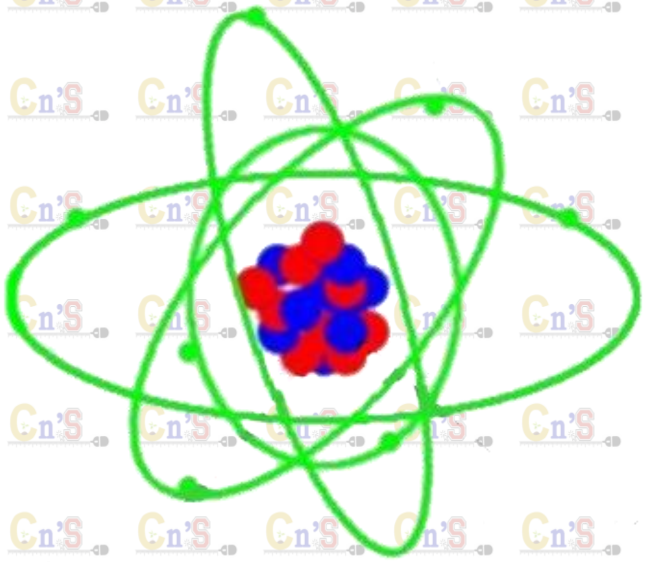
ATOMIC STRUCTURE
Presentation 2



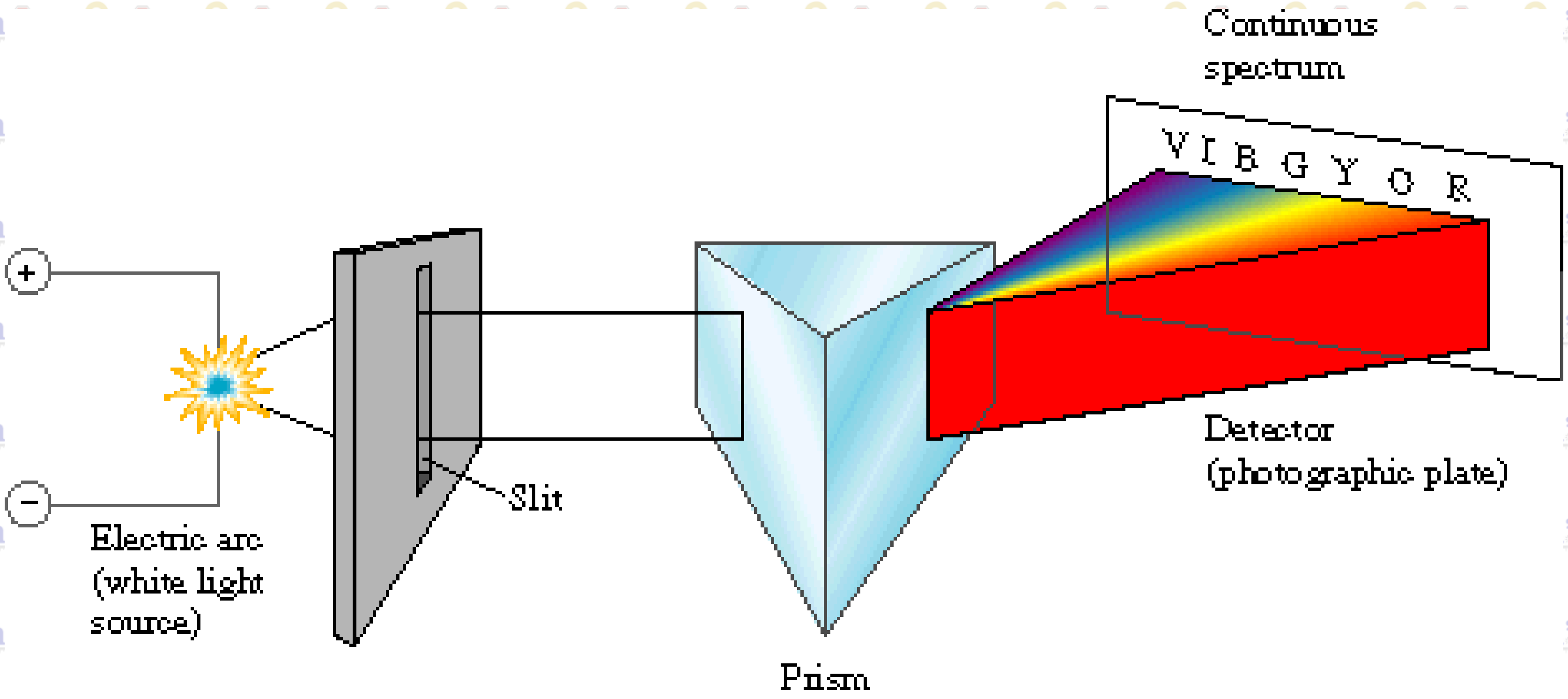
What is spectrum?

The dispersion of the components of white light, when it is passed through prism is called spectrum. The distribution among various wavelengths of the radiant energy emitted or absorbed by an object is also called spectrum.

Differentiate between continuous spectrum and line spectrum.

Continuous spectrum: A spectrum containing light of all wavelengths is called continuous spectrum. In this type of spectrum, the boundary line between the colours cannot be marked. The colours diffuse into each other. One colour merges into another without any dark space. The best example of continuous spectrum is rainbow.

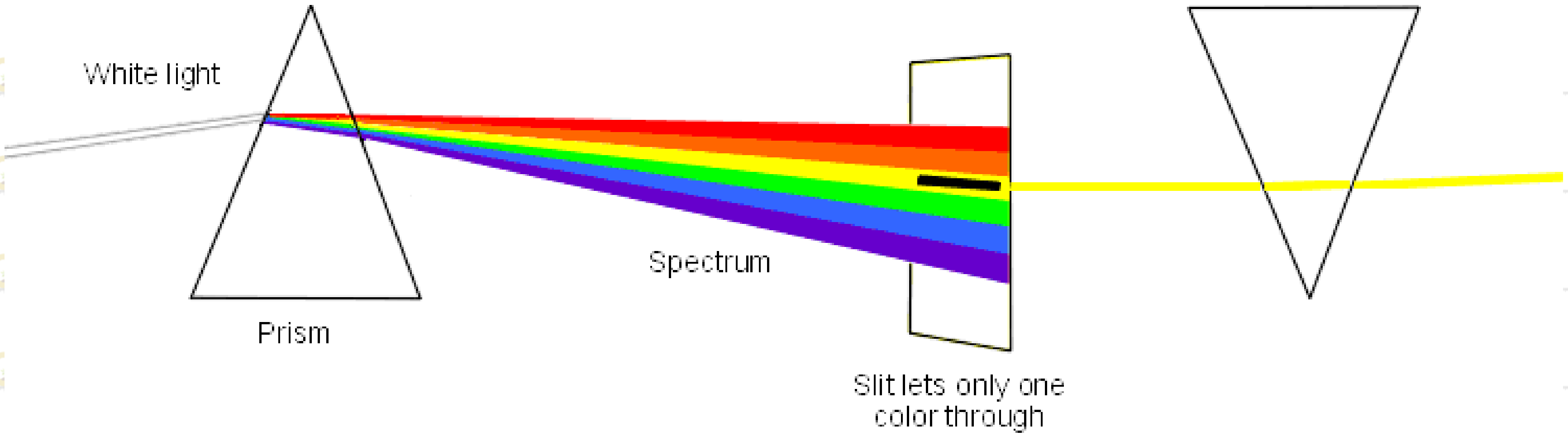
Original Studies Of Light Used Only One Prism



When a narrow band of light from a "white" light source is sent through a prism, a continuous spectrum containing all wavelengths of visible light is formed.

Newton's Contribution to Spectroscopy

Newton contributed more to spectroscopy than scientifically proving that sunlight traveling through a prism was always broken down into the components of the rainbow.

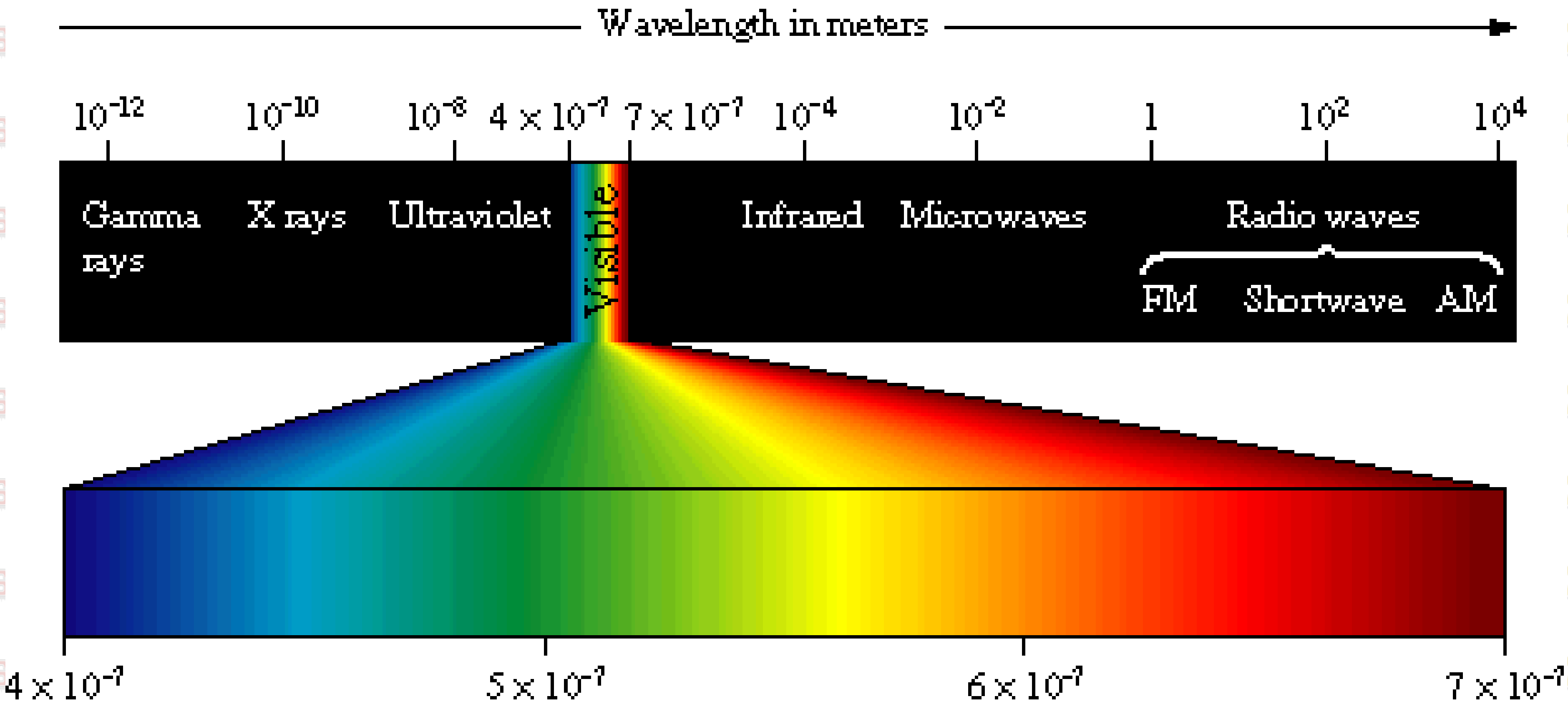


In fact, his main contribution was to show that after the sunlight had been broken down into its components by one prism, if a narrow ray of the light from the first prism was passed through another prism **no further breakdown** of light occurred.

Line spectrum:

When an element or its compound is volatilized on a flame and the light emitted is seen through, a spectrometer. We see distinct lines separated by dark spaces. This type of spectrum is called line spectrum. This is the characteristic of an atom.

Classification of Electromagnetic Radiation



The color components of light are separated along the visible range of light. The **visible range of light (400-700 nm)** is merely a small portion of the entire electromagnetic spectrum.

ELECTROMAGNETIC RADIATION

- subatomic particles (electron, photon, etc) have both PARTICLE and WAVE properties
- Light is electromagnetic radiation - crossed electric and magnetic waves:

Properties :

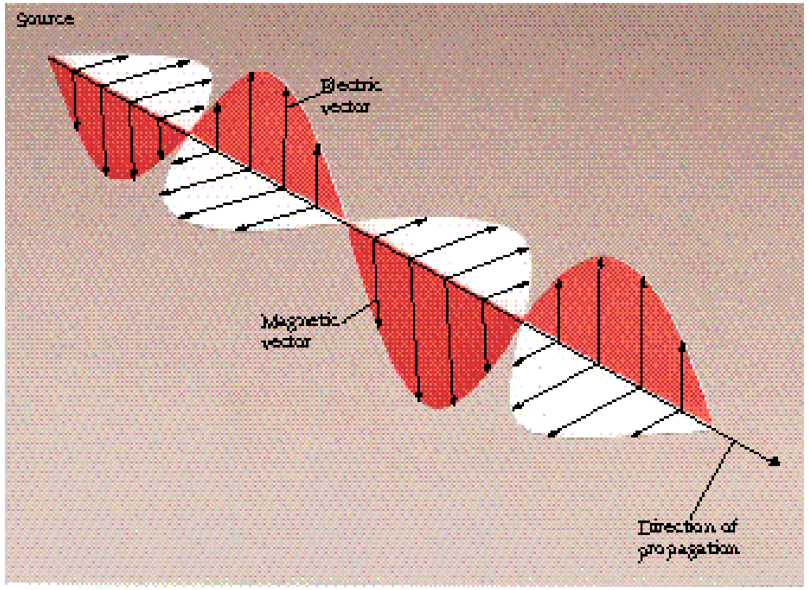
Wavelength, λ (nm)

Frequency, ν (s^{-1}, Hz)

Amplitude, A

constant speed. c

3.00×10^{8} $m.s^{-1}$

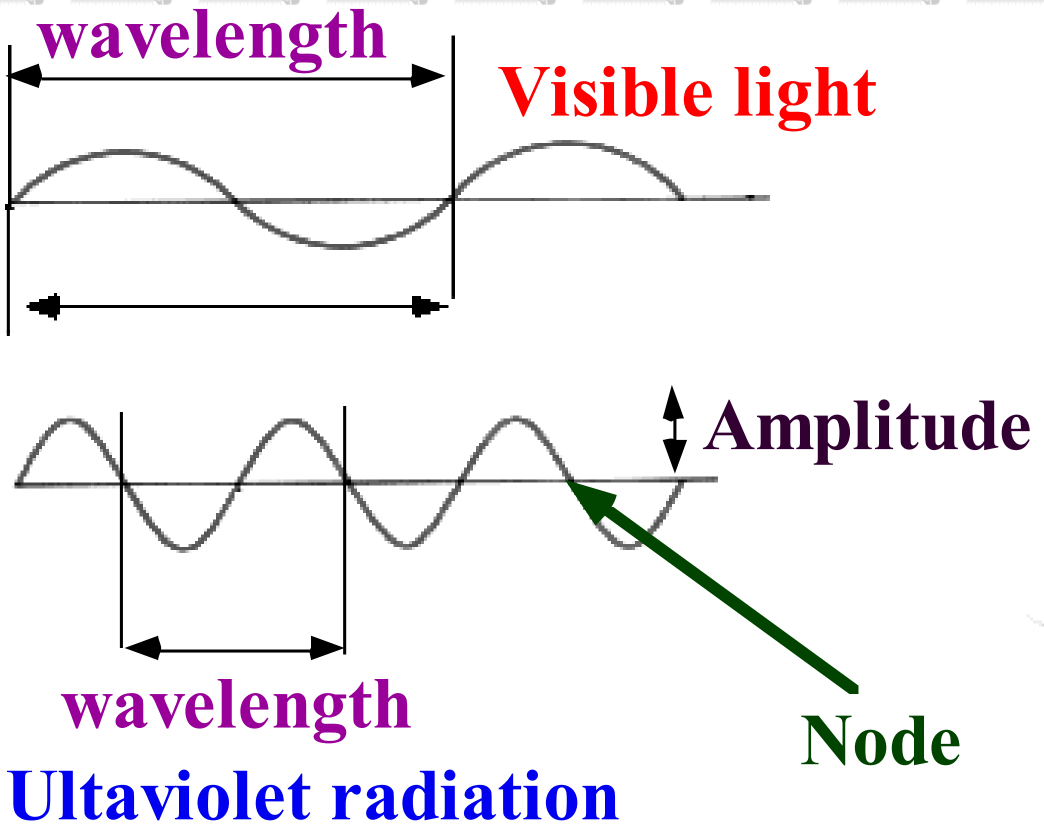


- Elements are radioactive when they emit radiation (radioisotopes)

Types of Radiation

- Particle radiation
  - refers to alpha and beta particles, protons, and neutrons given off by isotopes.
- Electromagnetic radiation
  - i.e. gamma rays and X-rays

Electromagnetic Radiation (2)



Electromagnetic Radiation (3)

- All waves have:

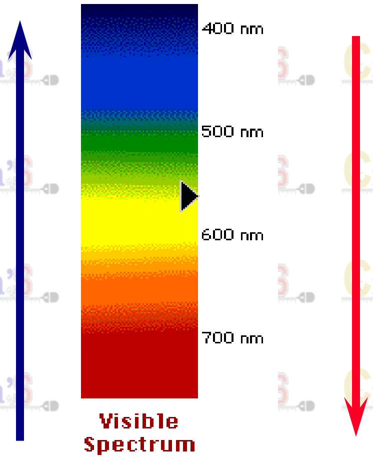
  frequency and **wavelength**

- symbol: ν (Greek letter "nu") λ (Greek "lambda")
- units: "cycles per sec" = Hertz "distance" (nm)

- **All radiation: $\lambda \bullet \nu = c$**

where c = velocity of light = 3.00×10^8 m/sec

Note: Long wavelength
$\rightarrow$ small frequency
Short wavelength
$\rightarrow$ high frequency



increasing frequency

increasing wavelength

Electromagnetic Radiation (4)

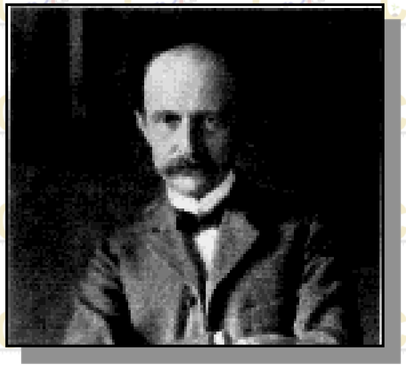
Example: Red light has λ = 700 nm.

Calculate the frequency, ν.

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^{8}\ \text{m/s}}{7.00 \times 10^{-7}\ \text{m}} = 4.29 \times 10^{14}\ \text{Hz}$$

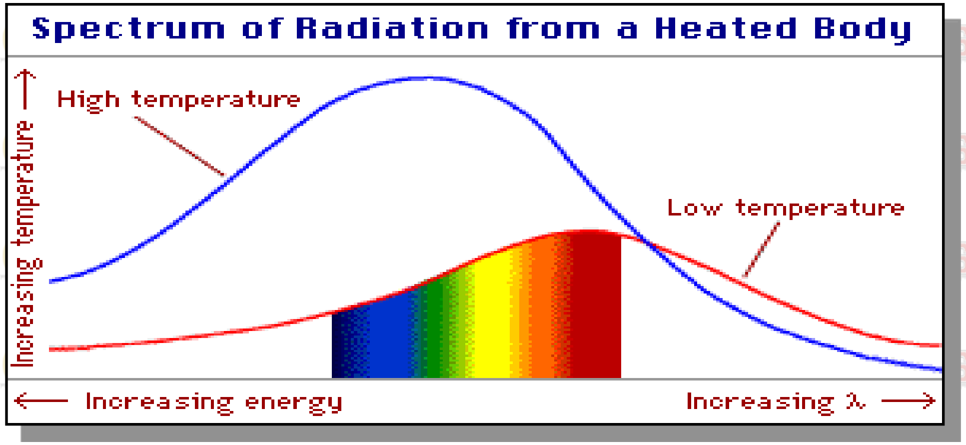
- **Wave nature of light is shown by classical wave properties such as**
  - ***interference***
  - ***diffraction***

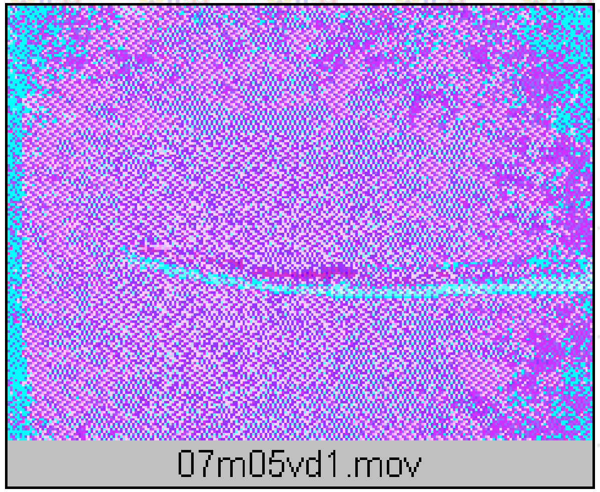
Quantization of Energy



Max Planck (1858-1947)

4-HOT_BAR.MOV





- **Planck's hypothesis: *An object can only gain or lose energy by absorbing or emitting radiant energy in QUANTA.***

Quantization of Energy (2)

Energy of radiation is proportional to frequency.

$$E = h \bullet \nu$$

where h = Planck's constant = 6.6262×10^{-34} J•s

Light with large λ (small ν) has a small E.

Light with a short λ (large ν) has a large E.

Quantum or Wave Mechanics

L. de Broglie
(1892-1987)

- Light has both wave & particle properties
- de Broglie (1924) proposed that all moving objects have wave properties.
- For light: $E = h\nu = hc / \lambda$
- For particles: $E = mc^2$ (Einstein)

Therefore, $\boxed{mc = h / \lambda}$

and for particles

$$\boxed{(\text{mass})\text{x}(\text{velocity}) = h / \lambda}$$

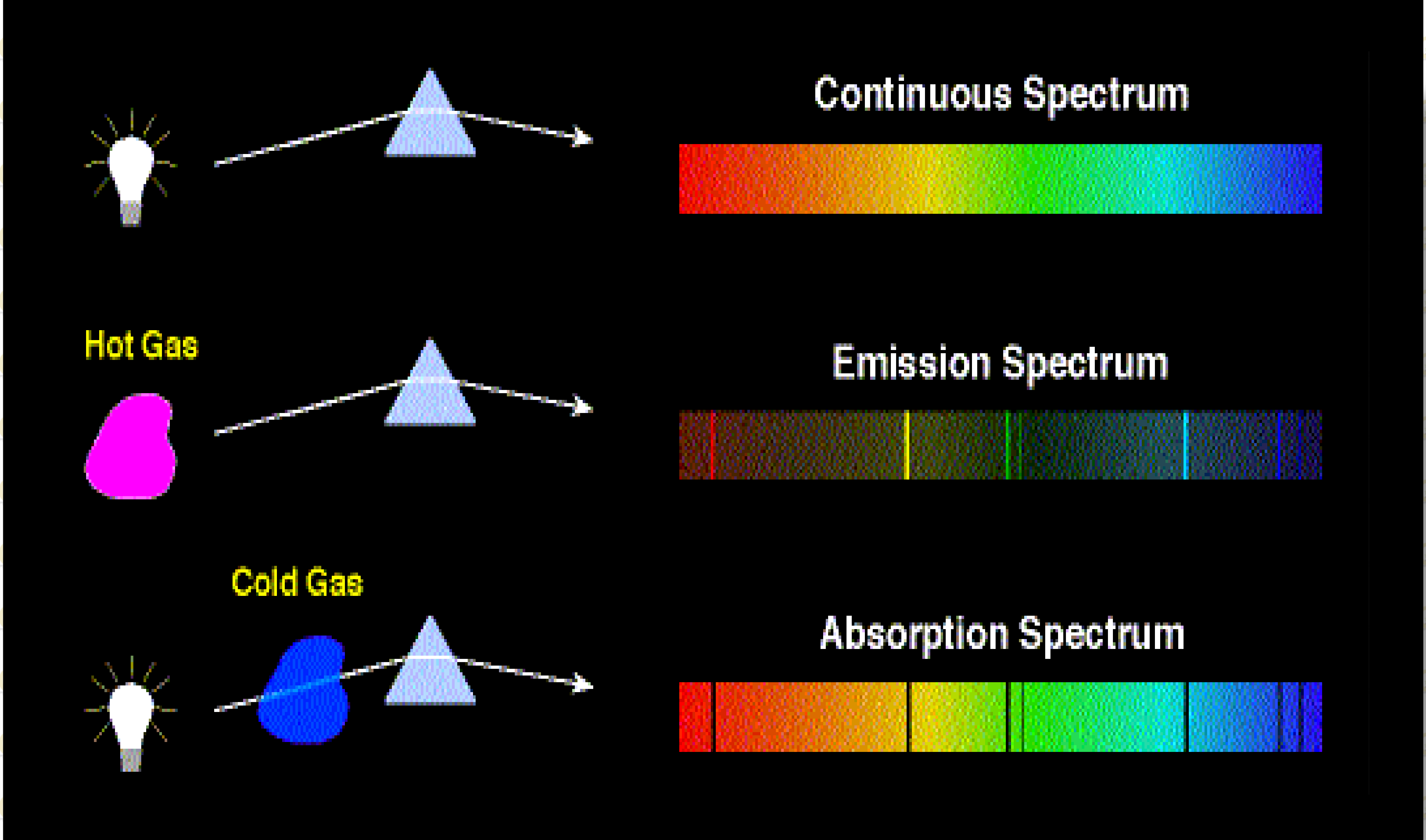
λ for particles is called the de Broglie wavelength

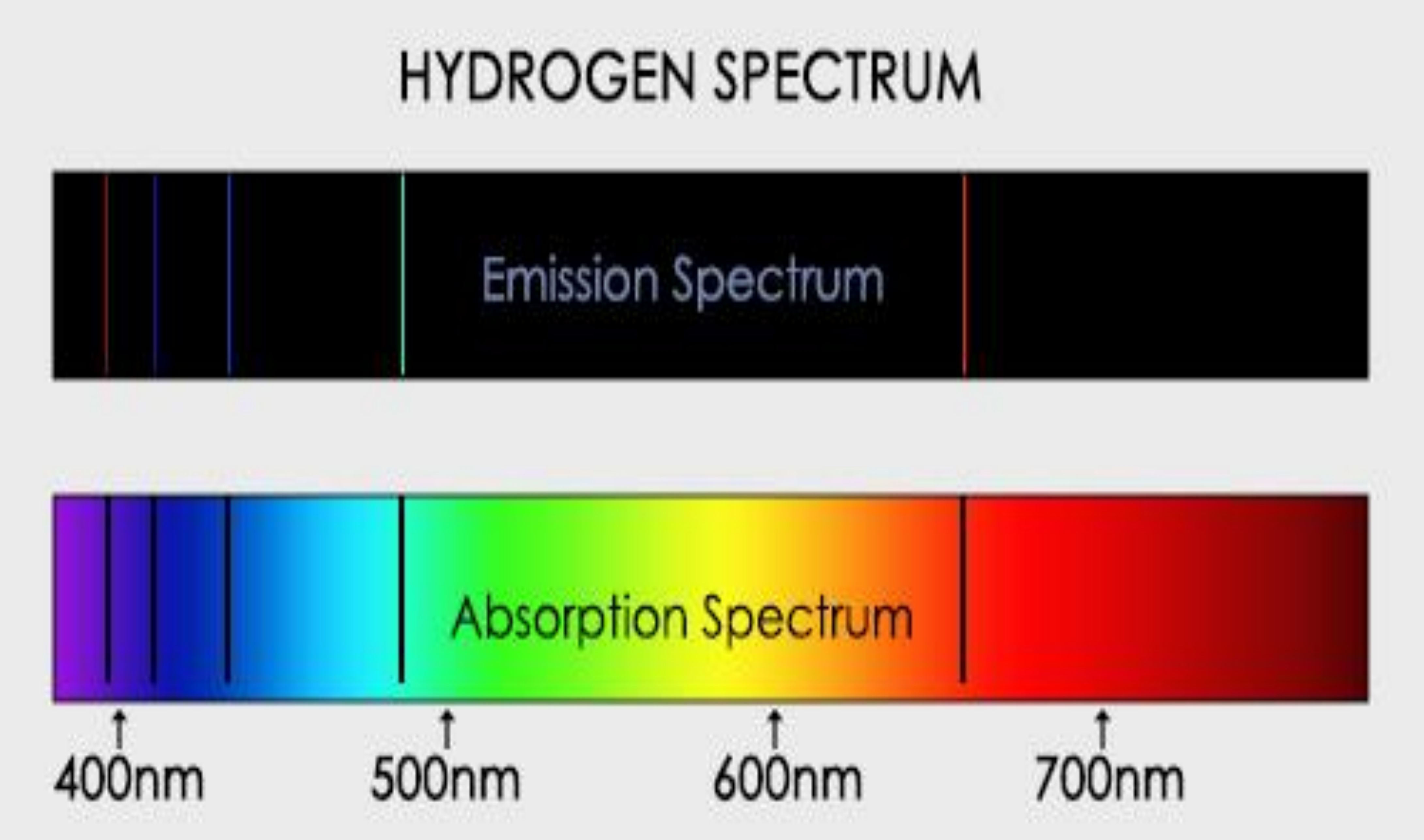
Line absorption spectrum

it is produced when radiations emitted by an excited substance are analysed in a spectroscope. 2. it is produced when white is passed through the gaseous element and the transmitted rays are analysed in a spectroscope.

Line emission spectrum

1. “An atomic spectrum which consists of bright lines against a dark background is called line emission spectrum. “1. “An atomic spectrum which consists of bright lines against a dark background is called line emission spectrum.

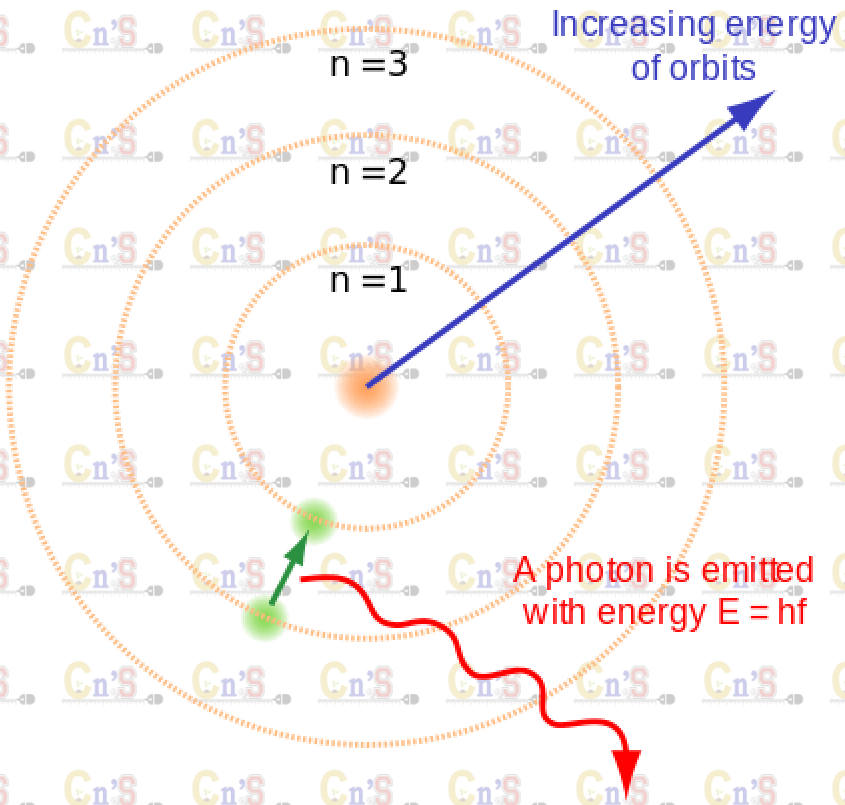

Continuous Spectrum
Hot Gas
Emission Spectrum
Cold Gas
Absorption Spectrum


HYDROGEN SPECTRUM
Emission Spectrum
Absorption Spectrum
400nm
500nm
600nm
700nm

Bohr's Atomic Model

Following the discoveries of hydrogen emission spectra and the photoelectric effect, the Danish physicist Niels Bohr (1885 - 1962) proposed a new model of the atom in 1915. Bohr proposed that electrons do not radiate energy as they orbit the nucleus, but exist in states of constant energy which he called **stationary states**. This means that the electrons orbit at fixed distances from the nucleus (see below). Bohr's work was primarily based on the emission spectra of hydrogen. This is also referred to as the planetary model of the atom. It explained the inner workings of the hydrogen atom. Bohr was awarded the Nobel Prize in physics in 1922 for his work.

Bohr explained that electrons can be moved into different orbits with the addition of energy. When the energy is removed, the electrons return back to their ground state, emitting a corresponding amount of energy - a quantum of light, or photon. This was the basis for what later became known as **quantum theory**. This is a theory based on the principle that matter and energy have the properties of both particles and waves.



According to the Bohr model, often referred to as a **planetary model**, the electrons encircle the nucleus of the atom in specific allowable paths called orbits. When the electron is in one of these orbits, its energy is fixed. The ground state of the hydrogen atom, where its energy is lowest, is when the electron is in the orbit that is closest to the nucleus. The orbits that are further from the nucleus are all of successively greater energy. The electron is not allowed to occupy any of the spaces in between the orbits. An everyday analogy to the Bohr model is the rungs of a ladder. As you move up or down a ladder, you can only occupy specific rungs and cannot be in the spaces in between rungs. Moving up the ladder increases your potential energy,

while moving down the ladder decreases your energy.

Summary

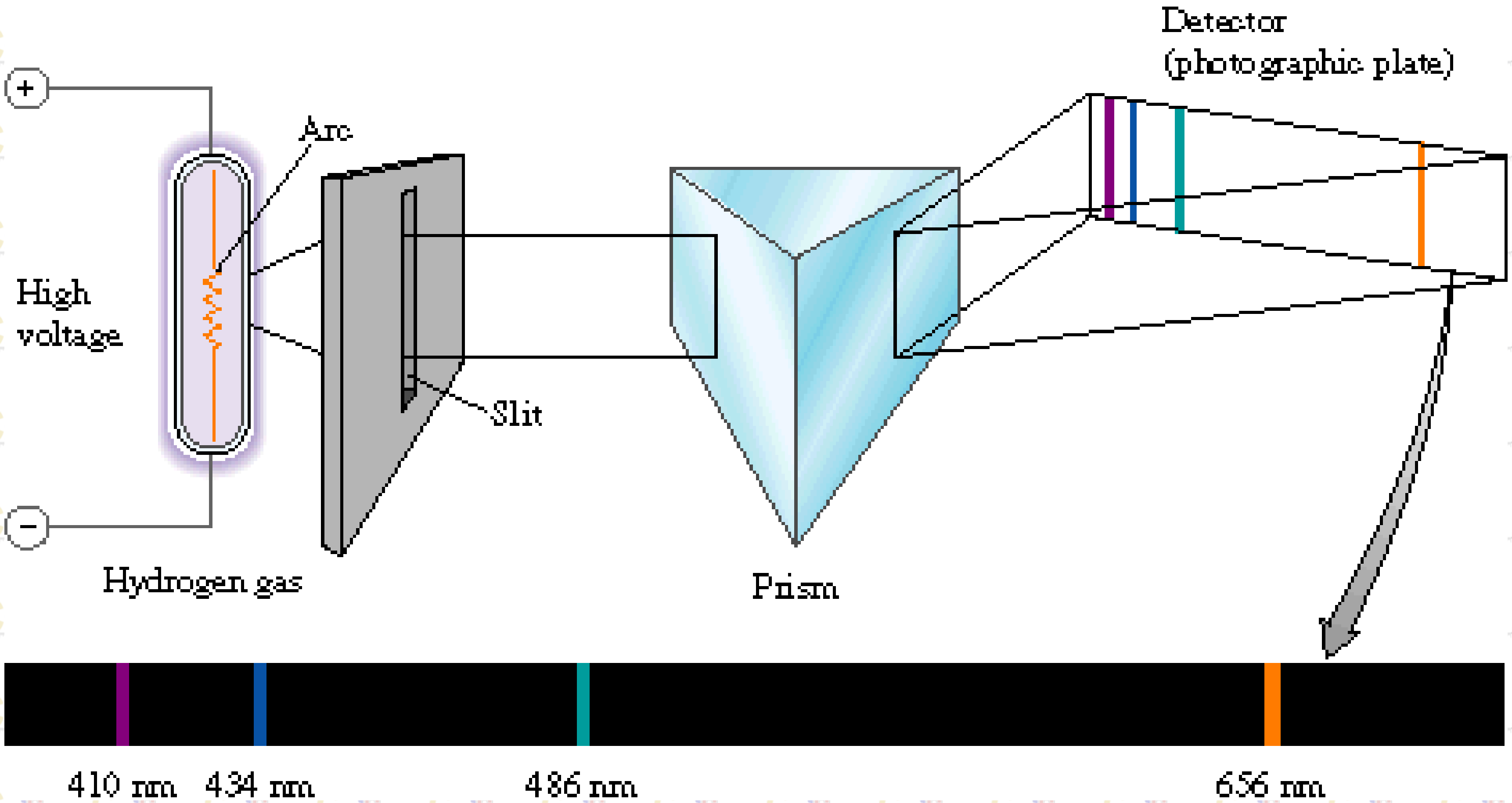
The Bohr model postulates that electrons orbit the nucleus at fixed energy levels. Orbits further from the nucleus exist at higher energy levels. When electrons return to a lower energy level, they emit energy in the form of light

Quantization and photons

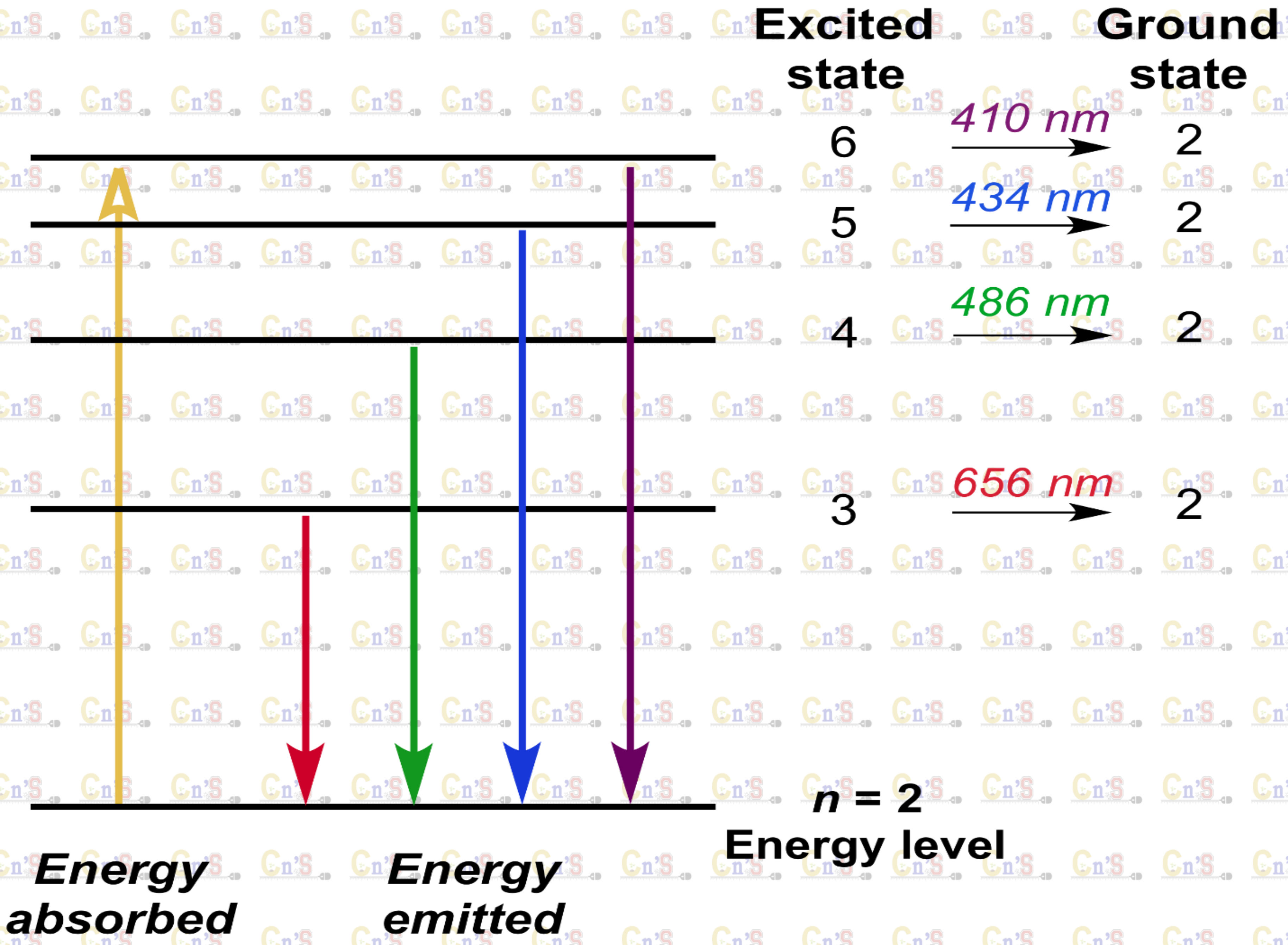
Physicists Max Planck and Albert Einstein theorized that electromagnetic radiation not only behaves like a wave, but also sometimes like particles called *photons*. Planck studied the electromagnetic radiation emitted by heated objects, and he proposed that the emitted electromagnetic radiation was "quantized" since the energy of light could only have values given by the following equation: E{photon} = *nhv* Q**uantized**, means that only specific values are allowed, such as when playing a piano…………….F F*

Atomic line spectra are another example of quantization. When an element or ion is heated by a flame or excited by electric current, the excited atoms emit light of a characteristic color. The emitted light can be refracted by a prism, producing spectra with a distinctive striped appearance due to the emission of certain wavelengths of light.

PART A: Record Hydrogen line spectrum with a Scanning Spectrophotometer.



The hydrogen line spectrum contains only a few discrete wavelengths.
In the visible region, there are **only four** wavelengths.


Excited state
Ground state
410 nm
434 nm
486 nm
656 nm
6
5
4
3
2
2
2
2
n = 2
Energy level
Energy absorbed
Energy emitted

Hydrogen Spectrum – The Balmer Series

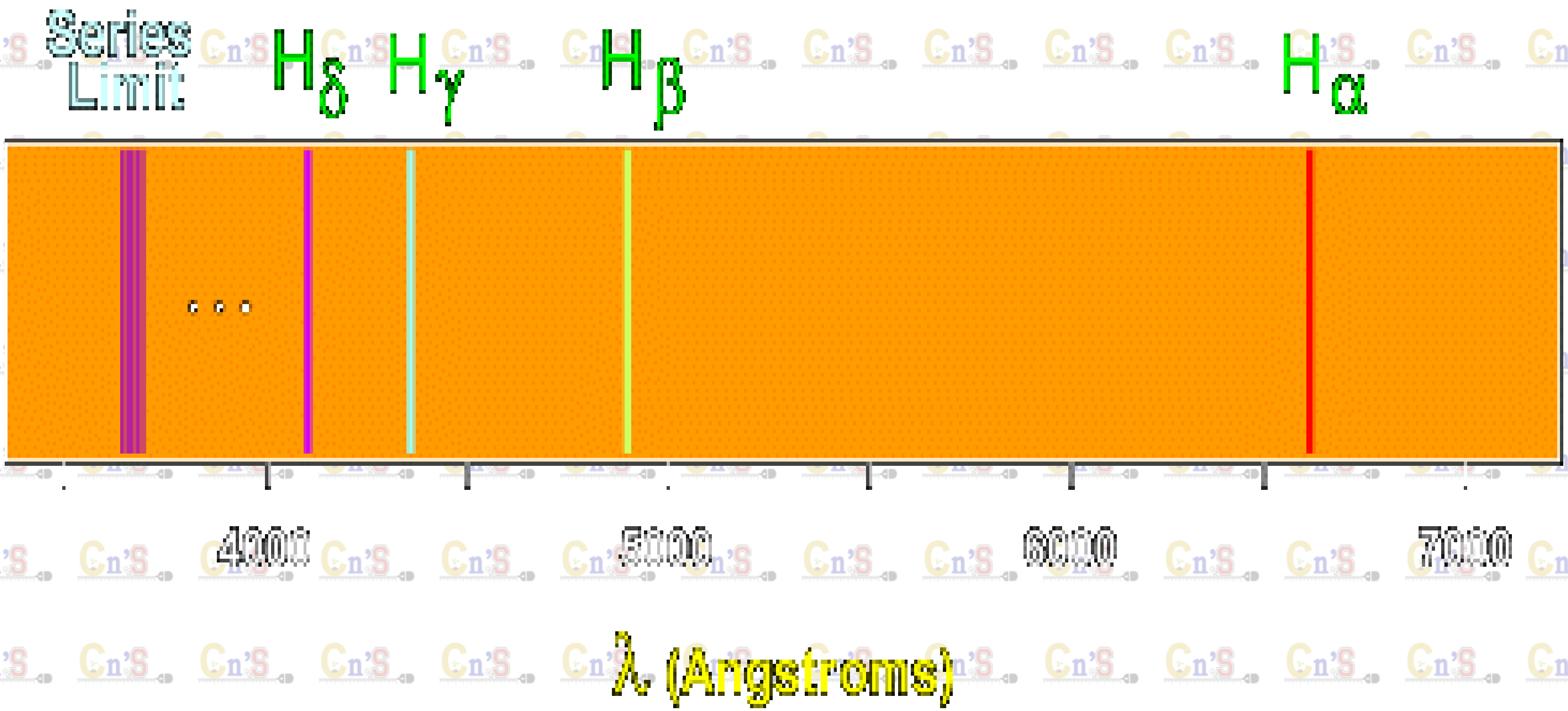
In 1885, Johann Jakob Ballmer

Swiss Mathematician & Honorary Physicist

analyzed the hydrogen spectrum and found that hydrogen emitted four bands of light within the visible spectrum

Wavelength (nm)	Color
656.2	red
486.1	blue
434.0	Indigo (blue-violet)
410.1	violet

The Balmer Series is in the visible spectrum.



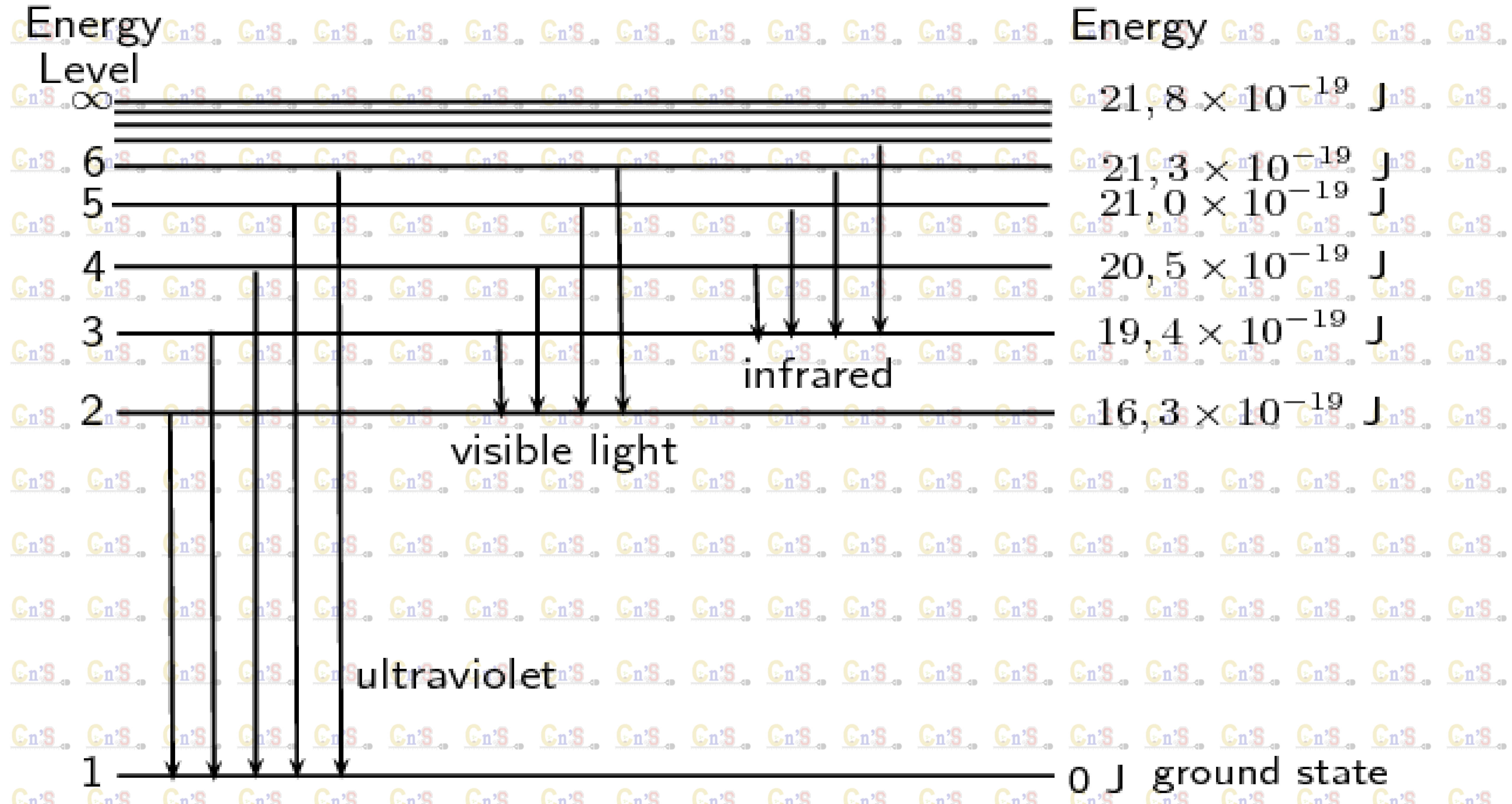
The Lyman Series is in the the UV.

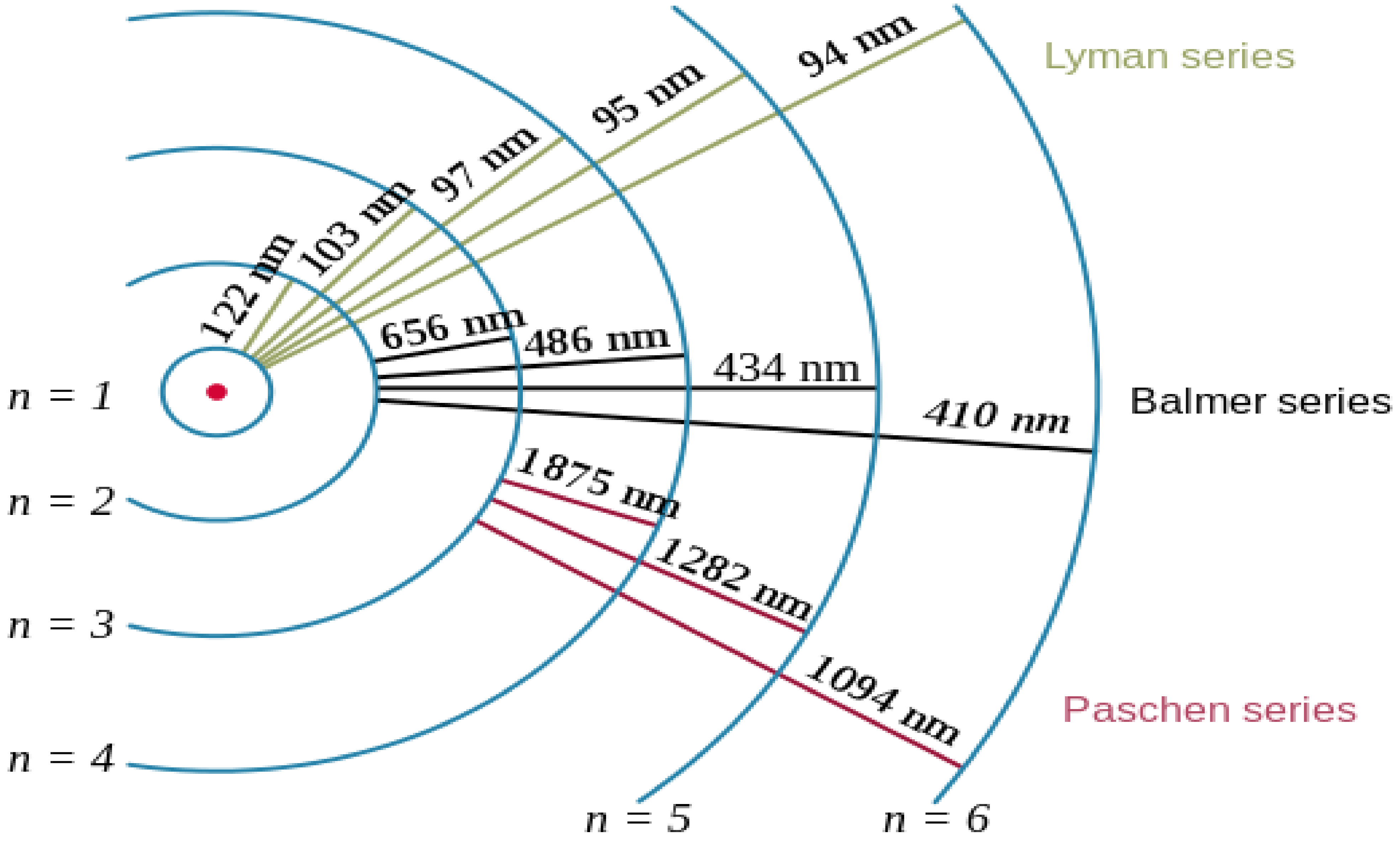
The Barmer series, First 4 lines are in the Visible region & Others are in the the UV.

The Paschen Series are in the the IR.

The Brakket Series are in the the IR.

The Pfund Series are in the the IR.


Energy Level
∞
6
5
4
3
2
1
Energy
$21,8 \times 10^{-19}$ J
$21,3 \times 10^{-19}$ J
$21,0 \times 10^{-19}$ J
$20,5 \times 10^{-19}$ J
$19,4 \times 10^{-19}$ J
$16,3 \times 10^{-19}$ J
0 J ground state
infrared
visible light
ultraviolet

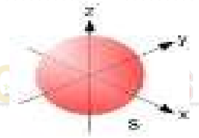
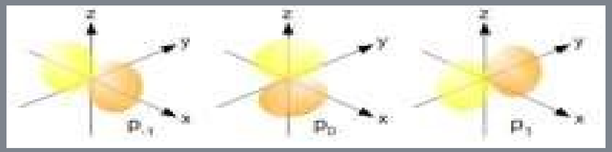

Lyman series
94 nm
95 nm
97 nm
103 nm
122 nm
656 nm
486 nm
434 nm
410 nm
Balmer series
1875 nm
1282 nm
1094 nm
Paschen series
n = 1
n = 2
n = 3
n = 4
n = 5
n = 6

● The **emission spectrum** of a gas is represented by a collection of separate colored lines, with dark spaces between them. The lines are the parts of the spectrum where emission occurs and photons are emitted, while the dark spaces are the parts where there is no emission, hence the darkness. The difference in colors is due to the variation of the energy levels of the electrons.

● The **absorption spectrum** of an element is represented by a continuous band of colors with separate dark lines between them. The entire band represents the total light that is focused on the element. The dark lines are the parts of the spectrum where the electrons absorb light photons, hence, there is absence of light at these parts. The remaining colored parts of the spectrum represent the parts of the incident light that has not been absorbed, and hence, appear as wavelength-specific colors.

Electron Orbitals

- Is a probability function that describes where electrons may be found.
- Are not orbits rather they are electron density clouds which describe the highest probability distribution of electrons.
  - Four types of orbitals: s, p, d, and f

s	Is a sphere, The simplest orbital and can hold up only two electrons	
p	The subshell consists of six three dimensional lobes and can hold up to six electrons.	
d	can hold up to ten electrons. Is made of five subshells which can hold two electrons each.	

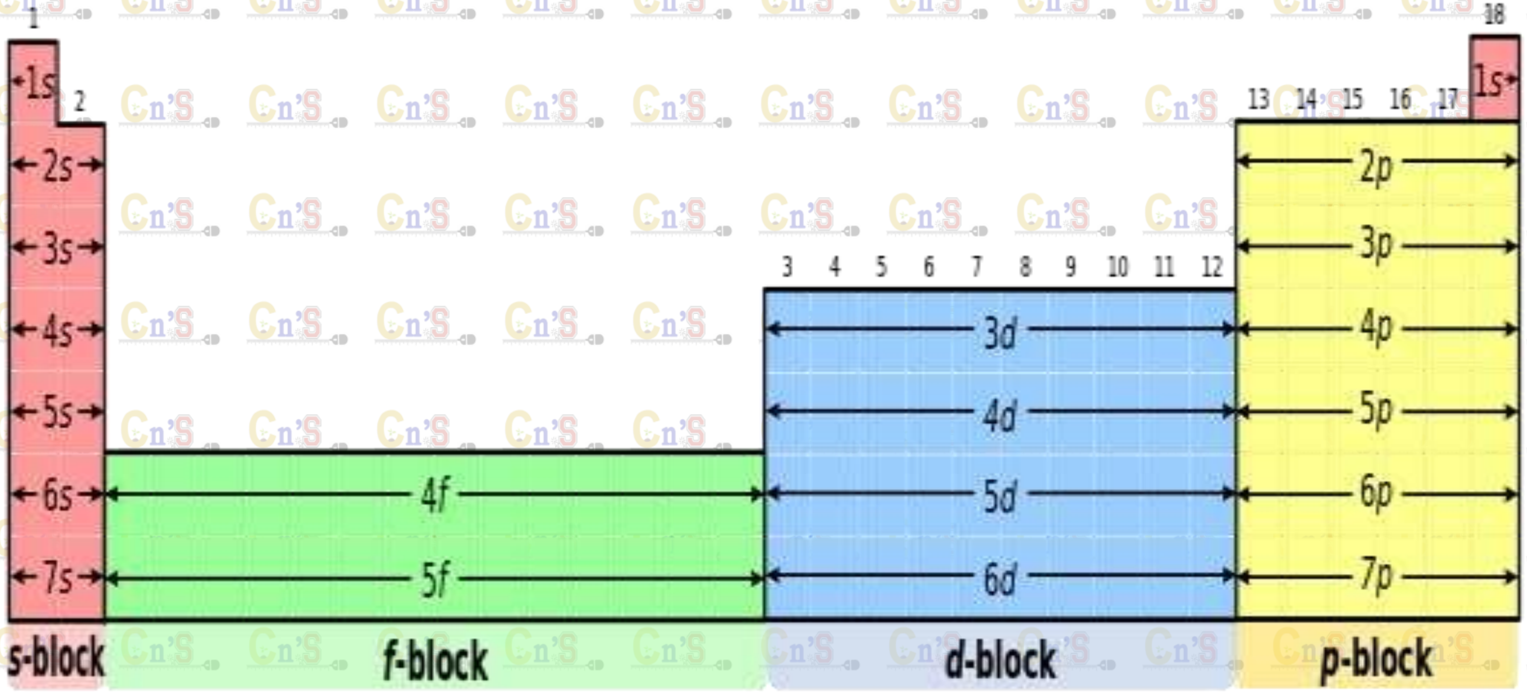
Electron Configuration

- ❖The distribution of electrons among orbitals in an atom.

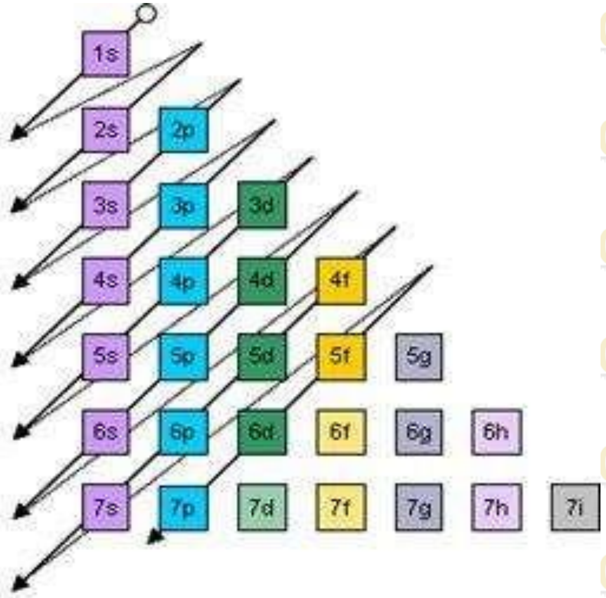
- ❖An atom can be configured without drawing or illustrating any of the s, p, d, and f orbitals.

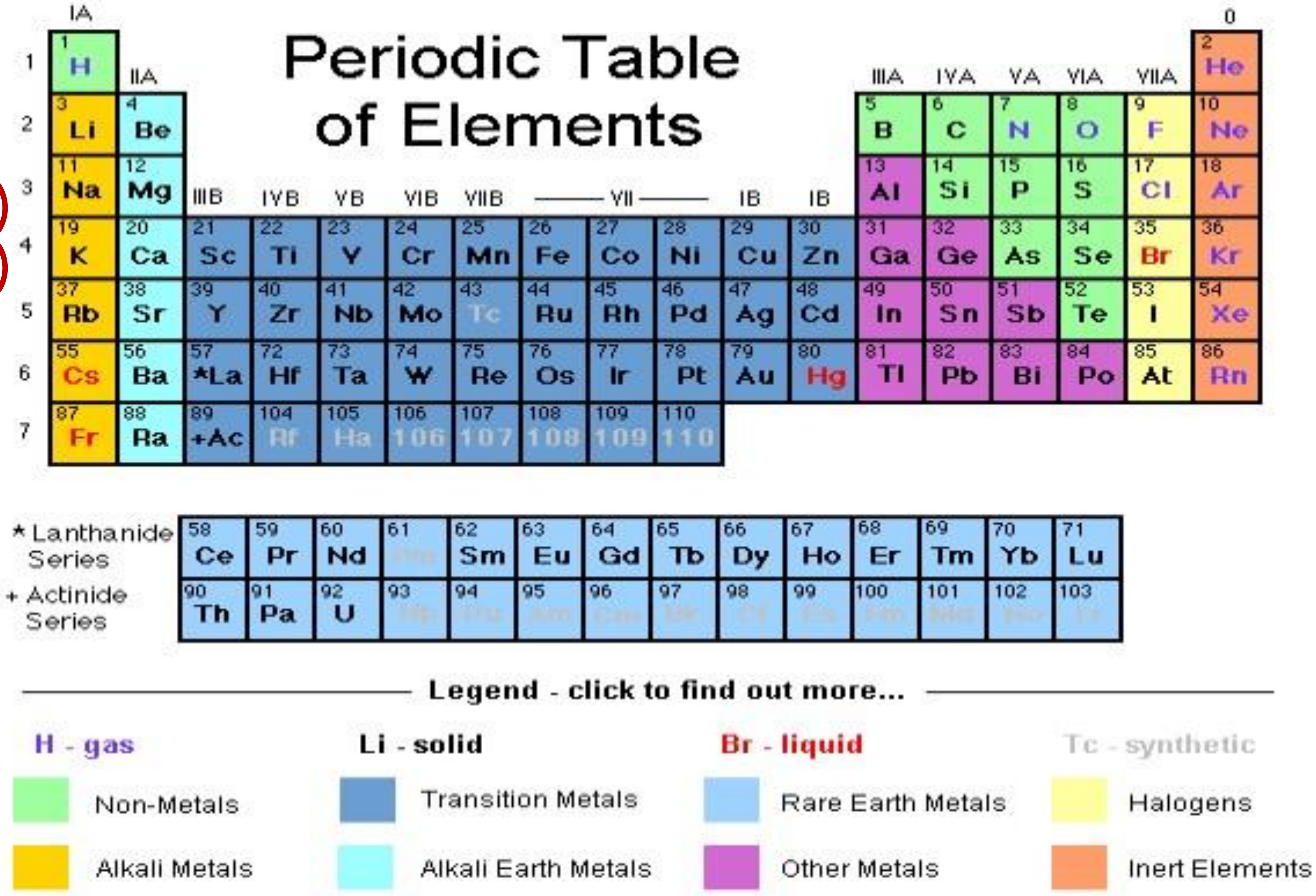
TWO WAYS OF DETERMINING HOW ELECTRONS ARE ARRANGED IN AN ATOM.

1. Periodic Table



2. Diagonal Rule





Example Na

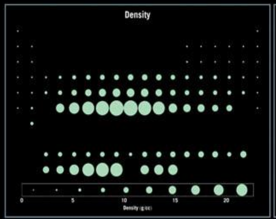
(2 ,8,1) and

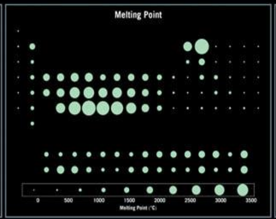
K (2,8,8,1)

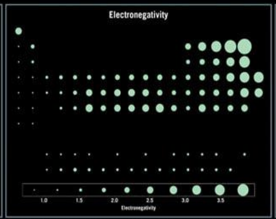
are both in

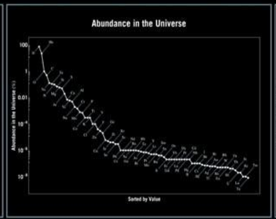
Group 1

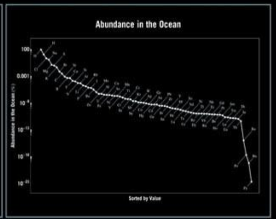
THE PERIODIC TABLE OF THE ELEMENTS

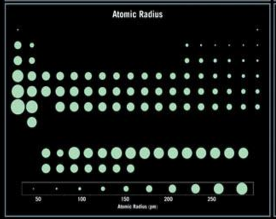


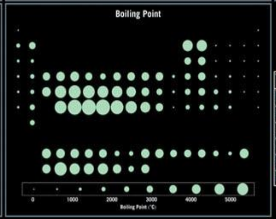




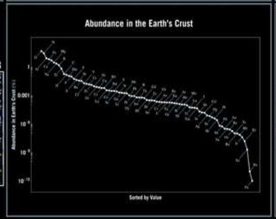


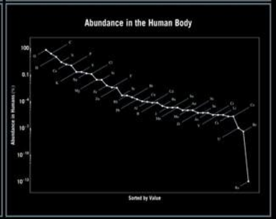






Ball plots show periodic trends. Abundances are shown on log scales. Properties are given for the most common allotrope except where noted. In the tiles below, the scale on the right shows the solid, liquid, and gas ranges graphically, and the scale on the bottom shows the density. Electronegativity is the Pauling value.





Symbol	Number	Atomic Weight	Density	Melting Point	Boiling Point	Name
H	1	1.0079	0.0899	-259.14°C (-434.45°F)	-252.87°C (-423.17°F)	Hydrogen
He	2	4.0026	0.1785	Never solidifies at standard pressure	-268.93°C (-452.07°F)	Helium
Li	3	6.941	0.535	180.54°C (356.97°F)	1342°C (2448°F)	Lithium
Be	4	9.0122	1.848	1287°C (2349°F)	2470°C (4478°F)	Beryllium
B	5	10.811	2.460	2075°C (3767°F)	4000°C (7232°F)	Boron
C	6	12.011	2.260	3550°C (6422°F)	4027°C (7281°F)	Carbon
N	7	14.007	1.251	-210.1°C (-346.2°F)	-195.79°C (-320.42°F)	Nitrogen
O	8	15.999	1.429	-218.3°C (-360.9°F)	-182.9°C (-297.2°F)	Oxygen
F	9	18.998	1.696	-219.6°C (-363.3°F)	-188.12°C (-306.62°F)	Fluorine
Ne	10	20.180	0.900	-248.59°C (-415.46°F)	-246.08°C (-410.94°F)	Neon
Na	11	22.990	0.968	97.72°C (207.9°F)	883°C (1621°F)	Sodium
Mg	12	24.305	1.738	650°C (1202°F)	1090°C (1994°F)	Magnesium
Al	13	26.982	2.7	660.32°C (1220.6°F)	2519°C (4566°F)	Aluminum
Si	14	28.085	2.330	1414°C (2577°F)	2900°C (5250°F)	Silicon
P	15	30.974	1.823	44.2°C (111.6°F)	280.5°C (536.9°F)	Phosphorus
S	16	32.065	1.960	115.21°C (239.38°F)	444.72°C (832.50°F)	Sulfur
Cl	17	35.453	3.214	-101.5°C (-150.7°F)	-34.04°C (-29.27°F)	Chlorine
Ar	18	39.948	1.784	-189.3°C (-308.7°F)	-185.8°C (-302.4°F)	Argon
K	19	39.098	0.856	63.38°C (146.1°F)	759°C (1398°F)	Potassium
Ca	20	40.078	1.550	842°C (1548°F)	1484°C (2703°F)	Calcium
Sc	21	44.956	2.985	1541°C (2806°F)	2830°C (5126°F)	Scandium
Ti	22	47.867	4.507	1668°C (3034°F)	3287°C (5949°F)	Titanium
V	23	50.942	6.110	1910°C (3470°F)	3407°C (6165°F)	Vanadium
Cr	24	51.996	7.140	1907°C (3465°F)	2671°C (4840°F)	Chromium
Mn	25	54.938	7.470	1246°C (2275°F)	2061°C (3742°F)	Manganese
Fe	26	55.845	7.874	1538°C (2800°F)	2861°C (5182°F)	Iron
Co	27	58.933	8.9	1495°C (2723°F)	2927°C (5301°F)	Cobalt
Ni	28	58.693	8.908	1455°C (2651°F)	2913°C (5275°F)	Nickel
Cu	29	63.546	8.920	1084.62°C (1984.32°F)	2927°C (5301°F)	Copper
Zn	30	65.409	7.140	419.53°C (787.15°F)	907°C (1665°F)	Zinc
Ga	31	69.723	5.904	29.76°C (85.57°F)	2204°C (3999°F)	Gallium
Ge	32	72.64	5.323	938.3°C (1721°F)	2820°C (5108°F)	Germanium
As	33	74.922	5.727	817°C (1503°F)	614°C (1137°F)	Arsenic
Se	34	78.96	4.819	221°C (430°F)	685°C (1265°F)	Selenium
Br	35	79.904	3.120	-7.3°C (19°F)	59°C (138°F)	Bromine
Kr	36	83.798	3.75	-157.36°C (-251.25°F)	-153.22°C (-243.80°F)	Krypton
Rb	37	85.468	1.532	39.31°C (102.76°F)	688°C (1270°F)	Rubidium
Sr	38	87.62	2.630	777°C (1431°F)	1382°C (2520°F)	Strontium
Y	39	88.906	4.472	1526°C (2779°F)	3345°C (6053°F)	Yttrium
Zr	40	91.224	6.511	1855°C (3371°F)	4409°C (7968°F)	Zirconium
Nb	41	92.906	8.570	2477°C (4491°F)	4744°C (8571°F)	Niobium
Mo	42	95.94	10.280	2623°C (4753°F)	4639°C (8382°F)	Molybdenum
Tc	43	[98]	11.5	2157°C (3915°F)	4265°C (7709°F)	Technetium
Ru	44	101.07	12.370	2334°C (4233°F)	4150°C (7502°F)	Ruthenium
Rh	45	102.91	12.450	1964°C (3567°F)	3695°C (6683°F)	Rhodium
Pd	46	106.42	12.023	1554.9°C (2830.8°F)	2963°C (5365°F)	Palladium
Ag	47	107.87	10.490	961.78°C (1763.2°F)	2162°C (3924°F)	Silver
Cd	48	112.41	8.650	321.07°C (609.93°F)	767°C (1413°F)	Cadmium
In	49	114.82	7.310	156.6°C (313.9°F)	2072°C (3762°F)	Indium
Sn	50	118.71	7.310	231.93°C (449.47°F)	2602°C (4716°F)	Tin
Sb	51	121.76	6.697	630.63°C (1167.1°F)	1587°C (2889°F)	Antimony
Te	52	127.60	6.240	449.51°C (841.12°F)	988°C (1810°F)	Tellurium
I	53	126.90	4.940	113.7°C (236.7°F)	184.3°C (363.7°F)	Iodine
Xe	54	131.29	5.9	-111.8°C (-169.2°F)	-108°C (-162°F)	Xenon
Cs	55	132.91	1.879	28.44°C (83.19°F)	671°C (1240°F)	Cesium
Ba	56	137.33	3.510	727°C (1341°F)	1870°C (3398°F)	Barium
Hf	72	178.49	13.310	2233°C (4051°F)	4603°C (8317°F)	Hafnium
Ta	73	180.95	16.650	3017°C (5463°F)	5458°C (9856°F)	Tantalum
W	74	183.84	19.250	3422°C (6192°F)	5555°C (10,031°F)	Tungsten
Re	75	186.21	21.020	3186°C (5767°F)	5596°C (10,105°F)	Rhenium
Os	76	190.23	22.59	3033°C (5491°F)	5012°C (9054°F)	Osmium
Ir	77	192.22	22.56	2466°C (4471°F)	4428°C (8002°F)	Iridium
Pt	78	195.08	21.090	1768.3°C (3214.9°F)	3825°C (6917°F)	Platinum
Au	79	196.97	19.3	1064.18°C (1947.52°F)	2856°C (5173°F)	Gold
Hg	80	200.59	13.534	-38.83°C (-37.89°F)	356.73°C (674.11°F)	Mercury
Tl	81	204.38	11.850	304°C (579°F)	1473°C (2683°F)	Thallium
Pb	82	207.2	11.340	327.46°C (621.43°F)	1749°C (3180°F)	Lead
Bi	83	208.98	9.780	271.3°C (520.3°F)	1564°C (2847°F)	Bismuth
Po	84	[209]	9.196	254°C (489°F)	962°C (1764°F)	Polonium
At	85	[210]	Unknown	302°C (576°F)	Unknown	Astatine
Rn	86	[222]	9.73	-71°C (-96°F)	-61.7°C (-79°F)	Radon
Fr	87	[223]	Unknown	Unknown	Unknown	Francium
Ra	88	[226]	5.0	700°C (1290°F)	1737°C (3159°F)	Radium
Rf	104	[261]	Unknown	Unknown	Unknown	Rutherfordium
Db	105	[262]	Unknown	Unknown	Unknown	Dubnium
Sg	106	[266]	Unknown	Unknown	Unknown	Seaborgium
Bh	107	[264]	Unknown	Unknown	Unknown	Bohrium
Hs	108	[277]	Unknown	Unknown	Unknown	Hassium
Mt	109	[268]	Unknown	Unknown	Unknown	Meitnerium
Ds	110	[281]	Unknown	Unknown	Unknown	Darmstadtium
Rg	111	[272]	Unknown	Unknown	Unknown	Roentgenium
Uub	112	[285]	Unknown	Unknown	Unknown	Ununbium
Uut	113	[284]	Unknown	Unknown	Unknown	Ununtrium
Uuq	114	[289]	Unknown	Unknown	Unknown	Ununquadium
Uup	115	[288]	Unknown	Unknown	Unknown	Ununpentium
Uuh	116	[292]	Unknown	Unknown	Unknown	Ununhexium
Uus	117	Unknown	Unknown	Unknown	Unknown	Ununseptium
Uuo	118	[294]	Unknown	Unknown	Unknown	Ununoctium

Lanthanides

This space should contain elements 57-71 but in order to make the table fit in a reasonable shape they are customarily shown on a separate line below.

Actinides

As with the Lanthanides elements 89-103 are shown in a separate row below purely for space reasons, chemically they belong here.

Symbol	Number	Atomic Weight	Density	Melting Point	Boiling Point	Name
La	57	138.91	6.146	920°C (1688°F)	3464°C (6267°F)	Lanthanum
Ce	58	140.12	6.689	798°C (1468°F)	3360°C (6080°F)	Cerium
Pr	59	140.91	6.640	931°C (1708°F)	3290°C (5954°F)	Praseodymium
Nd	60	144.24	7.010	1021°C (1870°F)	3100°C (5610°F)	Neodymium
Pm	61	[145]	7.264	1100°C (2010°F)	3000°C (5430°F)	Promethium
Sm	62	150.36	7.353	1072°C (1962°F)	1803°C (3277°F)	Samarium
Eu	63	151.96	5.244	822°C (1512°F)	1527°C (2781°F)	Europium
Gd	64	157.25	7.901	1313°C (2395°F)	3250°C (5882°F)	Gadolinium
Tb	65	158.93	8.219	1356°C (2473°F)	3230°C (5846°F)	Terbium
Dy	66	162.5	8.551	1412°C (2574°F)	2567°C (4653°F)	Dysprosium
Ho	67	164.93	8.795	1474°C (2685°F)	2700°C (4892°F)	Holmium
Er	68	167.26	9.066	1497°C (2727°F)	2868°C (5194°F)	Erbium
Tm	69	168.93	9.321	1545°C (2813°F)	1950°C (3542°F)	Thulium
Yb	70	173.04	6.570	819°C (1506°F)	1196°C (2185°F)	Ytterbium
Lu	71	174.97	9.841	1663°C (3025°F)	3402°C (6156°F)	Lutetium
Ac	89	[227]	10.070	1050°C (1922°F)	3200°C (5792°F)	Actinium
Th	90	232.04	11.724	1750°C (3182°F)	4820°C (8708°F)	Thorium
Pa	91	231.04	15.370	1572°C (2862°F)	4000°C (7232°F)	Protactinium
U	92	238.03	19.050	1135°C (2075°F)	3927°C (7101°F)	Uranium
Np	93	[237]	20.450	644°C (1191°F)	4000°C (7230°F)	Neptunium
Pu	94	[244]	19.816	640°C (1184°F)	3230°C (5846°F)	Plutonium
Am	95	[243]	Unknown	1176°C (2149°F)	2011°C (3652°F)	Americium
Cm	96	[247]	13.510	1345°C (2453°F)	3110°C (5630°F)	Curium
Bk	97	[247]	14.780	1050°C (1922°F)	Unknown	Berkelium
Cf	98	[251]	15.1	900°C (1652°F)	Unknown	Californium
Es	99	[252]	Unknown	860°C (1580°F)	Unknown	Einsteinium
Fm	100	[257]	Unknown	1527°C (2781°F)	Unknown	Fermium
Md	101	[258]	Unknown	827°C (1521°F)	Unknown	Mendelevium
No	102	[259]	Unknown	827°C (1521°F)	Unknown	Nobelium
Lr	103	[262]	Unknown	1627°C (2961°F)	Unknown	Lawrencium

Photographs show samples of the pure or nearly pure element except as follows: At, Rn, Fr, Ac, Pa, and Np show radioactive minerals containing minute traces of the element. Po, Ra, Pm, Pu, and Am show artificial objects containing invisible amounts of the element. Technetium shows a Tc-99 bone scan. Hydrogen shows a Hubble Space Telescope image of the Eagle Nebula, which is mostly hydrogen. 96-111 show the person or place after which the element is named. 112-118 had not been named yet in 2009.

Poster and photography by Theodore W. Gray and Nick Mann.

Technical data from *Mathematica*™'s ChemicaData information service by Wolfram Research, Inc.

Other sizes of this poster: periodictable.com
Real samples like these: element-collection.com

On the other side of this poster you will find a version with the pictures filling the whole space of each element and only minimal text to get in the way.

More images and complete technical data can be found at periodictable.com.

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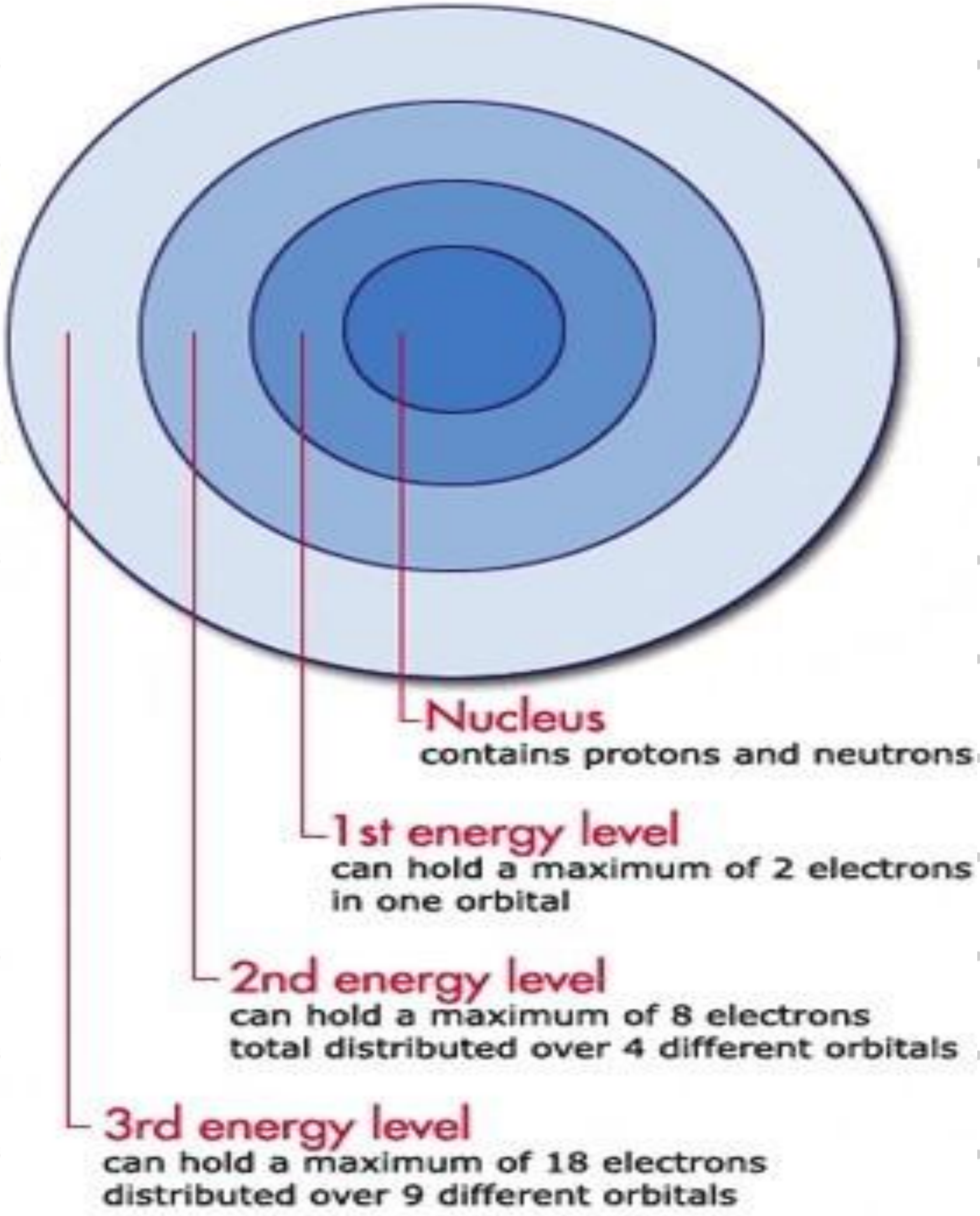
Distribution of Electrons

- Electrons revolve around the nucleus of an atom in a definite pattern. Groups of electrons maintain definite average distances from the nucleus forming ***shell*** or ***energy levels*** of electrons surrounding the nucleus. Each shell is capable of containing a definite number of electrons, the number increasing in distance from the nucleus increases. Letters – k, l, m, n, o, and p starting with the shell nearest the nucleus, designates the shells. The maximum number of electrons in any shell can be calculated from the relationship:

Number = 2s Where:

- Number = maximum number of electrons possible in the shell
- S = the number of the shell (K=1, l=2. m=3, etc)



Nucleus
contains protons and neutrons
1st energy level
can hold a maximum of 2 electrons in one orbital
2nd energy level
can hold a maximum of 8 electrons total distributed over 4 different orbitals
3rd energy level
can hold a maximum of 18 electrons distributed over 9 different orbitals

Sublevels

The energy levels are further subdivided into sublevels designated by the letters **s, p, d, f, g …** the number of which corresponds to the number of the energy level.

- Each sublevel has a set of orbital, which are of equal energy.

K (n=1) one sublevel: 1s

L (n=2) two sublevels: 2s and 2p

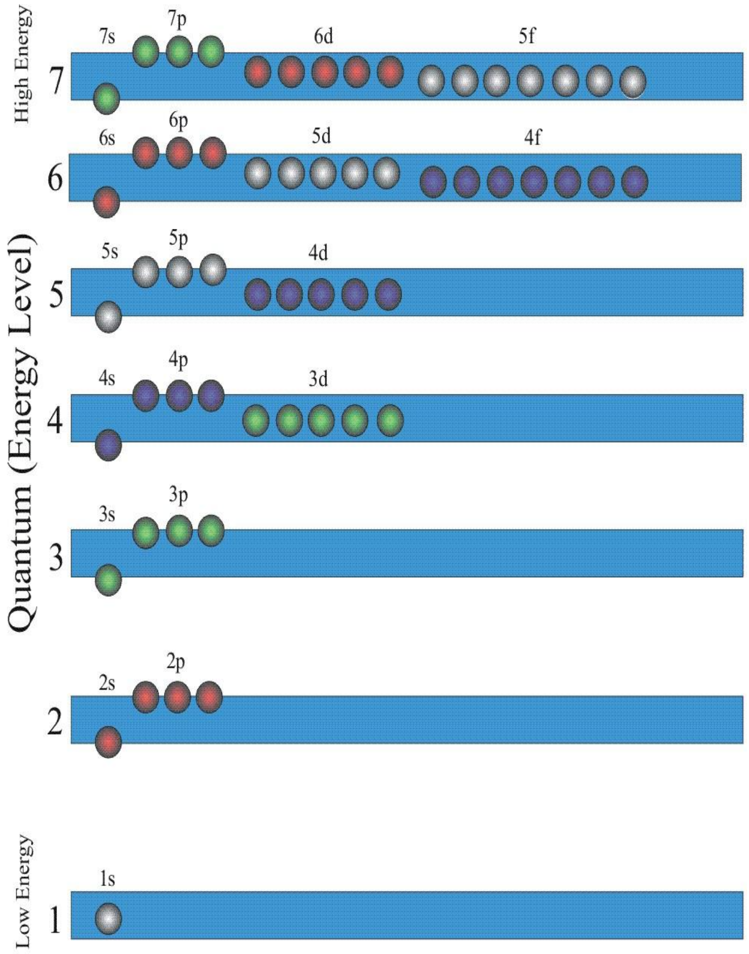
M (n=3) three sublevels: 3s, 3p, and 3d

s = 2 electrons

p = 6 electrons

d = 10 electrons

f = 14 electrons



High Energy
Low Energy
Quantum (Energy Level)
7
7s
7p
6d
5f
6
6s
6p
5d
4f
5
5s
5p
4d
4
4s
4p
3d
3
3s
3p
2
2s
2p
1
1s

Orbital is the home of the electrons or the region of space where the probability of finding the electrons is greatest.

Orbital diagram for oxygen

Rules in building up the electronic configuration

- The number of electrons entering the atom must be equal to its atomic number, z, and the number of protons. Thus, the atom is neutral.
- No more than 2 electrons with opposite spins can enter any single orbital (Pauli's Exclusion Principle.)
- When there are orbitals of the same kind of energy, the electrons occupy the equivalent

orbitals singly to the maximum and with their spins parallel (Hund's Rule).

- The opposite spin may be represented by arrows pointing upwards and downwards.

Wave Mechanics/Quantum Mechanics or Orbital Theory

Light exhibits dual wave-particle properties. Interference and diffraction patterns formed when light passes through slits can only be explained by the addition of waves. Discontinuous emission of light from hot bodies can only be explained by particle-like photons of emitted light. *Louis de Broglie* reasoned that if light can exhibit wave and particle propertie

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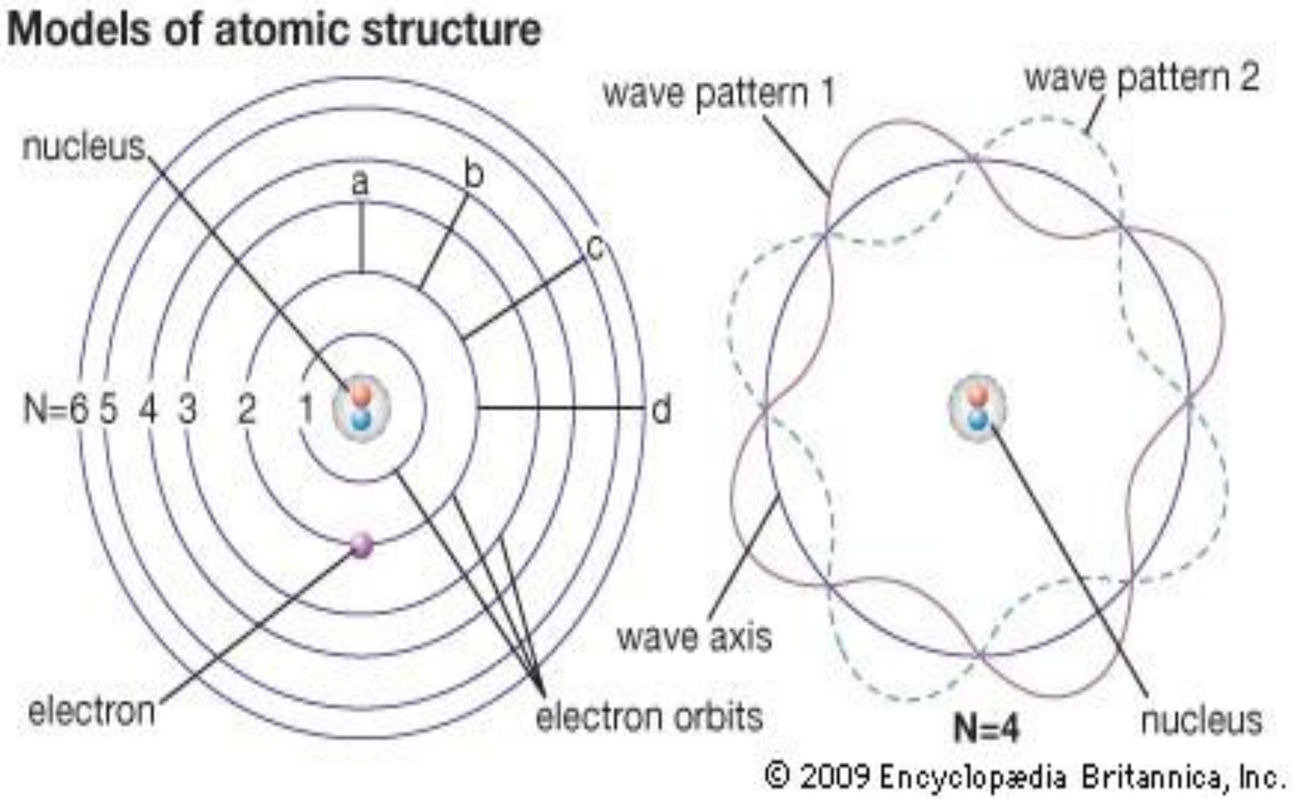
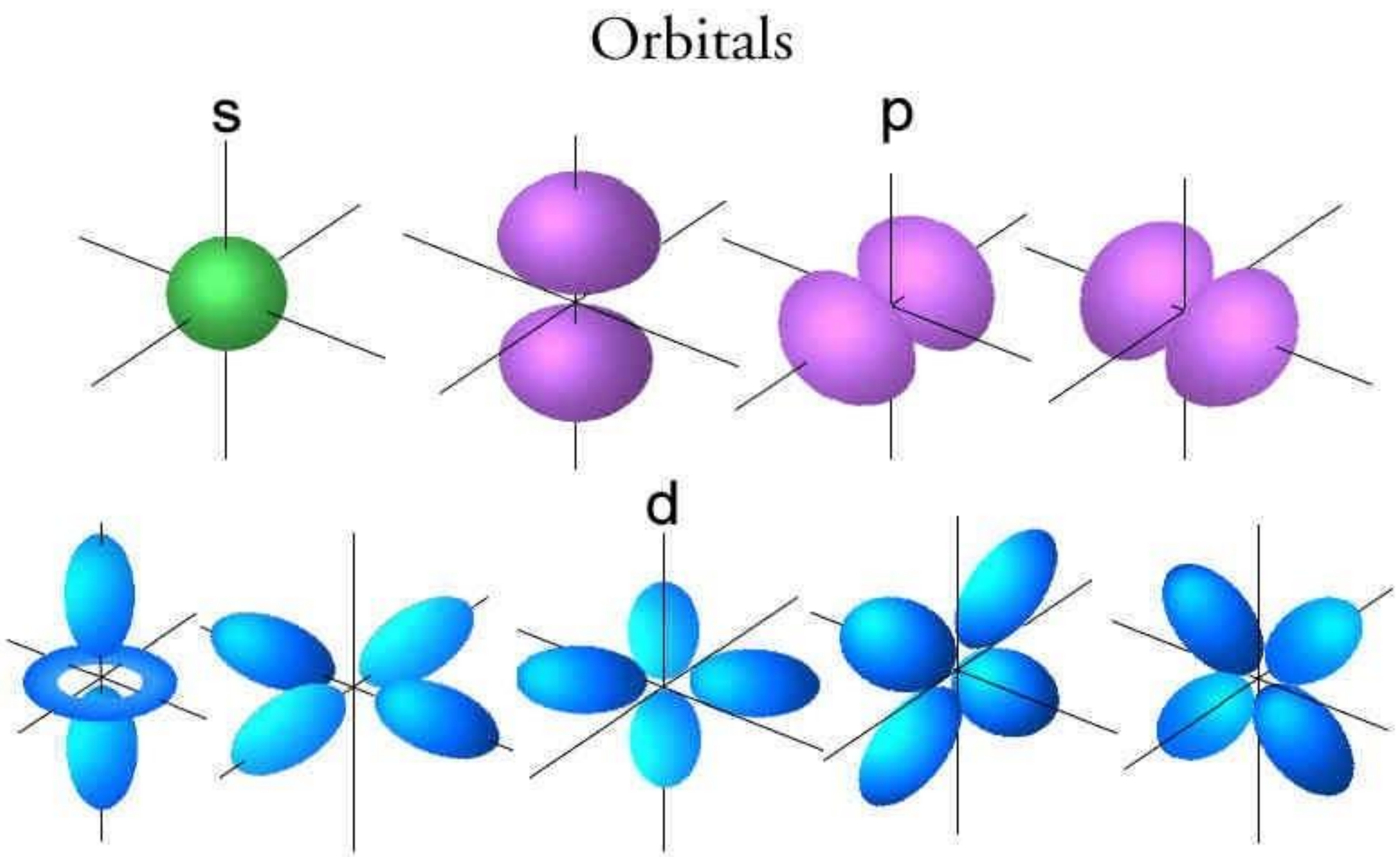


Quantum Numbers

- The **Principal Quantum Number (n)** is associated with the distance of the electron from the nucleus and it determines the gross energy of the electron.
- The **Second Quantum Number** *(Azimuthal Quantum Number)* **(l)** gives the shape of the orbital. It has integral values from 0 to n-1.
- The **Third Quantum Number** *(Magnetic Quantum Number)* **(m)** describes the orientation of the orbital in space. The integral values may be l_1, l-1, l-2 down to -l. Positive values of m describe orientation in the direction of applied magnetic field while negative values refer to orientation in the opposite direction.
- The **Fourth Quantum Number** is the electron spin quantum number (m). It describes the spinning of the electron on its axis. It can have a clockwise spin or counter clockwise spin.
- ***Pauli's Exclusion Principle*** states that no two electrons can have the same set of four quantum numbers.

Shapes of Orbitals



Quantum Numbers

n	l	m_l	Orbital	Elements		Shell
$n = 1$	0	0	1s	2	2	K
$n = 2$	0	0	2s	2	8	L
	1	−1, 0, 1	2p	6		
$n = 3$	0	0	3s	2	18	M
	1	−1, 0, 1	3p	6		
	2	−2, −1, 0, 1, 2	3d	10		
$n = 4$	0	0	4s	2	32	N
	1	−1, 0, 1	4p	6		
	2	−2, −1, 0, 1, 2	4d	10		
	3	−3, −2, −1, 0, 1, 2, 3	4f	14		
$n = 5$	0	0	5s	2	32	O
	1	−1, 0, 1	5p	6		
	2	−2, −1, 0, 1, 2	5d	10		
	3	−3, −2, −1, 0, 1, 2, 3	5f	14		
	4	−4, −3, −2, −1, 0, 1, 2, 3, 4	5g	18	Unknown Corresponding Elements	
$n = 6$	0	0	6s	2	18	P
	1	−1, 0, 1	6p	6		
	2	−2, −1, 0, 1, 2	6d	10		
	3	−3, −2, −1, 0, 1, 2, 3	6f	14	54 Unknown Corresponding Elements	
	4	−4, −3, −2, −1, 0, 1, 2, 3, 4	6g	18		
	5	−5, −4, −3, −2, −1, 0, 1, 2, 3, 4, 5	6h	22		
$n = 7$	0	0	7s	2	8	Q
	1	−1, 0, 1	7p	6		
	2	−2, −1, 0, 1, 2	7d	10	90 Unknown Corresponding Elements	
	3	−3, −2, −1, 0, 1, 2, 3	7f	14		
	4	−4, −3, −2, −1, 0, 1, 2, 3, 4	7g	18		
	5	−5, −4, −3, −2, −1, 0, 1, 2, 3, 4, 5	7h	22		
	6	−6, −5, −4, −3, −2, −1, 0, 1, 2, 3, 4, 5, 6	7i	26		