CHAPTER 13

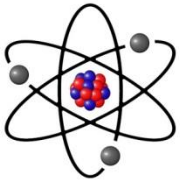
Electrons in Atoms

Atomic Models Review

- Dalton – thought atoms were solid and indivisible
- JJ Thomson – discovered the electron, and made the plum-pudding model

Atomic Models Review

- Rutherford – discovered the nucleus and protons
- Chadwick discovered the neutron
- Millikan determined the charge of an electron

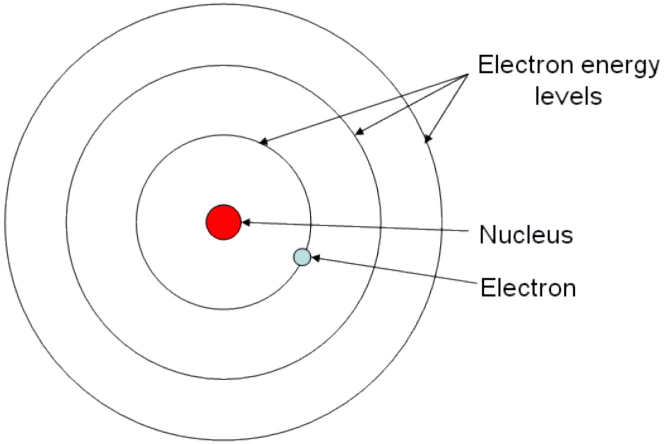


Bohr's Model

- Bohr – proposed that electrons have a fixed energy and move in energy levels around the nucleus – which is why they don't fall into the nucleus

Bohr's Model

- The energy levels are like the rungs of a ladder – electrons cannot be in between levels, and need a specific amount of energy to move from one to another



Chapter 13.1

Quantum Mechanical Model

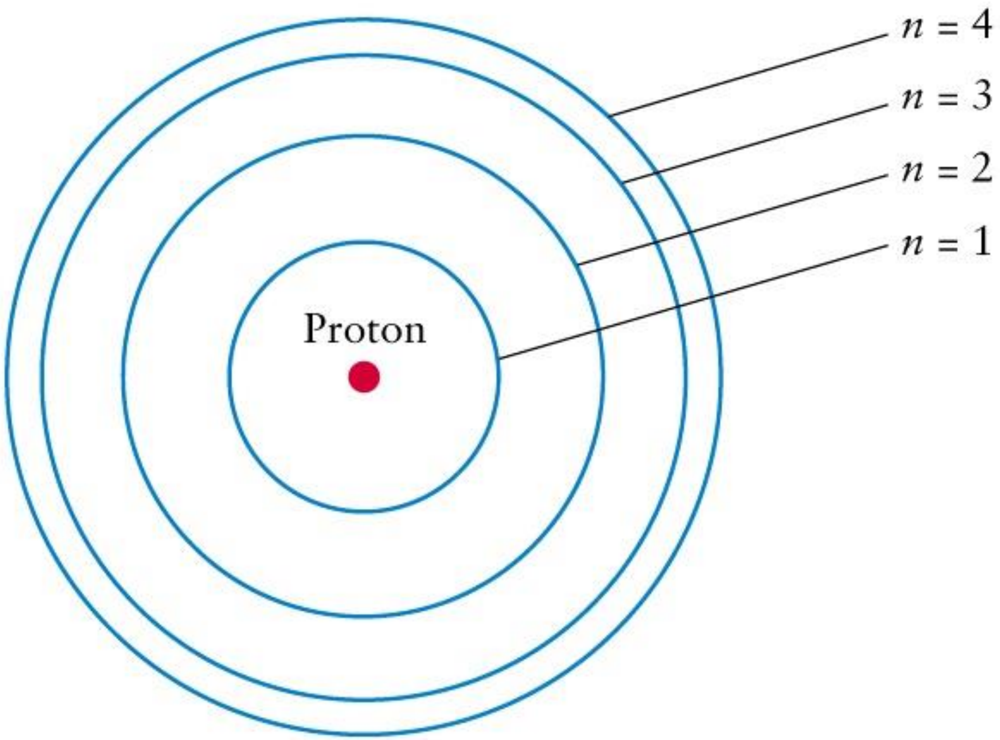
The Current Model

- The newest model of the atom is called the Quantum Mechanical Model
- It is similar to Bohr's model, with some key exceptions

Quantum Mechanical Model

- The energy levels are not equally spaced like a ladder – they get closer the farther from the nucleus you go
- The higher the energy of the e^-, the easier it leaves the atom

Quantum Mechanical Model



Quantum Mechanical Model

- There is no exact path for an e^- to take
- Instead, there are areas of high probability to find an e^-, and areas with a low probability to find an e^-

Quantum Mechanical Model

- The probability of finding an e^- is represented by a fuzzy cloud
- The cloud shows where the electron is ~90 % of the time

Atomic Orbitals

The principle (main) energy levels are assigned numbers according to their energy

- $n = 1, 2, 3, 4$, etc
- 1 has lowest energy

Atomic Orbitals

Each of the principle energy levels has 1 or more sublevels

(The number of sublevels = the energy level #)

- First level (1) has 1 sublevel called the 1s
- Second level (2) has 2 sublevels called 2s & 2p

Atomic Orbitals

- Third level (3) has 3 sublevels called 3s, 3p, and 3d
- Fourth level (4) has 4 sublevels called 4s, 4p, 4d, and 4f

Atomic Orbitals

- Each s sublevel has only one orbital
- Each p sublevel has 3 orbitals: x, y and z
- Each d sublevel has 5 orbitals, xy, xz, yz, x^2-y^2 and z^2
- Each f sublevel has 7 orbitals

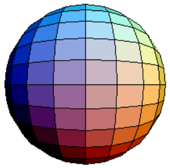
Atomic Orbitals

Energy Level	# of Sublevels							
n=1	1	s						
n=2	2	s	p					
n=3	3	s	p	d				
n=4	4	s	p	d	f			
n=5	5	s	p	d	f	g		
n=6	6	s	p	d	f	g	h	
n=7	7	s	p	d	f	g	h	i

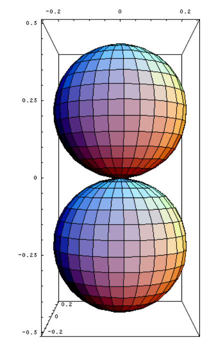
Atomic Orbital Visualizations

Atomic Orbitals

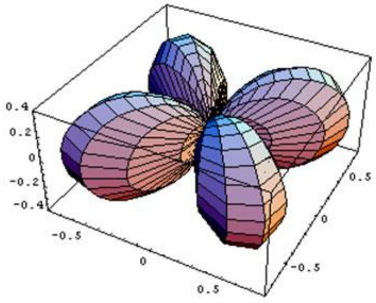
- Each s orbital is spherical



- Each p orbital is dumbbell shaped

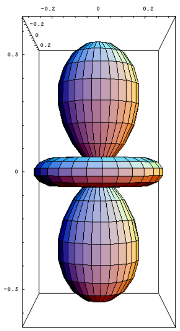


- Four d orbitals are clover shaped



Atomic Orbitals

- One d orbital is shaped like this:



- Each orbital can only hold 2 e^- maximum

Orbital Visualizations

http://winter.group.shef.ac.uk/orbitron/AOs/2s/index.html

http://www.uwosh.edu/faculty_staff/gutow/Orbitals/Cl/Cl_AOs.shtml

Atomic Orbitals

Each orbital may only contain a maximum of 2 electrons

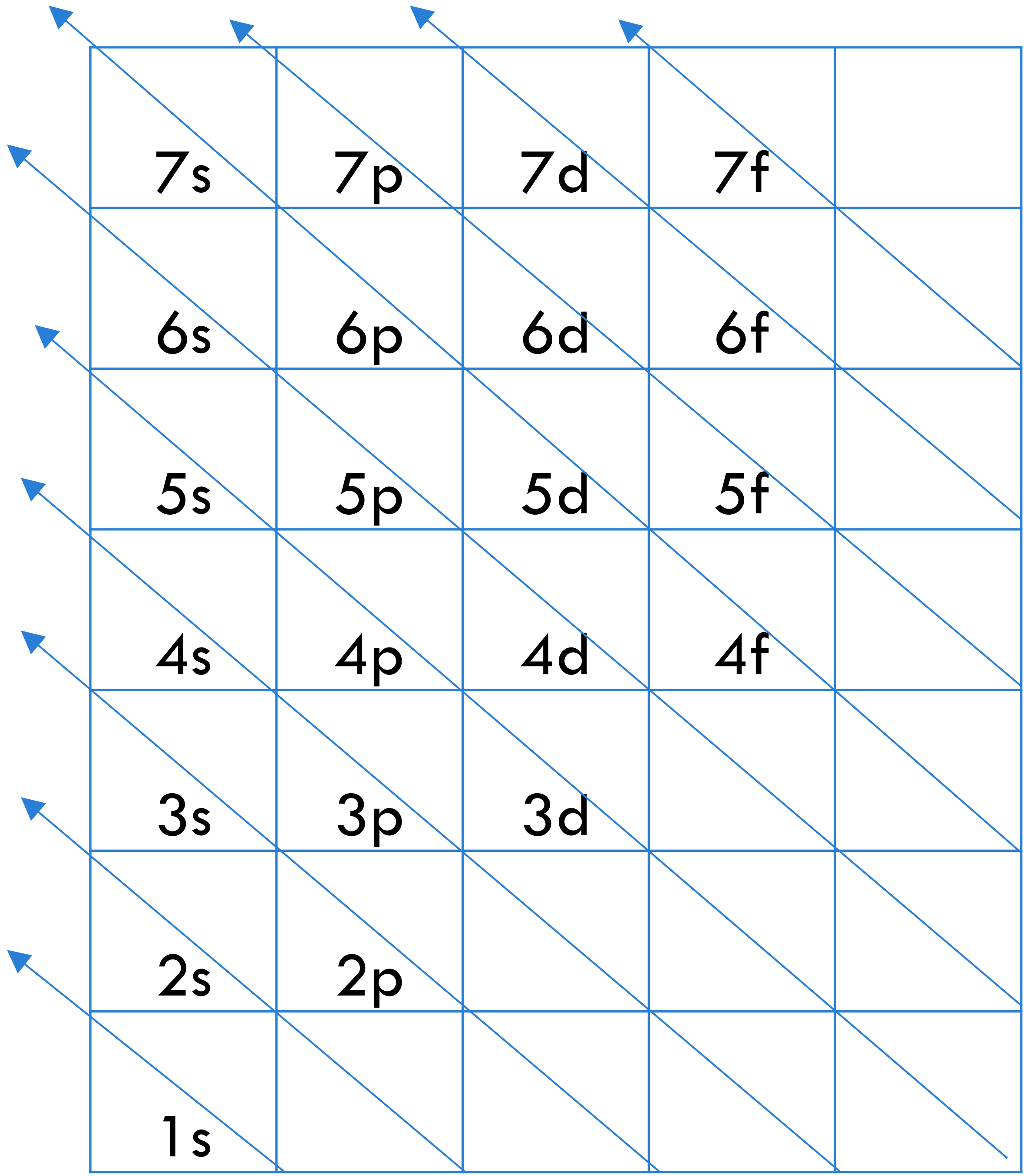
- n = 1 has only 1 sublevel, “s”
- The maximum number of electrons for this level is therefore 2

Atomic Orbitals

Energy Level	# of Sublevels					Total # of Electrons
n=1	1	s				2
n=2	2	s	p			8
n=3	3	s	p	d		18
n=4	4	s	p	d	f	32

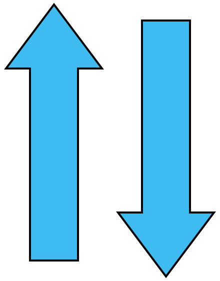
Electrons in orbitals

- **Aufbau Principle:** states that electrons (e^-) enter the lowest energy orbitals first
- The next slide shows the order in which orbitals fill


7s
7p
7d
7f
6s
6p
6d
6f
5s
5p
5d
5f
4s
4p
4d
4f
3s
3p
3d
2s
2p
1s

Electrons in orbitals

- **Pauli Exclusion Principle:** states an orbital can hold a maximum of 2 electrons
- They must have opposite "spins"



Electrons in orbitals

- **Hund's Rule:** States that when orbitals have equal energy, one e^- goes into each orbital before any are filled – each would have 1 e^- with parallel spins

Electrons in orbitals

Hund's rule applies to p and d orbitals:

$\uparrow$	$\uparrow$	$\uparrow$
p_x	p_y	p_z

$\uparrow$	$\uparrow$	$\uparrow$	$\uparrow$	$\uparrow$
d	d	d	d	d

Electrons in orbitals

↑↓	↑↓	↑↓
p_x	p_y	p_z

↑↓	↑↓	↑↓	↑↓	↑↓
d	d	d	d	d

Electron Configuration

To summarize, there are three rules we have to follow:

- Aufbau
- Pauli-Exclusion
- Hund's

Draw orbitals according to the 3 rules

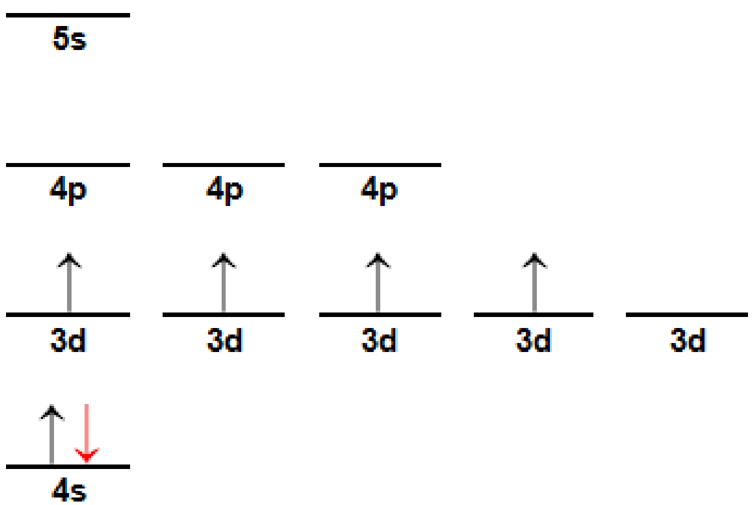
There are a few exceptions to the rules:

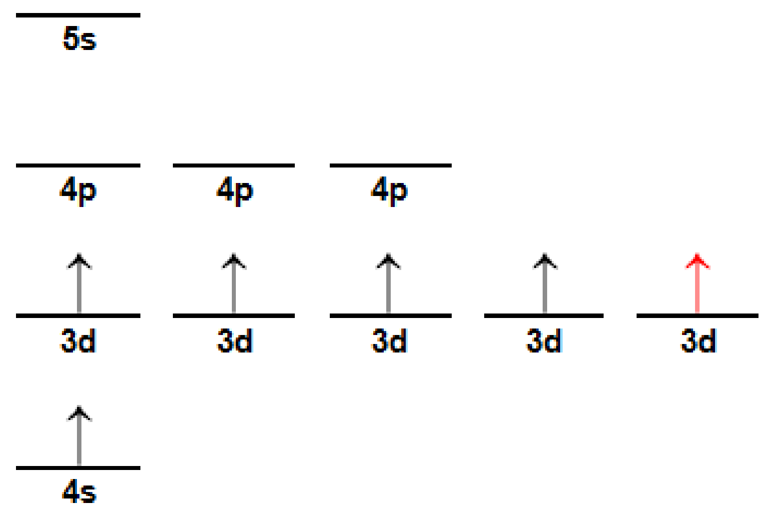
- Cr, Cu, & Mo

Exceptions to Aufbau Rule:

- ^{24}Cr, the aufbau principle predicts the an electron configuration of [Ar]$3d^4 4s^2$ but experimentally we find it to be [Ar]$3d^5 4s^1$

Chromium e^- Configuration

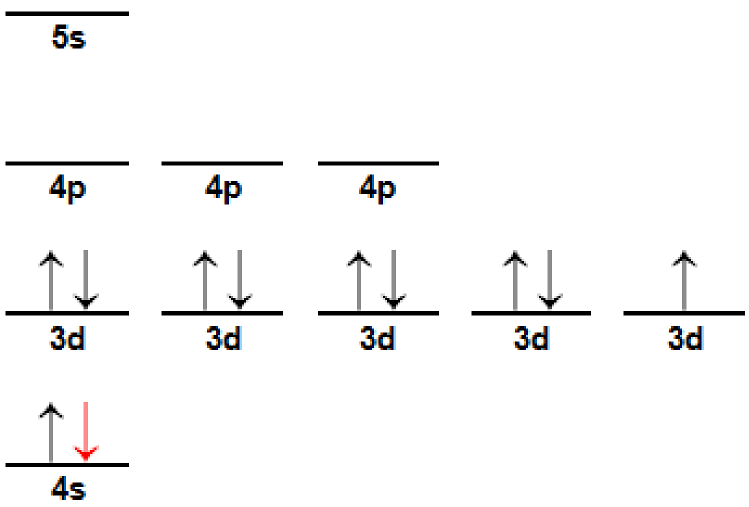


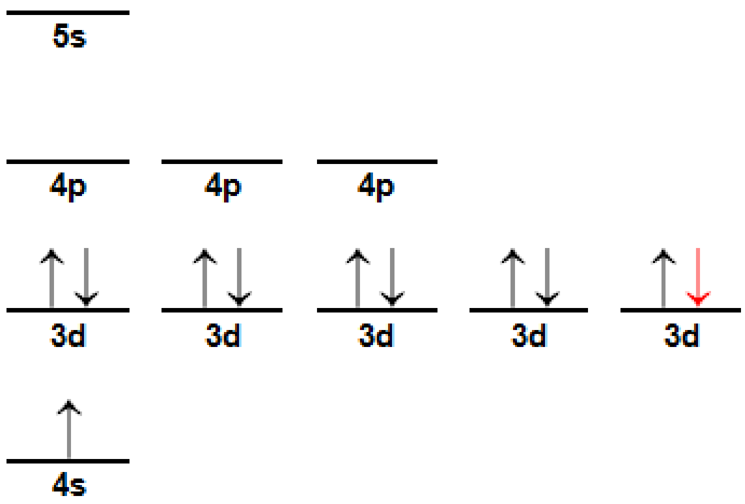


Exceptions to Aufbau Rule:

- ^{29}Cu, the predicted electron configuration is $[Ar]3d^9 4s^2$ but experimentally we find it to be $[Ar]3d^{10} 4s^1$

Copper e^- Configuration

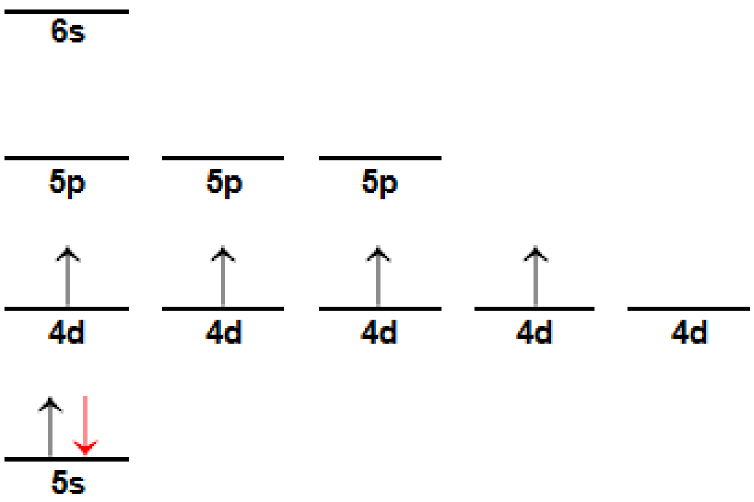


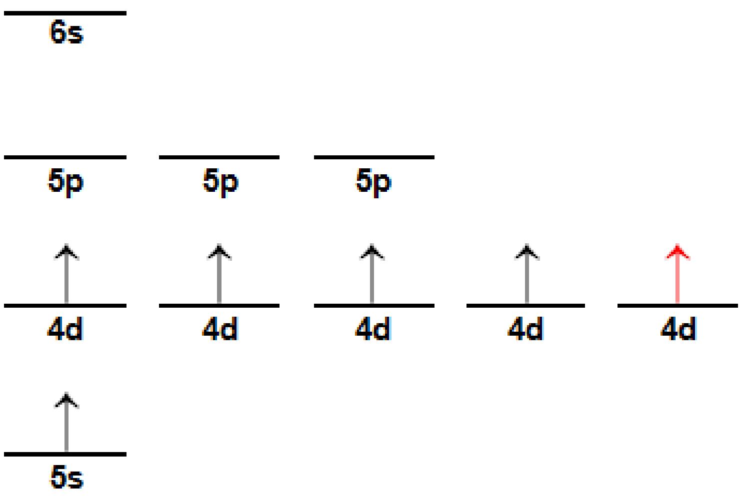


Exceptions to Aufbau Rule:

- ^{42}Mo, the predicted electron configuration is $[Kr]4d^4 5s^2$ but experimentally we find it to be $[Kr]4d^5 5s^1$

Molybdenum e^- Configuration





Electron Configuration

Three ways to represent configurations:

- Orbital Notation (Draw orbital notation)
- Electron Configuration Notation (1 line)
- Noble Gas Notation (as in Periodic Table)

ELECTRONS AND LIGHT

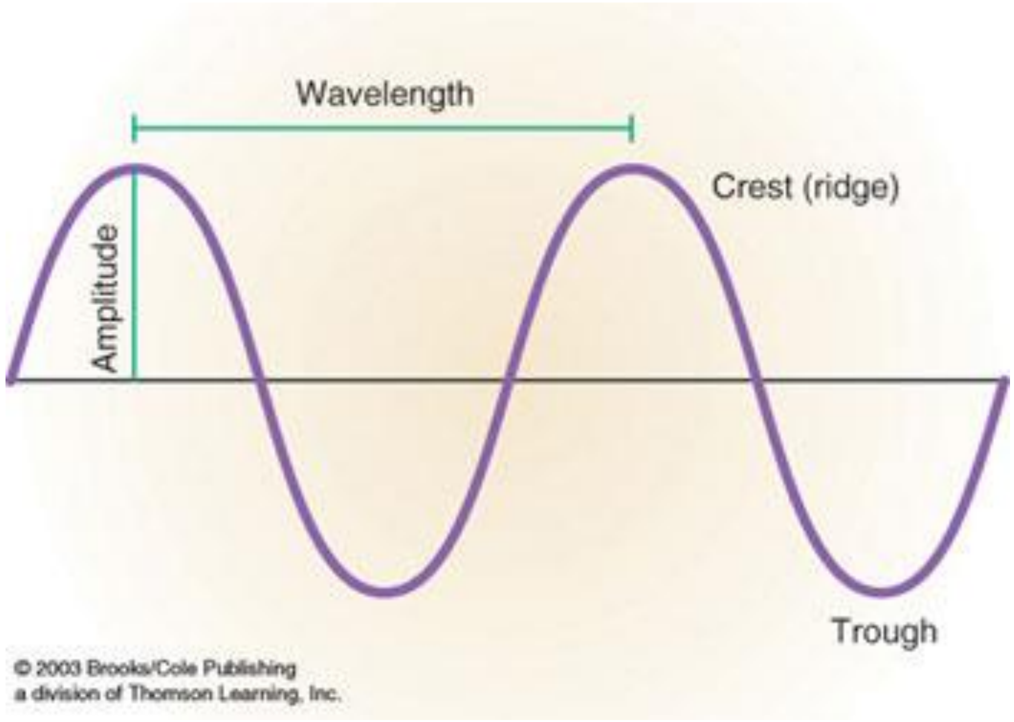
Chapter 13.3

Light & Atomic Spectra

- According to the wave model, light consists of electromagnetic waves
- Includes microwaves, UV rays, visible light, etc.
- Speed = 3.0×10^{10} cm/s

Wave model of light

- Light is a wave with a frequency, speed, and wavelength
- Units of frequency is hertz (Hz) or 1/sec or cycles per second



Light & Atomic Spectra

- Properties of waves: amplitude – wave height, wavelength – distance between the crests of two waves, frequency - # of wave cycles per unit time
- Frequency & wavelength are inversely related (↑↓)

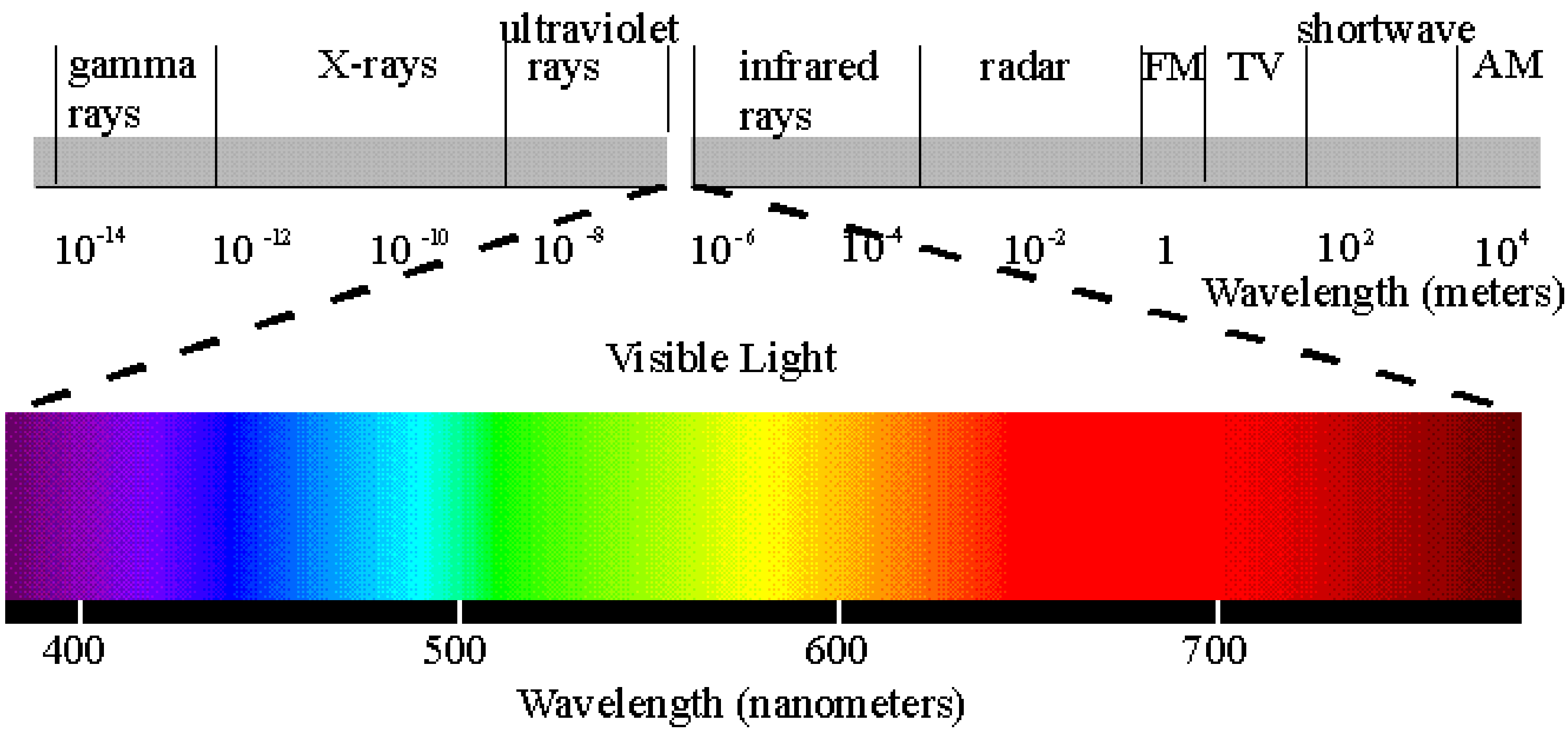
Light & Atomic Spectra

- Speed = wavelength x frequency

$$C = \lambda \nu$$

- Units of frequency are hertz (Hz) or 1/sec or cycles per second

Electromagnetic spectrum



Wave Model of Light

$$c = \lambda \times \nu$$

Speed	Wavelength	Frequency
m/s	m	Hz (hertz) 1/s

c = Speed of light = 3.00 x 10^8 m/s

Light & Atomic Spectra

Sunlight consists of a continuous range of λ and ν

- A prism separates sunlight into a **spectrum** of colors

Each color has a specific λ & ν

- Red = lowest ν, longest λ
- Blue = highest ν, shortest λ

Light & Atomic Spectra

- Elements emit light when "electrocuted" in gaseous form
- Atoms absorb energy, then lose it and emit light
- The light is passed through a prism and an <u>atomic emission spectrum</u> is obtained

Light & Atomic Spectra

- Atomic emission spectra are not continuous like sunlight
- Each line represents one distinct frequency and wavelength
- Every element has a unique emission spectrum

Light & Atomic Spectra

- In lab, you observed (or will observe) the emission spectra of some common gases : helium, nitrogen, hydrogen, argon, neon, bromine, krypton

Max Planck

- Planck said the energy of a body changes only in small discrete units
- The amount of energy is proportional to the frequency (ν)

Planck

$$E = h\,\nu$$

h = Planck's Constant = 6.63×10^{-34} J x s

A small energy change involves an emission or absorption of radiation

Photons

Einstein proposed that light can be described as quanta of energy, called photoelectrons (or photons), that behave as particles

Photoelectric Effect

- When light shines on metals, they emit electrons called photoelectrons
- It takes a specific frequency of light to make certain metal eject a photoelectron

The Hydrogen Spectrum

- Hydrogen has 1 electron
  - In its lowest energy level, it is said to be at its ground state
  - $n = 1$ ground state
- When excited:
  - $n = 2, 3, 4, 5, 6, 7$

The Hydrogen Spectrum

- It absorbs a specific amount of energy, and emits that same amount
- energy is emitted when the e^- falls back to ground state: emission of light

The Hydrogen Spectrum

- There are three groups of lines in the hydrogen spectrum
- Balmer series, is visible
- Lyman series = UV
- Paschen series = IR

The Hydrogen Spectrum

- Lyman series = when e^- fall back to the n = 1 energy level
- Balmer series = when e^- fall to the n = 2 energy level
- Paschen series = when e^- fall to the n = 3 energy level

Quantum Mechanics

De Broglie's Equation:

$$\lambda = h/mv$$

- h = Planck's constant
- m = mass of the particle
- v = velocity of the particle

Quantum vs. Classical

- Classical mechanics predicts the motion of large bodies – energy in any amount
- Quantum mechanics describes motion of subatomic particles – energy in specific quanta

Uncertainty Principle

- Heisenberg Uncertainty Principle – you cannot tell where (position) an e^- is and how fast (velocity) it is going at the same time
- More obvious with small bodies like electrons not baseballs