

Ch. 5 Notes - ELECTRONS IN ATOMS

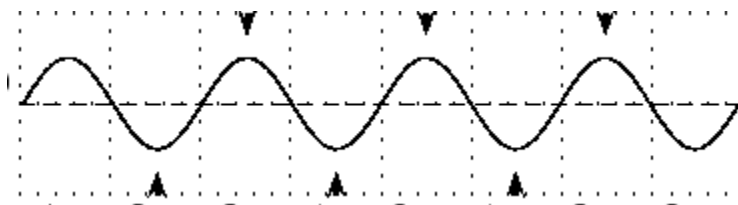
NOTE: Vocabulary terms are in **boldface and underlined**. Supporting details are in *italics*.

5.1 Notes

I. Light and Quantized Energy

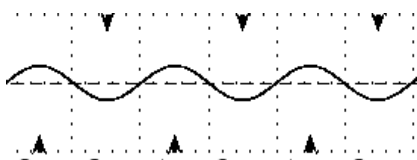
A. The Wave Nature of Light

- 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)
 - a. the number of *cycles per unit time*
 - b. *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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$c = \text{speed of light } (3.0 \times 10^8 \text{ m/s}); \quad \lambda = \text{wavelength}; \quad \nu = \text{frequency}$

E 1) A wave has a frequency of 6.6×10^6 Hz. What is its wavelength?

$$c = \lambda \nu \quad \lambda = \frac{c}{\nu} = \frac{3.0 \times 10^8 \text{ m/s}}{6.6 \times 10^6 \text{ 1/s}} = \boxed{45 \text{ m}}$$

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

$$c = \lambda \nu \quad \nu = \frac{c}{\lambda} = \frac{3.0 \times 10^8 \text{ m/s}}{4.47 \times 10^{-4} \text{ m}} = \boxed{6.7 \times 10^{11} \text{ Hz}} \quad (1/\text{s} = \text{Hz})$$

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) <i>energy of a quantum or photon: E = h v</i>

<i>E = energy; h = Planck's constant (6.626 x 10⁻³⁴ Js); v = frequency</i>

E 3) A photon in a wave has a frequency of 6.6 x 10⁶ Hz. What is its energy?

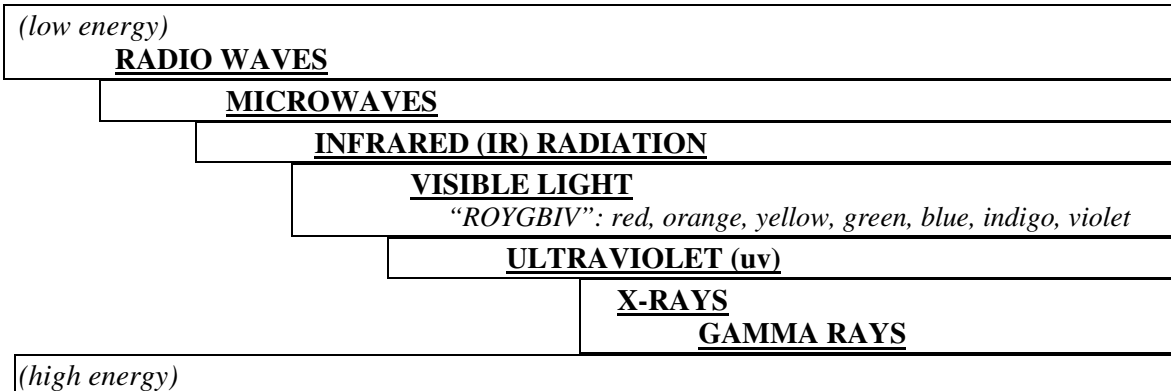
$$E = h v \quad E = (6.6 \times 10^6 \text{ 1/s}) (6.626 \times 10^{-34} \text{ Js}) = \boxed{4.4 \times 10^{-27} \text{ J}}$$

E 4) A quantum has a measured energy of 1.333 x 10⁻²⁰ J. What is its frequency?

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

II. Spectra

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



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*

5.2 Notes

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 n⁰
 - 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 sec. 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); model proposed 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)

5.3 Notes

- IV. Electron configurations—Distribution of Electrons in Energy Levels
- A. **electron configuration**—specific e⁻ arrangement in orbitals
 - B. **atomic orbital**—area of high probability (~95%) of finding electrons

C. **principal quantum number** = n

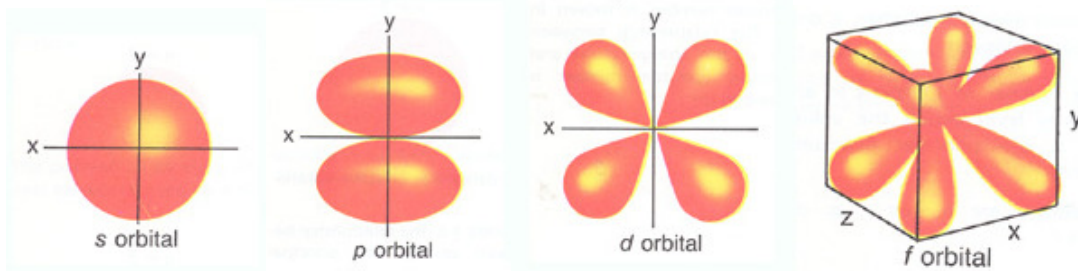
- 1) *electron energy level or “shell” number*
- 2) $n = 1, 2, 3, 4, 5, 6, 7$

D. energy **sublevels**—*divisions within a level*

- 1) many similar energy states grouped together in a level
- 2) different shapes: spherical, dumbbell, cloverleaf
- 3) **sublevels: s, p, d, f**
(s p d f stand for sharp, principal, diffuse, fundamental)
- 4) *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 – where the new elements are being placed
n = 7	as of now: 7s, 7p

SHAPES OF ORBITALS from <http://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

- 1) *orbital = one “box;” can also be a line or a small circle*
- 2) *electrons = “arrows”*
- 3) **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:
 $1s^2 2s^2 2p^6 3s^2 3p^2$ = read “1 s 2, 2 s 2, 2 p 6, 3 s 2, 3 p 2” total $e^- = 14$ (Si)
- 6) *the sum of the superscripts equals the atomic number of the element*
- 7) *superscripts are NOT exponents* (nothing is being squared, etc.)
- 8) *** *valence configurations will be s OR s and p* ***
- 9) ways to do electron configurations: Aufbau diagram or periodic table location
- 10) Examples with the Aufbau diagram (see last page of notes for Aufbau diagram)

E5) Write out the electron configuration for nitrogen.

E6) Write the electron configuration for cadmium.

ANSWERS: E5) $1s^2 2s^2 2p^3$ E6) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$

5.3/6.2 Notes (The rest of Ch. 6 was already covered before Ch. 5)

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 gas configuration** = $s^2 p^6$
- 6) **Pseudo-Noble gas configuration** (an exception; see VI. C) = $s^2 p^6 d^{10}$

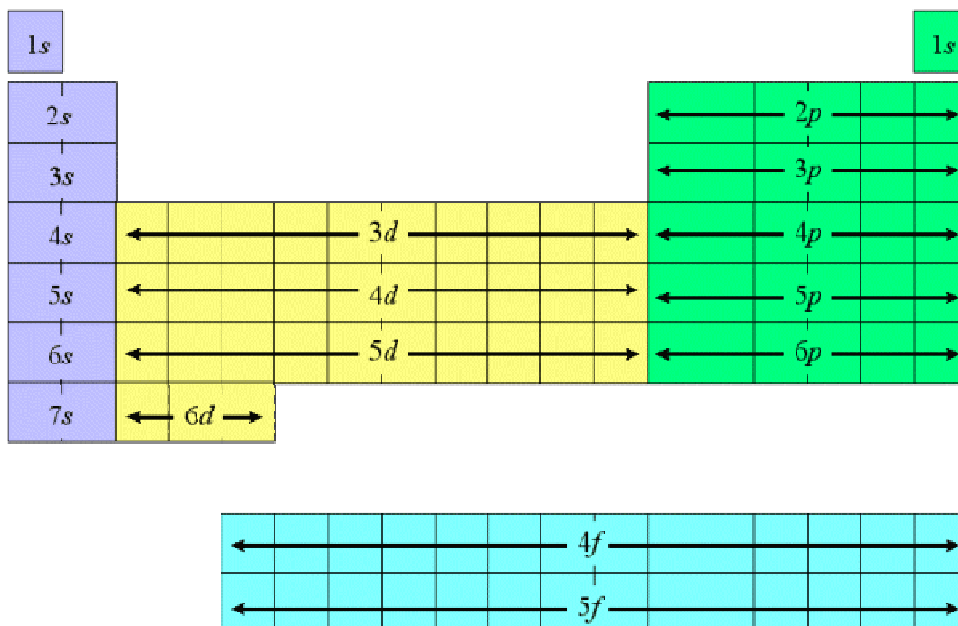
B. Complete vs. valence electron configurations

- 1) *complete*—lists all orbitals (superscripts match the atomic number)
- 2) *valence*—only lists the outer shell (matches the dot diagram)
 DOTS: “right, left, up, down, top all the way around.”

C. Why are valence configurations s OR s/p and not normally include d or f?

- 1) *d orbitals* are “step-downs” – they are one level below the valence
- 2) *f orbitals* are “two step-downs” – they are two levels below the valence

ELECTRON CONFIGURATIONS AND THE PERIODIC TABLE (from FSU)



GROUP #	VALENCE #	SAMPLE VALENCE CONFIG.	# DOTS
IA (1)	1	$3s^1$	1
IIA (2)	2	$3s^2$	2
IIIA (13)	3	$3s^2 3p^1$	3
IVA (14)	4	$3s^2 3p^2$	4
VA (15)	5	$3s^2 3p^3$	5
VIA (16)	6	$3s^2 3p^4$	6
VIIA (17)	7	$3s^2 3p^5$	7
VIIIA (18)	8	$3s^2 3p^6$	8

D. Condensed (Noble Gas, Abbreviated) Electron Configurations

- 1) use the previous Noble Gas as the starting point in brackets, then finish the configuration
- 2) Example 7: Indium #49
 - complete: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 \underline{5s^2} 4d^{10} \underline{5p^1}$
 - condensed: $[\text{Kr}] \underline{5s^2} 4d^{10} \underline{5p^1}$
 - valence $5s^2 5p^1$
- 3) Example 8: Titanium #22
 - complete: $1s^2 2s^2 2p^6 3s^2 3p^6 \underline{4s^2} 3d^2$
 - condensed: $[\text{Ar}] \underline{4s^2} 3d^2$
 - valence $4s^2$

E. *Exceptions* to electron configurations

- 1) *to fill a sublevel = stable*
- 2) *to half-fill a sublevel = stable*
- 3) new unpredicted configurations are observed
 - Cu, expected: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$
 - Cu, observed: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$

Electron Configurations for Ions – covered in chapter 7