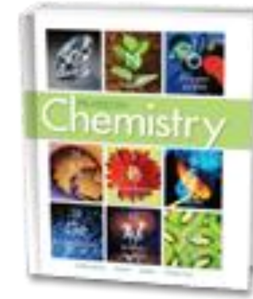
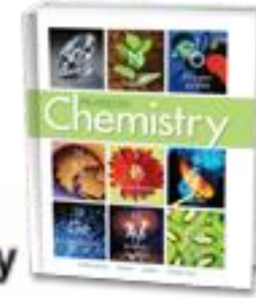


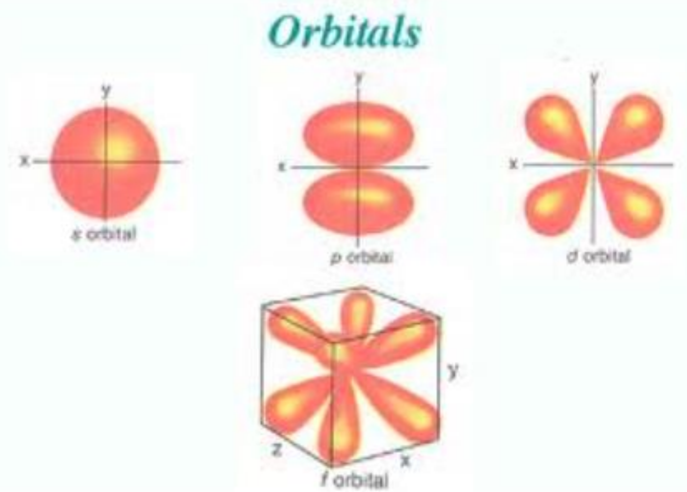
Chapter 5
Electrons In Atoms

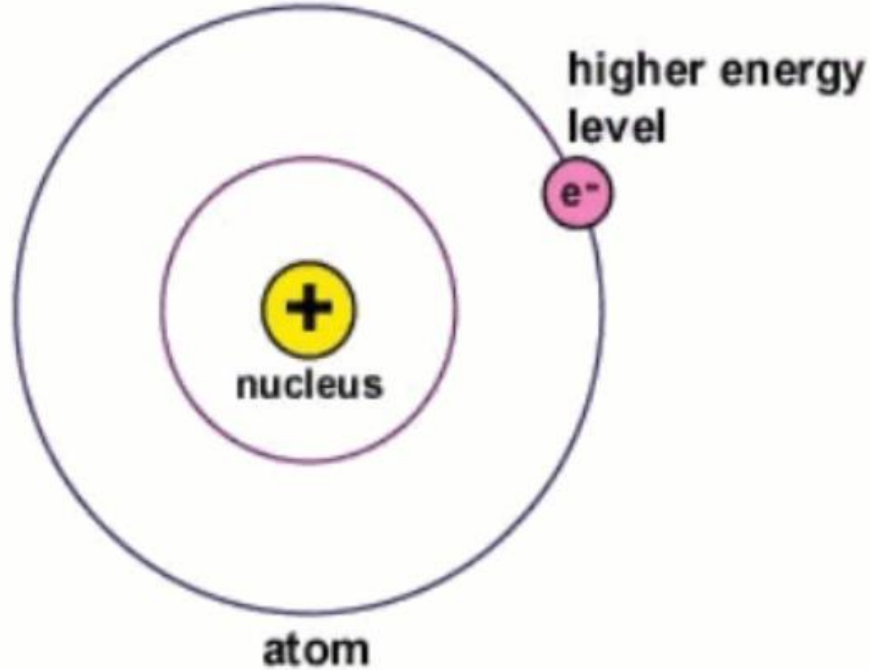
Revising the Atomic Model

Electron Arrangement in Atoms

Atomic Emission Spectra and the Quantum Mechanical Model







Topics:

1. Modern Atomic Theory: Electrons in Atoms

Objectives:

1. Describe contributions to the revised atomic theory (Bohr, DeBroglie, Shroedinger, Heisenberg, wave-particle duality, Photoelectric effect, Absorption & Emission Spectra)
2. Explain and calculate the relationship of wavelength, frequency and energy of emitted light related to changes in electron energies.
3. Understand Quantum Mechanics model of the atom and write electron configurations of elements. Give the 4 quantum numbers of elements 1 -11.

WARM-UP

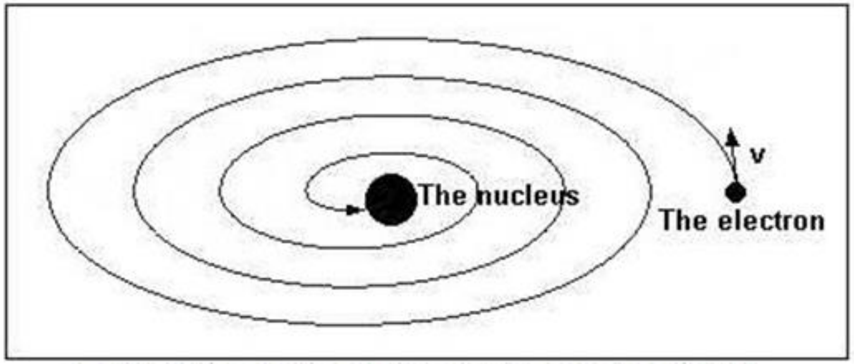
“Models of the Hydrogen Atom”

showing the History of Atomic Theory.

https://screencast-omatic.com/watch/cD6ZXZj5Ma

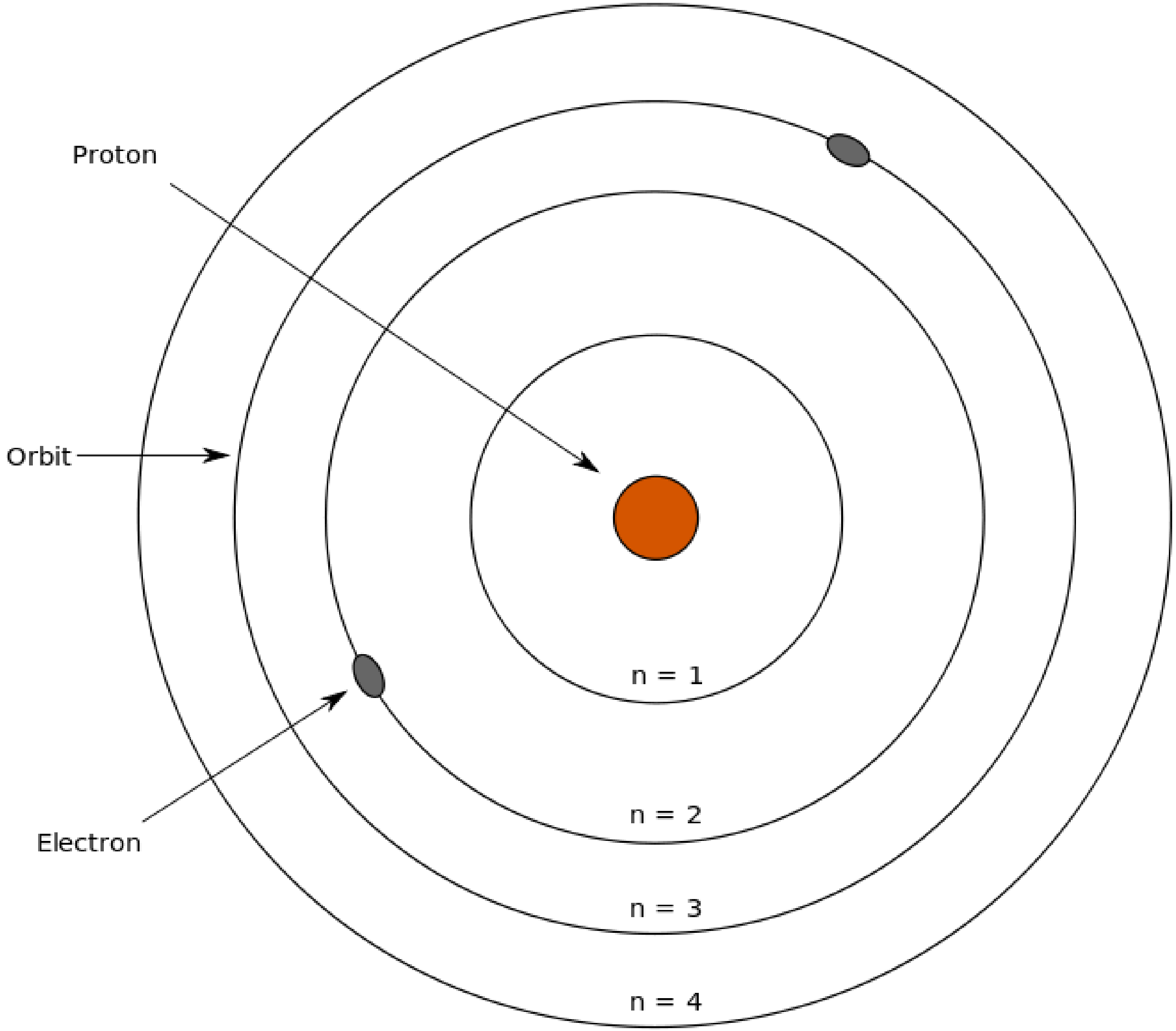
Limitations of Rutherford's Atomic Model

- Electrons are charged particles (unlike planets).
- An accelerating electric charge would steadily lose energy and spiral in, toward the positively charged nucleus, colliding with it in a fraction of a second.



- *Rutherford's model could not explain the highly peaked emission and absorption spectra of atoms that were observed.*

How was the modern understanding of the atom developed?



THINK ABOUT IT

Determine and explain the commonality of the following and how each may relate to Atomic Structure:

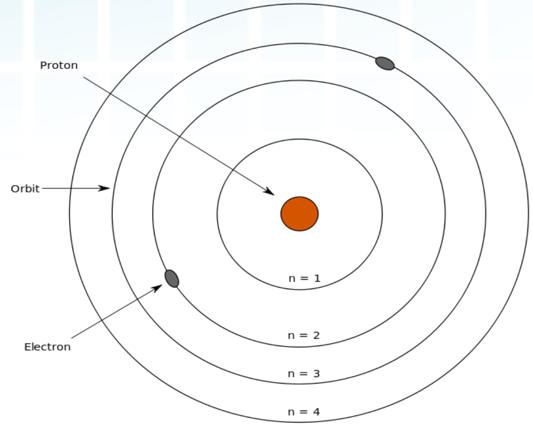
Staircase or ladder or bleacher

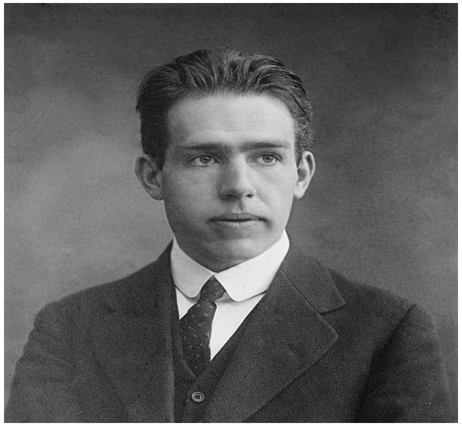
Pitches in a major scale

Radio stations on the AM or FM scale

You come up with your own example that fits

The Bohr Model





In 1913, Niels Bohr (1885–1962), a young Danish physicist and a student of Rutherford, developed a new atomic model.

Bohr proposed that an electron is found only in specific circular paths, or **orbits**, around the nucleus.

Bohr's model only worked for the simple Hydrogen atom and his perspective was still "Newtonian" based (electrons are particles).

The Bohr Model

shows that atoms and molecules can only exist in certain energy states so he designated the electron orbitals as **energy levels** or **quantum levels**.

A change in the energy level of such a system involves the absorption or emission of a definite amount (QUANTA) of energy.

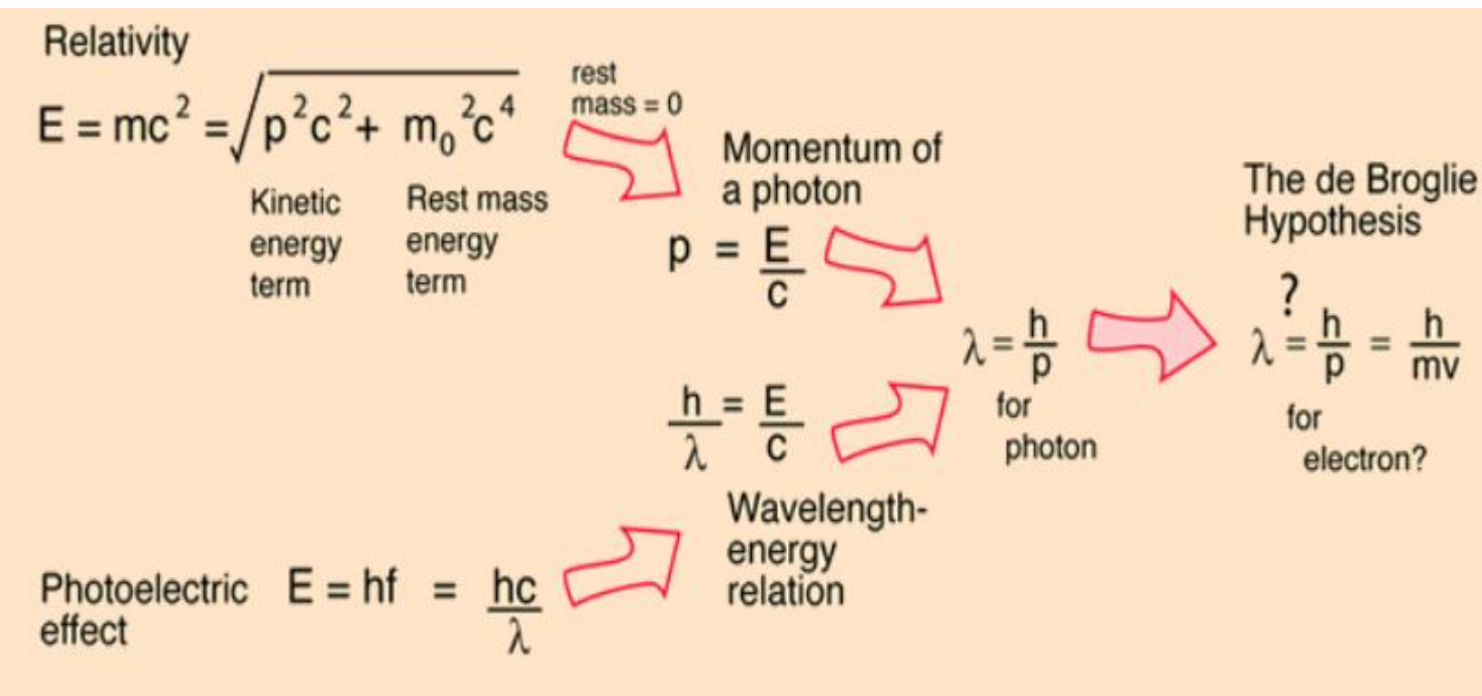
The rungs on this unusual ladder are somewhat like the energy levels in Bohr's model of the atom.

One can only stand ON the rungs of a ladder. ***Similarly, the electrons in an atom cannot exist between energy levels.***

Louis DeBroglie

Proposed that all matter and slow moving particles (e.g. electrons) have a dual nature.

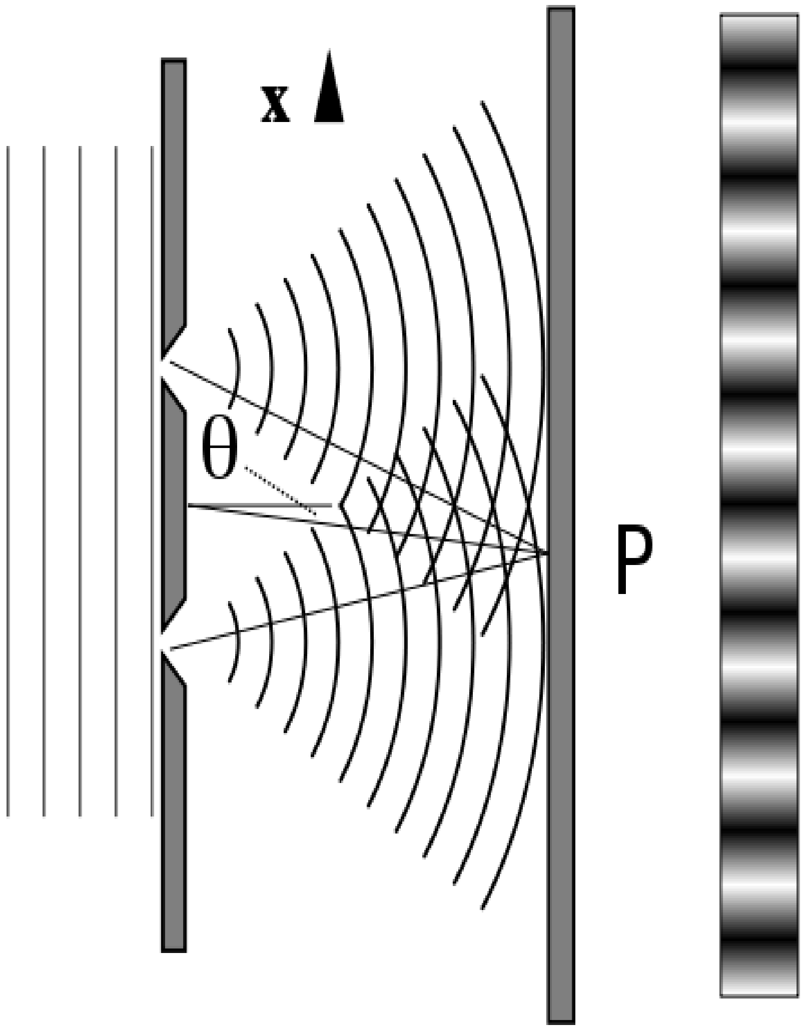
Electrons can behave as **waves (wavelength, reflection, refraction, diffraction**) or **particles (momentum)**.



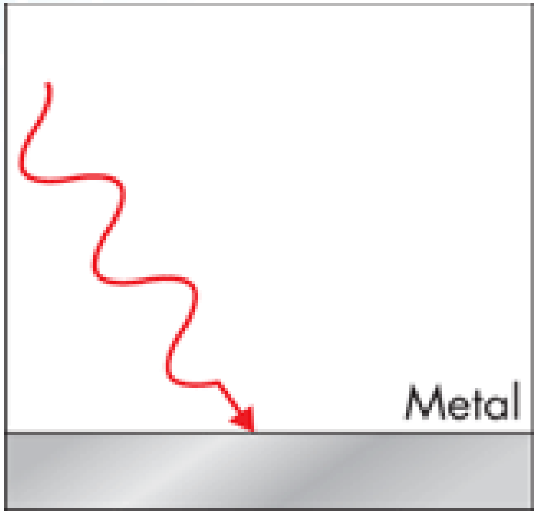
Modification: Wave-Particle Duality of Light

Major Theories of Light:

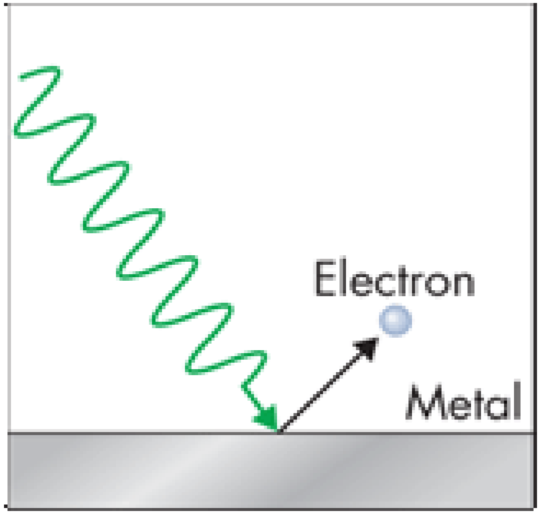
- Newton (1704): Light behaves as a particle.
  - *With mass, acceleration, action/reaction.*
- Young (*early 1800s*): Light behaves as a wave.
  - *Reflection, refraction,* ***diffraction in double-slit experiment*** →



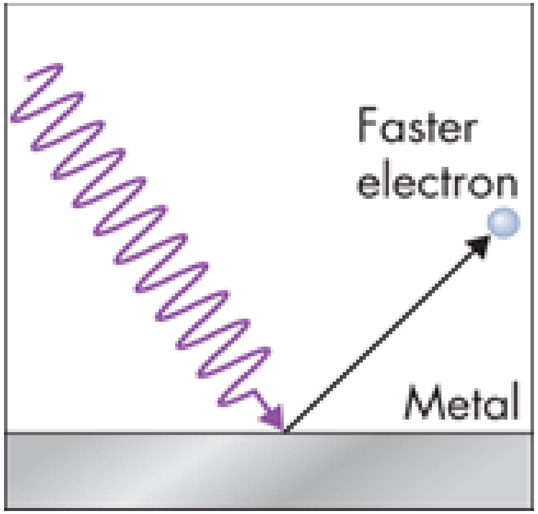
The Photoelectric Effect → Shows "Quanta"



No electrons are ejected because the frequency of the light is below the threshold frequency.



If the light is at or above the threshold frequency, electrons are ejected.



If the frequency is increased, the ejected electrons will travel faster.

e.g. garage door opener; remote controls

Photoelectric Effect: Discrete Particle & Wavelength

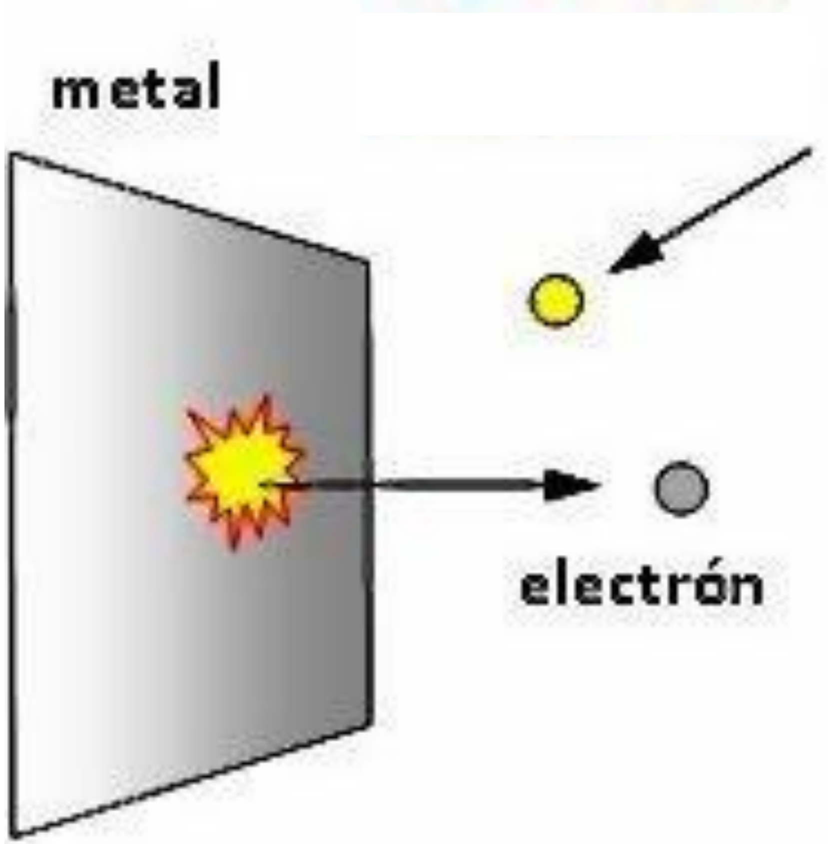
Einstein explained the photoelectric effect incorporating Planck's particle view (quantum) while proposing that light being a wave behaves as "Photons" or bundle of energy.

Every Photon has a quantized amount of energy as described by $\boldsymbol{E = hv}$.

E = energy of photon

h = Planck's constant

v = frequency



THINK ABOUT IT

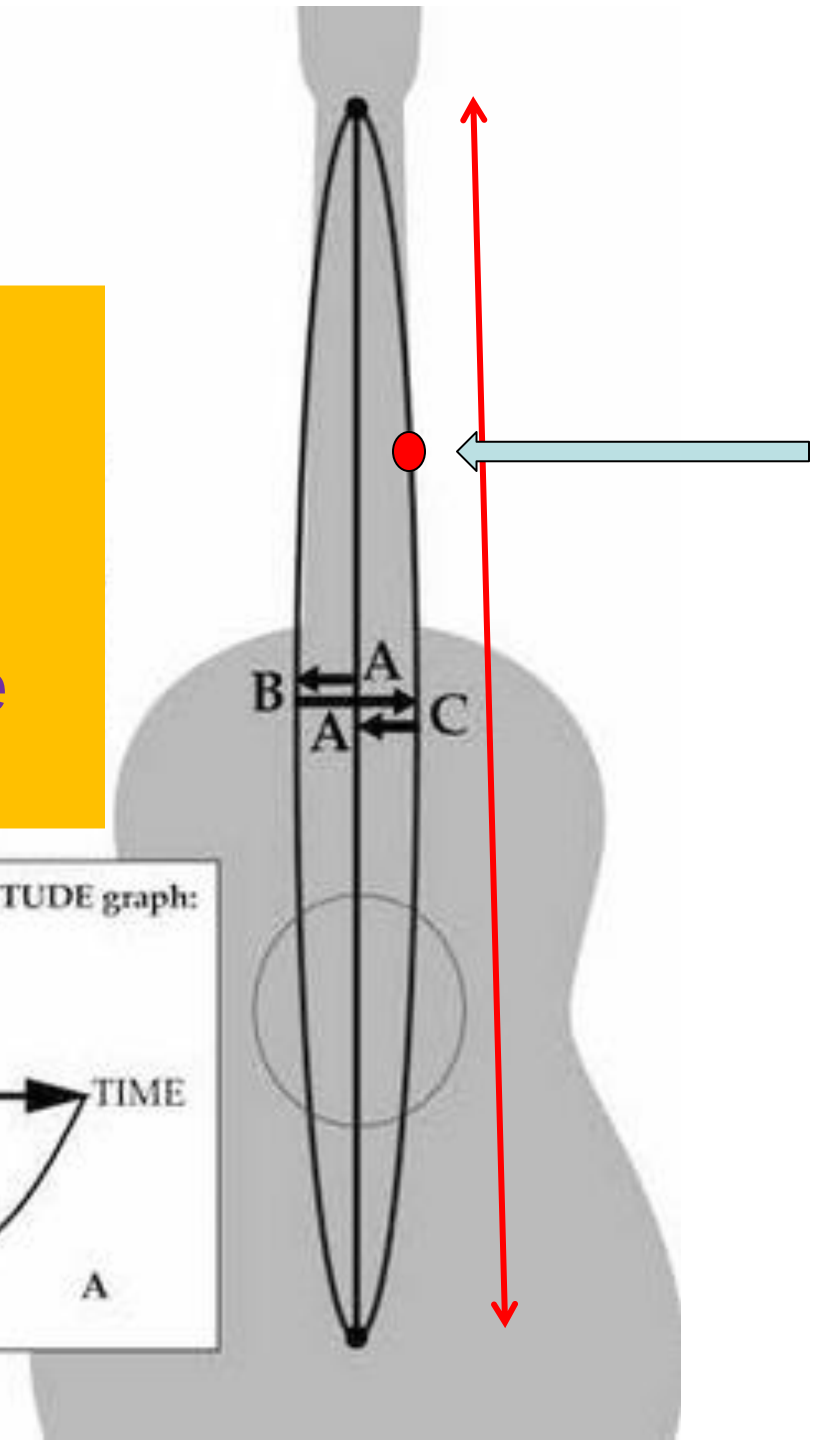
Consider a Guitar

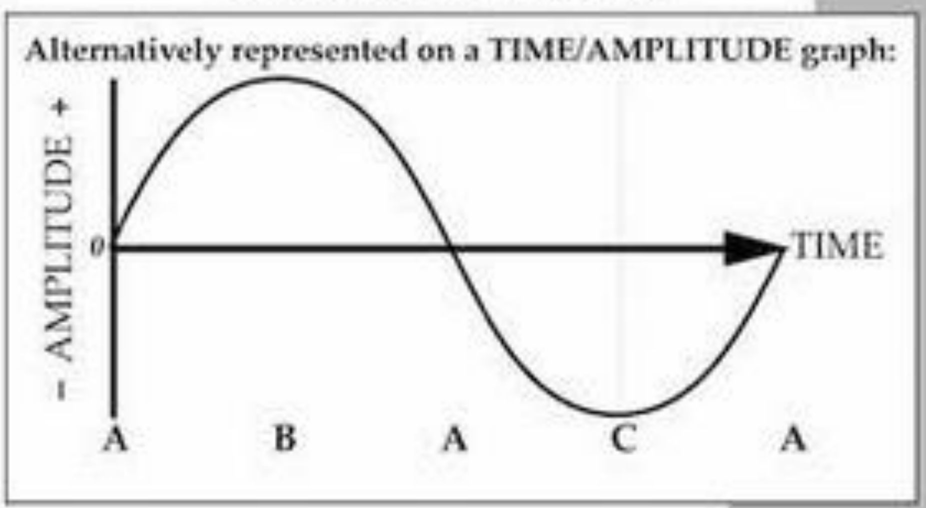
Wavelength → each string represents ½ wavelength from the bridge

Particle → the pulse sent on the string represents the particle nature of sound which travels back and forth on the string (*longitudinal wave*)

Guitar: Wave – Particle Duality

- One string represent λ/2 (wave)
- The sound travels as a pulse back & forth on the string (particle)

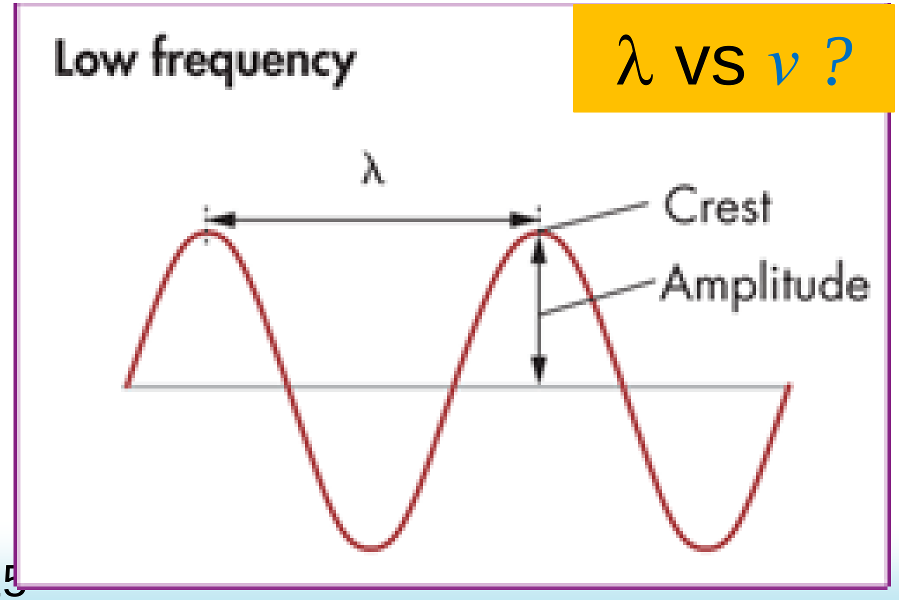


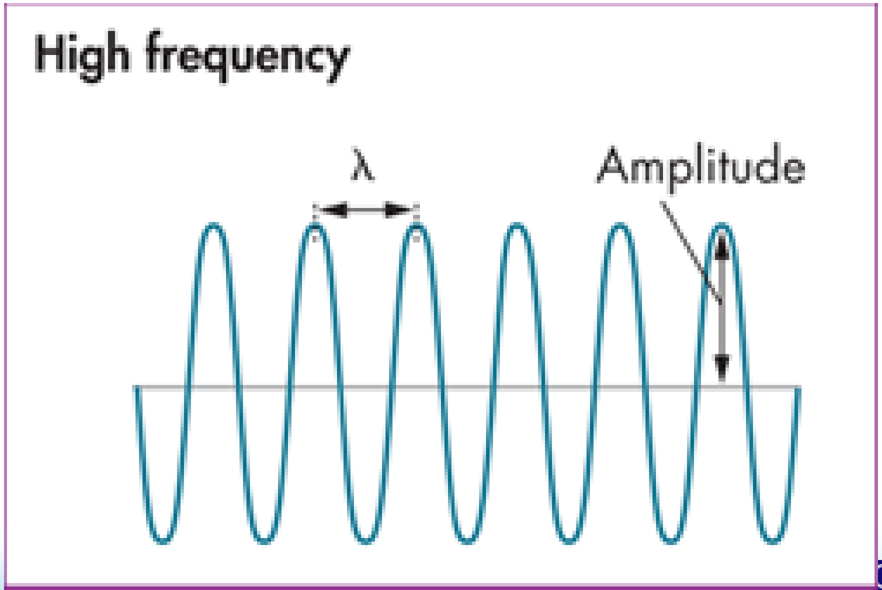


Wavelength (**λ**) of light corresponds to its color; red light has the longest λ, 700 nm, & lowest frequency; violet, at 380 nm, has the shortest λ & highest frequency. units: meters

Frequency (*ν*) corresponds to energy; wave cycles to pass a given point per unit of time; units: hertz (hz, sec^{-1})

Amplitude corresponds to brightness of the light.





Wavelength, λ, and **frequency, ν**, are ***inversely*** proportional.

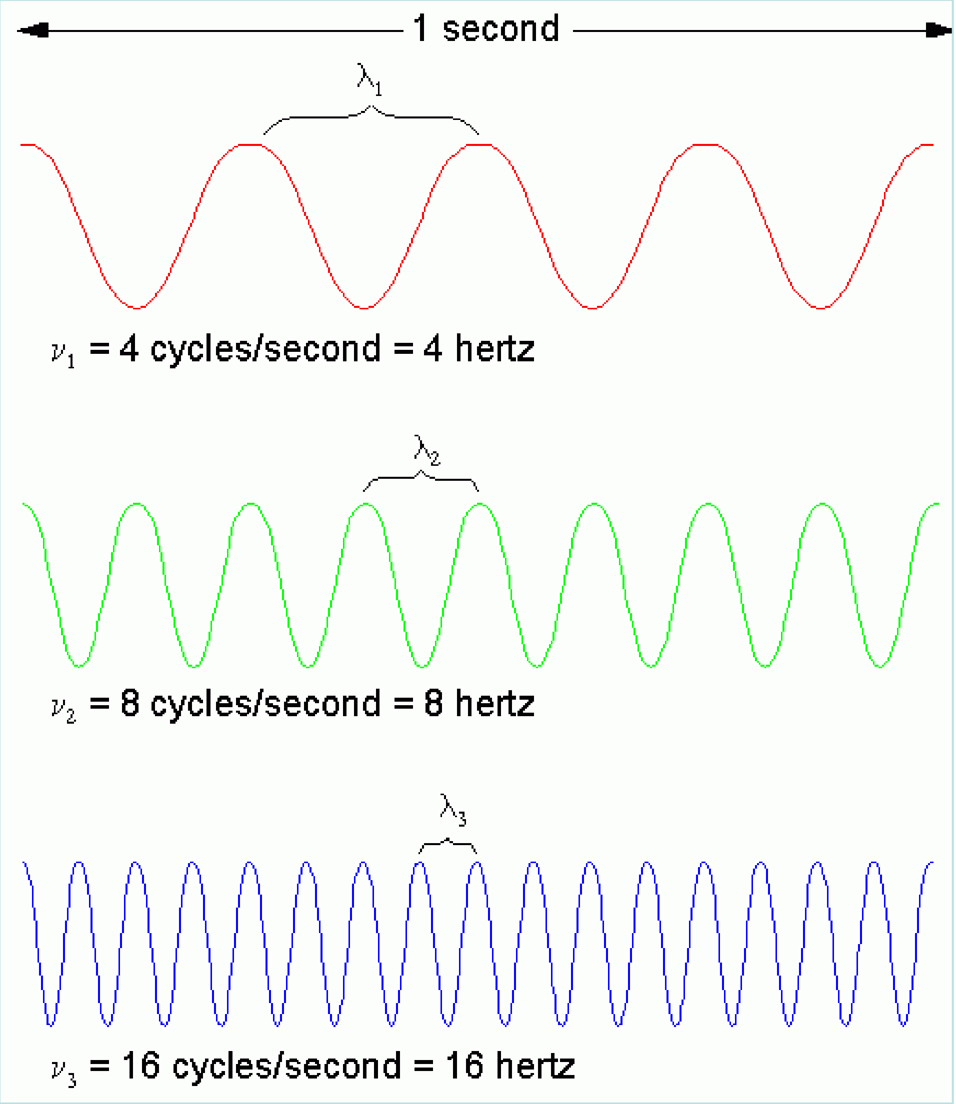
$$c = f\lambda$$

Speed of light

Frequency, ν, and **energy** are ***directly*** proportional. The higher the frequency, the higher the energy.

$$E = h\nu$$

$h = 6.626 \times 10^{-34}$ J·s
Planck's constant.



How does the energy of the higher energy levels of an atom compare with the energy of the lower energy levels of the atom?

TRY IT

A. They are greater in magnitude than those of lower energy levels.

B. They are lesser in magnitude than those of lower energy levels.

C. There is no significant difference in the magnitudes.

Which variable is directly proportional to frequency in relation to the speed of light?

wavelength   position   velocity   energy

What are quanta of light called?

quarks   excitons   muons   photons

Who predicted that all matter can behave as waves as well as particles?

Einstein   Schrodinger   Planck   Louis de Broglie

How does the energy of the higher energy levels of an atom compare with the energy of the lower energy levels of the atom?

A. They are greater in magnitude than those of lower energy levels.

~~**B. They are lesser in magnitude than those of lower energy levels.**~~

~~**C. There is no significant difference in the magnitudes.**~~

Which variable is directly proportional to frequency in relation to the speed of light?

wavelength ~~**position velocity energy**~~

What are quanta of light called?

~~**quarks excitons muons**~~ **photons**

Who predicted that all matter can behave as waves as well as particles?

~~**Einstein Schrodinger Planck**~~ **Louis de Broglie**

TRY IT

Calculating the Energy of a Photon

TRY IT

Calculate the energy of a photon of red light with a wavelength of 5.77×10^{-5} cm. $h = 6.626 \times 10^{-34}$ Js $c = 2.998 \times 10^{8}$ m/s

Calculating the Energy of a Photon

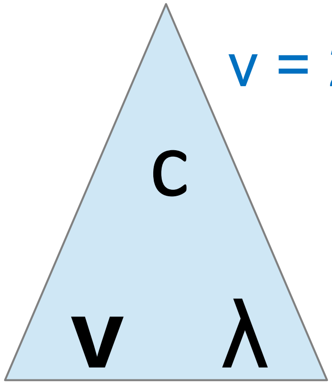
TRY IT

Calculate the energy of a photon of red light with a wavelength of 5.77×10^{-5} cm. $h = 6.626 \times 10^{-34}$ Js $c = 2.998 \times 10^{8}$ m/s

NOTICE: there are two unknown variables (E, frequency). Therefore, we need to do a step in-between … $c = v\lambda$ … $v = c/\lambda$

Change to common units: $\lambda = 5.77 \times 10^{-5}$ cm x 1 m/100cm = 5.77×10^{-7} m

$v = 2.998 \times 10^{8}$ m/s / 5.77×10^{-7} m = 5.20×10^{14} 1/s



Calculating the Energy of a Photon

TRY IT

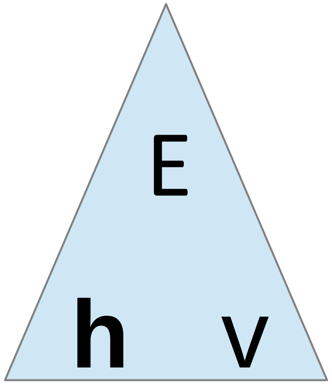
Calculate the energy of a photon of red light with a wavelength of 5.77×10^{-5} cm. $h = 6.626 \times 10^{-34}$ Js $c = 2.998 \times 10^{8}$ m/s

SUBSTITUTE frequency into the Energy equation:

$$E = h\nu$$

$$E = 6.626 \times 10^{-34} \text{ Js} \times 5.20 \times 10^{14} \text{ 1/s}$$

$$E = 3.44 \times 10^{-19} \text{ J}$$



You may use $c = 3.000 \times 10^{8}$ m/s

Calculating the Energy of a Photon

Calculate the energy of a photon of red light with a wavelength of 5.77×10^{-5} cm. $h = 6.626 \times 10^{-34}$ Js $c = 2.998 \times 10^{8}$ m/s

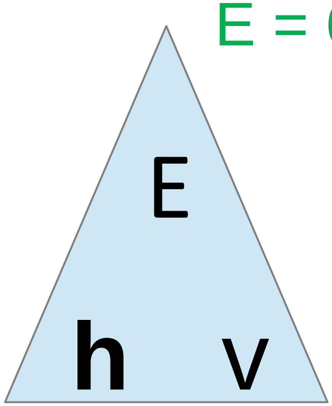
Change to common units: $\lambda = 5.77 \times 10^{-5}$ cm x 1 m/100cm = 5.77×10^{-7} m

The easier way to do this: since $v = c/\lambda$, replace "v" with "c/λ"

$$E = hv = hc/\lambda$$

$$E = 6.626 \times 10^{-34}\ \text{Js} \times 2.998 \times 10^{8}\ \text{m/s} \,/\, 5.77 \times 10^{-7}\ \text{m}$$

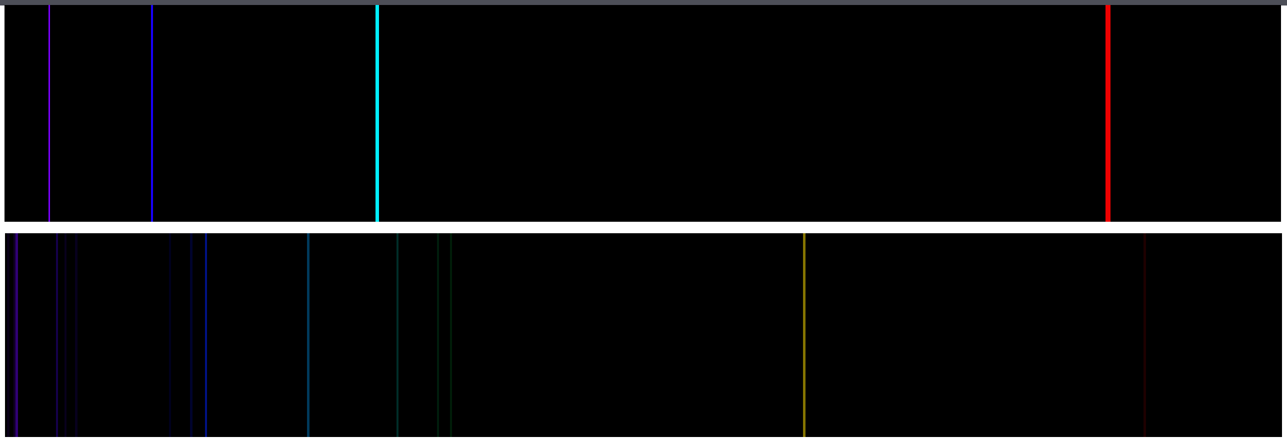
$$E = 3.44 \times 10^{-19}\ \text{J}$$



Evidence of "Quanta" of Energy

Hydrogen (g)

Helium (g)



Johannes Rydberg studied **emission spectra**.

- **Emission spectrum**: a visible light spectrum in which wavelengths of light emitted by a substance show up as bright, colored lines
- Emission spectra for some metals produced discrete lines (e.g. quanta), not continuous or gradual.
- *Determined a **DIRECT relationship** between frequency and energy.*

Elements give off characteristic Emission Spectra (colors of light), as electrons transition between energy levels.


strontium


sodium


lithium

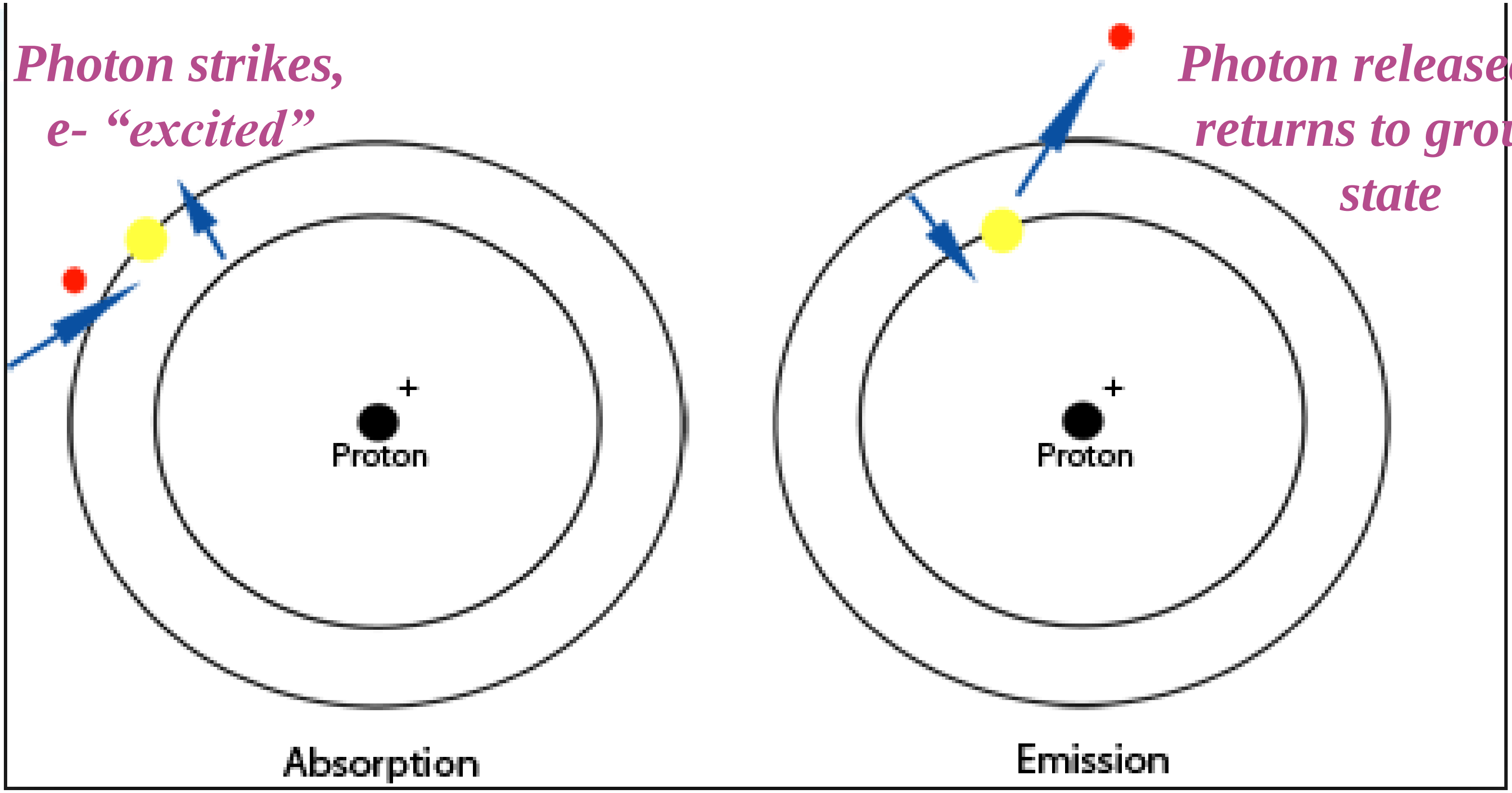

potassium


copper

Absorption & Emission Spectra

Photon strikes, e- "excited"

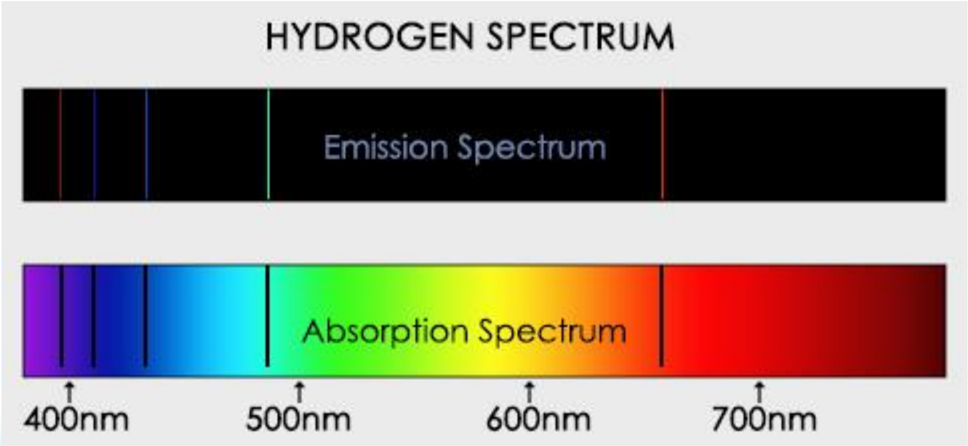
Photon released, e- returns to ground state



Notice the movement of electrons based on the photon.

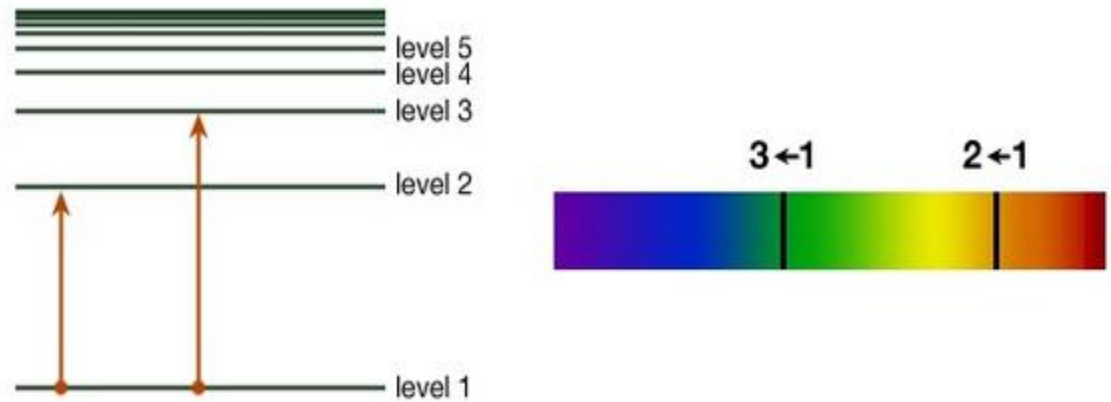
Absorption & Emission Spectra

- The energy **absorbed** by an electron to move from its current energy level to a higher energy level.
  - is identical to the energy of the light emitted by the electron as it **drops back to its original energy level (Emission)**.
- Emission Spectra are like "**fingerprints**" … no two elements have the same spectra.

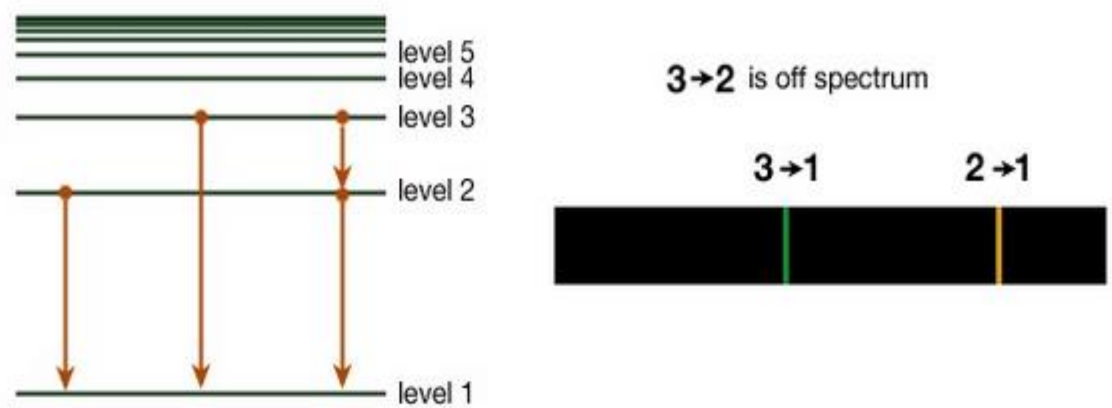


5.3 Atomic Emission Spectra and the Quantum Mechanical Model >

Electrons absorb heat or electrical energy to reach the EXCITED STATE → Absorption, dark line spectra

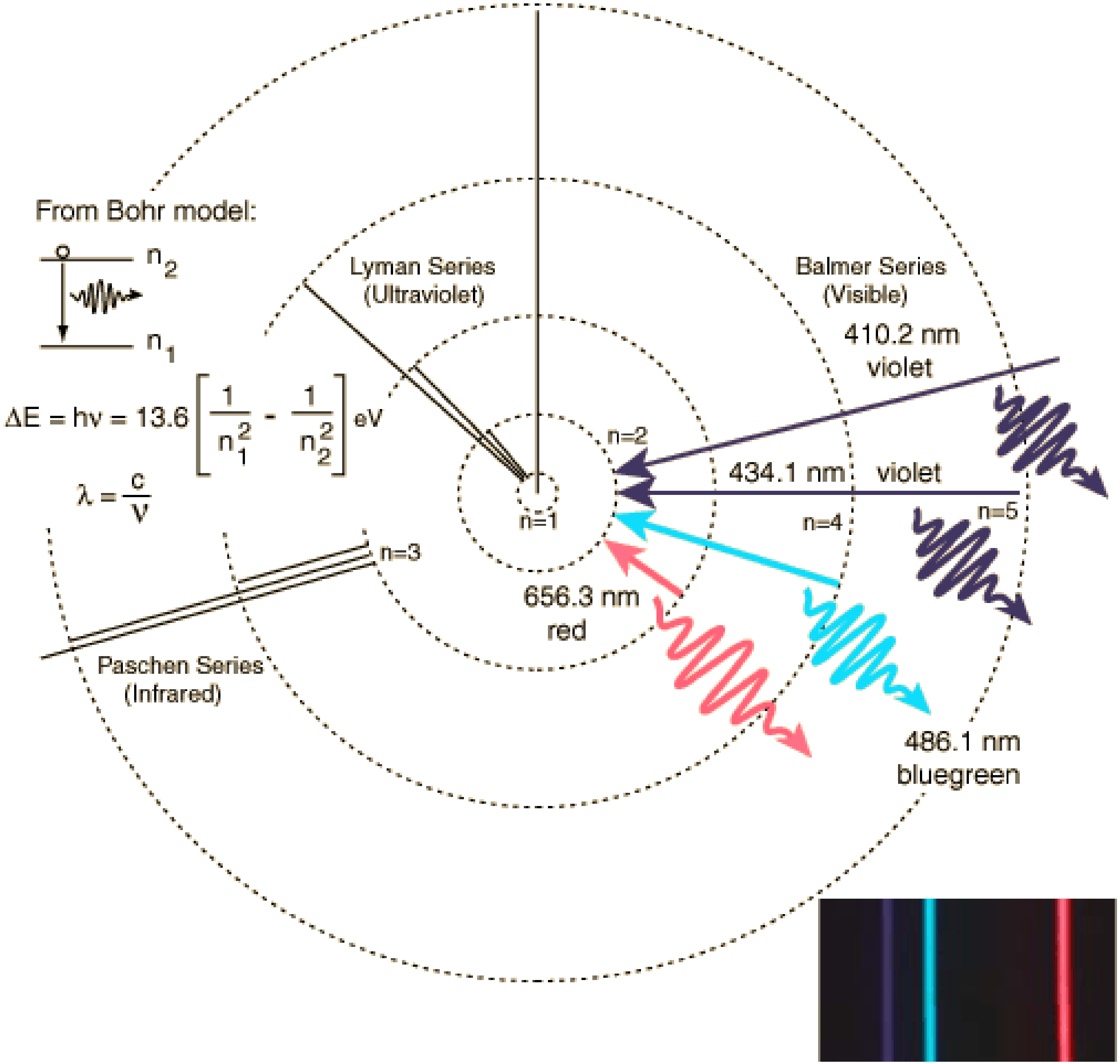


Electrons return to the GROUND STATE (*most stable energy state*) → Bright-Line Spectra



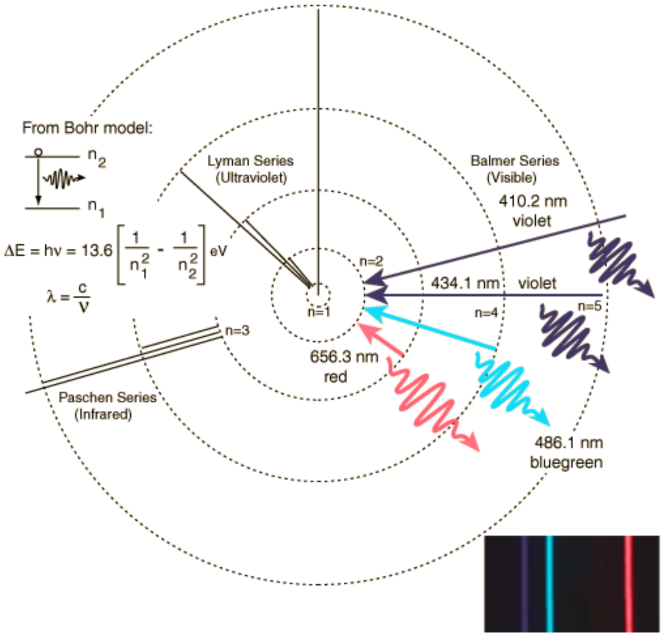
Energy is DISCRETE or QUANTIZED (*like stairs*)

QUICK CHECK



Explain this diagram in terms of energy, electrons & spectra

PEARSON



Electrons begin in the ground state (lowest energy level).

Energy is "absorbed" so electrons get "excited" to a higher energy level → (Absorption Spectra)

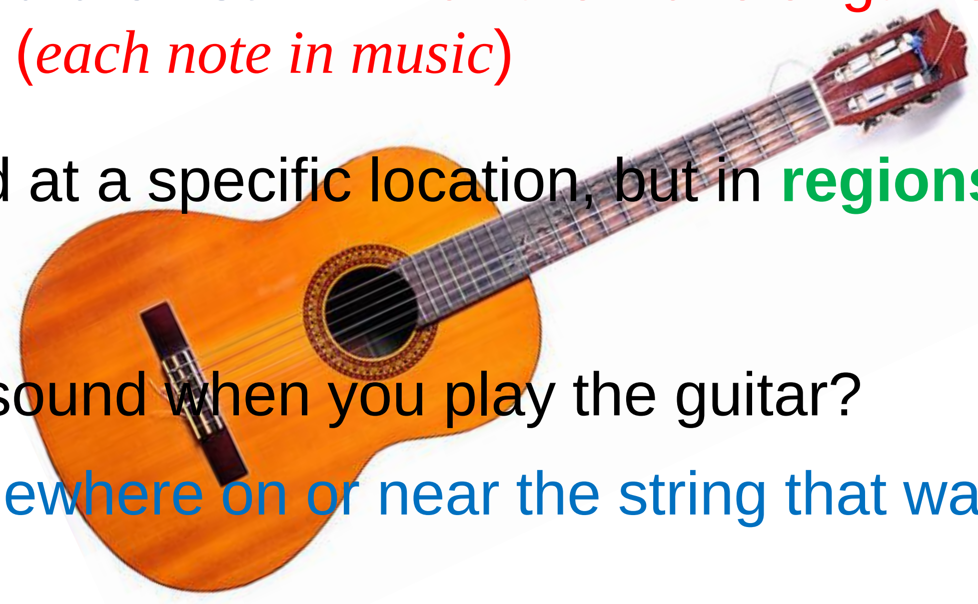
The "excited" state is UNSTABLE so the electrons will return to the ground state by giving off energy in the form of light (color) → Emission Spectra.

Energy absorbed" or "emitted" is in discrete bundles (quanta), not gradual.

"O, where oh where has my Electron gone?"

Erwin Schrödinger (1887–1961) worked from the premise that the electron was a wave and a particle. Therefore, its **location** could be statistically determined using previous diffraction techniques [*Thomson Double Slit Diffraction pattern*]

- **Only certain energies could exist** in which the wavelength form → "**STANDING WAVES**" (*each note in music*)
- Electrons cannot be found at a specific location, but in **regions of high probability**.
- E.g. where exactly is the sound when you play the guitar?
- ~90% of the time it is somewhere on or near the string that was plucked.

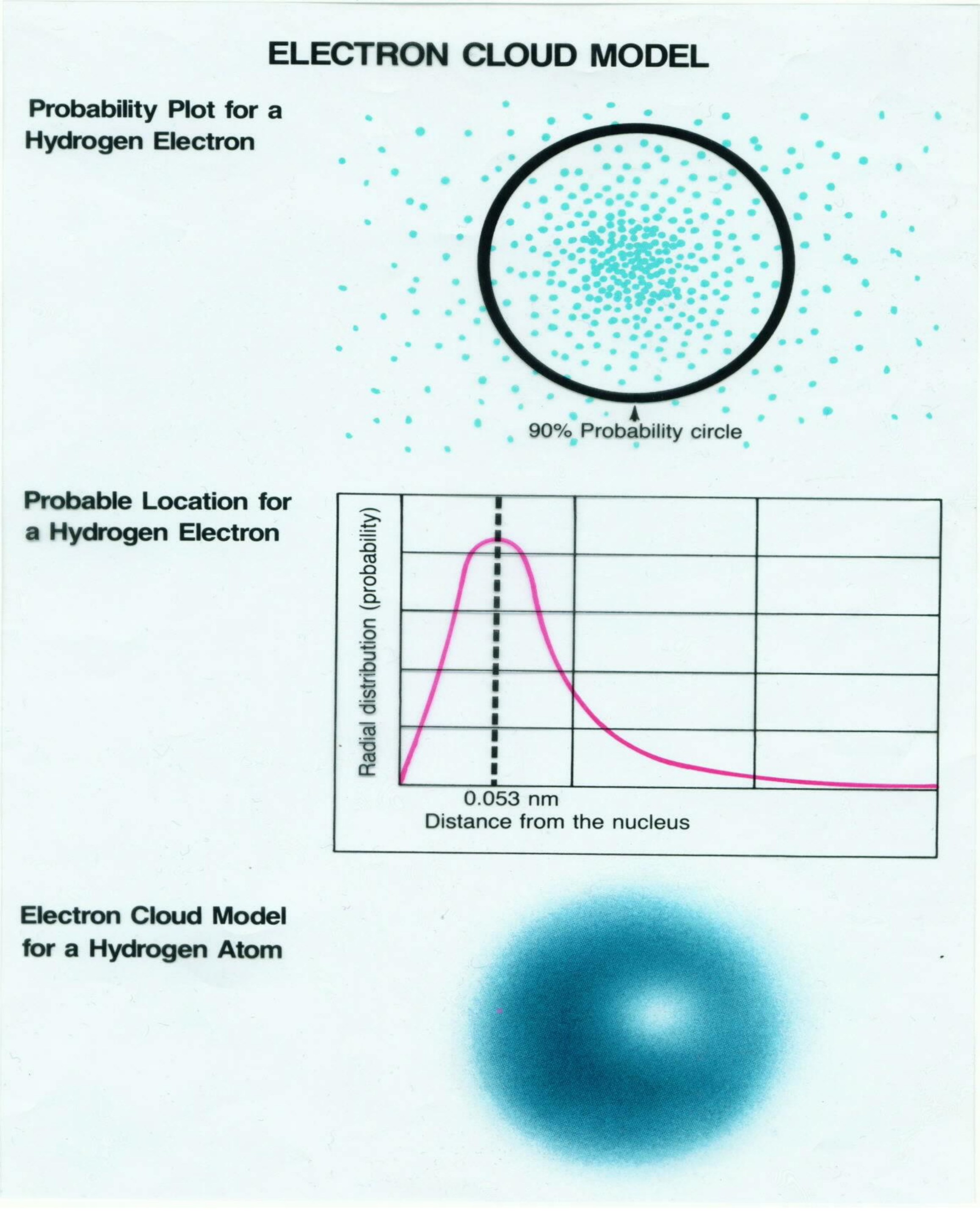


Electron Cloud

Schrödinger

Region of **high probability** (90%) for finding an electron

Developed the **quantum numbers** to describe the location of the electrons in the atom.



Alice in Wonderland

Electron Configuration Song (3:24)

https://screencast-o-matic.com/watch/cq6nYuuIbb

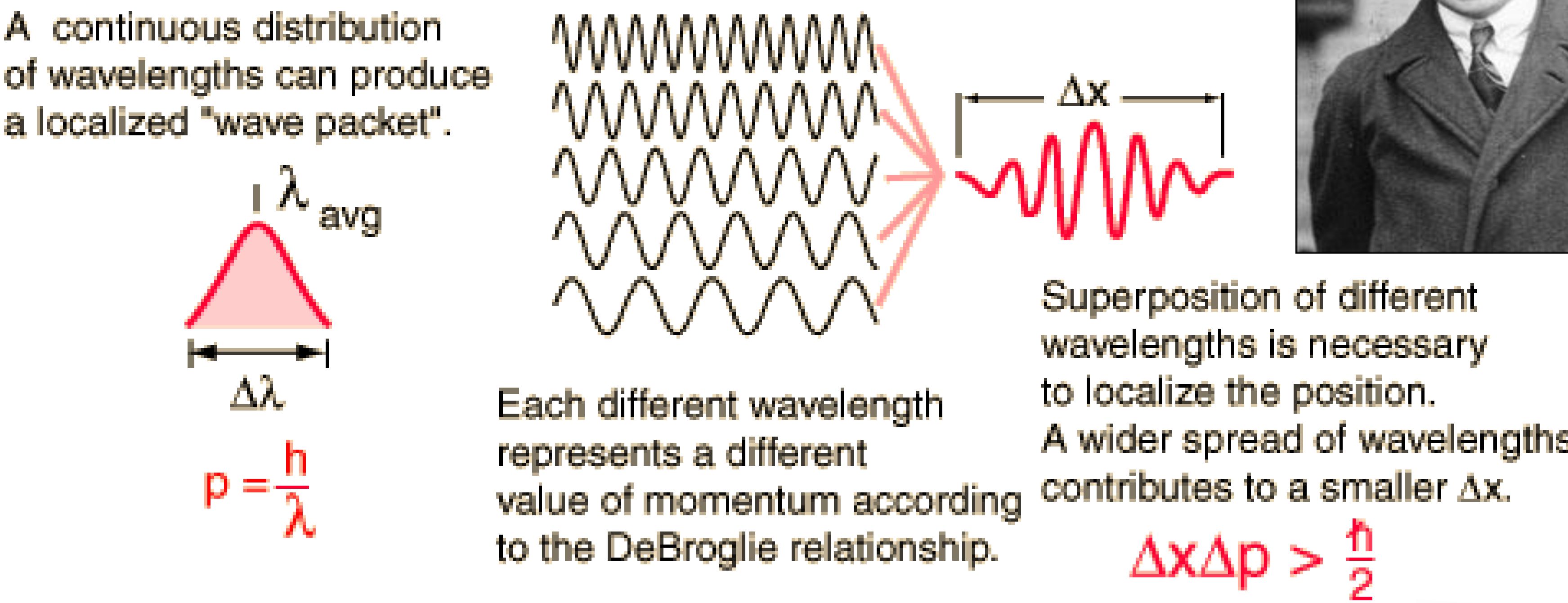
The Heisenberg Uncertainty Principle

- Heisenburg noted and worked with the uncertainties in the **position (ΔX)** and **momentum (Δρ)** of small particles
- He used the **PROBABILITY** of finding an electron between two points of a **monochromatic wave** (*integrals of calculus*)
- He needed to use a "Pulse Wave" → constructive addition of many waves of various wavelength (*antinodes*) and destructive interference at the nodes

$$\Delta X \times \Delta \rho = \sim h/2$$

h → Planck's constant → 6.6 x 10^{-34} *js*

5.3 Atomic Emission Spectra and the Quantum Mechanical Model



The Heisenberg Uncertainty Principle

- *The length of the guitar string ~* **ΔX**
- *The momentum of the sound "pulse" ~***Δρ**

"You observed me speeding? Are you familiar with the Heisenberg uncertainty principle?"

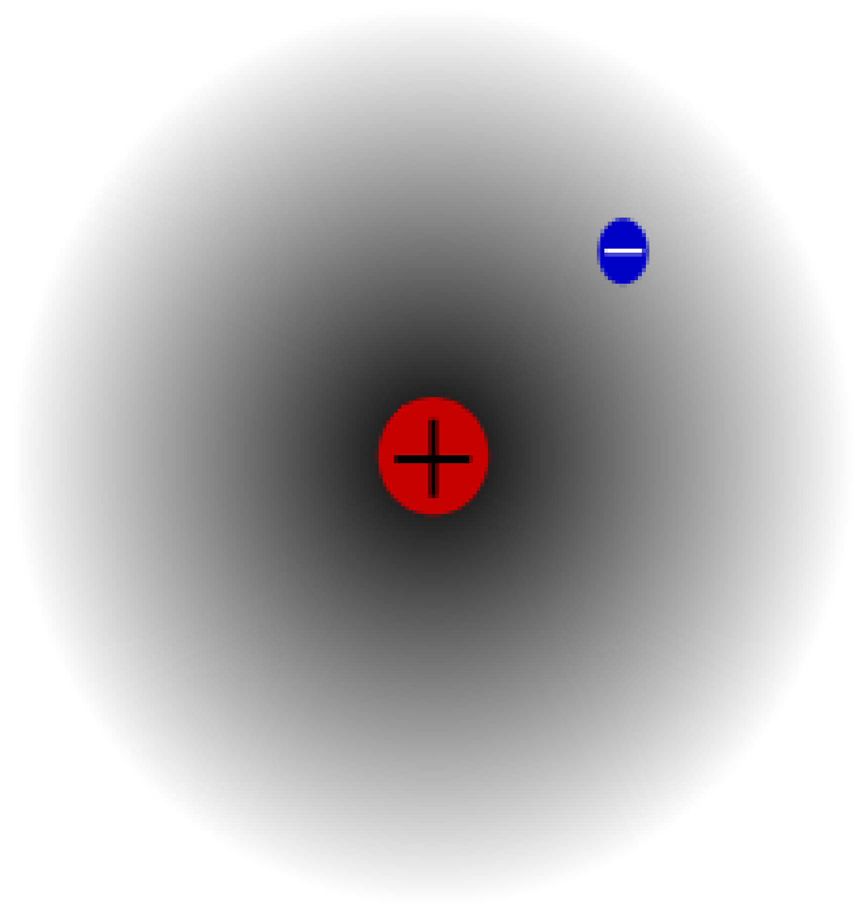
Quantum Mechanics: Electron Cloud Model

The modern description of the electrons in atoms, the **quantum mechanics model**, came from the mathematical solutions (the Schrödinger equation).

$$i\hbar \frac{\partial \Psi}{\partial t} = -\frac{\hbar^2}{2m}\frac{\partial^2 \Psi}{\partial x^2} + V(x)\Psi.$$

Electron Cloud model:

- *Electrons in a cloud have regions of high probability (uncertain location).*
- *Electron clouds have different energy levels that are discrete.*
- *Cannot know the exact position of the electron.*



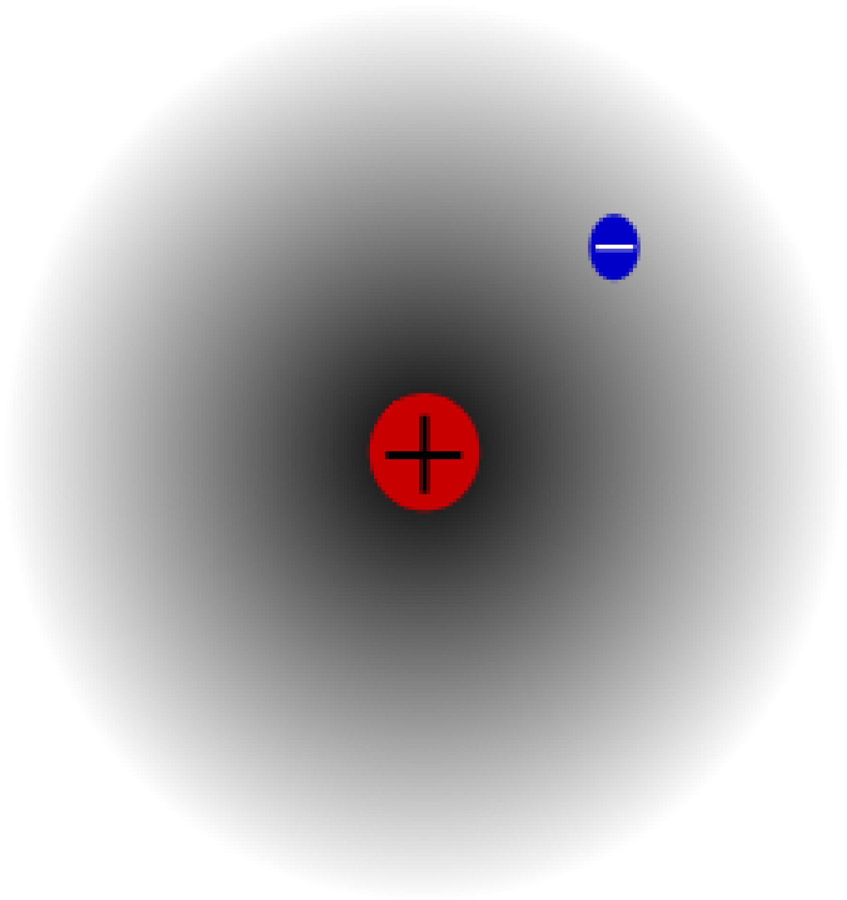
Quantum Mechanics: Electron Cloud Model

Demonstration:

http://somup.com/crjT2YriIi

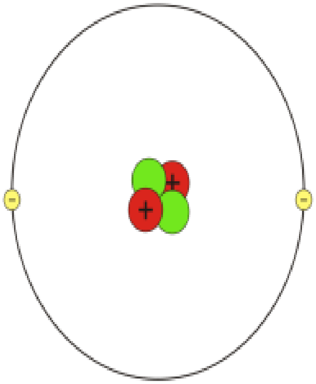
Uncertainty Principle with Pennies (1:28)

$$i\hbar \frac{\partial \Psi}{\partial t} = -\frac{\hbar^2}{2m}\frac{\partial^2 \Psi}{\partial x^2} + V(x)\Psi.$$

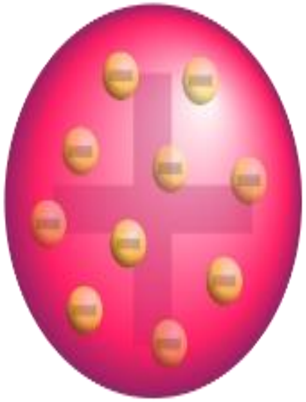


Understanding Atomic Structure

What scientist suggested each of the models shown below? Which best represents the modern understanding of the structure of the atom?



Model A



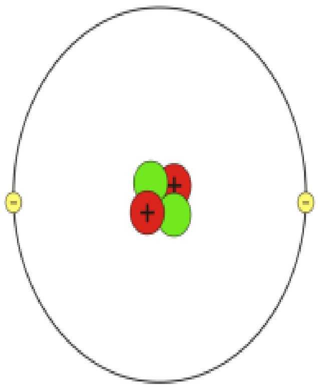
Model B

Model C

Understanding Atomic Structure

QUICK CHECK

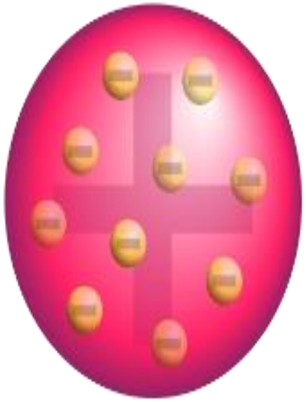
What scientist suggested each of the models shown below? Which best represents the modern understanding of the structure of the atom?



Model A

Rutherford & Bohr

Nucleus with orbiting electrons



Model B

Thomson

Plum Pudding

Model C

Electron Cloud

Schroedinger

Subshell	n	l	Maximum No. of Electrons
1s	1	0	2
2s	2	0	2
2p	2	1	6
3s	3	0	2
3p	3	1	6
3d	3	2	10
4s	4	0	2
4p	4	1	6
4d	4	2	10
4f	4	3	14

LESSON QUESTION

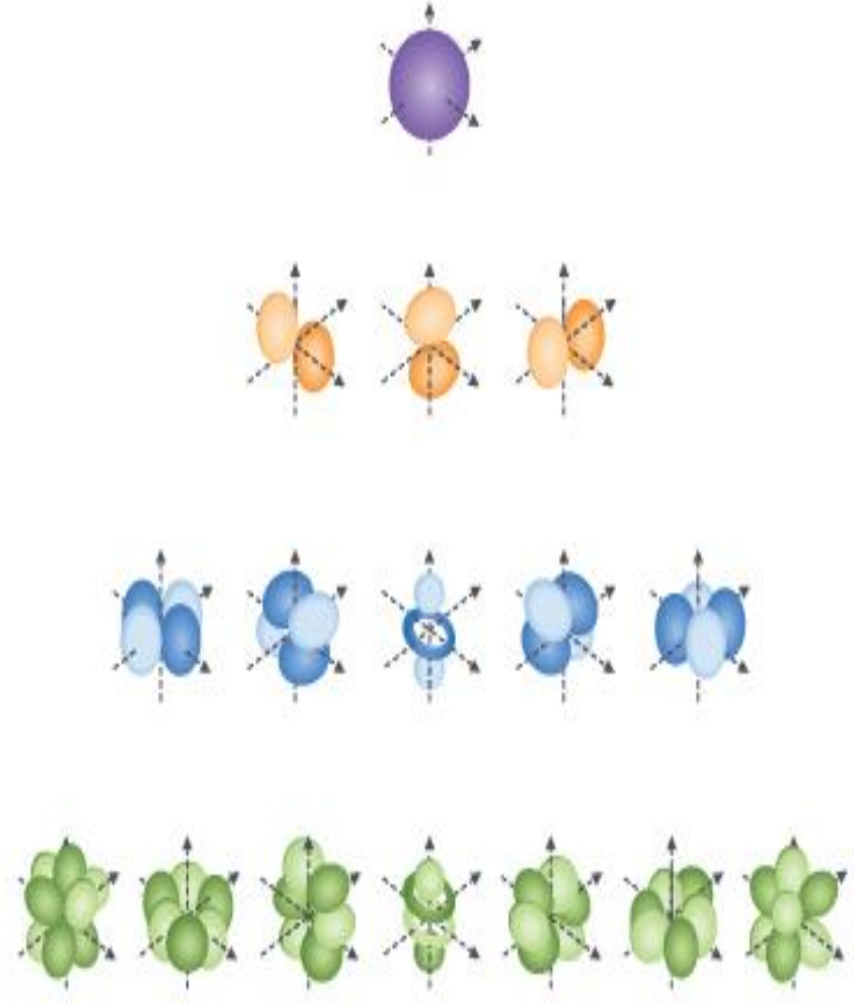
How can scientists describe the arrangement of electrons in an atom?

Atomic Orbitals

- An atomic orbital is represented pictorially as a region of space in which there is a high probability of finding an electron.
- **Every electron in an atom is assigned a QUANTUM NUMBER** described by the Schrödinger equation - a mathematical expression
- **Quantum numbers indicate different energy states of electrons in an atom**
- Every **electron** can be described by **FOUR quantum numbers** and NO two electrons have the same 4 numbers.

Atomic Orbitals → Quantum Numbers (4)

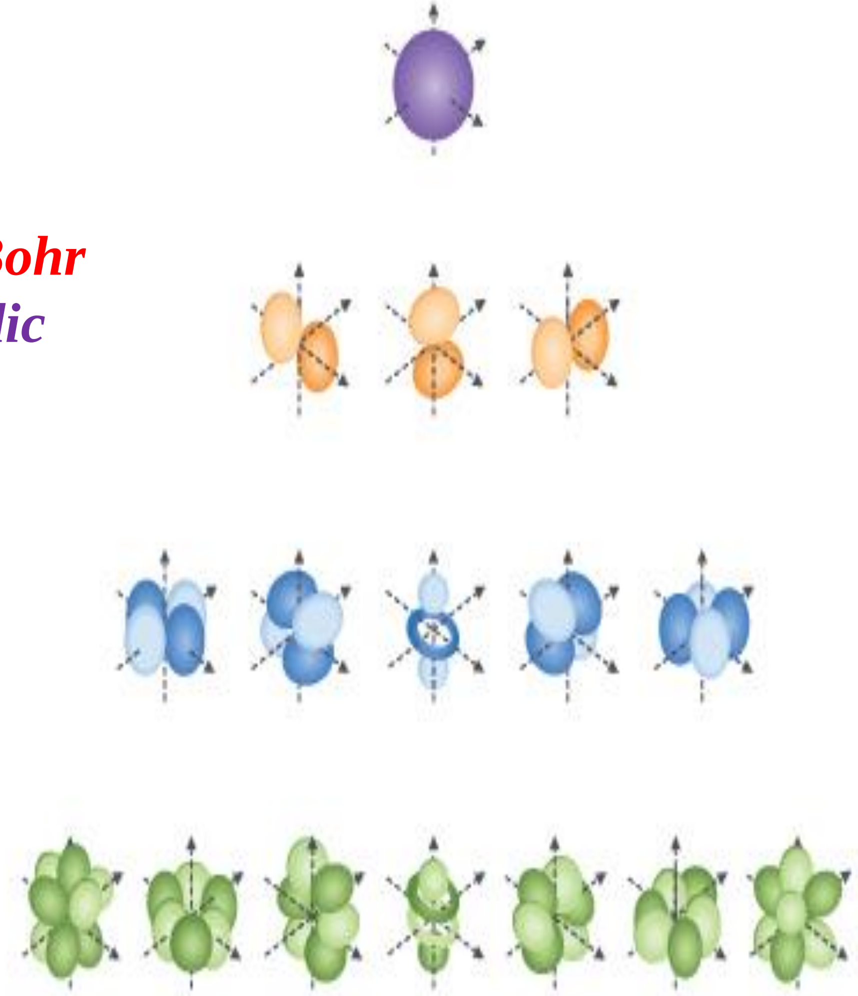
Quantum Numbers Made Easy ctr (11:22)



Atomic Orbitals → Quantum Numbers (4)

1 *n* (**principal quantum number**) = size of the e- cloud

Corresponds to energy levels in the Bohr Model of the atom (7 Rows on Periodic Table)

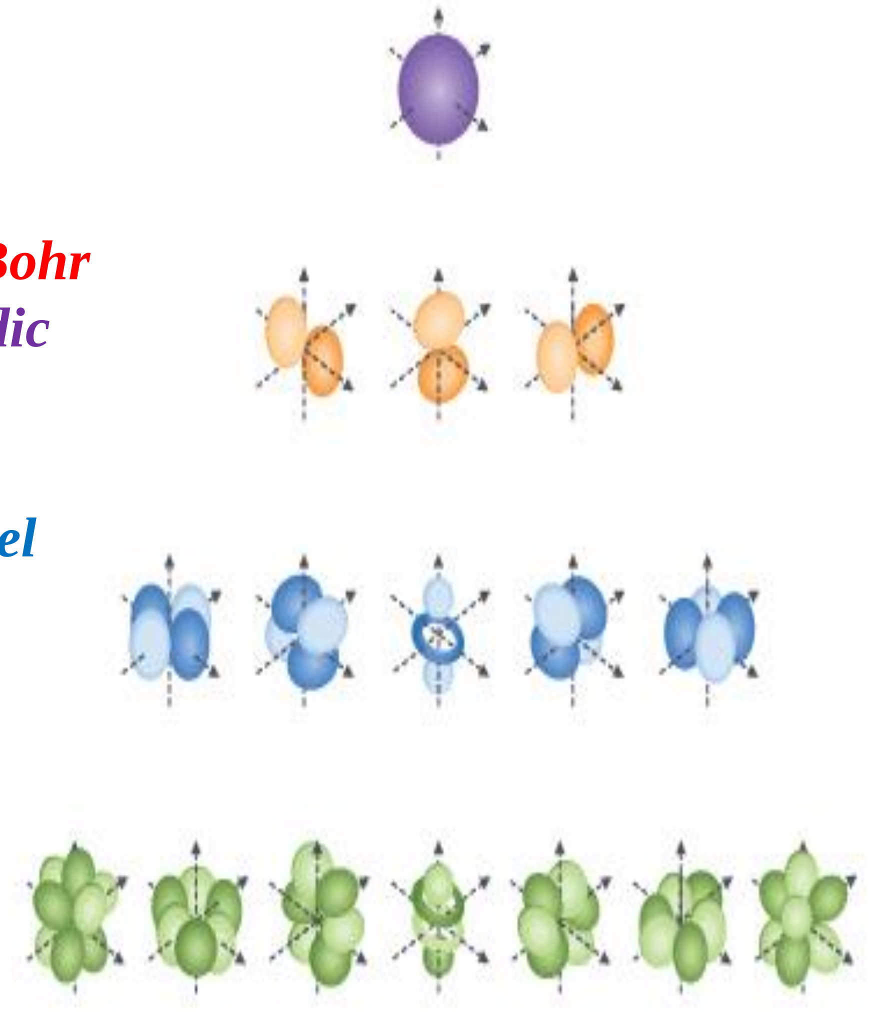


Atomic Orbitals → Quantum Numbers (4)

1 *n* (**principal quantum number**) = size of the e- cloud

Corresponds to energy levels in the Bohr Model of the atom (7 Rows on Periodic Table)

2 ***sublevel*** = ***shape of the sublevel***



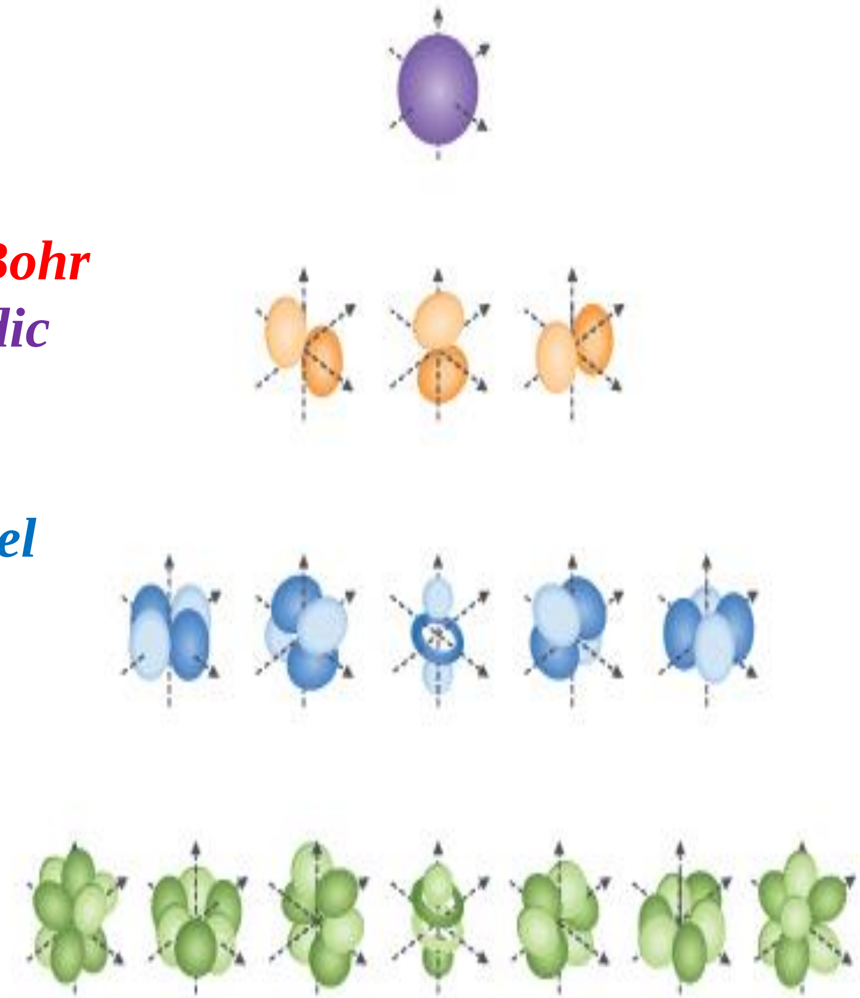
Atomic Orbitals → Quantum Numbers (4)

1 *n* (**principal quantum number**) = size of the e- cloud

Corresponds to energy levels in the Bohr Model of the atom (7 Rows on Periodic Table)

2 ***sublevel*** = ***shape of the sublevel***

3 *Orbital* = ***orbital orientation***



Atomic Orbitals → Quantum Numbers (4)

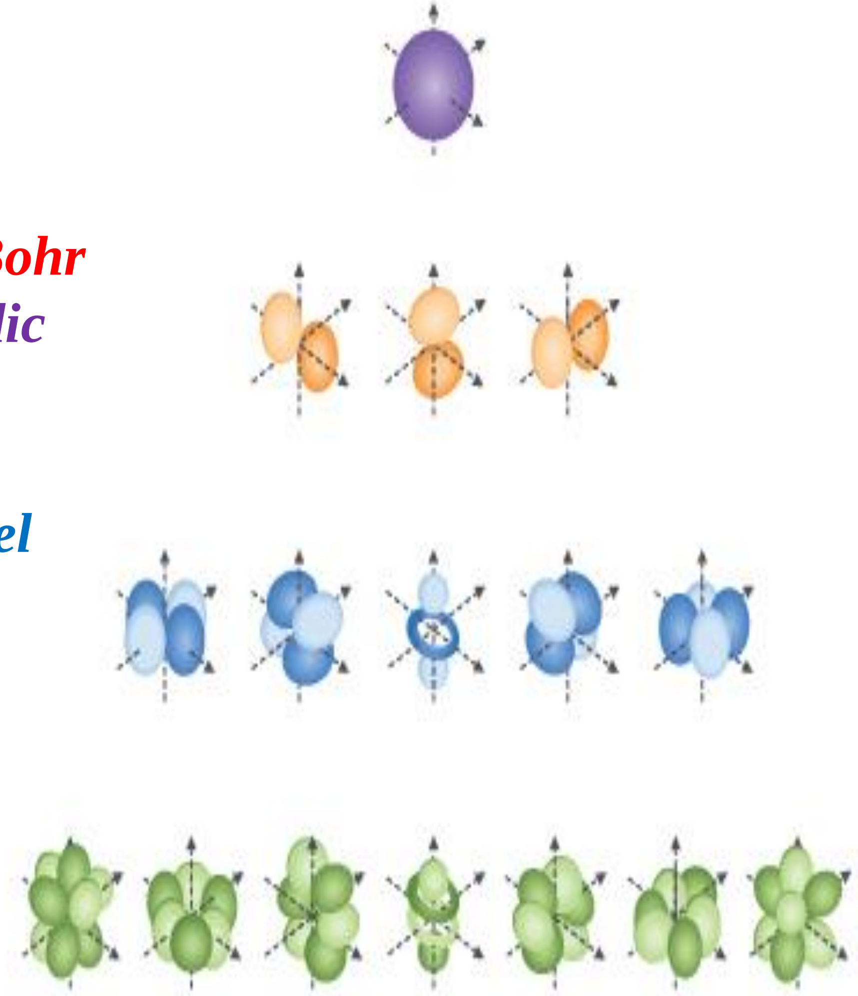
1 *n* (**principal quantum number**) = size of the e- cloud

Corresponds to energy levels in the Bohr Model of the atom (7 Rows on Periodic Table)

2 ***sublevel*** = ***shape of the sublevel***

3 *Orbital* = ***orbital orientation***

4 *s* (**spin**)



Watch these Tutorials

Quantum Numbers Introduction (4:38)

https://screencast-o-matic.com/watch/cFQX3uqmHp

Principal Quantum Number (5:12)

https://screencast-o-matic.com/watch/cFQ2orqKA5

Second Quantum Number (5:22)

https://screencast-o-matic.com/watch/cFQ2DZqKCL

Watch these Tutorials

Third Quantum Number (6:25)

http://somup.com/cFQ2YOVSNt

Fourth Quantum Number (4:57)

https://screencast-o-matic.com/watch/cFQol4q7xH

Pauli Exclusion Principle & Hund's Rule (4:41)

https://screencast-o-matic.com/watch/cFQ2bQqKmJ

PRINCIPAL QN

The numbers and types of atomic orbitals depend on the **principal energy level, n**.

Summary of Principal Energy Levels and Sublevels			
Principal energy level (1 – 7)	**Number of sublevels**	**Type of sublevel**	**Maximum number of electrons ($2n^2$)**
$n = 1$	1	1*s* (1 orbital)	2
$n = 2$	2	2*s* (1 orbital), 2*p* (3 orbitals)	8
$n = 3$	3	3*s* (1 orbital), 3*p* (3 orbitals), 3*d* (5 orbitals)	18
$n = 4$	4	4*s* (1 orbital), 4*p* (3 orbitals), 4*d* (5 orbitals), 4*f* (7 orbitals)	32

PRINCIPAL QN

The principal quantum number (1), n, always equals the number of sublevels (2) within that principal energy level.

The number of orbitals (3) in a principal energy level is equal to n^2.

A maximum of two electrons can occupy an orbital.

Therefore, the maximum number of electrons that can occupy a principal energy level (1) is given by the formula $2n^2$.

Electron Sublevels – Second Quantum Number

An **electron sublevel (2)** is a set of orbitals with the same principal quantum number (1), n, and the same quantum number (3)

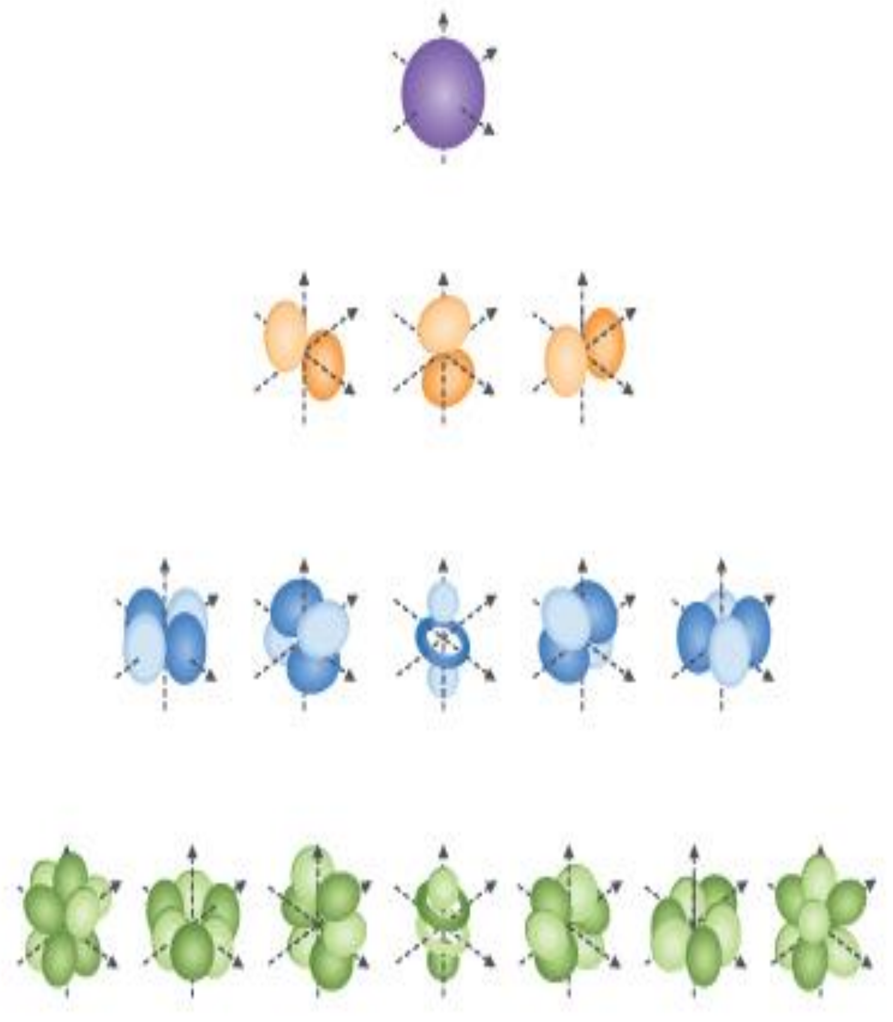
- Represented by letters

  s

  p

  d

  f

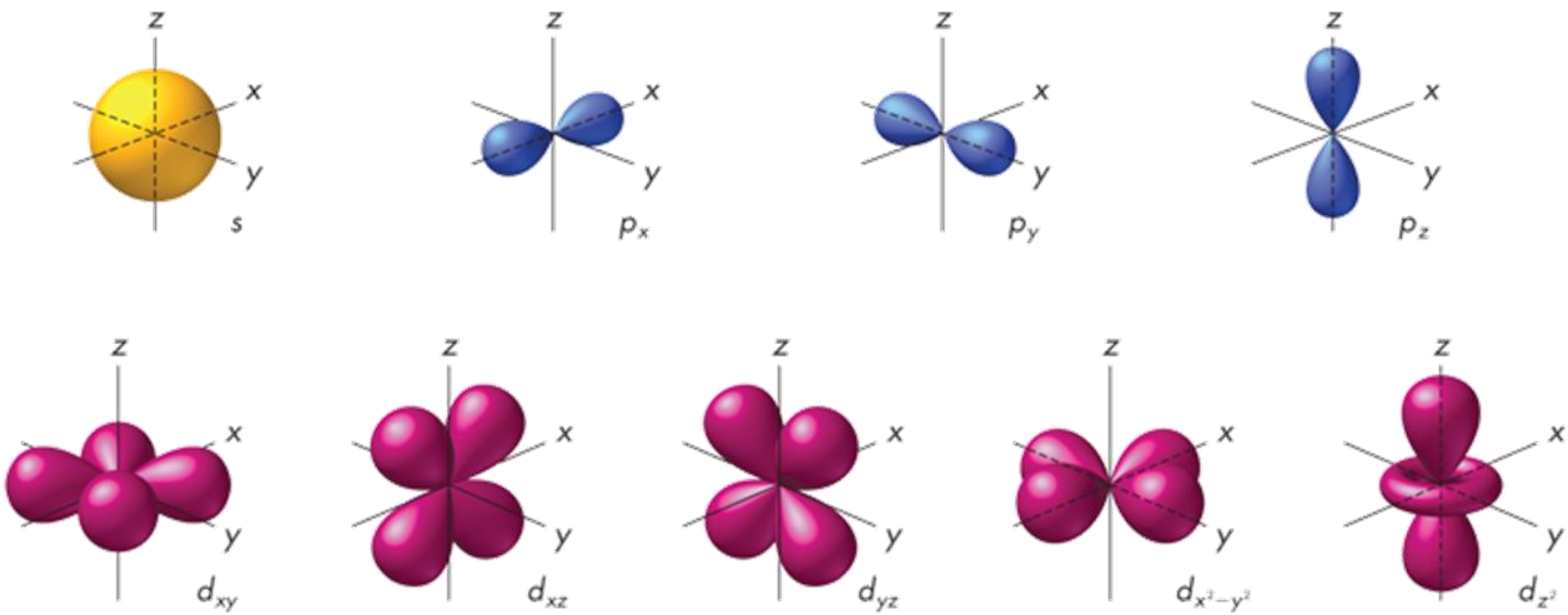


5.2 Electron Configurations

Second QN

Each energy sublevel corresponds to one or more orbitals of different **shapes**.

The *s* orbitals are spherical. The *p* orbitals are dumbbell-shaped. *d* & f orbitals are very complex.



Energy Level	Sublevels	Actual
n = 1	s	1 s
n = 2	s, p	2 s 2 p
n = 3	s, p, d	3 s 3p 3 d
n = 4	s, p, d, f	4s 4 p 4 d 4 f

Recap of Quantum Numbers 1 & 2 (2:36)

http://somup.com/cFQ2FRVSLt

Third Quantum Number (3)

- **Corresponds to the ORBITAL orientation in space**
- **A region of high probability for finding electrons & the direction in space of each orbital**
- **Each orbital can contain a maximum of 2 electrons**
- **n^2 indicates the number of orbitals in a particular sublevel**

There are TWO electrons in each orbital (last column below)

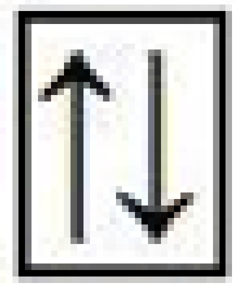
Energy Level	Sublevel # of orbitals	Total # of Orbitals (n^2)	Max. e– ($2n^2$)	Orbital Diagram s p d f
n = 1	s^1	1	2	0
n = 2	s^1, p^3	4	8	0 000
n = 3	s^1, p^3, d^5	9	18	0 000 00000
n = 4	s^1, p^3, d^5, f^7	16	32	0 000 00000 0000000

Fourth QN

Pauli Exclusion Principle

An atomic orbital may contain at most two electrons. To occupy the same orbital, the two electrons must have opposite spins; that is, the electron spins must be paired.

Spin (4)



- “clockwise” or “counterclockwise”
- A vertical arrow indicates an electron and its direction of spin (↑ or ↓).
- **Values of “s” range from + ½ to – ½**

A shorthand method for showing the electron configuration of an atom involves writing the energy level and the symbol for every sublevel occupied by an electron.

The number of electrons occupying that sublevel is indicated with a superscript.

For hydrogen, with one electron in a 1*s* orbital, the electron configuration is written $1s^1$.

For oxygen, with two electrons in a 1*s* orbital, two electrons in a 2*s* orbital, and four electrons in 2*p* orbitals, the electron configuration is $1s^2 2s^2 2p^4$.

The sum of the superscripts equals the number of electrons in the atom.

Do the electron configurations (4 QN):

- Energy Level (1)
- Sublevel (2)
- Orientation (orbitals) (3)
- Spin (4)

$_{1}H^{1}$

$_{2}He^{4}$

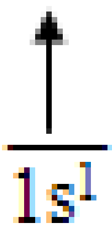
$_{3}Li^{7}$

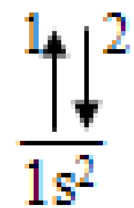
$_{4}Be^{9}$

Do the electron configurations (4 QN):

- Energy Level (1) → *"1" and "2" before "s"*
- Sublevel (2) → *"s"*
- Orientation (orbitals) (3) → *"lines"*
- Spin (4) → *arrows*

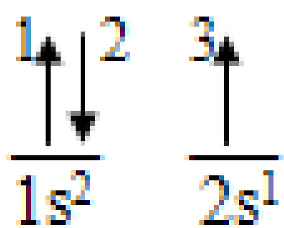
$_1H^1$

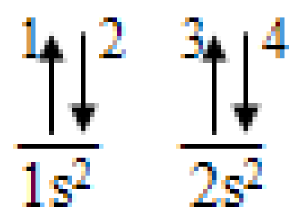




$_2He^4$

$_3Li^7$





$_4Be^9$

Hund’s Rule (Finish the e- configurations)

TRY IT

Boron:

Carbon:

Fluorine:

Hund's Rule (Finish the e- configurations)

TRY IT

Within any sublevel (p, d or f), all the orbitals of that sublevel must contain 1 electron BEFORE any orbital can be filled (2 e- max.)

Boron:

Carbon:

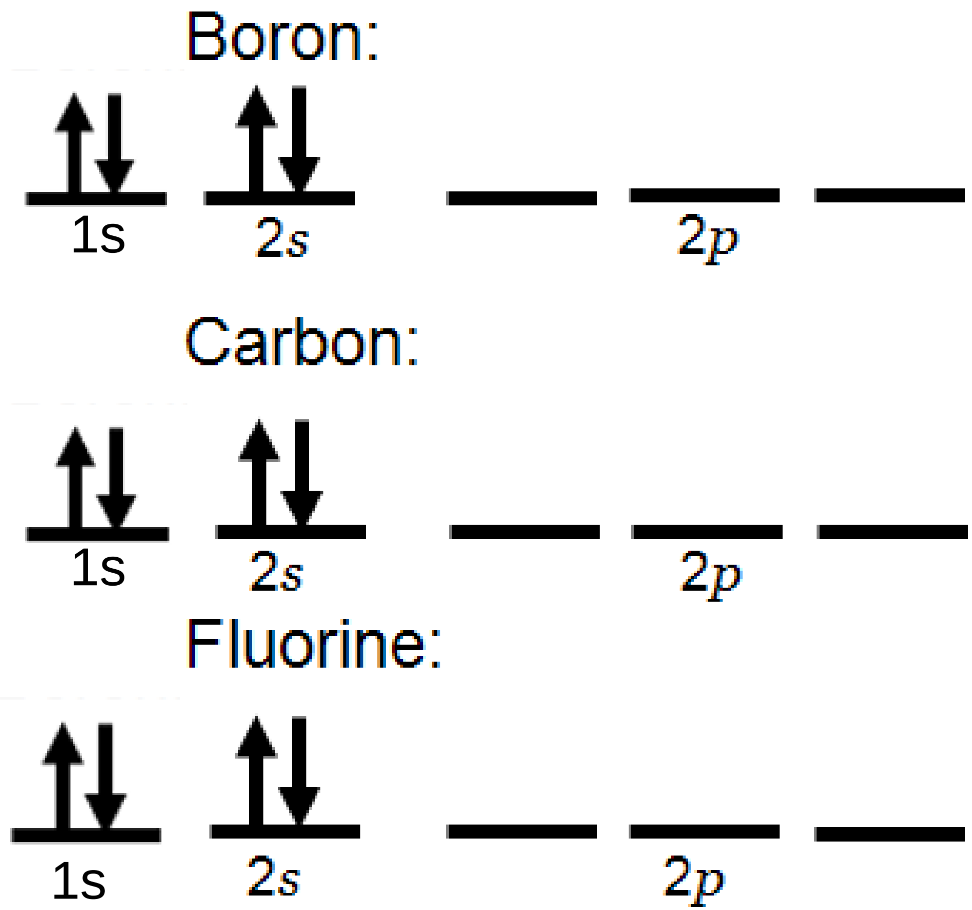
Fluorine:

Hund's Rule (Finish the e- configurations)

TRY IT

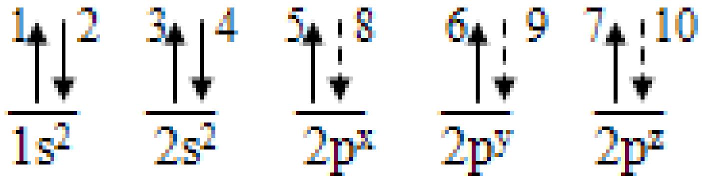
Within any sublevel (p, d or f), all the orbitals of that sublevel must contain 1 electron BEFORE any orbital can be filled (2 e- max.)

Electrons added to empty orbitals have the same spin quantum number.



Hund's Rule

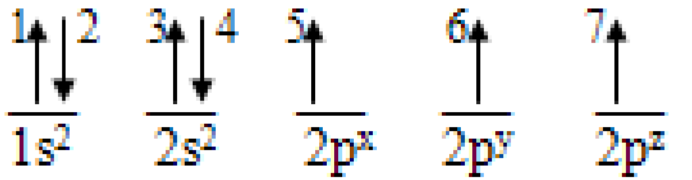
<u>Example</u> $_{10}Ne^{20}$ $\mathbf{1}\ s^2$ $\mathbf{2}\ s^2\ \mathbf{2}\ p^6$



e–
e– orientation

b. Example: Parents with children in terms of feeding them

<u>Example</u> $_{7}N^{14}$ $\mathbf{1}\ s^2$ $\mathbf{2}\ s^2\ \mathbf{2}\ p^3$

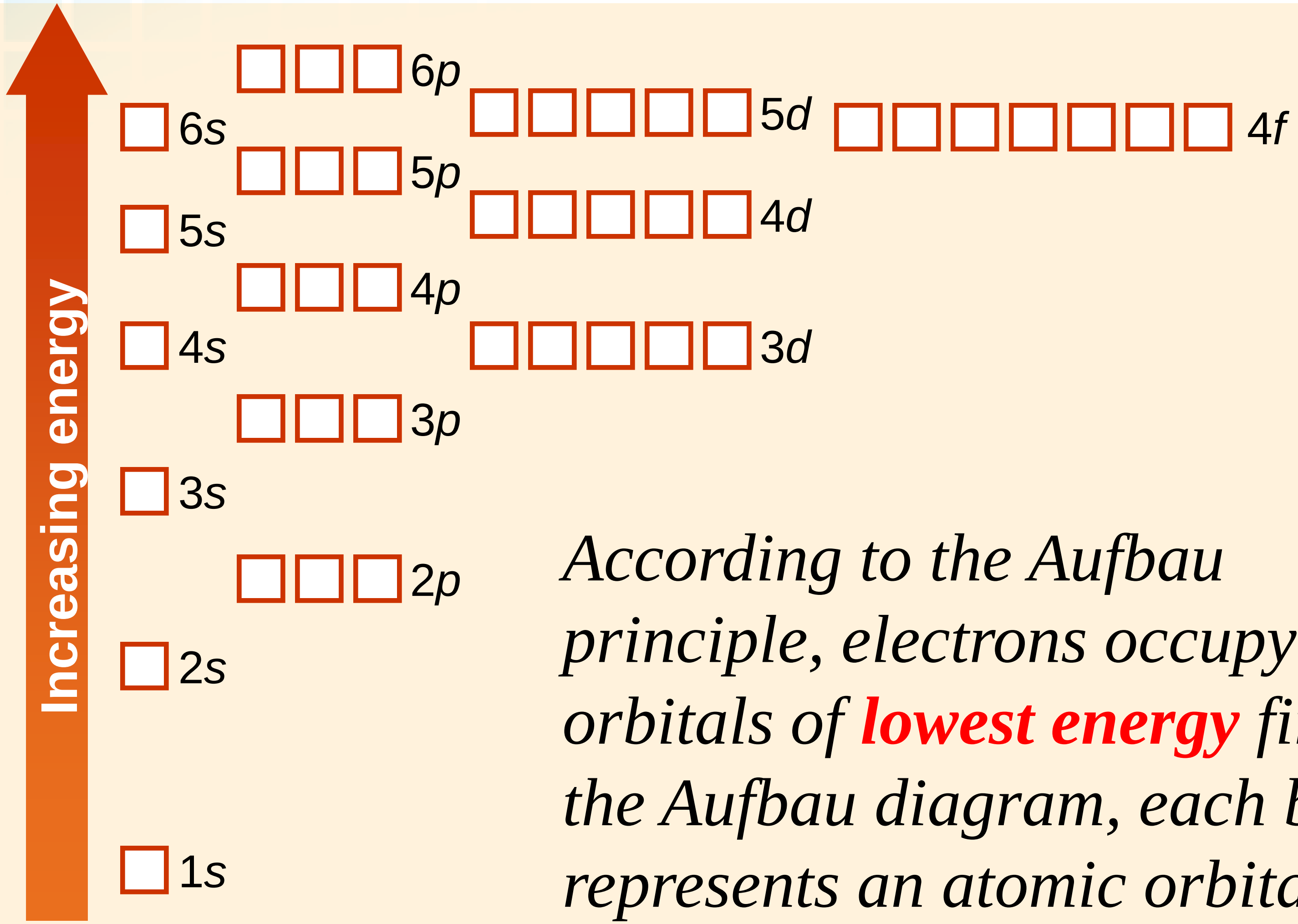


e–
e– orientation

1) triplets (*3 orbitals: p*) with 4 hamburgers
   - Each child (orbital) gets 1 hamburger (e-) before anyone has a 2nd

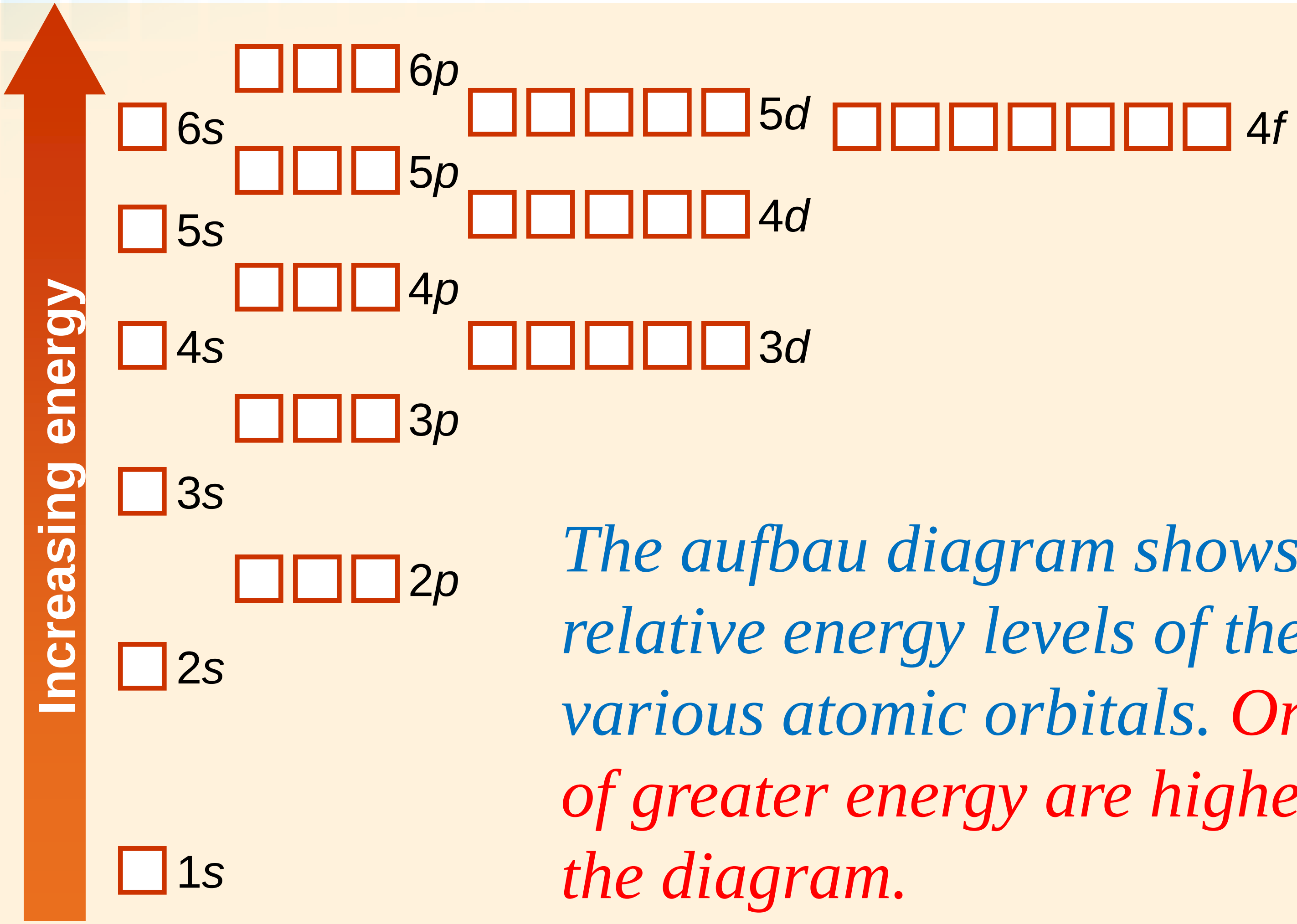
2) quintuplets (*5 orbitals: d*) with 6 hamburgers
   - Each child (orbital) gets 1 hamburger (e-) before anyone has a 2nd

Aufbau Principle



According to the Aufbau principle, electrons occupy orbitals of ***lowest energy*** *first. In the Aufbau diagram, each box represents an atomic orbital.*

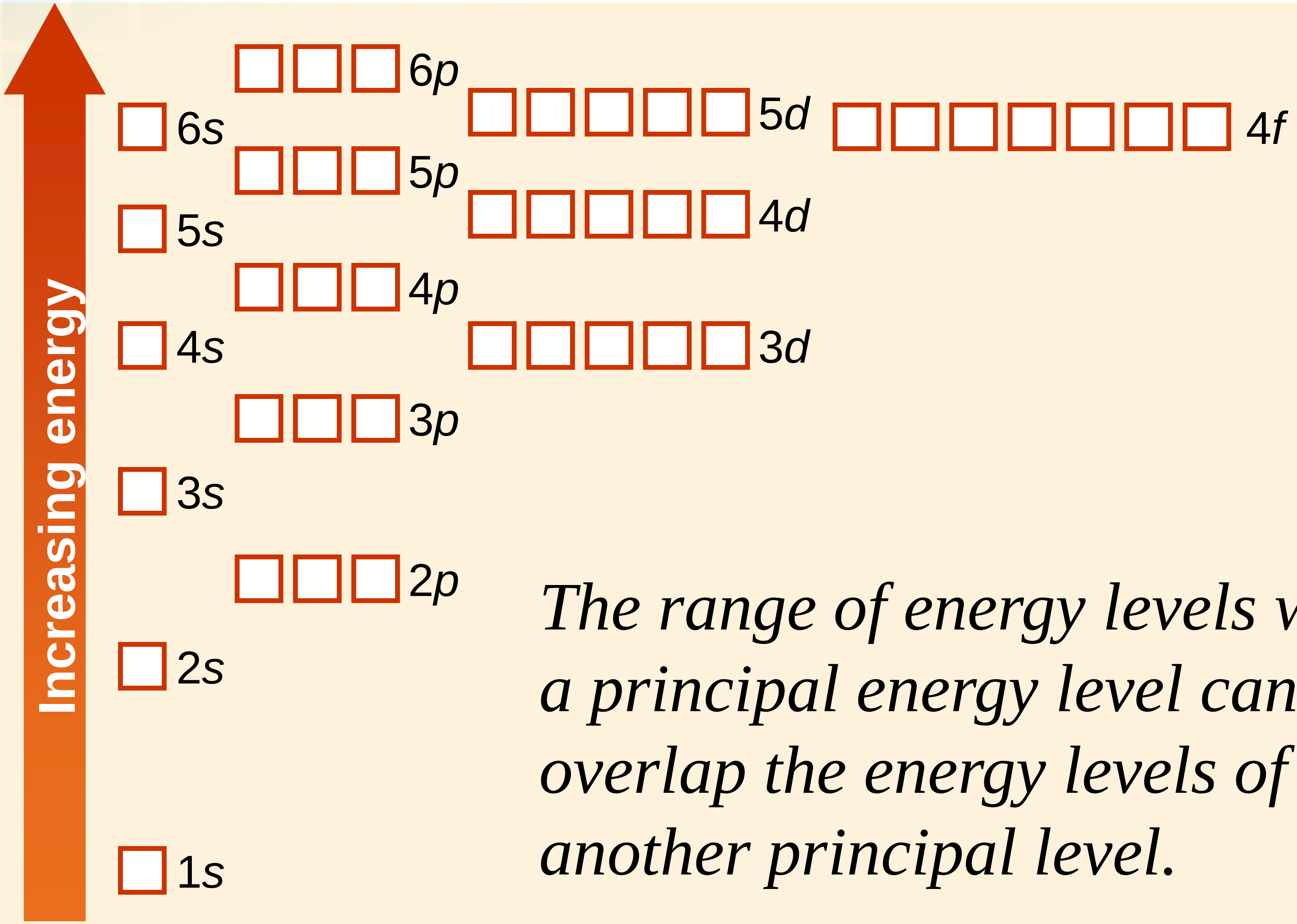
Aufbau Principle



The aufbau diagram shows the relative energy levels of the various atomic orbitals. Orbitals of greater energy are higher on the diagram.

Electron Configurations

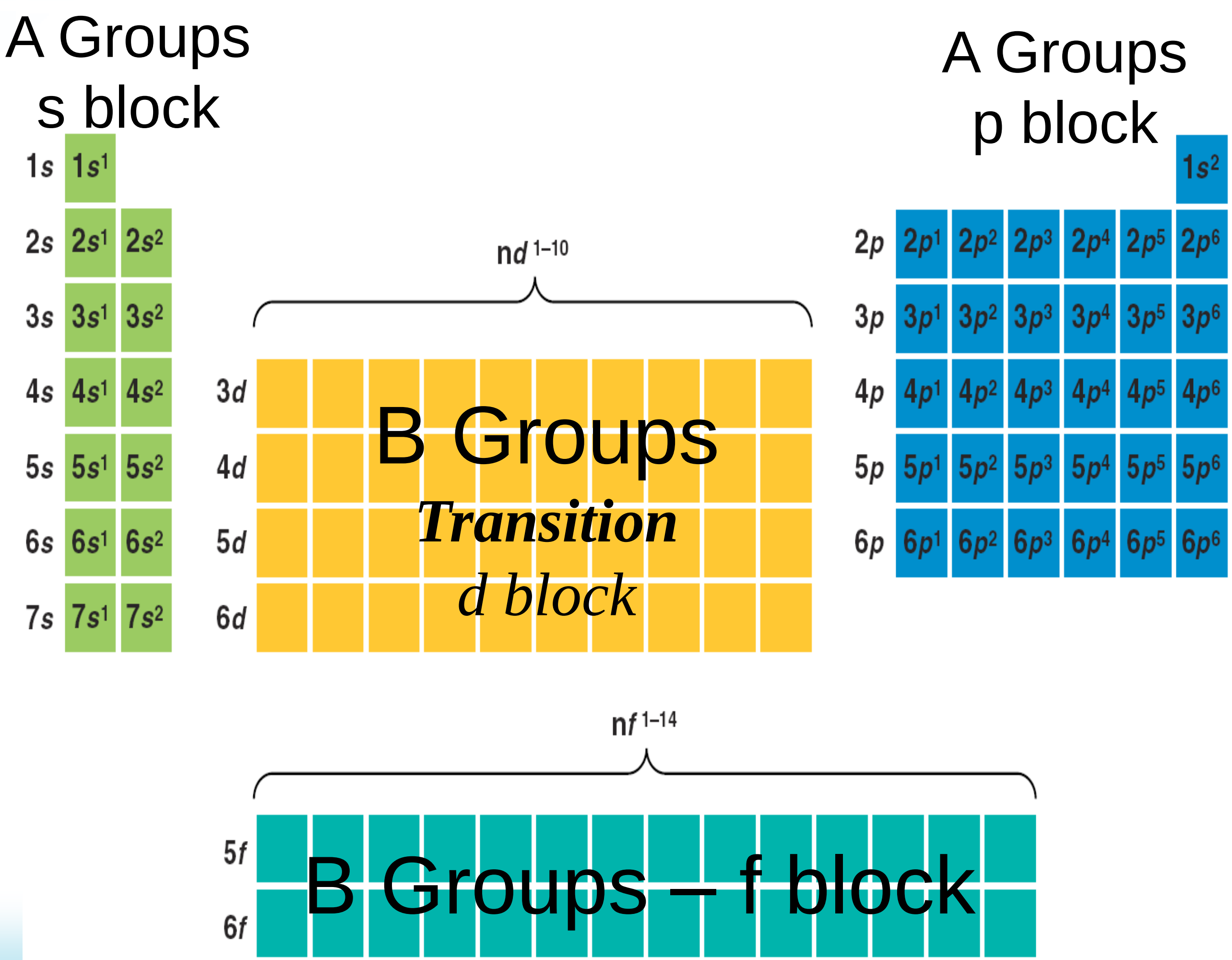
Aufbau Principle



The range of energy levels within a principal energy level can overlap the energy levels of another principal level.

Group "A" and "B" Elements

Create the electron configuration for each element by rows



Energetics of the Quantum Numbers (4:51)

http://somup.com/cFQoolVS98

Show 4 Quantum # (e- configuration) of elements 1 - 11

Electron Configurations of Selected Elements					
Element	**1s**	**2s**	**$2p_x$ $2p_y$ $2p_z$**	**3s**	**Electron configuration**
H					$1s^1$
He					$1s^2$
Li					$1s^2 2s^1$
C					$1s^2 2s^2 2p^2$
N					$1s^2 2s^2 2p^3$
O					$1s^2 2s^2 2p^4$
F					$1s^2 2s^2 2p^5$
Ne					$1s^2 2s^2 2p^6$
Na					$1s^2 2s^2 2p^6 3s^1$

One of these configurations will be on your test

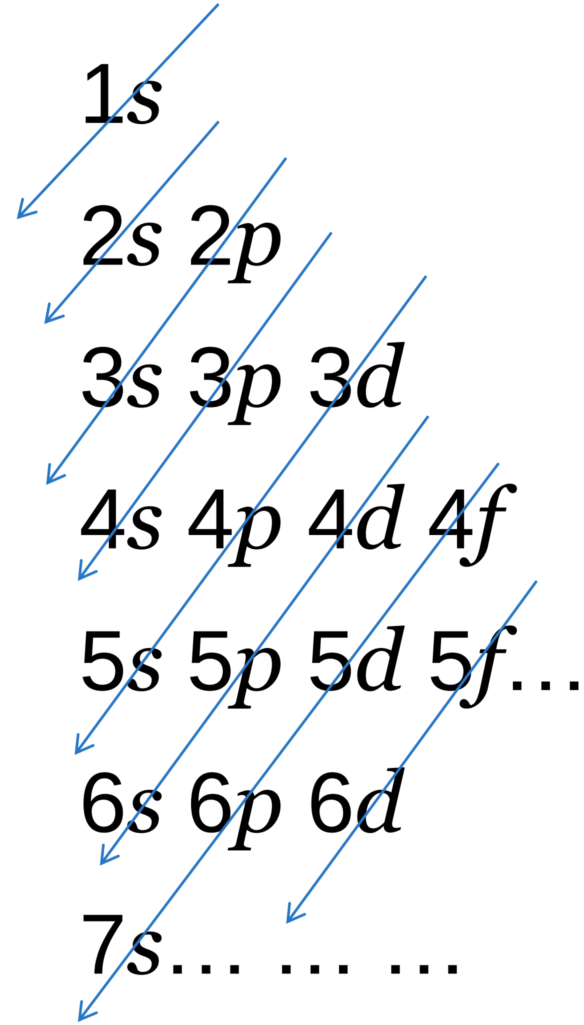
Electron Configurations of Selected Elements

Element	1s	2s	$2p_x$	$2p_y$	$2p_z$	3s	Electron configuration
H	↑						$1s^1$
He	↑↓						$1s^2$
Li	↑↓	↑					$1s^22s^1$
C	↑↓	↑↓	↑	↑			$1s^22s^22p^2$
N	↑↓	↑↓	↑	↑	↑		$1s^22s^22p^3$
O	↑↓	↑↓	↑↓	↑	↑		$1s^22s^22p^4$
F	↑↓	↑↓	↑↓	↑↓	↑		$1s^22s^22p^5$
Ne	↑↓	↑↓	↑↓	↑↓	↑↓		$1s^22s^22p^6$
Na	↑↓	↑↓	↑↓	↑↓	↑↓	↑	$1s^22s^22p^63s^1$

Using the Diagonal Rule to Write Electron Configurations [ENRICHMENT]

http://somup.com/cF6lFQnef7

Diagonal Rule & Electron Configuration Example (3:06)

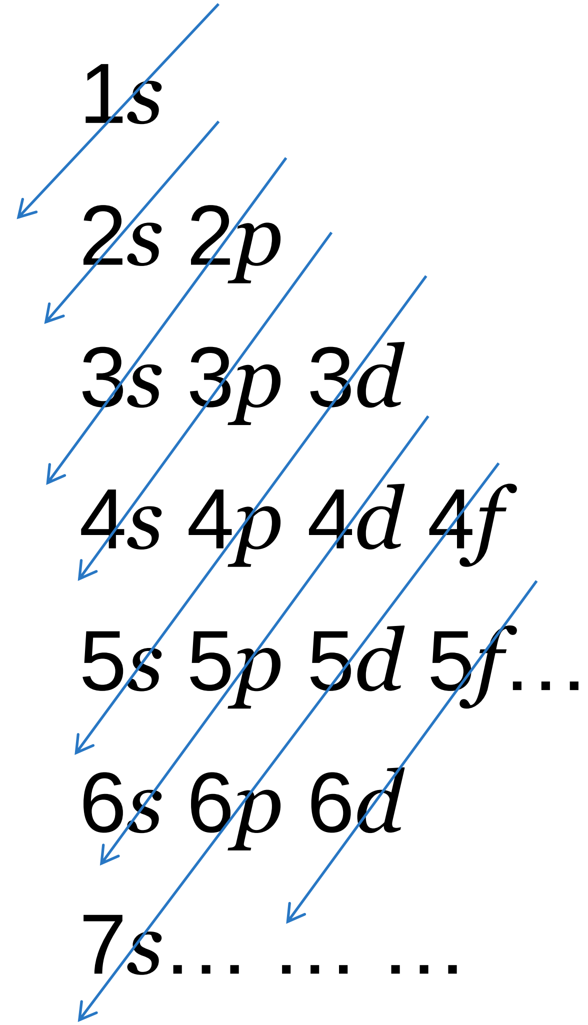


Using the Diagonal Rule to Write Electron Configurations [ENRICHMENT]

EXAMPLE

Write the complete electron configurations for cobalt (Co).

- Identify the number of electrons from the periodic table: ________

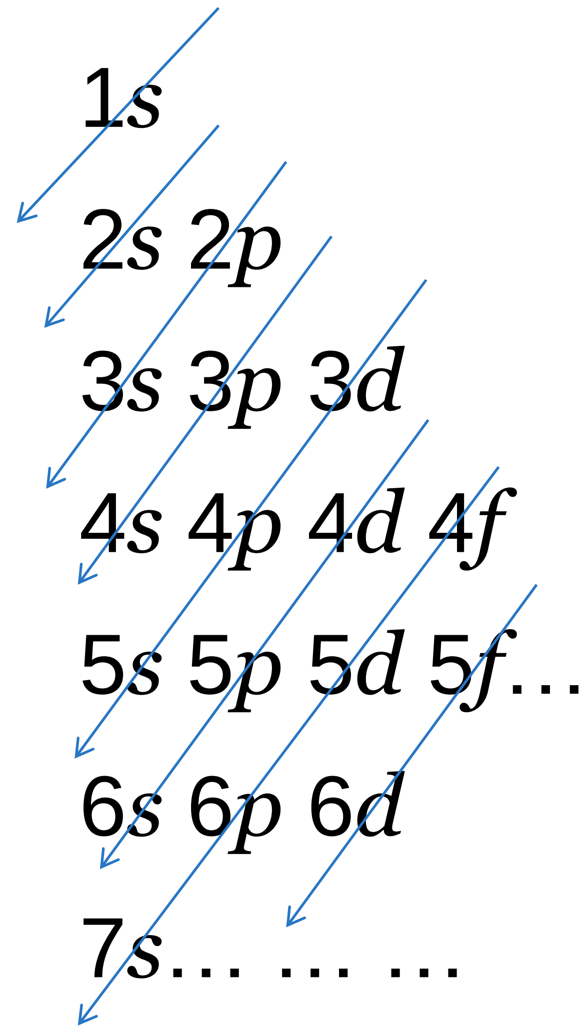


EXAMPLE

Using the Diagonal Rule to Write Electron Configurations [ENRICHMENT]

Write the complete electron configurations for cobalt (Co).

- Identify the number of electrons from the periodic table: ________
- Fill the subshells according to the aufbau principle until the correct number of electrons is reached.



EXAMPLE

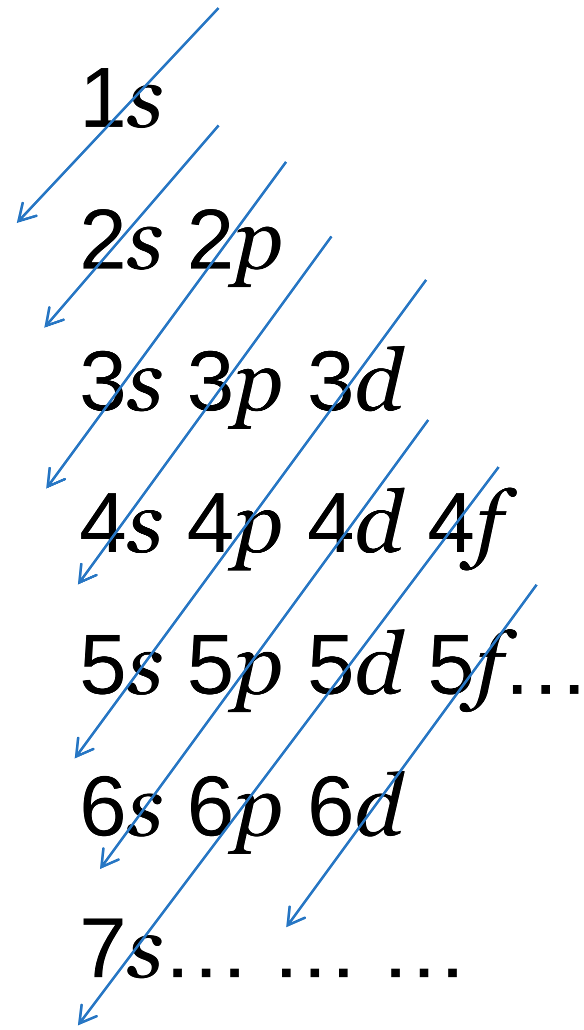
Using the Diagonal Rule to Write Electron Configurations [ENRICHMENT]

Write the complete electron configurations for cobalt (Co).

- Identify the number of electrons from the periodic table: ________
- Fill the subshells according to the aufbau principle until the correct number of electrons is reached.

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^7$

- Check your work.

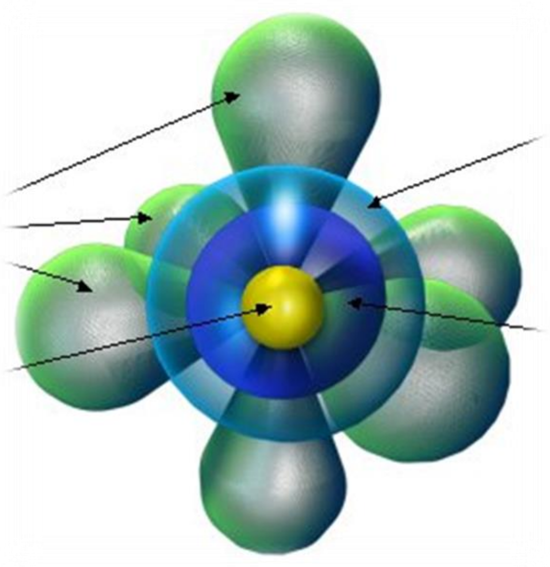


[ENRICHMENT]

Building Atoms: Electron Configuration

- All forms of matter try to stay in their lowest possible energy state.
- Ground state - The lowest possible energy state for a given substance
- Each orbital can hold two electrons
- Orbitals fill in order of increasing energy:

$1s^2$,
$2s^2$, $2p^6$,
$3s^2$, $3p^6$,
$4s^2$, $3d^{10}$, $4p^6$,
$5s^2$, $4d^{10}$, $5p^6$,
$6s^2$, $4f^{14}$, $5d^{10}$, $6p^6$,
$7s^2$, $5f^{14}$, $6d^{10}$, $7p^6$



atom

 [ENRICHMENT]

Exceptional Electron Configurations

Copper and Chromium are exceptions to the electron-filling rules.

The correct electron configurations are as follows:

Cr: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{4s^1}\ \underline{3d^5}$

Cu: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{4s^1}\ \underline{3d^{10}}$

These arrangements give chromium a half-filled *d* sublevel and copper a filled *d* sublevel.

Some actual electron configurations differ from those assigned using the aufbau principle because although half-filled sublevels are not as stable as filled sublevels, they are more stable than other configurations.

[ENRICHMENT]

QUICK CHECK

What is the correct electron configuration of a sulfur atom?

A. $1s^2 2s^2 2p^4 3s^2 3p^6$

B. $1s^2 2s^2 2p^6 3s^2 3p^3$

C. $1s^2 2s^2 2p^6 3s^2 3p^4$

D. $1s^2 2s^2 2p^6 3s^6 3p^2$

5.2 Electron Configurations

QUICK CHECK

What is the correct electron configuration of a sulfur atom?

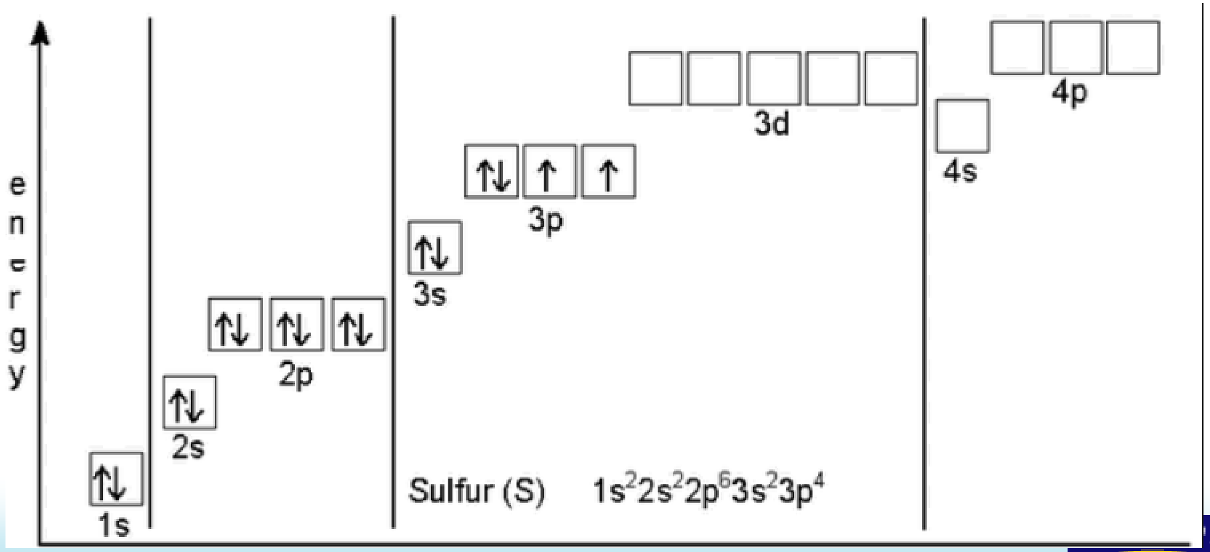
[ENRICHMENT]

A. $1s^2 2s^2 2p^4 3s^2 3p^6$

B. $1s^2 2s^2 2p^6 3s^2 3p^3$

C. $1s^2 2s^2 2p^6 3s^2 3p^4$

D. $1s^2 2s^2 2p^6 3s^6 3p^2$

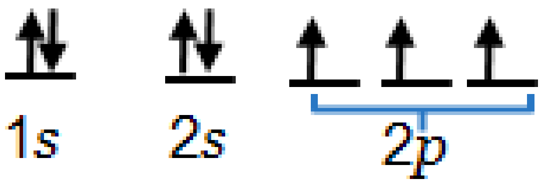


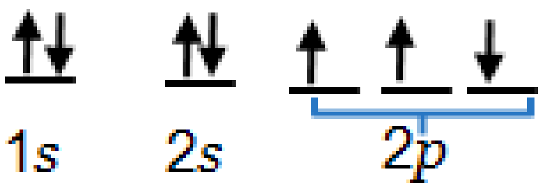
Represent Electron Configurations

TRY IT

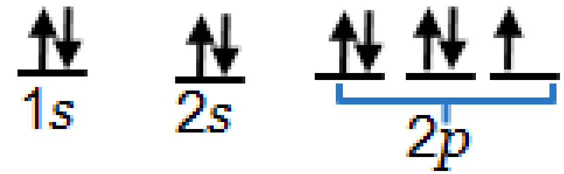
Use the periodic table to answer the questions below.

Which diagram shows the correct electron configuration for nitrogen (N)?





Which diagram shows the correct electron configuration for fluorine (F)?

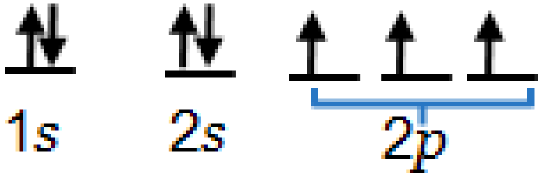


Represent Electron Configurations

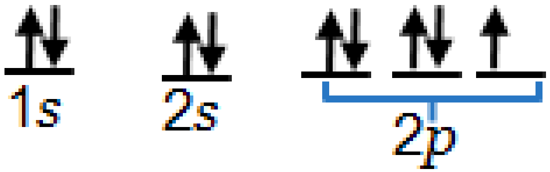
TRY IT

Use the periodic table to answer the questions below.

Which diagram shows the correct electron configuration for nitrogen (N)?



Which diagram shows the correct electron configuration for fluorine (F)?



Describe Electron Configurations

TRY IT

Use the periodic table to give the electron configuration for the element. Show all 4 quantum numbers. Give the nuclear symbol of each element.

Ne:

P:

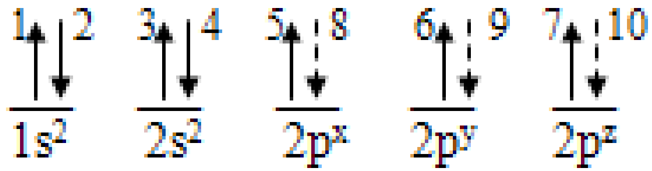
K:

Describe Electron Configurations

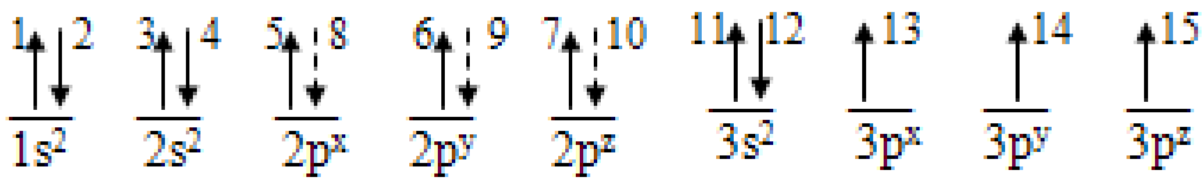
TRY IT

Use the periodic table to give the electron configuration for the element. Show all 4 quantum numbers.

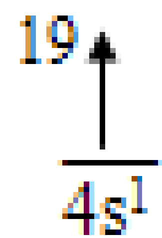
${}_{10}Ne^{20}$: $1s^2 2s^2 2p^6$



${}_{15}P^{31}$: $1s^2 2s^2 2p^6 3s^2 3p^3$



${}_{19}K^{39}$: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ [Ar]



Describe Electron Configurations

TRY IT

Use the periodic table [insert link to periodic table, tab 1] to fill in the numbers in the electron configurations shown below.

B: $1s^2\,2s^{A}2p^{B}$

A =

B =

Na: $1s^2 2s^{C}p^{D}3s^{E}$

C =

D =

E =

Nb: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^{F} 4d^{G}$

F =

G =

Ni: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^{H} 3d^{i}$

H =

I =

Describe Electron Configurations

TRY IT

Use the periodic table [insert link to periodic table, tab 1] to fill in the numbers in the electron configurations shown below.

B: $1s^{2}\,2s^{\mathbf{A}}2p^{\mathbf{B}}$

A = 2

B = 1

Na: $1s^{2}2s^{\mathbf{C}}p^{\mathbf{D}}3s^{\mathbf{E}}$

C = 2

D = 6

E = 1

Nb: $1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}4s^{2}3d^{10}4p^{6}5s^{\mathbf{F}}4d^{\mathbf{G}}$

F = 2

G = 3

Ni: $1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}4s^{\mathbf{H}}3d^{\mathbf{I}}$

H = 2

I = 8

TRY IT

The quantum mechanical model of the atom

a. Defines the exact path of an electron around the nucleus.

b. Was proposed by Neils Bohr.

c. Involves the probability of finding an electron in a certain position.

d. No longer requires the use of energy levels.

The maximum number of electrons in any orbital: ____.

Give the maximum number of electrons in any single energy level for the s sublevel: ____; p sublevel _____; d sublevel _____; f sublevel _____.

The letters s, p, d, f in an electron configuration indicate the ____.

a. Spin of an electron

b. Orbital shape

c. Principle energy level

d. Speed of an electron

Emission of light from an atom occurs when an electron

a. Drops to a lower energy level

b. Jumps to a higher energy level

c. Moves within its atomic orbital

d. Falls into the nucleus

Heisenberg's principle dealt with the uncertainty of ____ and _____ of electrons.

The quantum mechanical model of the atom

~~a. Defines the exact path of an electron around the nucleus.~~

~~b. Was proposed by Neils Bohr.~~

c. Involves the probability of finding an electron in a certain position.

~~d. No longer requires the use of energy levels.~~

The maximum number of electrons in any orbital: 2.

Give the maximum number of electrons in any single energy level for the s sublevel: 2; p sublevel 6; d sublevel 10; f sublevel 14.

The letters s, p, d, f in an electron configuration indicate the ____.

~~a. Spin of an electron~~

b. Orbital shape

~~c. Principle energy level~~

~~d. Speed of an electron~~

Emission of light from an atom occurs when an electron

a. Drops to a lower energy level

~~b. Jumps to a higher energy level~~

~~c. Moves within its atomic orbital~~

~~d. Falls into the nucleus~~

Heisenberg's principle dealt with the uncertainty of position & momentum (speed) of electrons.

Period	s-block 1 IA
1	1.00794 (+1, −1) **H** 1 $1s^1$

KEY

Atomic Mass → 12.0111

Selected Oxidation States → −4, +2, +4

Symbol → **C**

Atomic Number → 6

Electron Configuration → $1s^2 2s^2 2p^2$

Relative atomic masses are based on ^{12}C = 12.00000

s-block

GROUP

New Designation

Former Designation (prior to 1984 IUPAC decision)

d-block

Transition Elements

Period	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIII	9 VIII	10 VIII	
2	6.941 (+1) **Li** 3 $1s^2 2s^1$	9.01218 (+2) **Be** 4 $1s^2 2s^2$									
3	22.98977 (+1) **Na** 11 $[Ne]3s^1$	24.305 (+2) **Mg** 12 $[Ne]3s^2$									
4	39.0983 (+1) **K** 19 $[Ar]4s^1$	40.08 (+2) **Ca** 20 $[Ar]4s^2$	44.9559 (+3) **Sc** 21 $[Ar]3d^1 4s^2$	47.88 (+2, +3, +4) **Ti** 22 $[Ar]3d^2 4s^2$	50.9415 (+2, +3, +4, +5) **V** 23 $[Ar]3d^3 4s^2$	51.996 (+2, +3, +6) **Cr** 24 $[Ar]3d^5 4s^1$	54.9380 (+2, +3, +4, +7) **Mn** 25 $[Ar]3d^5 4s^2$	55.847 (+2, +3) **Fe** 26 $[Ar]3d^6 4s^2$	58.9332 (+2, +3) **Co** 27 $[Ar]3d^7 4s^2$	58.69 (+2, +3) **Ni** 28 $[Ar]3d^8 4s^2$	63.54 29 [Ar]3d
5	85.4678 (+1) **Rb** 37 $[Kr]5s^1$	87.62 (+2) **Sr** 38 $[Kr]5s^2$	88.9059 (+3) **Y** 39 $[Kr]4d^1 5s^2$	91.224 (+4) **Zr** 40 $[Kr]4d^2 5s^2$	92.9064 (+3, +5) **Nb** 41 $[Kr]4d^4 5s^1$	95.94 (+3, +6) **Mo** 42 $[Kr]4d^5 5s^1$	(98) (+4, +6, +7) **Tc** 43 $[Kr]4d^6 5s^1$	101.07 (+3) **Ru** 44 $[Kr]4d^7 5s^1$	102.906 (+3) **Rh** 45 $[Kr]4d^8 5s^1$	106.42 (+2, +4) **Pd** 46 $[Kr]4d^{10} 5s^0$	107.86 47 [Kr]4d
6	132.905 (+1) **Cs** 55 $[Xe]6s^1$	137.33 (+2) **Ba** 56 $[Xe]6s^2$	**La-Lu** 57 71	178.49 (+4) **Hf** 72 $[Xe]4f^{14} 5d^2 6s^2$	180.948 (+5) **Ta** 73 $[Xe]4f^{14} 5d^3 6s^2$	183.85 (+6) **W** 74 $[Xe]4f^{14} 5d^4 6s^2$	186.207 (+4, +6, +7) **Re** 75 $[Xe]4f^{14} 5d^5 6s^2$	190.2 (+3, +4) **Os** 76 $[Xe]4f^{14} 5d^6 6s^2$	192.22 (+3, +4) **Ir** 77 $[Xe]4f^{14} 5d^7 6s^2$	195.08 (+2, +4) **Pt** 78 $[Xe]4f^{14} 5d^9 6s^1$	196.96 79 [Xe]4f
7	(223) (+1) **Fr** 87 $[Rn]7s^1$	226.025 (+2) **Ra** 88 $[Rn]7s^2$	**Ac-Lr** 89 103	(261) **Unq*** 104	(262) **Unp** 105	(263) **Unh** 106	(262) **Uns** 107	**Uno** 108	**Une** 109		* The sys 103 wil

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	10	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 0
				10.81 **B** 5 +3 $1s^2 2s^2 2p^1$	12.0111 **C** 6 −4 +2 +4 $1s^2 2s^2 2p^2$	14.0067 **N** 7 −3 −2 −1 +1 +2 +3 +4 +5 $1s^2 2s^2 2p^3$	15.9994 **O** 8 −2 $1s^2 2s^2 2p^4$	18.998403 **F** 9 −1 $1s^2 2s^2 2p^5$	20.179 **Ne** 10 0 $1s^2 2s^2 2p^6$
				26.98154 **Al** 13 +3 [Ne]$3s^2 3p^1$	28.0855 **Si** 14 −4 +2 +4 [Ne]$3s^2 3p^2$	30.97376 **P** 15 −3 +3 +5 [Ne]$3s^2 3p^3$	32.06 **S** 16 −2 +4 +6 [Ne]$3s^2 3p^4$	35.453 **Cl** 17 −1 +1 +3 +5 +7 [Ne]$3s^2 3p^5$	39.948 **Ar** 18 0 [Ne]$3s^2 3p^6$
	58.69 **Ni** 28 +2 +3 [Ar]$3d^8 4s^2$	63.546 **Cu** 29 +1 +2 [Ar]$3d^{10} 4s^1$	65.39 **Zn** 30 +2 [Ar]$3d^{10} 4s^2$	69.72 **Ga** 31 +3 [Ar]$3d^{10} 4s^2 4p^1$	72.59 **Ge** 32 −4 +2 +4 [Ar]$3d^{10} 4s^2 4p^2$	74.9216 **As** 33 −3 +3 +5 [Ar]$3d^{10} 4s^2 4p^3$	78.96 **Se** 34 −2 +4 +6 [Ar]$3d^{10} 4s^2 4p^4$	79.904 **Br** 35 −1 +1 +5 [Ar]$3d^{10} 4s^2 4p^5$	83.80 **Kr** 36 0 +2 [Ar]$3d^{10} 4s^2 4p^6$
	106.42 **Pd** 46 +2 +4 [Kr]$4d^{10} 5s^0$	107.868 **Ag** 47 +1 [Kr]$4d^{10} 5s^1$	112.41 **Cd** 48 +2 [Kr]$4d^{10} 5s^2$	114.82 **In** 49 +3 [Kr]$4d^{10} 5s^2 5p^1$	118.71 **Sn** 50 +2 +4 [Kr]$4d^{10} 5s^2 5p^2$	121.75 **Sb** 51 −3 +3 +5 [Kr]$4d^{10} 5s^2 5p^3$	127.60 **Te** 52 −2 +4 +6 [Kr]$4d^{10} 5s^2 5p^4$	126.905 **I** 53 −1 +1 +5 +7 [Kr]$4d^{10} 5s^2 5p^5$	131.29 **Xe** 54 0 +2 +4 +6 [Kr]$4d^{10} 5s^2 5p^6$
	195.08 **Pt** 78 +2 +4 [Xe]$4f^{14} 5d^9 6s^1$	196.967 **Au** 79 +1 +3 [Xe]$4f^{14} 5d^{10} 6s^1$	200.59 **Hg** 80 +1 +2 [Xe]$4f^{14} 5d^{10} 6s^2$	204.383 **Tl** 81 +1 +3 [Xe]$4f^{14} 5d^{10} 6s^2 6p^1$	207.2 **Pb** 82 +2 +4 [Xe]$4f^{14} 5d^{10} 6s^2 6p^2$	208.980 **Bi** 83 +3 +5 [Xe]$4f^{14} 5d^{10} 6s^2 6p^3$	(209) **Po** 84 +2 +4 [Xe]$4f^{14} 5d^{10} 6s^2 6p^4$	(210) **At** 85 [Xe]$4f^{14} 5d^{10} 6s^2 6p^5$	(222) **Rn** 86 0 [Xe]$4f^{14} 5d^{10} 6s^2 6p^6$