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**Ch 5 - Electrons in the Atom**

[**The Bohr Model**](file:///H:\Kgray%20folder\Chemistry\chapter%204,5-atomic%20structure,%20electrons%20in%20atoms\Ch%205.1%20Bohr.ppt)**: (1913)**

* Electrons move only in specific orbits (Energy levels) around the nucleus
* **Quantum**- unit of energy
* **Quantum Jump**- energy needed to move electrons from 1 energy level to

the next

**Development of the Modern Atomic Model:**

* Grew from the study of light

**Louis de Broglie (1923):**

* Electrons – both wave and particle like properties ([Wave/particle duality)](https://www.youtube.com/watch?v=M4_0obIwQ_U)

[**Dr. Quantum video**](https://www.youtube.com/watch?v=DfPeprQ7oGc)

[**Properties of Light**](file:///H:\Kgray%20folder\Chemistry\chapter%204,5-atomic%20structure,%20electrons%20in%20atoms\Ch%205.3%20Light%20and%20Atomic%20Spectrum.flipchart)**:**

* Consists of waves
* **Photon**– quantum of light
* **Wavelength (λ)**- distance b/w crests (nm or m) (109 nm = 1m or 10-9 m = 1nm)
* **Frequency (ν)** - # of **λ** to pass a given point per a unit of time (Hz or s-1)
* **λ** ↑, **ν** ↓

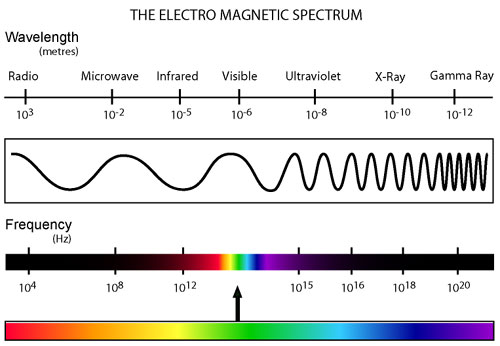
c =**λν** (c= speed of light 2.998 x 108 m/s)

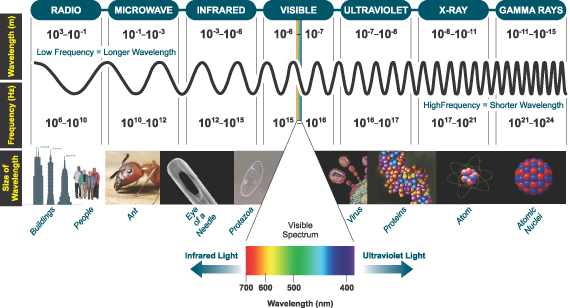
<http://www.astronomynotes.com/light/s3.htm>

* **Electromagnetic Radiation**- form of energy (photon) that exhibits wavelike behavior

(Radio, microwaves, infrared, [**visible light**](https://www.youtube.com/watch?v=Gf33ueRXMzQ), ultraviolet, x-rays, gamma rays)

**Electromagnetic Spectrum**



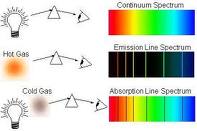


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[**Atomic Spectra**](https://www.youtube.com/watch?v=Gf33ueRXMzQ)**:**

* Emission Spectrum – frequencies of light separated into individual lines

(colors) (ex. Red –lowest **ν**, longest **λ**; violet – highest **ν**, shortest **λ**)

[](http://www.google.com/imgres?q=emission+spectrum&hl=en&sa=X&rlz=1T4ADSA_enUS430US431&biw=1024&bih=556&tbm=isch&prmd=ivns&tbnid=ZKapP2AB8BBciM:&imgrefurl=http://www.astro.cornell.edu/academics/courses/astro101/herter/lectures/lec09.htm&docid=hfJAFtcB-H7rmM&w=441&h=294&ei=GwZuTo7TKMrJgQe32ujiBQ&zoom=1&iact=rc&dur=156&page=7&tbnh=105&tbnw=157&start=90&ndsp=15&ved=1t:429,r:1,s:90&tx=91&ty=74)

**Explanation of Atomic Spectra:**

* Electrons absorbs energy = “quantum jump” to higher energy level – **excited state**
* Electrons lose energy = electrons transition back to lower energy level- **ground state** and emits light (photon)
* Each transition = specific frequency in the spectrum
* Amount of energy determines frequency (and therefore color emitted)
* Frequency directly proportional to the energy change

<http://astro.u-strasbg.fr/~koppen/discharge/nitrogen.html>

**Hydrogen spectrum** = produced 3 groups of lines based on electrons transitions:

**Lyman series** – down to the 1st energy level (UV range)

**Balmer series** – down to the 2nd energy level; (Visible range)

**Paschen series**- down to the 3rd energy level; (IR range)

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**Quantum Mechanics:**

**Werner Heisenberg (1925):**

* Matrix Mechanics
* Electrons- quantum jumps

**Erwin Schrodinger (1926):**

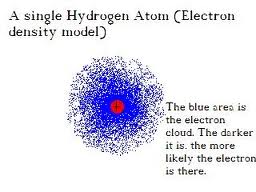
* Schrodinger Wave Equation
* Electrons- continuous waves of energy

**Heisenberg Uncertainty Principle (192**7**)**:

* Can’t tell the position and velocity (speed) of an electron at same time
* Led to Quantum Mechanical Model

[**Quantum Mechanical Model (Electron Cloud Model):**](file:///Z:\Chemistry\chapter%204,5-atomic%20structure,%20electrons%20in%20atoms\Ch%205.1%20The%20Quantum%20Mechanical%20Model.pptx)

* No exact path for electrons – Probability regions instead
* Mathematical expression

[Quantum leap lab results](file:///\\swain.local\storage-ns\homes\high\staff\kgray\Kgray%20folder\Chemistry%20II\quantum%20leap%20lab%20data.flipchart)

[**Atomic Orbital:**](file:///H:\Kgray%20folder\Chemistry\chapter%204,5-atomic%20structure,%20electrons%20in%20atoms\orbitals%20shape%20chart.flipchart)

* Principle Energy levels (n)= 1, 2, 3, 4, 5, 6, 7
* "Probability regions” –fuzzy clouds – regions where electrons can be found – most dense where probability is high (closest to the nucleus)
* Divided into sublevels and orbitals (max. of 2 electrons per)
  + - s– spherical shaped (1 orbital)
    - p - dumbbell shaped (3 orbitals)
    - d –varies (5 orbitals)
    - f - varies (7 orbitals)

[**Table 1: 1st 4 Principle Energy levels**](file:///\\swain.local\storage-ns\homes\high\staff\kgray\Kgray%20folder\Chemistry%20II\Table-%201st%204%20energy%20levels.flipchart)**:**

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Principle energy levels**  **(n)** | **Sublevels per each principle energy level** | **Number of orbitals per sublevel** | **Number of orbitals per principle energy level (n2)** | **Number of electrons per sublevel** | **Number of electrons per principle energy level (2n2)** |
| 1 | s | 1 |  |  |  |
| 2 | s  p | 1  3 |  |  |  |
|  |
| 3 | s  p  d | 1  3  5 |  |  |  |
|  |
|  |
| 4 | s  p  d  f | 1  3  5  7 |  |  |  |
|  |
|  |
|  |

**Bohr vs. Quantum Mechanical Model:**

* Both restricts the energy of electrons to certain values
* Bohr – electrons take an exact pathway around the nucleus
* Quantum – electrons do NOT have an exact pathway around the nucleus rather probability regions

**5.2 Electron Arrangement in Atoms:**

[**Electron Configurations**](file:///\\swain.local\storage-ns\homes\high\staff\kgray\Kgray%20folder\Chemistry%20II\Ch%205.2%20Electron%20Configurations.ppt)**:** [**electron energy level song**](http://www.youtube.com/watch?v=Vb6kAxwSWgU&list=PL65159266CFC74682&index=14&feature=plpp_video)

* Ways in which electrons are arranged in energy levels around the nucleus

**Aufbau Principle:**

* Electrons occupy the lowest energy level first (1 is lowest).
* s sub-level always the lowest within an energy level (n)

**Pauli Exclusion Principle:**

* Each orbital can have max. of 2 electrons
* Opposite spins in same orbital
* Vertical arrow indicates direction of spin (↑ or ↓)
* Orbital containing paired electrons written as ↑↓

**Hund’s Rule:**

* Electrons within the same sublevel (ex. 3p)- 1 electron enters each orbital until all orbitals contain electrons with same spin, then 2nd electrons can be added if necessary

Ex. Orbital Notation for C

2s ↑↓ 2p ↑\_ ↑\_ \_\_

1s ↑↓

[orbital notation filling order chart](H:\\Kgray folder\\Chemistry\\chapter 4,5-atomic structure, electrons in atoms\\orbital notation filling order.flipchart)

[**Writing Electron Configurations**](file:///\\swain.local\storage-ns\homes\high\staff\kgray\Kgray%20folder\Chemistry%20II\Periodic%20tables-%20electron%20configurations.flipchart)**:** [**electron energy level song**](http://www.youtube.com/watch?v=Vb6kAxwSWgU&list=PL65159266CFC74682&index=14&feature=plpp_video)

* **#** = principle energy level
* **Letter** = sublevel
* **Superscript** = # of electrons within that sublevel
* **Sum of superscript** = atomic # for that element

**Long Form:**

* Write entire configuration

Ex. Na = 1s22s22p63s1

**Short Form (abbreviated form):**

* Use Noble gas configurations (far right column)
* Put the Noble gas element symbol closest to the element you are

writing (the preceding row’s noble gas) in brackets and continue with

the rest of the configuration

Ex. Na = [Ne]3s1

Neon and sodium’s long form: Ne = 1s22s22p6

Na = 1s22s22p63s1

[**Periodic Table Rule**](file:///Z:\Chemistry\chapter%204,5-atomic%20structure,%20electrons%20in%20atoms\Electron%20config%20periodic%20table.doc)**:**

## Periods (rows)= energy levels

**Exceptions:** Period 4 - columns 3-12 = 3 energy level (3d)

Period 5 - column 3-12 = 4 energy level (4d)

Period 6 - column 3-12 = 5 energy level (5d)

Period 7 - column 3-12 = 6 energy level (6d)

## \* (bottom 2 rows of elements = 4f, 5f)

## Groups (Columns)= sublevels

**s** sublevel = column 1-2 + He

**p** sublevel = column 13-18

**d** sublevel = column 3-12

**f** sublevel = bottoms rows

\*\*Read across the periodic table to determine the configurations. Count the

elements per energy level (row) and sublevel (column) -that # will equal the

superscript ( # of electrons) for that energy and sublevel.

Ex. Oxygen = 1s22s22p4

1 s2 = energy level 1 (row 1), sublevel s (column 1+2), 2 electrons (2 elements)

2s2 = energy level 2 (row 2), sublevel s (column 1+2) 2 electrons (2 elements)

2p4 = energy level 2 (row 2), sublevel p (column 13-18) 4 electrons (4 elements)

[**Correct filling order**](file:///Z:\Chemistry\chapter%204,5-atomic%20structure,%20electrons%20in%20atoms\diagonal%20rule%20filling%20order%20chart.flipchart):

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

[blank periodic table](file:///\\swain.local\storage-ns\homes\high\staff\kgray\Kgray%20folder\Chemistry%20II\blank%20periodic%20table%20filled%20in.flipchart)

[periodic table](file:///Z:\Chemistry\chapter%204,5-atomic%20structure,%20electrons%20in%20atoms\periodic%20table.flipchart)

[Spdf periodic table](file:///\\swain.local\storage-ns\homes\high\staff\kgray\Kgray%20folder\Chemistry%20II\spdf%20periodic%20table.flipchart)

**Diagonal Rule:**

Correct filling order – after energy level 2, things get tricky!!

1s2

2 s2 2p6

3 s2 3p6 3d10

4 s2 4p6 4d10 4f14

5 s2 5p6 5d10 5f14

6 s2 6p6 6d10 6f14

7 s2 7p6 7d10 7f14

**Correct filling order**:

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