Chapter 4

Arrangement of Electrons in Atoms

The Development of the New Atomic Model

4.1

Objectives

- **Explain** the mathematical relationship among the speed, wavelength, and frequency of electromagnetic radiation.
- **Discuss** the dual wave-particle nature of light.
- **Discuss** the significance of the photoelectric effect and the line-emission spectrum of hydrogen to the development of the atomic model.
- **Describe** the Bohr model of the hydrogen atom

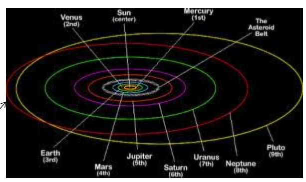
Intro

- Rutherford's model of the atom was better than the previous model but did not explain how e^- were distributed in the atom
  - Opposite charges should attract each other leading to e^- in the nucleus
- The new model used the absorption and emission of light

Misconception

- Electrons do NOT actually orbit like planets

NO



Light Characteristics

- Light was thought of as ONLY a wave but later it was discovered that light has ___________ like properties
- This lead to the discovery of the current model of the _____________

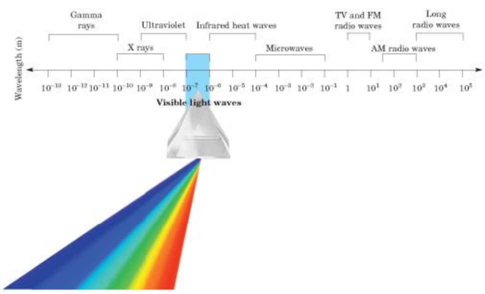
The Wave Description

- **Electromagnetic radiation** is a form of energy that exhibits wavelike behavior as it travels through space.

  Examples:

- Together, all the forms of electromagnetic radiation form the ______________________

Electromagnetic Spectrum



Properties of Waves

- ______________________ is the distance between corresponding points on adjacent waves.
- ______________________is defined as the number of waves that pass a given point in a specific time, usually one second.
- Move at ____________________in a vacuum
  - Slightly less in air because air is made of matter

Props cont.

- Using the frequency and wavelength, you can find the speed of the wave

$$c = \lambda v$$

c = ________________________

λ = ________________________

v = ________________________

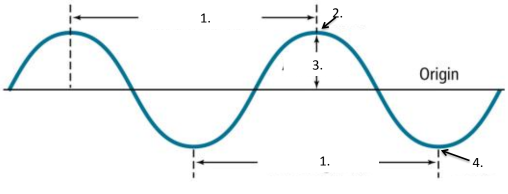
The Wave Nature of Light

- A light is found to have a wavelength of 5.3 x 10^{-7} m. What is the frequency of the light?

- An electromagnetic wave has a frequency of 2.6 x 10^{9} Hz. What is the wavelength?

Props Cont.

1. ____________ 3. _________
2. ____________ 4. _________



Quick Demo

- Super Slinky
- Move back and forth 1/sec
  - Observe Wavelength
- Move back and forth 2/sec
  - Observe wavelength
    - What is the difference?

Photoelectric Effect

- The ________________refers to the emission of electrons from a metal when light shines on the metal
  - Scientists thought that all light should result in electron emission, but this was not true
- It requires a minimum about of NRG to eject and electron
  - Different types of radiation (light) have different amounts of_____________!

Analogy

- You are trying to knock down 3 milk bottles at the county fair
- Would you choose 3 baseballs or 12 ping pong balls?

The Particle Description of Light

- Max Planck discovered light can be small units of NRG, called _______________
- A **quantum** of energy is the ___________ quantity of energy that can be lost or gained by an atom.

The Particle Description of Light

- A **photon** is a particle of electromagnetic radiation having _________mass and carrying a ______________of energy.
- The energy of a particular photon depends on the _______________of the radiation
- Planck went of to propose the following relationship

$$E_{photon} = h\nu$$

E_{photon} = ________________________

h =_____________________________

ν = ____________________________

Practice Problems

- Calculate the NRG of a photon with a frequency of 5.3 x 10^{-7}.
- What is the frequency of a photon that has 2.58 x 10^{-46}J of NRG?

The Particle Description of Light

- Albert Einstein proposed in 1905 that light has a _____________________.
- A beam of light has _________________ and __________________ properties.
- A photon is a particle of electromagnetic radiation with _______ mass that carries a ___________________ of energy.

H Atom Line Emission Spectrum

- The _________________ energy state of an atom is its **ground state.**
  - You can add NRG to an atom which results in an ________________state
- A state in which an atom has a ___________________ potential energy than it has in its ground state is an **excited state.**
  - When it returns to its ground state, it ______________ NRG in the form of ________

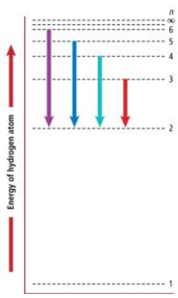
H Atom Line Emission Spectrum

- When investigators passed electric current through a vacuum tube containing hydrogen gas at low pressure, they observed the emission of a characteristic _______________ glow.
- When a narrow beam of the emitted light was shined through a prism, it was separated into _____________ specific colors of the visible spectrum.
- The four bands of light were part of what is known as hydrogen's _______________________

H Atom Line Emission Spectrum

- Scientist expected to see all light (continuous spectrum) but only got the specific spectrum
- **Continuous Spectrum** is a _______________ ____________________________________
- This led them to believe only __________ NRG exist in the H atom (hence the _____ colors)

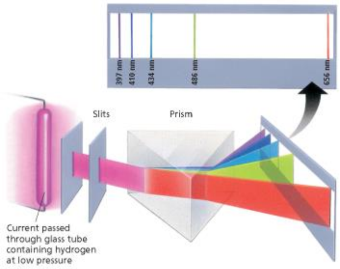
Bohr's Model of the Atom



Demo

- Atomic Line Spectrum for H…and more!

H Atom Line Emission Spectrum



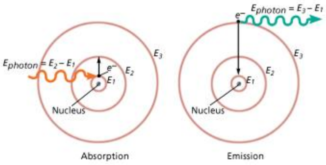
Excited hydrogen atoms emit a pinkish glow. When the visible portion of the emitted light is passed through a prism, it is separated into specific wavelengths that are part of hydrogen's line-emission spectrum.

Bohr Model of the H Atom

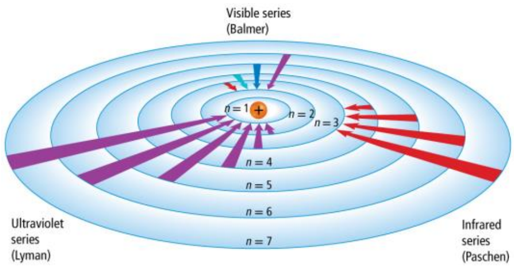
- Niels Bohr proposed a hydrogen-atom model that linked the atom's electron to ______________ emission.
- Bohr stated, the electron can circle the nucleus only in allowed paths, or ___________.
  - The electrons stayed in their ________ when at ____________ , leaving empty space between them and the ____________

Bohr Model of the H Atom

- When the e^- absorbed NRG it moved to a ____________ orbit and them back down to a ____________orbit
  - Ladder Ex.
    - Higher up = More NRG



Other Series



The Quantum Model of the Atom

4.2

Objectives

- **Discuss** Louis de Broglie's role in the development of the quantum model of the atom.
- **Compare and contrast** the Bohr model and the quantum model of the atom.
- **Explain** how the Heisenberg uncertainty principle and the Schrödinger wave equation led to the idea of atomic orbitals.
- **List** the four quantum numbers and **describe** their significance.
- **Relate** the number of sublevels corresponding to each of an atom's main energy levels, the number of orbitals per sublevel, and the number of orbitals per main energy level.

e^- have wave-like props

- French scientist Louis de Broglie suggested that electrons be considered waves confined to the space around an atomic nucleus.
- It followed that the electron waves could exist only at ________________frequencies.
- According to the relationship __________, these frequencies corresponded to specific energies—the quantized energies of Bohr's orbits.

e^- have wave-like props

- Electrons, like light waves, can be bent, or _________________.
- _____________________refers to the bending of a wave as it passes by the edge of an object or through a small opening.
- Electron beams, like waves, can interfere with each other.
- _________________occurs when waves overlap
- Fig 2.1 on page 99

de Broglie equation

- The **de Broglie equation** predicts that all moving particles have wave characteristics.

$$\lambda = \frac{h}{m\nu}$$

λ = ____________________
h = ____________________
m = ____________________
ν = ____________________

- Planck's constant = ____________________

Practice

- An electron has a frequency of 8.2 x 10^{14}. What is the wavelength for this electron?
- An electron has a wavelength of 2.66 x 10^{-7}, what is the electrons frequency?

Heisenberg uncertainty principle

- Theoretical physicist Werner Heisenberg showed it is impossible to take any measurement of an object without disturbing it.
- The **Heisenberg uncertainty principle** states that it is fundamentally __________________ ______________________________________ ______________________________________

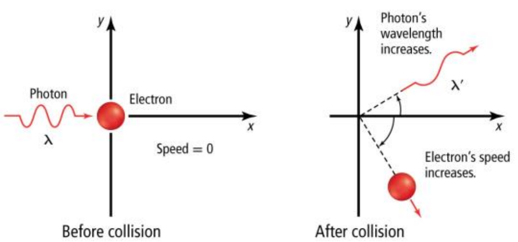
Heisenberg uncertainty principle

- Imagine trying to find a helium filled balloon in a dark room and the room has a large fan running. You run around until you touch it, but this would cause it to move. So…
  - You know where is WAS, not where it IS
- After you strike the balloon, you could measure its velocity, but you do not know the velocity BEFORE you struck it.

Heisenberg uncertainty principle

- Now, you are trying to find an electron!!
- When you "see" and electron, you are seeing light (a photon) that has reflected off of the electron
- The problem is that when the photon hits the electron it changes the position and the velocity.
  - Like the balloon
- Hence, Heisenberg's uncertainty principle

Heisenberg uncertainty principle



The Schrödinger Wave Equation

- In 1926, Austrian physicist Erwin Schrödinger developed an equation that treated electrons in atoms as _________________.
- Together with the Heisenberg uncertainty principle, the Schrödinger wave equation laid the foundation for ________________________.
- **Quantum theory** describes mathematically the wave properties of electrons and other very small particles.

The Schrödinger Wave Equation

- Electrons do not travel around the nucleus in _______________, as Bohr had postulated.
- Instead, they exist in certain regions called ___________________.
- An **orbital** is a _________________________ ______________________________________ ______________________________________

Analogy

- Think of a planes propeller..
  - When it is not moving you can tell where it is
  - When is starts moving, its harder to tell
  - Eventually it is virtually impossible and the propeller looks like a disc
- Electrons
  - Move so fast they are considered to be in a cloud
    - Areas where you <u>could</u> find an electron

Atomic Orbitals and Quantum #s

- ____________________specify the properties of atomic orbitals and the properties of electrons in orbitals.
- The ___________________________, symbolized by ______, indicates the main energy level occupied by the electron.
- The ______________________________ , symbolized by ______, indicates the shape of the orbital.

Atomic Orbitals and Quantum #s

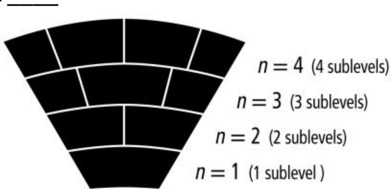
- The ________________________, symbolized by *m,* indicates the orientation of an orbital around the nucleus.
- The ________________________has only two possible values—(+1/2 , −1/2)—which indicate the two fundamental spin states of an electron in an orbital.

Hydrogen's Atomic Orbitals

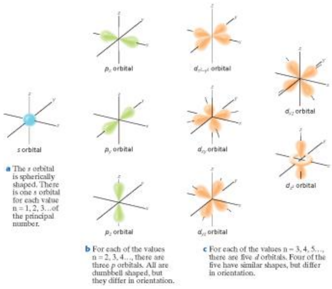
- So, '________' indicates the atoms NRG level
- *'n'* specifies the atom's major energy levels, called the ________________________.
- 'n' can range from _____ to ______, _____ being the lowest possible principle NRG level
  - Periodic Table NRG Level

Hydrogen's Atomic Orbitals

- Each principle NRG level has NRG sublevels
- The sublevels are labeled, ___, ____, ____, and ____



Shapes of s, p, and d Orbitals



Hydrogen's Atomic Orbitals

- Principle NRG level 1 consists of a ____ orbital
- Principle NRG level 2 consists of a ____ orbital and ____ orbitals
- Principle NRG level 3 consists of a ____ orbital, ____ orbitals, and ____ orbitals
- Principle NRG level 4 consists of a ____ orbital, ____ orbitals, ____ orbitals, and ____ orbitals

Hydrogen's Atomic Orbitals

- The most electrons that an orbital can hold is ____
- The s orbitals can have a max of ____ electrons
- The p orbitals can have a max of ____ electrons
- The d orbitals can have a max of ____ electrons
- The f orbitals can have a max of ____ electrons

Principle NRG Levels

Principal energy level	Sublevels available	Number of orbitals in sublevel $(2\ell + 1)$	Number of electrons possible in sublevel $[2(2\ell + 1)]$	Total electrons possible for energy level $(2n^2)$
1				2
2	*s* *p*			
3			2 6 10	
4		1 3 5 7		

What shape is a p orbital?

A. Dumbell
B. Sphere
C. Box
D. None of the above

How many sublevels are there when the principle quantum number is 3

A. 1
B. 2
C. 3
D. 4
E. 7
F. 9
G. 14
H. None of the above

How many orbitals can there be when the principle quantum number is 3

A. 1
B. 2
C. 3
D. 5
E. 9
F. 11
G. 14
H. None of the above

What orbitals are electrons found in for a non excited atom when the principle quantum number = 2?

A. S
B. S and p
C. S, p and d
D. S, p, d, and f
E. None of the above

Electron Configuration

4.3

Objectives

- **List** the total number of electrons needed to fully occupy each main energy level.
- **State** the *Aufbau principle,* the *Pauli exclusion principle,* and *Hund's rule.*
- **Describe** the electron configurations for the atoms of any element using *orbital notation, electron-configuration notation,* and, when appropriate, *noble-gas notation.*

e^- configurations

- **Electron configuration** is the arrangement of e^- in an atom
- Electrons tend to be in the ____________ NRG state possible because this is the ___________________
- This fact will help in determining where the electrons are in an atom
- The most stable, lowest NRG state is called the __________________electron configuration

Electron Configuration

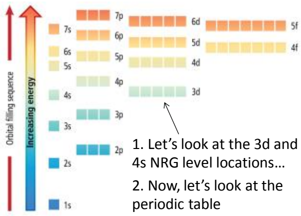
- There are 3 rules that determine where electrons are found:
  1. ________________________________
  2. ________________________________
  3. ________________________________

Electron Configuration

- The ____________ states that each electron occupies the lowest energy orbital available.
- Each box represents an atomic orbital
- Remember…

  2 electrons

  Per orbital



1. Let's look at the 3d and 4s NRG level locations…
2. Now, let's look at the periodic table

Electron Configuration

Features of the Aufbau Diagram	
All orbitals related to an NRG sublevel have equal NRG	All three "2p" orbitals have the same NRG
In multi-electron atoms, the NRG sublevel within a principal NRG level have different energies.	The three "2p" orbitals have more NRG than the "2s" orbitals
In order of increasing NRG, the sequence of NRG sublevels within a principal NRG level is s, p, d, and f	If n = 4, the sequence of NRG sublevels is 4s, 4p, 4d, 4f
Orbitals related to NRG sublevels within one principal NRG level can overlap orbitals related to NRG sublevels within another principal level.	The "4s" sublevel has less NRG than the "3d" orbitals.

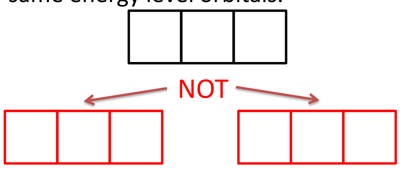
Electron Configuration

- The ___________________ states that a maximum of two electrons can occupy a single orbital, but only if the electrons have _______________ spins.
- An arrow pointing up represents an electron with an "upward" spin
- An unoccupied orbital is an open box (no arrows)

☐ ☐ OR ☐ NOT ☐ ☐

Electron Configuration

- **Hund's rule** states that single electrons with the ________________must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same energy level orbitals.



Practice – Electron Configuration

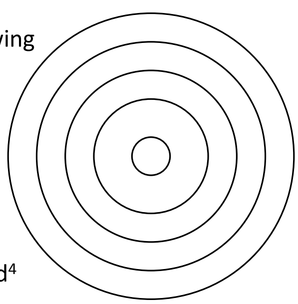
Element	Atomic #	Orbital Diagram	e^- configuration notation
H	1	[↑] [ ] [ ][ ][ ]	$1s^1$
He	2	[↑↓] [ ] [ ][ ][ ]	$1s^2$
Li	3	[↑↓] [↑] [ ][ ][ ]	$1s^2\ 2s^1$
Be		[ ] [ ] [ ][ ][ ]	
F		[ ] [ ] [ ][ ][ ]	

You try Be and F!

Practice

Where are the following Located?
n= 1,2,3,4,5

Place e^- for…
H, He, N, Al, Ca, Br
$1s^2 2s^2 2p^4$
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$



Electron Configuration Notation

How to write electron configuration notation

(NRG Level)(Sublevel)$^{\text{(Electron in sublevel)}}$

Example

H = $1s^1$

He = $1s^2$

Li = $1s^2 2s^1$ Shorthand = [He] $2s^1$

Be = $1s^2 2s^2$ Shorthand = [He] $2s^2$

Electron Configuration Notation (Shorthand)

Element	Atomic #	e^- configuration notation	e^- configuration notation (shorthand)
B	5	$1s^2 2s^2 2p^1$	[He] $2s^2 2p^1$
Mg	12		
	16		
		$1s^2 2s^2 2p^6 3s^2 3p^3$	
			[Ar] $4s^2$

Electron Configuration

- Noble gas notation uses noble gas symbols in brackets to shorten inner electron configurations of other elements.

Table 5.5 Electron Configurations for Elements 11–18

Element	Atomic Number	Complete Electron Configuration	Electron Configuration Using Noble Gas
Sodium	11	$1s^2 2s^2 2p^6 3s^1$	[Ne]$3s^1$
Magnesium	12	$1s^2 2s^2 2p^6 3s^2$	[Ne]$3s^2$
Aluminum	13	$1s^2 2s^2 2p^6 3s^2 3p^1$	[Ne]$3s^2 3p^1$
Silicon	14	$1s^2 2s^2 2p^6 3s^2 3p^2$	[Ne]$3s^2 3p^2$
Phosphorus	15	$1s^2 2s^2 2p^6 3s^2 3p^3$	[Ne]$3s^2 3p^3$
Sulfur	16	$1s^2 2s^2 2p^6 3s^2 3p^4$	[Ne]$3s^2 3p^4$
Chlorine	17	$1s^2 2s^2 2p^6 3s^2 3p^5$	[Ne]$3s^2 3p^5$
Argon	18	$1s^2 2s^2 2p^6 3s^2 3p^6$	[Ne]$3s^2 3p^6$ or [Ar]

Electron Configuration

- The electron configurations (for chromium, copper, and several other elements) reflect the increased stability of half-filled and filled sets of s and d orbitals.

Ex. Cr = [Ar] $4s^2$ $3d^4$ is actually $4s^1$ $3d^5$
Cu = [Ar] $4s^2$ $3d^9$ is actually $4s^1$ $3d^{10}$

This goes for the entire column

Practice

- Page 107
- Page 115
- Page 116

**Side note… the # of arrows needs to match the # of e^- for a neutral atom

The electron configuration for nitrogen is…

1. $1s^1$ $2s^2$ $2p^3$
2. $1s^2$ $2s^2$ $2p^3$
3. $1s^2$ $2s^2$ $2p^5$
4. $1s^2$ $2s^5$
5. Non of the above

The atomic orbital diagram for oxygen is…

1. A [↓↑] [↓↑] [↓↑][↓↑][ ]
2. B [↓↓] [↓↑] [↓↑][↓↑][ ]
3. C [↓↑] [↓↑] [↓↑][↑][↓]
4. None of the above

The electron configuration for Mg is [He] $2s^2$

1. Yes
2. No