

# Chapter 4

## Outline

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### Main Ideas

Q: What prevents the negative electrons from being drawn into the positive nucleus?

A: We need a better atomic model, one based on the absorption & emission of light by matter.

pp 97-103

### I. New Atomic Model

#### A. Properties of Light

##### 1. Electromagnetic radiation

a. Exhibits wavelike behavior as it travels thru space

1. Repetitive nature characterized by *wavelength* ( $\lambda$ ) & *frequency* ( $\nu$ ) in Hz (1/sec)

2.  $c = \lambda \nu$ , where  $c$  = speed of light

b. Travels at a constant speed of  $3.00 \times 10^8$  meters/sec thru a vacuum

c. Spectrum, from highest  $\nu$  to lowest:  $\gamma$  rays, X rays, UV, visible, IR, microwave & radio

#### B. Photoelectric Effect

##### 1. Emission of electrons when light shines on metal

a. Noted by Max Planck (1900) that when metals are heated, light is emitted.

b. Planck proposed energy is emitted in *quanta*, or small, specific packets of energy

c. Planck concluded that  $E = h \nu$ , where  $E$  = energy &  $h$  = Planck's constant =  $6.626 \times 10^{-34}$  Js

##### 2. Requires a minimum frequency *regardless* of the light's intensity

a. This could not be explained by simple wave dynamics.

##### 3. Albert Einstein proposed light had a dual nature: wave & particle (dubbed a "*photon*")

a.  $E_{\text{photon}} = h \nu$

b. The minimum energy needed to eject electrons corresponded to a minimum frequency.

c. Because different metals bind their electrons to varying degrees, each has own minimum  $\nu$ .

#### C. Hydrogen Atom Emission-line Spectrum

1. When current is passed thru vacuum tube of H gas at low pressure, the tube has a pinkish glow.

2. When a narrow beam of this was shined thru a prism, it separated into specific wavelengths.

a. Balmer series: 4 specific colors of the visible spectrum

b. Lyman series: 5 specific  $\lambda$  in the UV spectrum

c. Paschen series: 3 specific  $\lambda$  in the IR spectrum

3. Classical theories, predicting a *continuous spectrum*, failed to explain these results.

4. Neils Bohr (1913) developed atomic model linking emission spectrum to electrons in *orbits*.

a. *Orbits* were discrete paths of a definite, fixed energy.

b. *Ground state* = lowest energy = electron closest to the nucleus

c. Energy levels were fixed, like rungs on a ladder.

d. When energy is added to raise an electron to higher level, *absorption* occurs.

e. When electron falls from excited state back to lower level, a photon is emitted = *emission*.

f. Energy of *absorption* = *emission* =  $E_{\text{photon}} = h \nu$

5. All of this is great; however, it did **not** apply to other atoms nor explain their chemical behavior.

Q: If light had a dual wave/particle nature, did electrons themselves (and anything else) have that too?

A: We need even a BETTER model, one dealing with quanta of energy.

pp 104-110

### II. Quantum Model of the Atom

#### A. Electrons as Waves

1. Louis de Broglie (1924) proposed electrons have wavelike properties.

# Chapter 4

## Outline

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- a. Electron beams can be *diffracted*, or bent, passing by the edge of X or thru a small opening.
  - b. Electron beams can interfere with each other; *interference* occurs when waves overlap.
- B. Heisenberg Uncertainty Principle
1. Werner Heisenberg (1927) realized photons of light detects the presence of electrons, but the energy of photons & electrons are so close that the photon itself knocks the electron away.
  2. He concluded it was impossible to know *simultaneously* both position & energy of electrons.
- C. Schrödinger Wave Equation
1. Erwin Schrödinger (1926) developed an equation treating electrons as waves.
  2. Whereas Bohr assumed quantization, Schrödinger's equations reconciled data with theory.
  3. He & Heisenberg laid the foundation for the modern *Quantum theory*.
- D. Atomic Orbitals & Quantum Numbers
1. The 1<sup>st</sup> 3 Quantum numbers result directly from solutions to the Schrödinger equation.
    - a. Principle quantum # =  $n$  = main energy level occupied by the electron (1, 2, 3, 4...)
      - The total # of orbitals =  $n^2$
    - b. Angular momentum quantum # =  $l$  = shape of orbital = 0,  $n-1$  [0 =  $s$ , 1 =  $p$ ; 2 =  $d$ ; 3 =  $f$ ]
    - c. Magnetic quantum # =  $m$  = orientation of orbital = 0,  $\pm l$  [ $s$  has 1,  $p$  has 3,  $d$  has 5,  $f$  has 7]
  2. The 4<sup>th</sup> Quantum number describes its spin state, either clockwise or counterclockwise.
    - a. Spin quantum # =  $s$  = either  $+\frac{1}{2}$  or  $-\frac{1}{2}$
    - b. A single orbital holds a max of 2 electrons, but they must have opposite spin states.

Q: How do we know where the electrons exist in any given atom?

A: We follow a few simple rules, with noted quantum # relationships above, & figure it out!

pp 111-122

### III. Electron Configurations

#### A. Rules

1. Aufbau principle
  - a. Electrons occupy the lowest-energy orbital possible.
  - b. Orbitals, in order: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d
2. Pauli exclusion principle
  - a. No 2 electrons in the same atom can have the exact same set of 4 quantum numbers.
  - b. So, you can only have 2 electrons per orbital/shape/orientation & these are opposite spins.
3. Hund's rule
  - a. Orbitals of equal energy keep unpaired electrons as long as possible, all with parallel spins.
  - b. This means, for the  $p$  orbital,  $p_x$ ,  $p_y$ , &  $p_z$ , that each gets 1 electron before any gets 2.
  - c. Same applies for  $d$  orbitals,  $d_{x^2-y^2}$ ,  $d_{xy}$ ,  $d_{yz}$ ,  $d_{xz}$ , &  $d_{z^2}$ .
  - d. This is also true for the  $f$  orbitals, of course.

#### B. Representing Electron Configuration (3 Methods)

##### 1. Orbital Notation

- a. Unoccupied orbitals are represented by a line, \_\_\_\_, with orbitals name written under it. \_\_\_\_
- b. An orbital containing 1 electron is shown as \_\_\_\_, & 2 electrons as \_\_\_\_\_. H= \_\_\_\_; He= \_\_\_\_

##### 2. Electron Configuration Notation

- a. The # of electrons in a sublevel is shown as a superscript to the sublevel designation.
- b. For example, H =  $1s^1$ ; He =  $1s^2$ ; Na =  $1s^2, 2s^2, 2p^6, 3s^1$ , etc...
  - Elements of the 2<sup>nd</sup> Period

# Chapter 4

## Outline

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1. The *highest-occupied energy level* =  $e^-$  containing main  $\mathcal{E}$  level with the highest  $n$  value
  2. The *inner-shell electrons* = those that are not in the highest-occupied energy level
  3. B-Ne:  $1s^2, 2s^2, 2p_x^1.. 2p_x^1 2p_y^1.. 2p_x^1 2p_y^1 2p_z^1.. 2p_x^2 2p_y^1 2p_z^1.. 2p_x^2 2p_y^2 2p_z^1.. 2p_x^2 2p_y^2 2p_z^2$  or  $2p^6$
  4. Group 18 (Group VIII) elements have 8 electrons in outermost shell, i.e., a full *octet* <not He>
3. Noble-Gas Notation
- a. After Ne's full octet, all the electrons of the 3<sup>rd</sup> Period (Na-Ar) have their 1<sup>st</sup> 10  $e^-$  the same.
  - b. Rather than rewrite Ne's electron configuration each time, it is abbreviated by [Ne]...
    - [Ne] =  $1s^2, 2s^2, 2p^6$
    - Na = [Ne]3s<sup>1</sup>; Mg = [Ne]3s<sup>2</sup>; Al = [Ne]3s<sup>2</sup>3p<sup>1</sup>, ... Ar = [Ne] 3s<sup>2</sup>3p<sup>6</sup>
  - c. Elements of the 4<sup>th</sup> Period begin with 4s then go to 3d (transition elements) & then to 4p
    - Exceptions!!!! Cr = [Ar]4s<sup>1</sup>3d<sup>5</sup> and Cu = [Ar]4s<sup>1</sup>3d<sup>10</sup> ... *and we don't know why!*
  - d. Elements of the 6<sup>th</sup> period: La adds an  $e^-$  to 5d but with Ce-Lu, electrons add to the f orbital.
    - Because 5d & 4f are so close in energy, deviations from "rules" are numerous.

### IV. Blocks

- A. The periodic table, subdivided into 'blocks', aids in determining configurations. (See below.)
- B. You never need to memorize an element's configuration, except Cr & Cu, just follow the table!  
(Really, it's that easy.)

