons in each

Shell	Number of Subshells	Names of Subshells	Number of Orbitals (<i>per Subshell</i>)	Number of Electrons (<i>per Subshell</i>)	Total Electrons (<i>per Shell</i>)
1	1	<i>1s</i>	1	2	2
2	2	<i>2s</i> and <i>2p</i>	1, 3	2, 6	8
3	3	<i>3s</i> , <i>3p</i> , and <i>3d</i>	1, 3, 5	2, 6, 10	18
4	4	<i>4s</i> , <i>4p</i> , <i>4d</i> , and <i>4f</i>	1, 3, 5, 7	2, 6, 10, 14	32

Why these patterns in this detail?

- Because it explains the nature of the atoms and Periodic Table
- It explains the **periods** (rows) and the **groups** (columns)
- This will be an exercise in memorization of these patterns of matter, nature, and the atom

Atomic Structure Glossary

Term	Definition	Symbol/Example
Shell	A major energy level in an atom, defined by principal quantum number n . All electrons with same n are in same shell	$n = 1, 2, 3, \dots$
Subshell	Subdivision of shell defined by the azimuthal quantum number l . Determine the shape of the orbital	$l = 0$ (s), $l = 1$ (p), $l = 2$ (d), $l = 3$ (f)
Energy Level	Used synonymously with shell (but can refer to quantized energy associated with electron's position in atom)	Can also refer to ground state vs excited state
Energy Sublevel	Informal term for subshell and it emphasizes the energy hierarchy within a shell	$3s < 3p < 3d$ (within shell $n = 3$)
Orbital	A region of space where there is high probability of finding an electron. Defined by (n, l, m_l)	Examples: $2p_x$, $3d_{xy}$

Quantum Number Relationships

More info for emphasis

Shell → defined by n

Subshell → defined by n and l

Orbital → defined by n, l, m_l

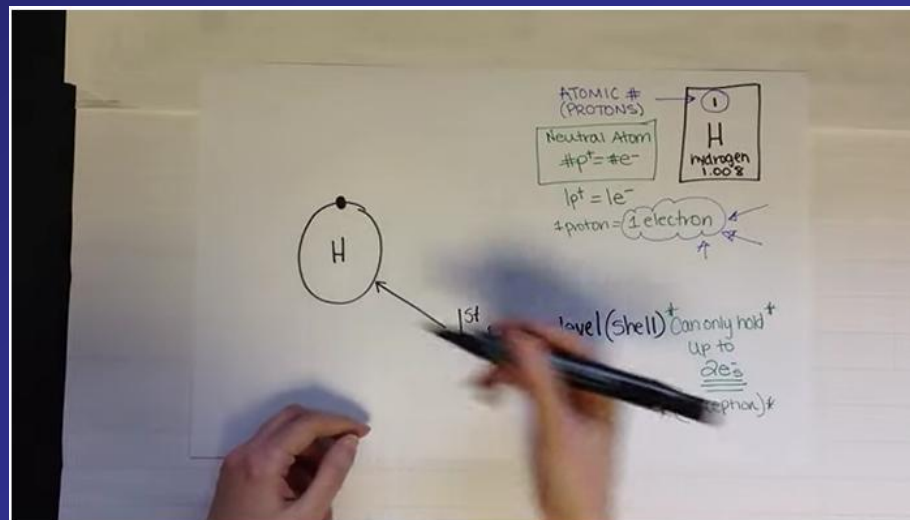
Electron spin → defined by m_s , either $+\frac{1}{2}$ or $-\frac{1}{2}$

This use of many terms is indicated to you because a 9-minute video in your book also uses certain terms

Visual Analogy

Think of it like a building:

- **Shell** = floor number
- **Subshell** = room type (s, p, d, f)
- **Orbital** = specific room
- **Electron** = person in the room, with spin direction like facing left or right



Video 3.3.1: Energy levels, sublevels and orbitals.

Orbital Filling

Aufbau Principle

In “filling” the atom’s orbitals so that electron’s are added in order of lowest to higher energy, use the $n + l$ Aufbau Principle. The lower value gets filled first (“Aufbau” is “build up” in German)

WAIT!!

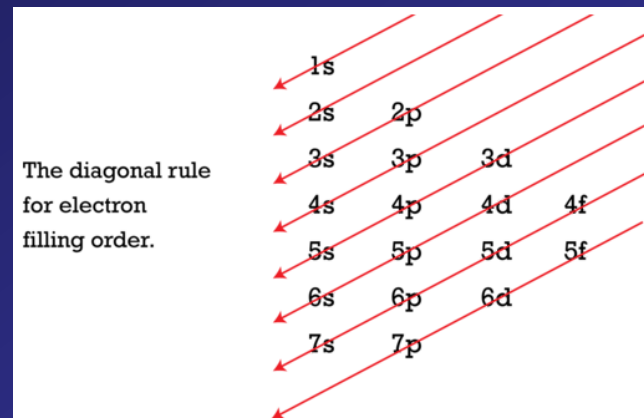
Shouldn't $n = 4$ have higher energy level than $n = 3$, so why fill $4s$ before $3d$!!??

Well...it's really $n + l$

$$4s = 4 + 0 = 4$$

$$3d = 3 + 2 = 5$$

So $4s$ gets filled first



shape	l
s	0
p	1
d	2
f	3

Quantum Number Uniqueness

Pauli Exclusion Principle

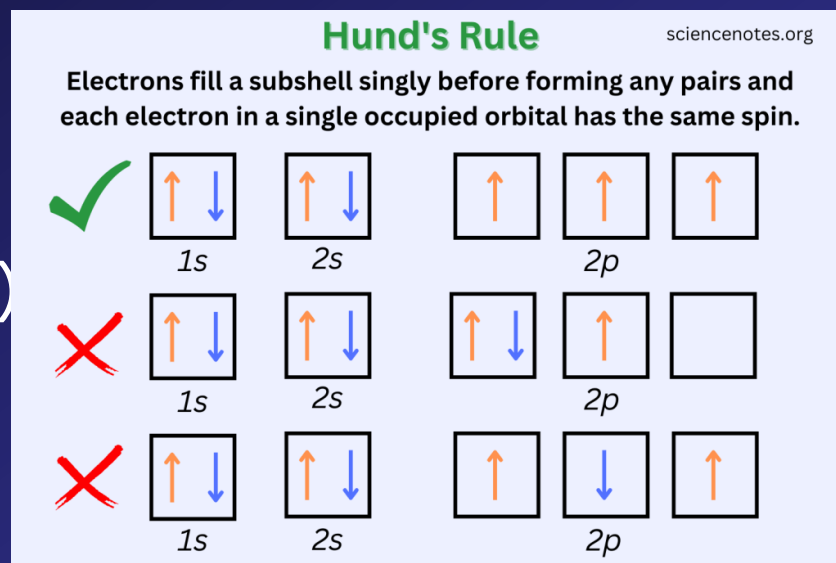
Remember those four quantum numbers that indicate the “identity” of an electron in the atom?

The Pauli Exclusion Principle is just a rule that states that no two electrons in an atom can have the same four quantum numbers, the same “identity”

Unpaired vs Paired Electrons

Hund's Rule

- Electrons can be “paired” into orbitals (2 e⁻ per orbital)
They must have opposite spins
- But pairing requires an energy input
- For orbitals of the same energy level (degenerate orbitals), fill the orbitals FIRST with one electron, and as filling proceeds, then pair them
- This applies to *p* (3 degenerate orbitals), *d* (5), and *f* (7). The *s* orbital has only one energy level



Electron Configurations

- The electrons that will complete a particular atom are presented in a particular format called an **electron configuration**
- The format is

<n-number><l-letter><# of electrons superscript>

The **n**-number will be shell number: 1, 2, 3, 4, 5, 6, 7

The **l**-letter will be subshell designation: **s, p, d, f**

The # of electrons in superscript are the range of electrons possible for the subshell: **s** $\rightarrow 1 \times 2 = 2$, **p** $\rightarrow 3 \times 2 = 6$, **d** $\rightarrow 5 \times 2 = 10$, **f** $\rightarrow 7 \times 2 = 14$

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4	4	4s, 4p, 4d, and 4f	1, 3, 5, 7	2, 6, 10, 14	32

