

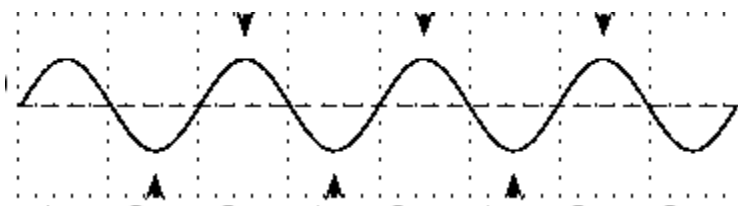
Ch. 4 Notes - ELECTRONS IN ATOMS

NOTE: Vocabulary terms are in **boldface and underlined**. Supporting details are in *italics*.

I. Light and Quantized Energy

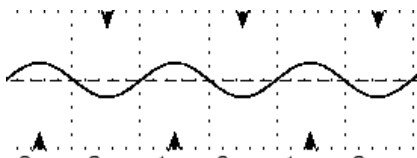
A. The Wave Nature of Light

- 1) *the wave nature of light: light acts as a moving wave*
- 2) wave anatomy and characteristics
 - a) **origin**, zero line, base line—flat line running horizontally, cutting the wave in two; line of zero movement
 - b) **crest** (“peak”)
 - c) **trough** (“valley”)
 - d) **amplitude**—vertical distance from the origin to the crest, or from the origin to the trough
 - e) **wavelength** (λ)—horizontal distance between two equivalent points on a wave, such as between two crests or two troughs
 - f) **wave height**—vertical distance from crest level to trough level; (2 x amplitude)



- g) **frequency** (ν , lowercase Greek letter nu... not a regular v)
 - the number of *cycles per unit time*
 - *measured in Hertz (Hz)—cycles per second, cps, 1/sec, sec⁻¹*

Slow wave with low energy, low frequency, small amplitude, and large wavelength



Fast wave with high energy, high frequency, large amplitude, and small wavelength



3) <i>em wave relationship equation:</i> $c = \lambda \nu$
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$c = \text{speed of light } (3.00 \times 10^8 \text{ m/s}); \quad \lambda = \text{wavelength}; \quad \nu = \text{frequency}$
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EXAMPLE 1) A wave has a frequency of 6.6×10^6 Hz. What is its wavelength?

$$c = \lambda \nu \quad \lambda = \frac{c}{\nu} = \frac{3.00 \times 10^8 \text{ m/s}}{6.6 \times 10^6 \text{ 1/s}} = \boxed{45 \text{ m}}$$

EXAMPLE 2) A wave is estimated to have a wavelength of 4.47×10^{-4} m. What is its frequency?

$$c = \lambda \nu \quad \nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{4.47 \times 10^{-4} \text{ m}} = \boxed{6.71 \times 10^{11} \text{ Hz}} \quad (1/\text{s} = \text{Hz})$$

B. The Particle Nature of Light

- 1) *the particle nature of light: light acts as a stream of moving bundles of light energy called **photons***
- 2) **wave-particle duality of nature**—waves can act as particles, AND particles can act as waves
- 3) **quantum**—a bundle of energy; the minimum amount that can be gained or lost by an atom
- 4) **photoelectric effect**—electrons (photoelectrons) being ejected from a metal when the metal is exposed to high-frequency light
- 5) **E = energy** of a wave, usually measured in Joules (J)

6) energy of a quantum or photon: $E = h \nu$

$E = \text{energy}; \quad h = \text{Planck's constant } (6.626 \times 10^{-34} \text{ Js}); \quad \nu = \text{frequency}$

EXAMPLE 3) A photon in a wave has a frequency of 6.6×10^6 Hz. What is its energy?

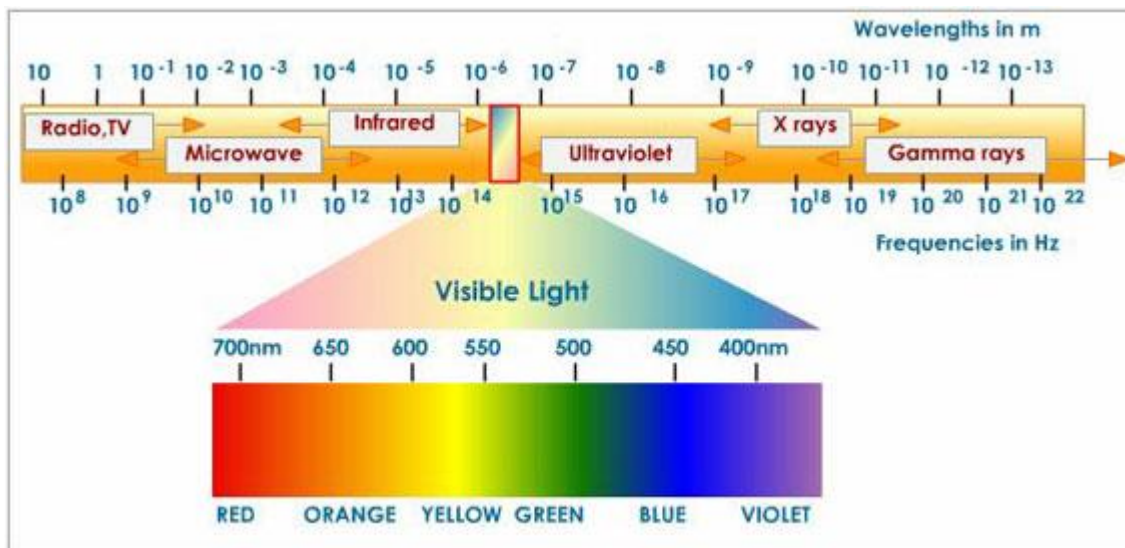
$$E = h \nu \quad E = (6.6 \times 10^6 \text{ 1/s}) (6.626 \times 10^{-34} \text{ Js}) = \boxed{4.4 \times 10^{-27} \text{ J}}$$

EXAMPLE 4) A quantum has a measured energy of 1.333×10^{-20} J. What is its frequency?

$$E = h \nu \quad \nu = \frac{E}{h} = \frac{1.333 \times 10^{-20} \text{ J}}{6.626 \times 10^{-34} \text{ Js}} = \boxed{2.012 \times 10^{13} \text{ Hz}} \quad (1/\text{s} = \text{Hz})$$

II. Spectra

A. **Electromagnetic (em) spectrum**—broad radiation spectrum



(source: Redmcolsoqo)

<p>(low energy) RADIO WAVES ... MICROWAVES ... INFRARED (IR) RADIATION ... VISIBLE LIGHT “ROYGBIV”: red, orange, yellow, green, blue, indigo, violet ... ULTRAVIOLET (uv) ... X-RAYS ... GAMMA RAYS ... COSMIC RAYS (high energy)</p>

B. **atomic emission spectrum**

- 1) array of colors from a heated element, separated by a prism
- 2) set of frequencies of em waves emitted by atoms of an element

III. Electrons and Light

A. evidence for **energy levels**: explanation of atomic spectra

- 1) **ground state**—*electron(s) at their lowest possible energy level*
- 2) **excited state**
 - a) *electron(s) at higher energy than normal*
 - b) *e- absorb energy and jump to higher levels where there is room*

B. Niels Bohr (1885 - 1962); *Bohr model* proposed in 1913

- 1) **quantum number** = n (“ring” number 1-7)
- 2) nucleus as the center, composed of p^+ and n^0
- 3) *e- orbit the nucleus; similar to planetary motion*
- 4) e^- in an orbit have a fixed energy level
- 5) lowest energy levels are closest to the nucleus
- 6) **quantum** (see section I B)—*a bundle of energy* needed to make an electron “jump” to a higher level, which is a **quantum leap**

C. **spectroscopy**—the study of substances exposed to continuous energy

- 1) *when exposed to intense energy, atoms absorb energy and become excited*
- 2) *when atoms are in their excited state, their electrons jump to higher energy levels*
- 3) *when the electrons eventually return to their normal (ground) state, energy is given off (emitted)*

D. quantum mechanical model of the atom

- 1) electrons as waves
- 2) Louis deBroglie (1892-1987)
 - a) all moving particles have wave characteristics

b) particle - em wave relationship equation: $\lambda = \frac{h}{mv}$

λ = wavelength; h = Planck's constant (6.626×10^{-34} Js); m = mass; v = velocity (speed with magnitude and direction) mv = momentum
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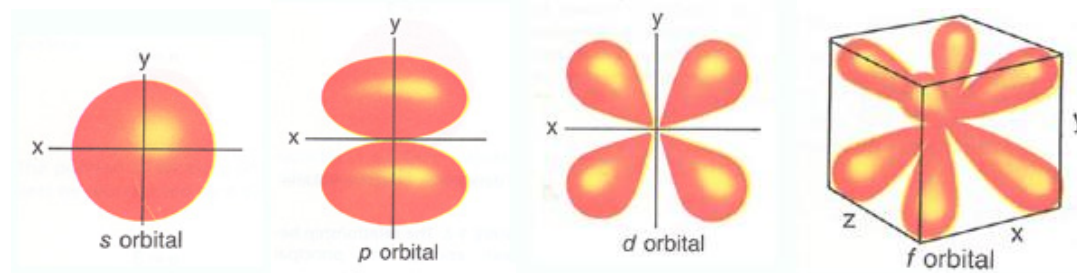
- 3) **Heisenberg Uncertainty Principle**—*it is not possible to know both the exact position and velocity of an object simultaneously*
 - a) *if you measure the position, you have to stop it (technically), so you cannot find the velocity or momentum*
 - b) *if you measure the velocity, it is moving, so you cannot pinpoint one position to measure*
- 4) *Erwin Schrödinger* (1887-1961); quantum mechanical model 1926
 - a) *based on probability of e^- location, not exact path*
 - b) e^- cloud model; “boundary surface diagram”
 - c) equation:
$$\frac{\rho^2 \psi}{\rho x^2} + \frac{\rho^2 \psi}{\rho y^2} + \frac{\rho^2 \psi}{\rho z^2} + \frac{8\pi^2 m}{h^2} (E-V)\psi = 0$$
 - d) e^- are mathematically treated as a wave
 - e) wave function—amplitude (ψ)

- f) solved the equation for a series of wave functions (definite energy states of an e^- and the probability of finding it there)

- IV. Electron configurations—Distribution of Electrons in Energy Levels
- electron configuration**—specific e^- arrangement in orbitals
 - atomic orbital**—area of high probability (~95%) of finding electrons
 - principal quantum number** = n
 - electron energy level or “shell” number
 - $n = 1, 2, 3, 4, 5, 6, 7$
 - energy **sublevels**—divisions within a level
 - many similar energy states grouped together in a level
 - different shapes: spherical, dumbbell, cloverleaf
 - sublevels: s, p, d, f**
(s p d f stand for sharp, principal, diffuse, fundamental)
 - maximum number of e^- in a principal energy level = $2n^2$

<u>n</u>	<u>SUBLEVELS</u>
n = 1	1s
n = 2	2s, 2p
n = 3	3s, 3p, 3d
n = 4	4s, 4p, 4d, 4f
n = 5	5s, 5p, 5d, 5f
n = 6	6s, 6p, 6d
n = 7	7s, 7p

SHAPES OF ORBITALS (boomeria.org)



E. electron capacity of the sublevels

s can hold $2 e^-$ p can hold $6 e^-$ d can hold $10 e^-$ f can hold $14 e^-$

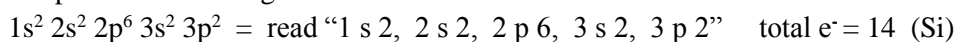
s	p	d	f

- orbital = one “box;” can also be a line or a small circle
- electrons = “arrows” (can be drawn as full \uparrow or half \uparrow arrows)
- antiparallel spin**—electrons in an orbital spin in opposite directions, so as not to repel one another (this is shown as arrows pointing in opposite directions)

4) rules

- **Aufbau Principle**— e^- enter low-energy orbitals first
- **Pauli Exclusion Principle**—no more than two e^- per orbital; the e^- must have compatible spins (shown by \uparrow and \downarrow)
- **Hund's Rule**—degenerate orbitals (orbitals of equal energy) have one e^- filling each orbital before pairs can exist

5) sample electron configuration:



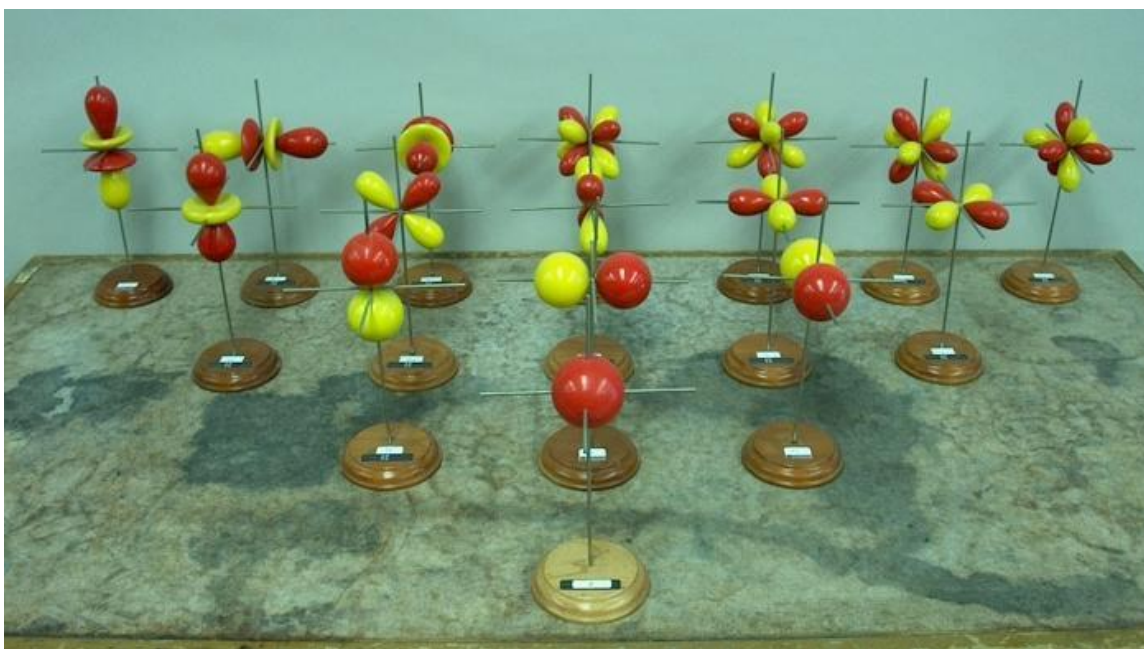
- 6) *the sum of the superscripts equals the atomic number of the element*
- 7) *superscripts are NOT exponents (nothing is being squared, etc.)*
- 8) *** *normal valence configurations will be s OR s and p ****
- 9) ways to do electron configurations: Aufbau (diagram or periodic table location)
Aufbau diagram = orbital-filling diagram, orbital notation
- 10) Examples with the Aufbau diagram (see last page of notes for Aufbau diagram)

EXAMPLE 5) Write out the electron configuration for nitrogen.

EXAMPLE 6) Write the electron configuration for cadmium.

ANSWERS: E 5) $1s^2 2s^2 2p^3$

E 6) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$



V. The Periodic Table and Atomic Structure

A. electron configurations according to location

- 1) Group IA (1) and IIA (2) = end in s
- 2) Group IIIA-VIIIA (13-18) = end in p
- 3) B Groups (Groups 3-12 transition metals) = end in d
- 4) Inner transition metals = end in f
- 5) *Noble gases configurations end in $s^2 p^6$*
- 6) *Pseudo-Noble gas configurations are exceptions = $s^2 p^6 d^{10}$*

