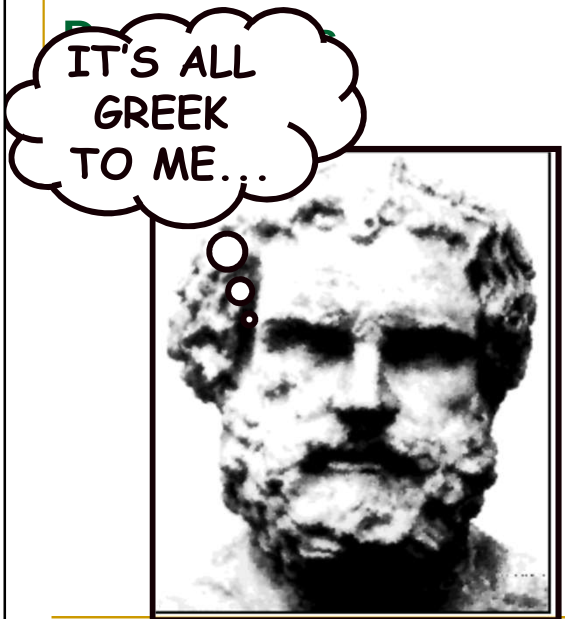
40S Chemistry

Development of Atomic Theory

Pioneers of Atomic Theory

- Democritus
- John Dalton
- John Joseph Thomson
- Ernest Rutherford
- Neils Bohr



- A Greek philosopher.
- All matter in the universe was made up of tiny particles.
- Tiny, indivisible
- Called “atomos” (indivisible).
- He had an opponent . . .

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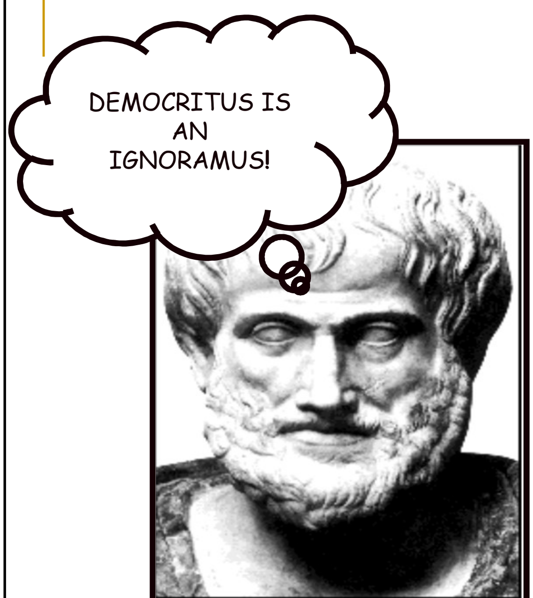
Aristotle



- He opposed idea of the atom.
- Believed that matter was continuous (was not made up of smaller parts).
- Was more popular than Democritus.
- Development of the atomic theory ceased for about 14 centuries!

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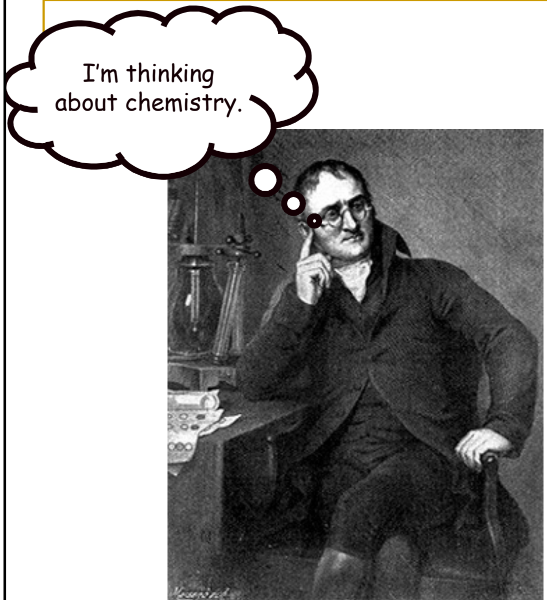

John Dalton

- 1808
- Based ideas on the work of other chemists.
- Developed Dalton's Atomic Theory . . .

  . . . All elements composed of tiny indivisible particles called atoms.

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. . . Atoms of the same element are identical and are different from those of any other element.

. . . Atoms from different elements combine with each other in simple whole number ratios.

. . . Chemical reactions occur when atoms are separated, joined or rearranged.

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- Some of Dalton's ideas were incorrect . . .

. . . Particles of gases are not close together.

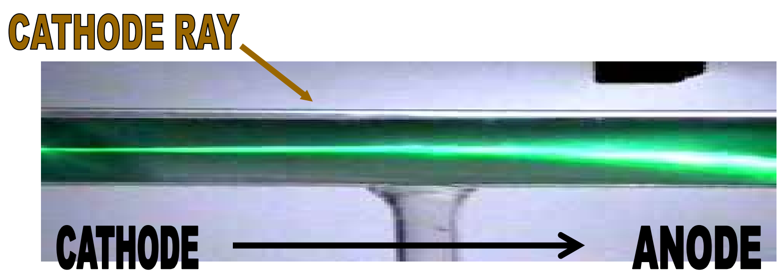
. . . Atoms of elements do not always combine in 1:1 ratios.

. . . Atoms are far from indivisible.

J. J. Thomson

- 1897
- Discovered that cathode rays could be deflected by magnets or by electrically charged plates.
- Cathode rays were negatively charged.

- The cathode ray traveled from cathode to anode when current was passed through the tube.
- The cathode is negatively charged, the anode is positively charged.



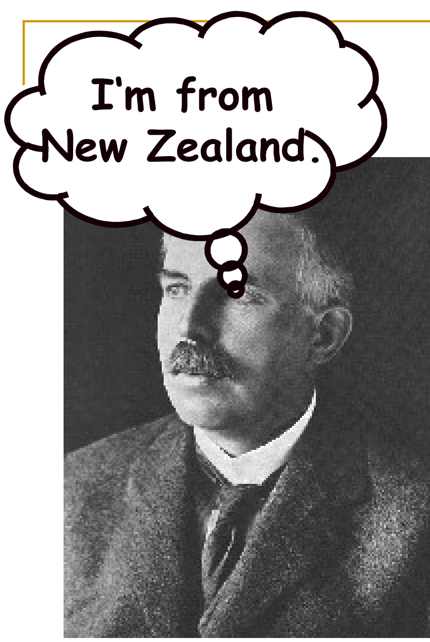
video

CATHODE RAY TUBE

- Thomson concluded that cathode rays were made up of negatively charged particles.
- Thomson found that the mass of each particle was identical, no matter what type of metal was used as a cathode.
- These particles were later called electrons.

- Thomson proposed the "plum pudding " model.
- Negative particles in an atom were scattered about an atom like chocolate chips in a cookie.
- The dough of the cookie represents positive charges.

Ernest Rutherford



- 1911
- Rutherford experimented to test Thomson's model of the atom.
- Performed the now famous "gold foil" experiment . . .

- Rutherford directed a narrow beam of alpha particles (α) at a thin film of gold foil.
- Alpha particles are helium atoms that have lost both electrons.
- They have two protons and a 2+ charge.

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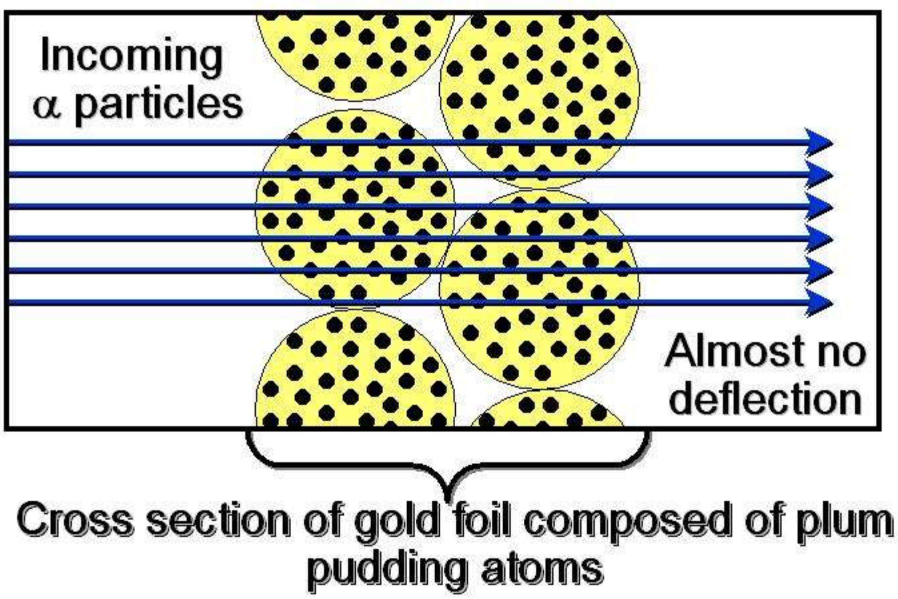
Hypothesis: expected result based on "plum pudding" model



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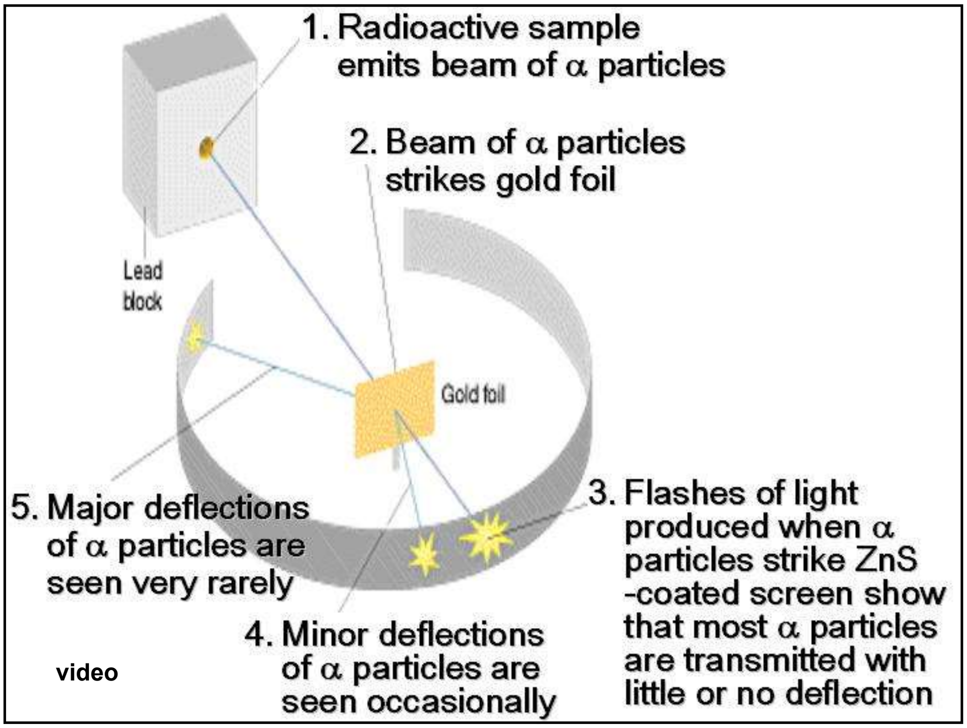
1. Radioactive sample emits beam of α particles
2. Beam of α particles strikes gold foil
Lead block
Gold foil
3. Flashes of light produced when α particles strike ZnS -coated screen show that most α particles are transmitted with little or no deflection
4. Minor deflections of α particles are seen occasionally
5. Major deflections of α particles are seen very rarely
video

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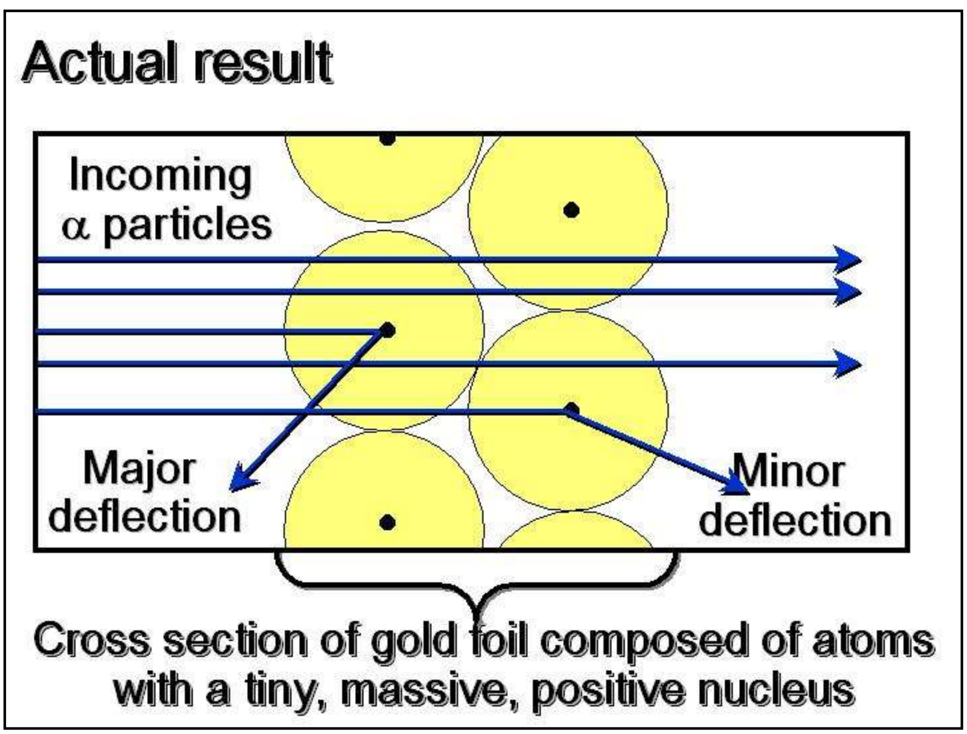
Actual result
Incoming α particles
Major deflection
Minor deflection
Cross section of gold foil composed of atoms with a tiny, massive, positive nucleus

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- Based upon his results, Rutherford proposed a different model of the atom.
- He proposed that most of the atom was empty space which surrounded a small, dense, positively charged core he called the nucleus.

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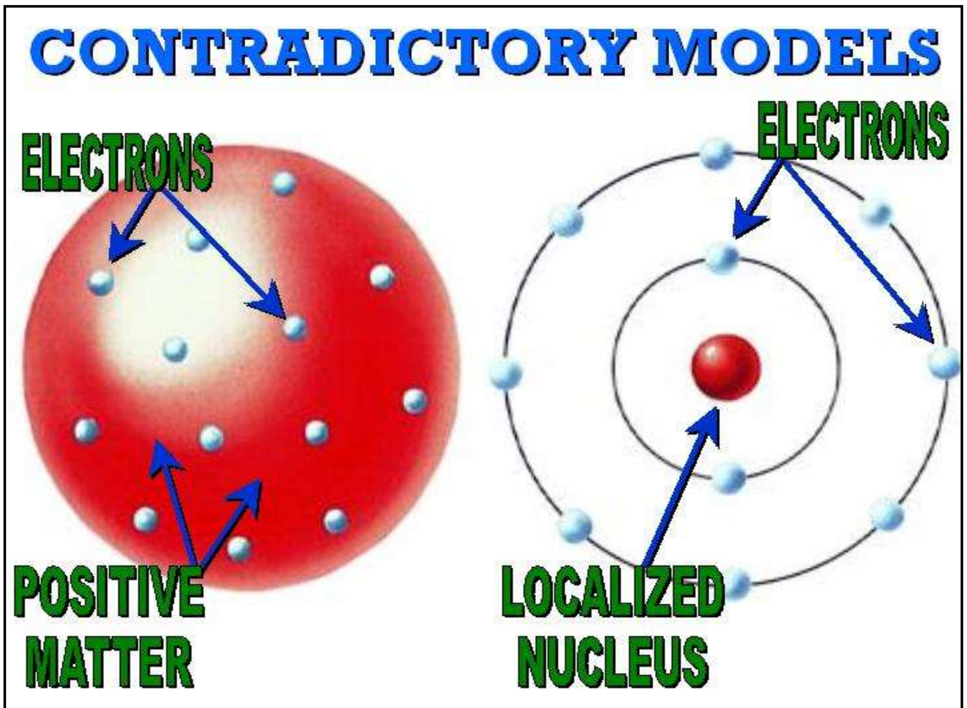

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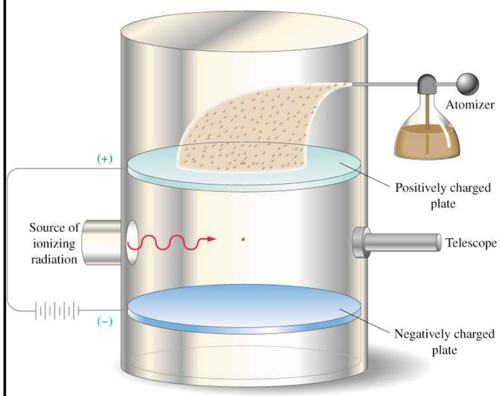
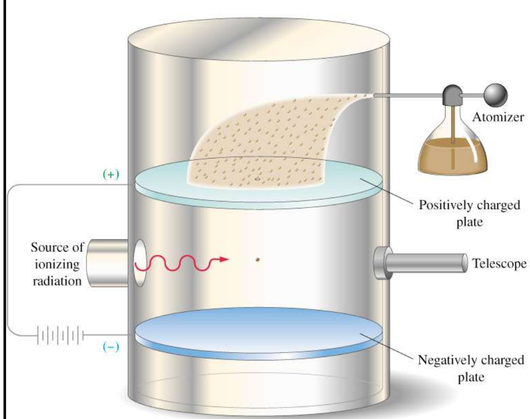
Millikan's Oil Drop Experiment



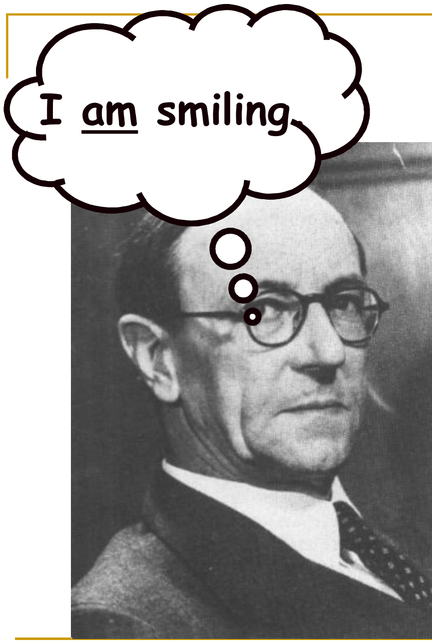
- The oil drop experiment was an experiment performed by Robert Millikan and Harvey Fletcher in 1909

video



- Measured the elementary electric charge (the charge of the electron).

video



James Chadwick

- 1932
- Confirmed the existence of another subatomic particle.
- The neutron was nearly the mass of a proton.
- Neutrons have no charge.

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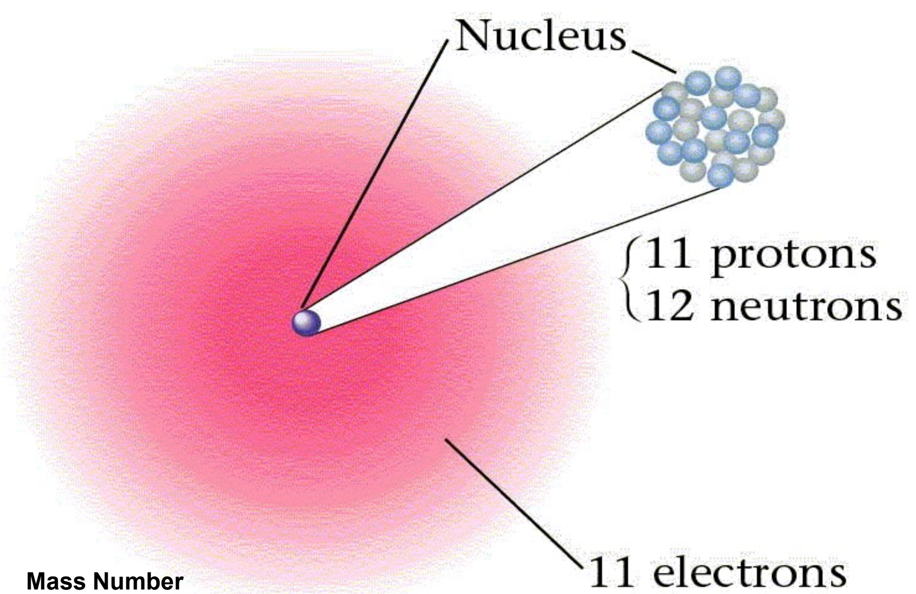

Mass Number

Atomic Number

$^{23}_{11}Na$

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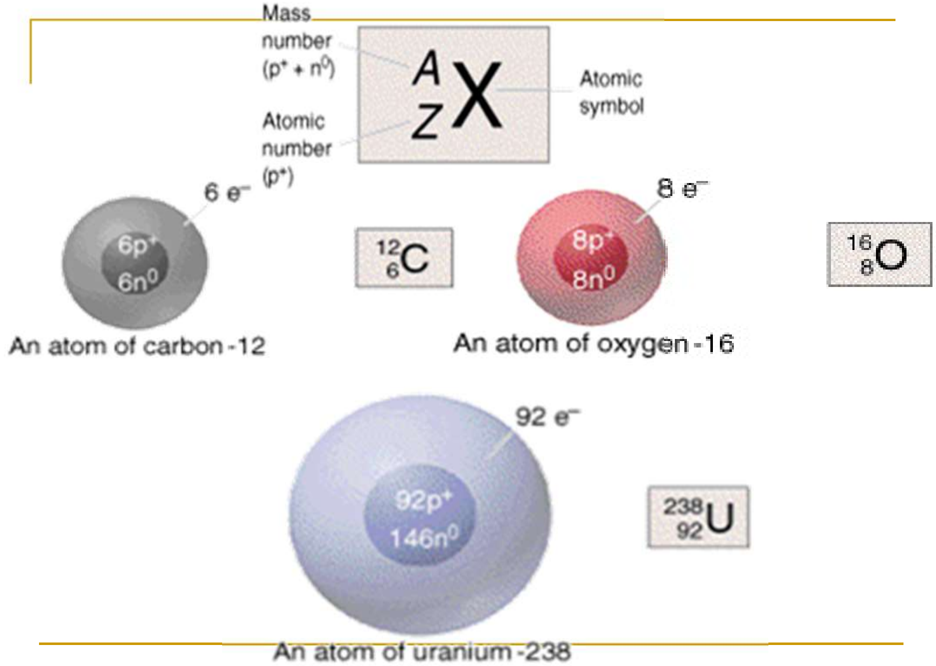
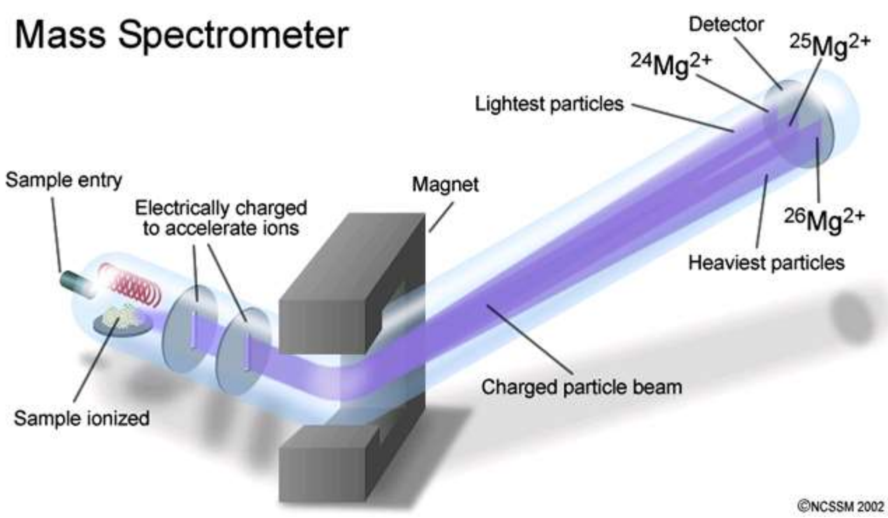


Isotopes

- Every Na atom has 11 protons.
- Not every Na atom has 12 neutrons.
- Atoms with the same number of protons but different numbers of neutrons are isotopes.
- Different isotopes of an atom have different masses.
- The atomic mass of an element in the periodic table is an average of the masses of an element's isotopes.

- Isotopes of an element are chemically alike because they have the same number of protons and electrons.
- The protons and electrons of an atom determine its chemical behavior.

video



Example of an Isotope

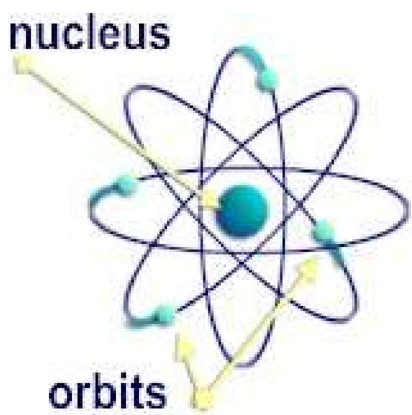
MASS NUMBER

$^{35}_{17}Cl$ $\quad$ $^{37}_{17}Cl$

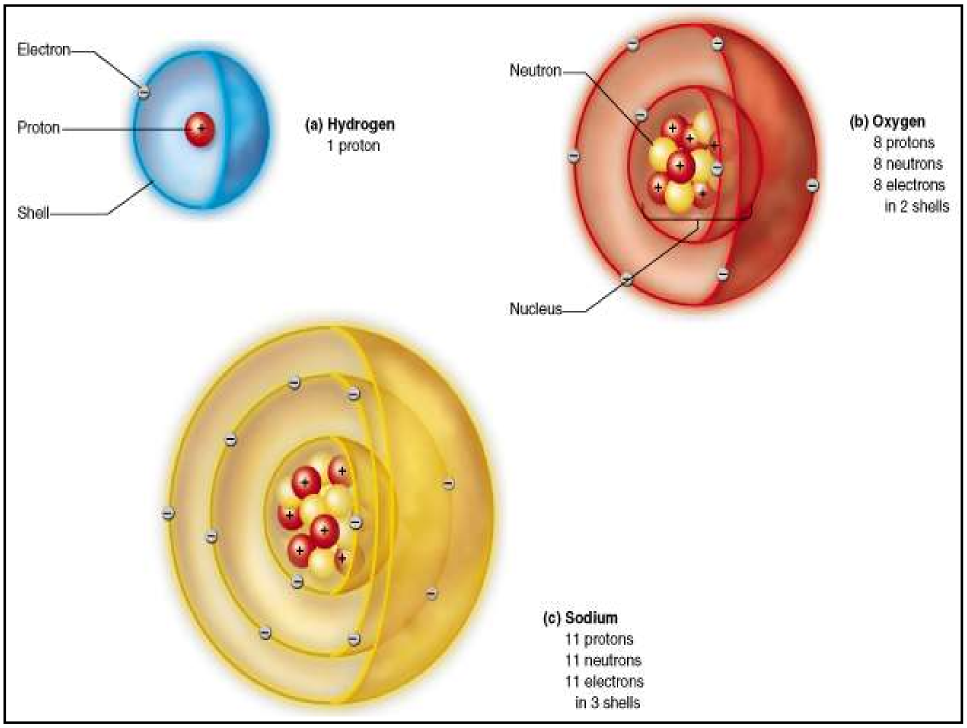
18 **NEUTRONS** $\quad$ **20** **NEUTRONS**

ATOMIC NUMBER

Niels Bohr



- 1915
- The Bohr Model is known as the "planetary model" of the atom.
- Electrons are arranged in concentric circles (orbits) around the nucleus.



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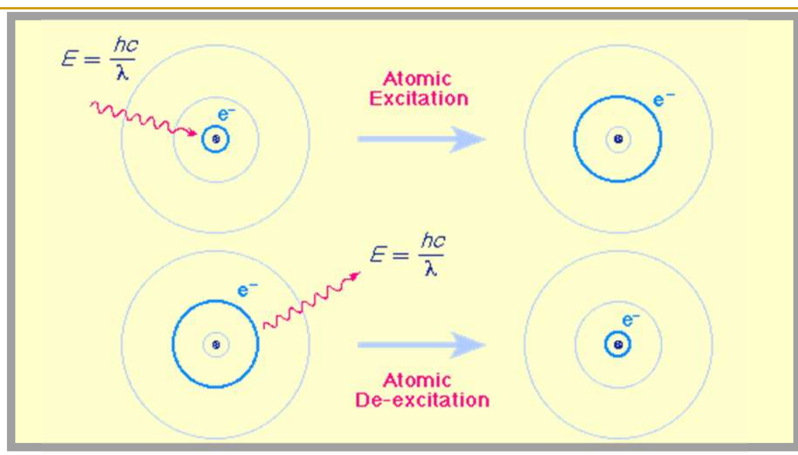

- The energy level of electrons is quantized.
- Electrons become"excited" when their energy level increases.
- They can move further away from the nucleus as a result, one energy level at a time.

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- The energy an electron gains or loses is light energy (a photon).
- A quantum of energy is the amount of energy required to move an electron form its present energy level to the next higher one.

http://www.colorado.edu/physics/2000/index.pl

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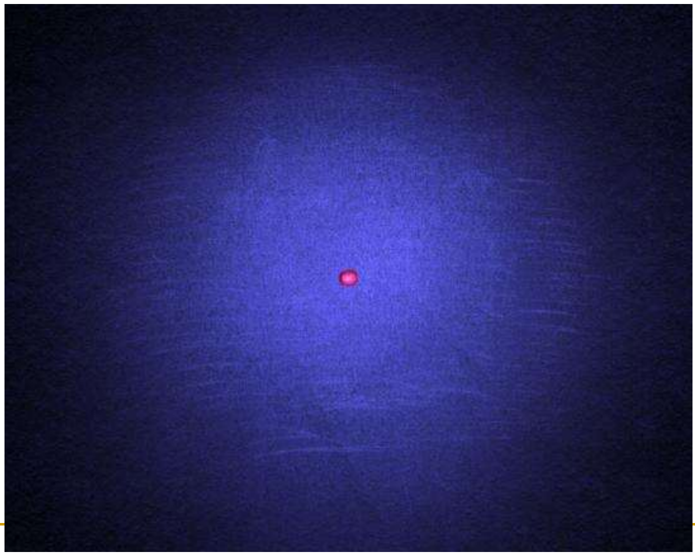
Current Atomic Model



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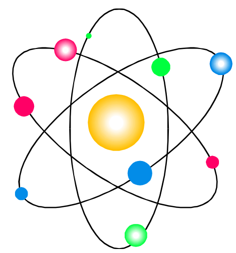
Current Atomic Model

- Electrons, Protons and Neutrons
- How Does the Electron Move Around the Atom?
- This is not what an atom looks like

Assignment

- Complete the *Atomic Structure* Test



40S Chemistry

Atomic Structure

Electron Configurations & Periodicity 35

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Introduction

- Atomic structure explains chemical properties and patterns of chemical reactivity.
- Chemical reactions involve electrons. Knowing where the electrons are, how many, and what their energy levels helps explain many chemical phenomena.
- Spectroscopy is used to explore atomic structure. Because of this, we start with a discussion of the nature of light.

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The Nature of Light

Particles or Waves??

- Light must be made of particles because it…
  - travels in a vacuum
  - reflects off of objects
  - exerts force (on the tails of comets)
- Light must consist of waves because it…
  - reflects like waves
  - refracts and diffracts
  - exhibits interference

Sir Isaac Newton

Light is Corpuscles

Christian Huygens

Light is Waves

By the end of the 19th century, scientists had concluded that light is composed of WAVES!

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Electromagnetic Radiation

- Visible light is just one form of electromagnetic radiation
- Light propagates in space as a wave
- In vacuum, speed of light is constant and given the symbol "c", $c = 3.00 \times 10^8$ m/s
- Light waves have amplitude, frequency, and wavelength.
- λ = wavelength - distance between consecutive crests, ν = frequency, # of crests that pass a given point in one second (SI Unit is s^{-1} or Hz)
- $c = \nu\lambda$ and $\nu = c/\lambda$

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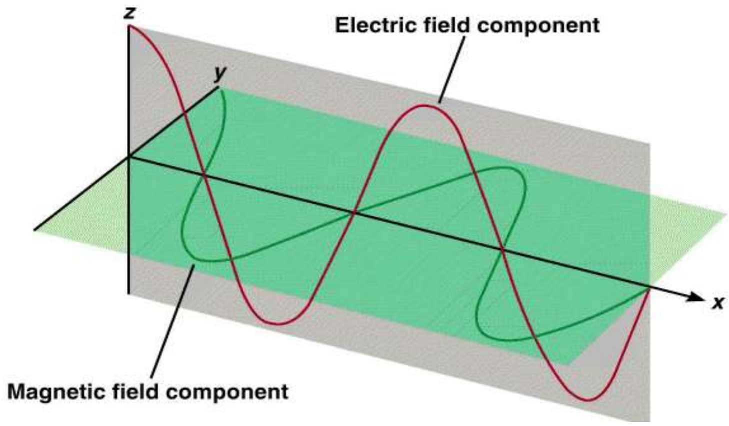
Components of an Electromagnetic Wave



James Maxwell proposed in the mid-1800's that light is composed of perpendicular electric and magnetic waves

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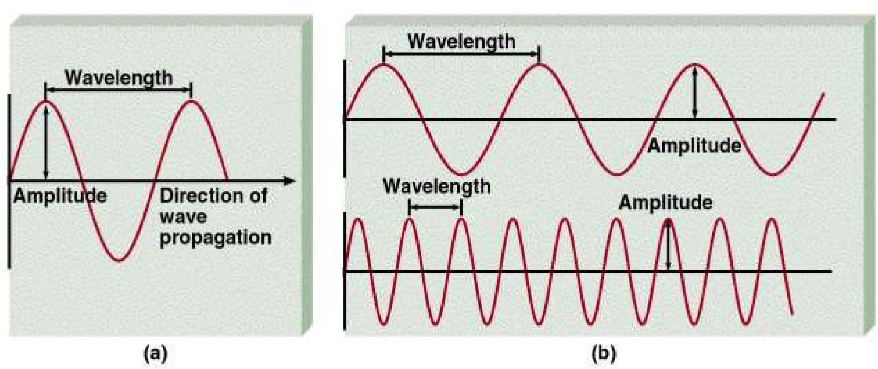
Wavelength and Amplitude

are distances – S.I. unit is the meter



Which of the above waves has the higher ***frequency****?*

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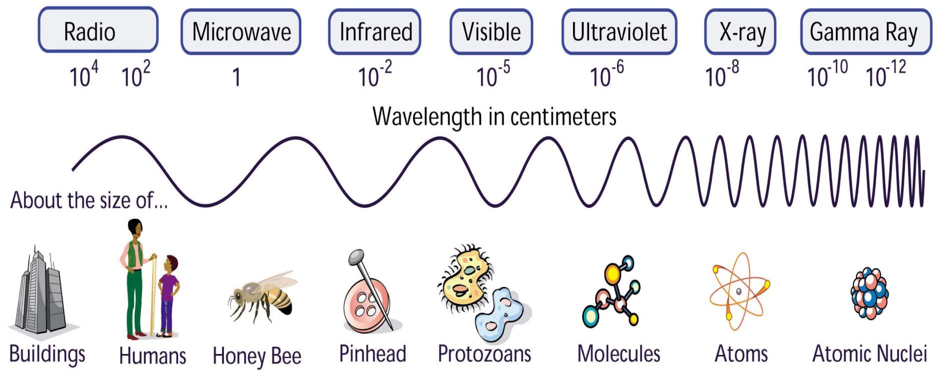
The 7 Regions of the E.M. Spectrum



Video

The World in UV video

The Wave Equation

$$\boxed{c = \lambda\nu}$$

c = Speed of light (3.0×10^8 m/s)

λ = wavelength (m)

ν = frequency (Hz)

Sample Calculation –
wavelength/frequency conversion

- Calculate the frequency of visible light having a wavelength of 485 nm?
  Remember to use S.I. units in your calculations!

  c = 3.00 x 10^8 m/s λ = 485 x 10^{-9} m
- c = $\lambda\nu$
- ν = c / λ

  = (3.00 x 10^8 m/s) ÷ (485 x 10^{-9} m)

  = 6.19 x 10^{14} s^{-1}
- What colour of visible light is this?

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Sample Calculation –
wavelength & frequency conversion

- CO_2 absorbs light with a wavelength of 0.018 mm. Which frequency is this?

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8\ ms^{-1}}{0.018 \times 10^{-3}\ m} = 1.7 \times 10^{13}\ s^{-1}$$

Remember that 1 Hertz (Hz) $\equiv$ 1 s^{-1}

- What is the wavelength of WFAE at 90.7 MHz?

$$\lambda = \frac{c}{\nu} = \frac{3.00 \times 10^8\ ms^{-1}}{90.7 \times 10^6\ s^{-1}} = 3.31\ m$$

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Particle Theory of Light

- In 1900, Max Planck turned the world of physics on its head by resurrecting the particle theory of light.
- Planck proposed that light is composed of particles (quanta) each carrying a fixed amount of energy.
- The amount of energy per quantum is directly proportional to the frequency of the light:

$$E = h\nu \quad \text{or} \quad E = \frac{hc}{\lambda} \quad , h = 6.626 \times 10^{-34}\ J \cdot s$$

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Planck's Equation – sample problem

- The wavelength of maximum visual acuity in humans is 550 nm. (green light)
  - What is the energy of a single photon having this wavelength?
    - Ans. 3.61×10^{-19} J per photon
  - What is the energy of a mole of photons at 550 nm?
    - Ans. 218 kJ/mol

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Energy, wavelength, frequency

a) X-Rays
b) Red light
c) Green light
d) Radio waves

- Rank the above in order of …
  - increasing wavelength
  - decreasing energy
  - increasing frequency

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Quantum Mechanics Video

- Part 1 – An Introduction to Modern Physics
- Part 2 – Modern Atomic Structure
- Part 3 – The Electron Shells
- Part 4 – Electron Spin

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Atomic Emission Spectra

"Line Spectra" vs "Continuous Spectra"

- Historic work (1800's) involving light emitted by pure elements in gas phase, subjected to very high voltages.
- The light was passed through a prism or diffraction grating to produce a spectrum.
- Each element emitted only certain wavelengths of light – a LINE SPECTRUM instead of the more familiar "continuous" spectrum seen in the rainbow.
- Balmer and Rydberg described the lines mathematically, for hydrogen.
- Hydrogen's emission spectrum has several lines in the visible region – called the "Balmer series"

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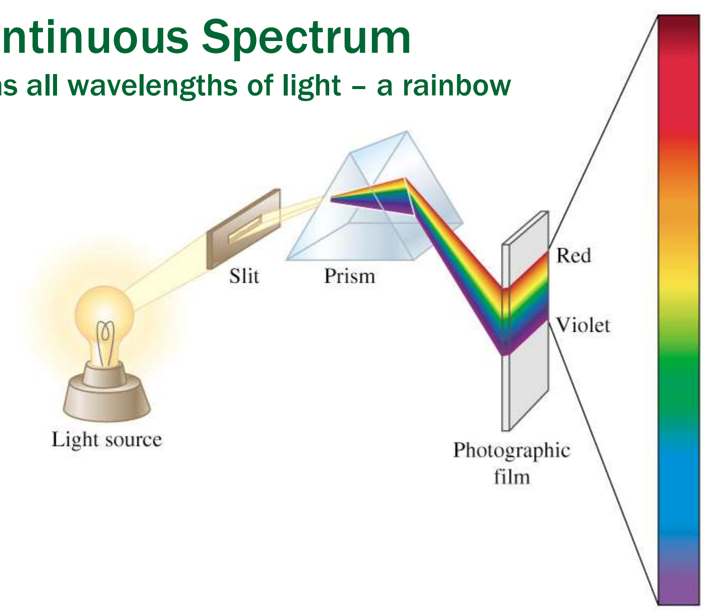
A Continuous Spectrum

contains all wavelengths of light – a rainbow



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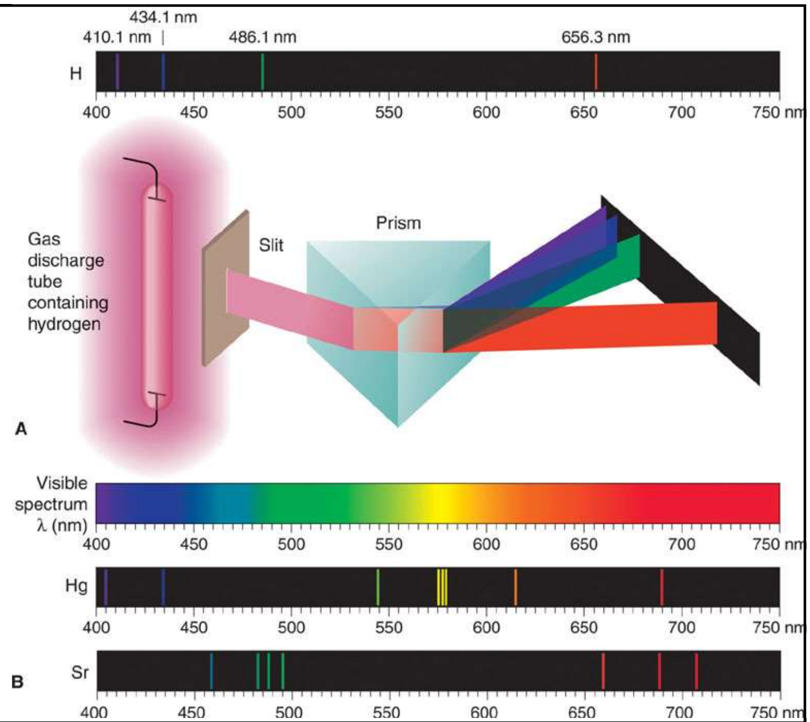
Line Spectra



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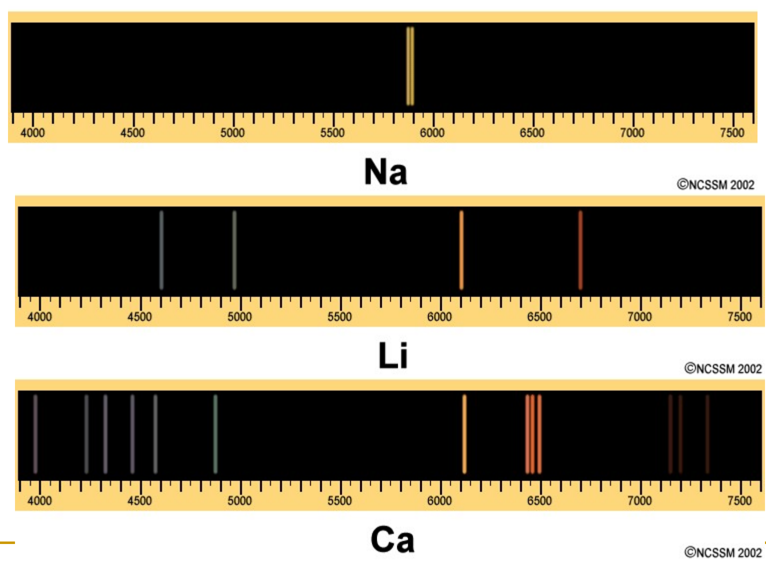
Emission Spectra



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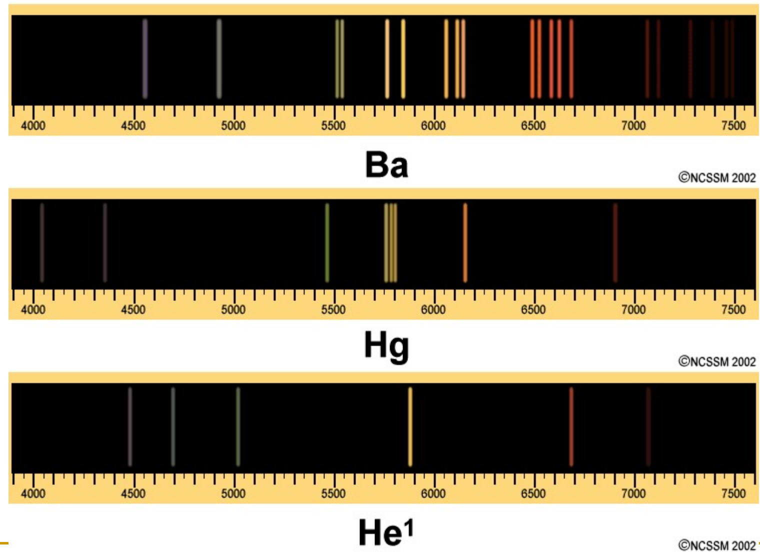
Emission Spectra



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Absorption Spectra

- An *absorption spectrum* occurs when light passes through a cold, dilute gas and atoms in the gas absorb at characteristic frequencies; since the re-emitted light is unlikely to be emitted in the same direction as the absorbed photon, this gives rise to dark lines (absence of light) in the spectrum.

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Emission and Absorption Spectra for Hydrogen

http://dosxx.colorado.edu/~bagenal/1010/SESSIONS/13.Light.html

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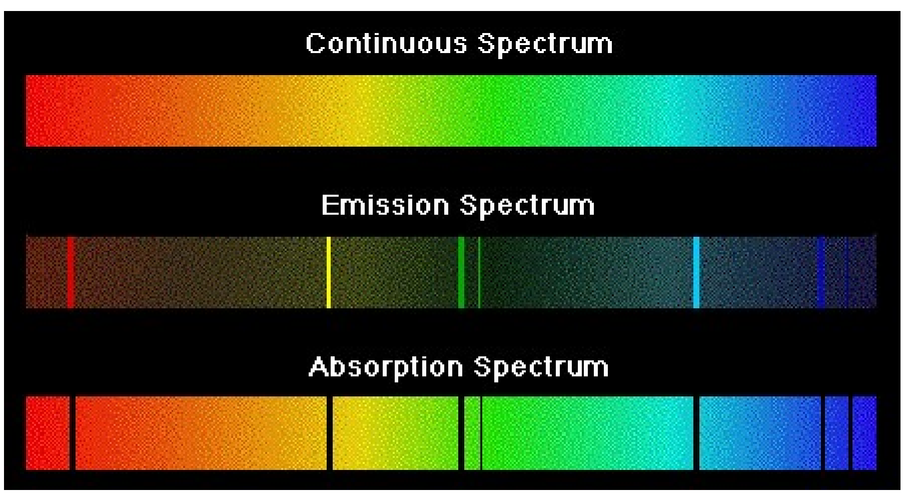

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Bohr's Explanation of Line Spectra

- Neils Bohr developed a mathematical model that could explain the observation of atomic line spectra.
- He proposed that electrons orbit the nucleus in certain "allowed" orbits - or *energy levels*. That is, the electron's energy is **QUANTIZED** (not continuous).
- Working on a model of <u>single-electron atoms</u>, Bohr derived an equation to calculate the energy of an electron in the **n**th orbit of such an atom:

$$E_n = -2.18 \times 10^{-18} \left(\frac{Z^2}{n^2} \right)$$

where Z = atomic number

and n = electron energy level

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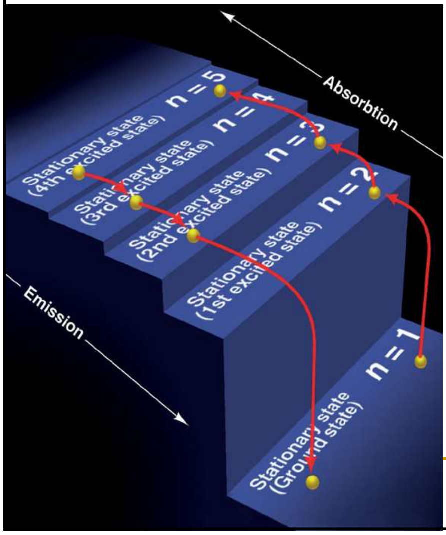
Bohr's Quantum Model for Hydrogen



- The electron in hydrogen occupies discrete energy levels.
- The atom does not radiate energy when the electron is in an energy level.
- When an electron falls to a lower energy level, a quantum of radiation is released with energy equal to the difference between energy levels.
- An electron can jump to higher energy levels if the atom absorbs a quantum of radiation with sufficient energy.

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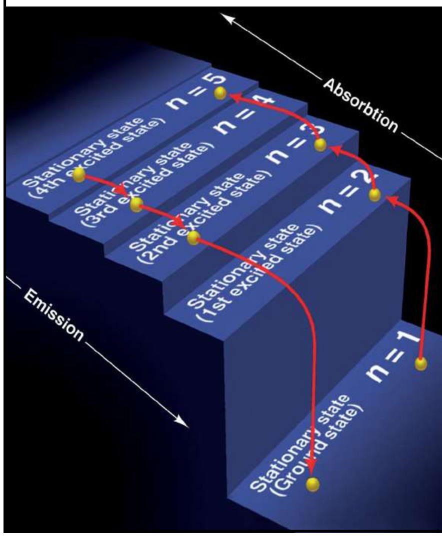
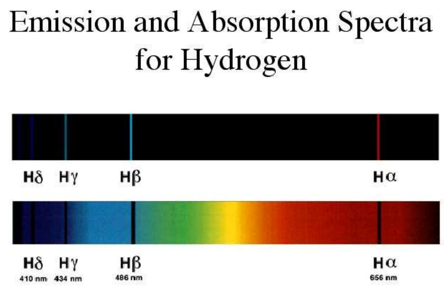
Bohr's Quantum Model for Hydrogen





http://dosxx.colorado.edu/~bagenal/1010/SESSIONS/13.Light.html

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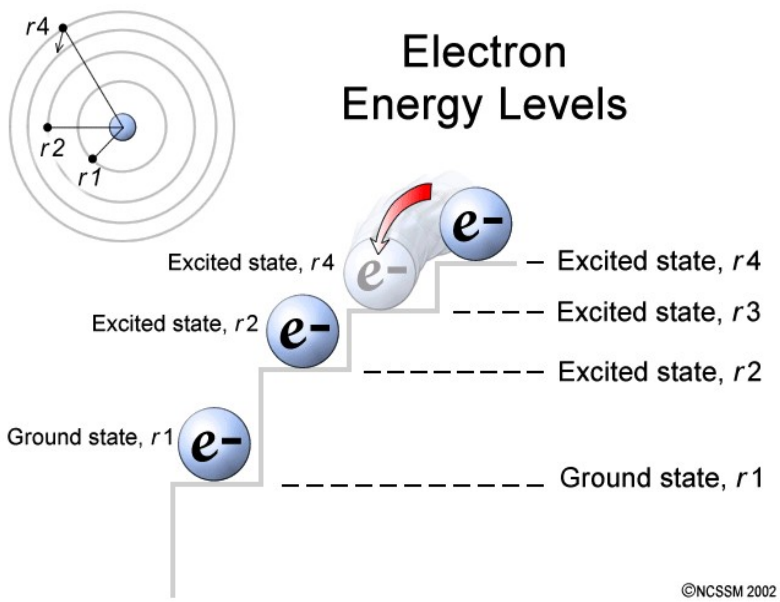

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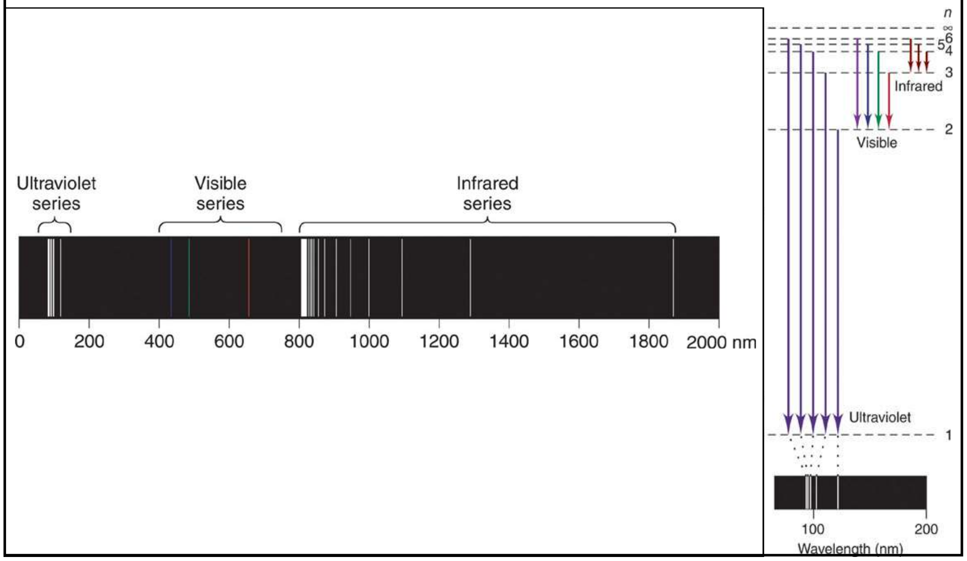
The Hydrogen Line Spectrum



What's Wrong with Bohr's Model??

- Although it works great for single-electron atoms, Bohr's model fails for atoms with 2 or more electrons!
- It was a huge leap forward, but was fundamentally flawed.
- Ultimately, the failure of Bohr's model lay in the fact that he treated the electron as a charged particle orbiting the nucleus like a planet around the sun.
- Electrons are more complicated …

Wave-Particle Dual Nature of Electrons

- Einstein's famous equation, $E = mc^2$, suggests that energy and mass are related – one can convert matter directly into energy.
- Louis de Broglie made the leap that if light can behave as wave/particles, then so can matter!
- He showed that the wavelength of a baseball is negligible (as expected), but that the wavelength of an electron was on the order of magnitude equal to that of electromagnetic radiation!
- This is what Bohr had missed – an atomic model must make use of the wave-nature of electrons to be complete!

video video

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Heisenberg's "Uncertainty Principle"

The more precisely the position is determined, the less precisely the momentum is known in this instant, and vice versa.

- It is impossible to determine BOTH the position and momentum (velocity) of an electron - the act of observing the electron changes it.
- The electron exists in nature as a wave-particle (whatever that *is*) until an experimenter observes it and imposes one of the two natures upon it!

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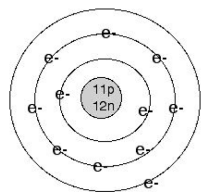
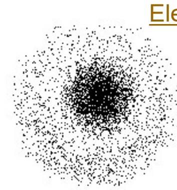
Heisenburg proposed that this inability to perceive reality on an atomic level is not the result of technological ignorance, but rather a fact of the laws of nature. In other words, Heisenburg believed that no matter how technologically advanced we become we would never be able to fully perceive and record atomic reality. This means that an electron’s orbital path around the nucleus of an atom can only be statistically approximated.

Electrons video



Bohr’s Atomic Model



Quantum Atomic Model

This is why the quantum atomic model has a cloud of electrons, not a few electrons following exact elliptical orbits as in Bohr's atomic model. Each electron in the quantum atomic model is a calculated probability, so the cloud describes not the exact orbit of millions of electrons, but rather the probability of the position of only a few electrons at any given time.

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The Photoelectric Effect

- Albert Einstein proposed the idea that light can be described as quanta of energy that behave as particles.
- Light quanta are called photons.
- Electrons called photoelectrons are ejected by metals when light shines on them.
- Alkali metals are affected the most.

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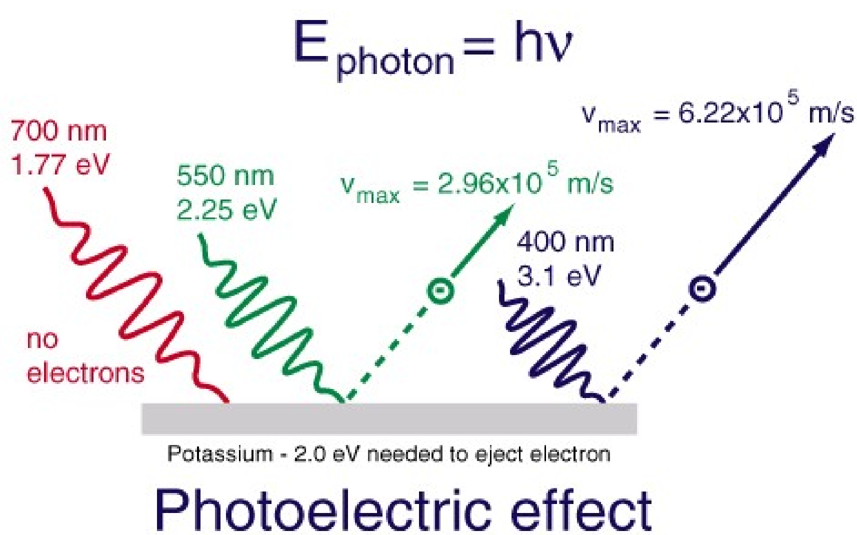
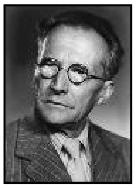
The Photoelectric Effect

- Not all light will cause this effect.
- If the frequency (energy) of the light is too low, photoelectrons will not be ejected.
- Increasing the light intensity causes more photoelectrons to be ejected but does not increase their speed.
- The more energy the light has, the faster the photoelectron travels as it leaves the metal surface.

The Photoelectric Effect



The Schrödinger Wave Equation



- Schrödinger applied de Broglie's concept of matter waves to the electron
- He explained the quantum energies of the electron orbits in the Bohr model of the atom as vibration frequencies of electron "matter waves" around the atom's nucleus.
- He provided a mathematical equation (wave function) to describe electron waves.
- Physicists had difficulty explaining the MEANING of the wave function itself, but it turns out that its square gives the probability of finding the electron in a particular region of space in the atom.

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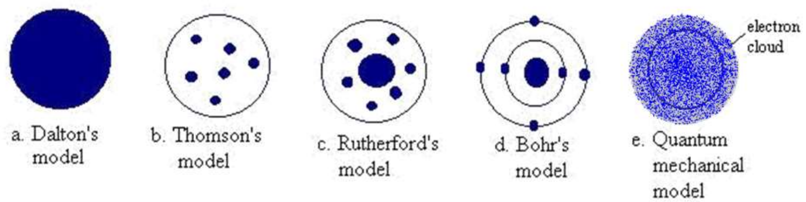
Schrodinger Equation, cont.

- These regions of probability are called **orbitals,** and this concept replaces that of the electron "orbit".
- Instead of electrons orbiting like planets around the sun, Schrodinger's picture shows an "electron cloud" surrounding the nucleus and doesn't state anything about the path or position of electrons within that cloud.
- Orbital "pictures" shown in texts represent regions of space where the probability of finding an electron is 90% or greater.



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Schrodinger Equation, cont.

Schrödinger's Equation

$$i\hbar \frac{\partial}{\partial t} \psi(\mathbf{r},t) = -\frac{\hbar^2}{2m} \nabla^2 \psi(\mathbf{r},t) + V(\mathbf{r},t)\psi(\mathbf{r},t)$$

i is the imaginary number, $\sqrt{-1}$.
$\hbar$ is Planck's constant divided by 2π: 1.05459×10^{-34} joule·second.
ψ ($\mathbf{r}$,t) is the wave function, defined over space and time.
m is the mass of the particle.
∇^2 is the Laplacian operator, $\frac{\partial^2}{\partial x^2} + \frac{\partial^2}{\partial y^2} + \frac{\partial^2}{\partial z^2}$.
V($\mathbf{r}$,t) is the potential energy influencing the particle.

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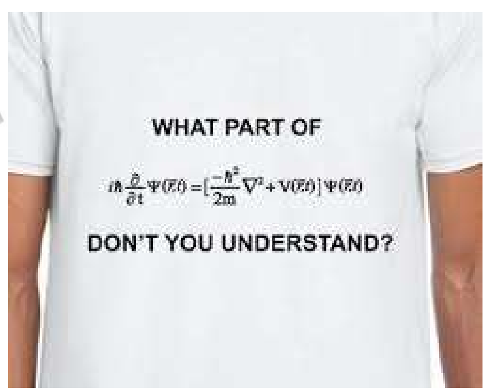
Schrodinger Equation, cont.



http://www.cafepress.ca/+schrodinger_equation_tshirt_ash_grey,1302907

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Schrodinger Equation, cont.

http://talklikeaphysicist.com/2008/quantum-physics-tattoos-schrodinger-equation/

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Schrodinger Equation, cont.

http://www.zazzle.ca/geek_cap_schrodinger_equation_sisal_embroidered_hat-233705118883062399

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Schrodinger’s Cat Videos:

Schrodinger's Cat - Minute Physics

Schrodinger’s Cat – Sixty Symbols

Schrodinger’s Cat – The Big Bang Theory

Schrodinger’s Cat – **IDTIMWYTIM**

Quantum Theory Documentary

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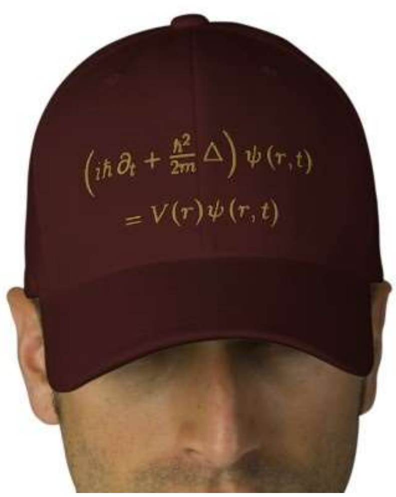
Schrodinger Equation, cont.



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Quantum Numbers

- Schrodinger's wave equation used three constants ("quantum numbers") that were used to describe orbitals – regions of space where an electron has a probability of being found.
- His equation used 3 **quantum numbers** to describe the size & energy, shape, and orientations of the orbitals in atoms.
- The three quantum numbers: n, l and m_l .
- This set of three numbers provides us with an "orbital address" for an electron within an atom – we can describe the orbital in which an electron is found if we know these three quantum numbers.

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Principal Quantum Number, n

- n can have values $\geq$ 1, integers only.
- For element atoms on the periodic table in their **ground states**, $1 \leq n \leq 7$
- As n increases, energy of electron increases and size of orbital increases
- Larger n means greater probability of greater distance from the nucleus.
- Electrons with same value of n are said to be in the same "electron shell"

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Principal Quantum Number, n

Contain electrons that are

- Close in energy
- Similar distance from nucleus
- Have values of n = 1, 2, 3, 4, 5, 6…..
- Maximum number of electrons = $2n^2$

n =1	$2(1)^2 = 2$
n =2	$2(2)^2 = 8$
n= 3	$2(3)^2 = 18$

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Shape Quantum Number, ℓ

- The values for ℓ (lower case letter L, script) are limited by the principal energy level – they range from 0 to "n–1"
- Example: If n = 3, ℓ can be 0, 1 or 2
- ℓ gives information about shape of orbital

$\ell = 0$	s orbital	(found on all principle energy levels)
$\ell = 1$	p orbitals	(found on level 2 and higher)
$\ell = 2$	d orbitals	(found on level 3 and higher)
$\ell = 3$	f orbitals	(found on level 4 and higher)

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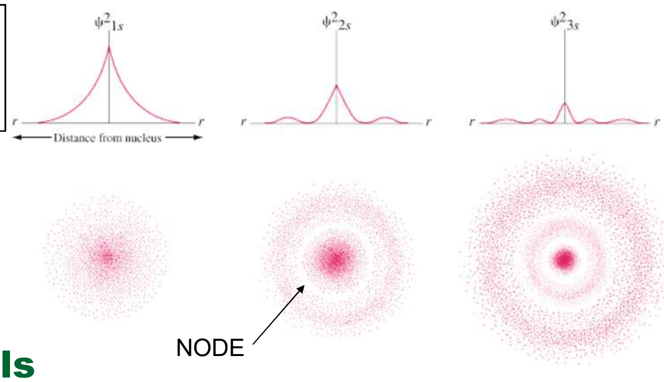
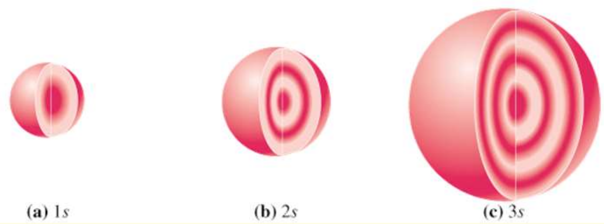
Graphs show Probability vs Distance from nucleus



s Orbitals
on three energy levels



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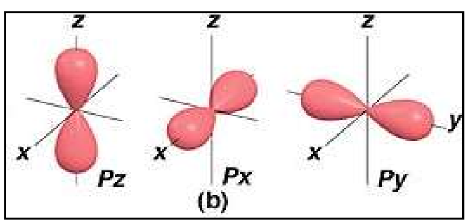
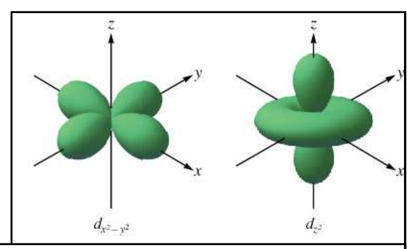
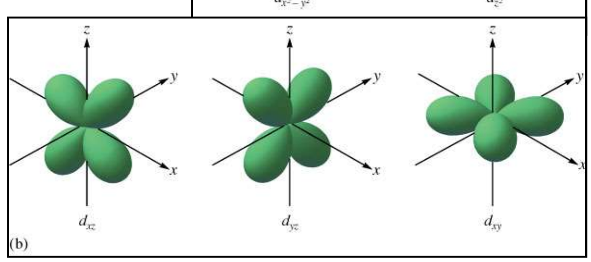
p and d Orbital shapes







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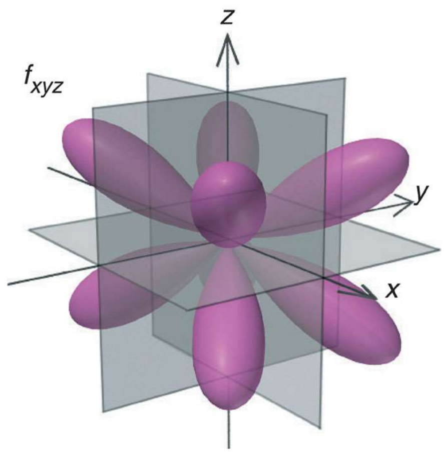

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Orientation Quantum Number, m_ℓ

- The values of m_l are limited by the value of l for a given orbital. They include positive and negative integers from $-\ell$ through $+\ell$, inclusive
- For a p orbital, ℓ = 1 , so m_ℓ can be: –1, 0, 1
- m_ℓ gives information about **orientation** of orbitals in space
- The # of allowed values gives the number of orbitals of this type on a given energy level. (3, in the case of p orbitals)

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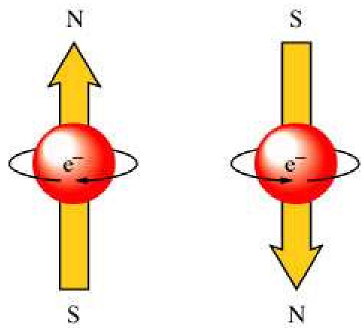
Electron Spin Quantum Number, m_s



A spinning negative charge creates a magnetic field. The direction of spin d determines the direction of the field.

- A fourth quantum number, m_s describes **electron spin** (either +½ or –½)
- **Each electron in atom has a unique set of these four quantum numbers.**
- Electrons in orbitals with same n and l values are said to be in the same **subshell.**
- Electrons with all three numbers the same, n, l , and m_l , are in the same orbital.

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Spin and Magnetic Fields

- Earth spins on its axis in a counter-clockwise rotation.
- The magnetic field generated by this direction of rotation yields a magnetic field that is the opposite of what most people expect . . .

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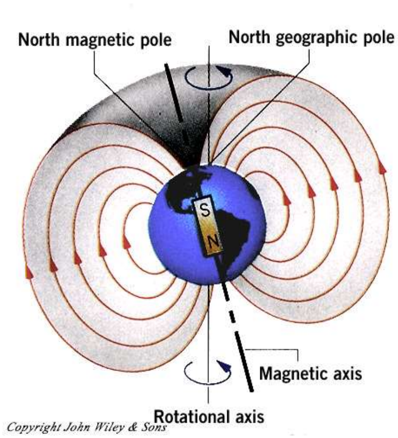

- Scientists believe that the magnetic field of the earth is oriented with its magnetic North Pole in the Southern Hemisphere.
- The south pole of this magnet is located in the northern Hemisphere!

Pauli Exclusion Principle

- Electrons have negative charge and repel each other. How are the electrons in an atom distributed?
- Wolfgang Pauli proposed that **no two electrons in a given atom can be described by the same four quantum numbers!**
- The first three quantum numbers determine an orbital – therefore spins must be opposite!
- **Practical result** is that each orbital can hold a maximum of two electrons, with opposite spins.

Electron Configurations

- Electrons orbitals are defined by their quantum numbers, n, l and m_l.
- Each electron in an atom has a unique set of 4 quantum numbers.
- No two electrons can have the same "address", i.e., the same 4 quantum #'s
- Rules define how multiple electrons will be distributed among the possible energy levels

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Aufbau Principle

- In the ground state, the electrons occupy the lowest available energy levels. An atom is in an excited state if one or more electrons are in higher energy orbitals.
- In atoms with more than one electron the lower energy orbitals get filled by electrons first!
- This is the Aufbau principle, which is named after the German word which means "to build up".
- When one describes the locations of the electrons in an atom, start with the lowest energy electron and work up to the highest energy electron.

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Hund's Rule

- A set of orbitals is said to be "degenerate" if the orbitals possess the same energies. For example, all three "2p" orbitals on energy level 2 are degenerate. All five "3d" orbitals on energy level 3 are also degenerate.
- When filling a set of degenerate orbitals, Hund said that electrons should be left unpaired as long as possible so as to minimize electron-electron repulsions within the orbitals.
- When each degenerate orbital has one electron, electrons will then pair, spins opposed, until that sub-shell is filled.

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Electronic Configurations

- An electron configuration is a way of describing the locations of electrons within an atom.
- Usually written for the ground state of an atom: when its electrons occupy orbitals giving the lowest possible overall energy for the atom
- Each subshell is designated by "n" and the type of orbital. Such as: 1s 3p 4d
- The number of electrons in an orbital is shown as a superscript: $1s^2$ $3p^3$ $4d^9$
- How does one determine the order of filling of the orbitals in an atom?? Which orbitals get filled first??

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"Valence Electrons"

- Valence electrons are those electrons in the highest principal energy level within an atom - the s and p electrons in the outermost shell of an atom in its ground state.
- These are the electrons involved in forming bonds with other atoms during chemical reactions.
- Most common oxidation states (ion charge) for the element can be derived from valence electrons
- Completely filled, half filled and empty sub-shells have special stability (not sure why!?)

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Valence Configurations

- The "valence configuration" is that part of an electron configuration that describes the valence electrons.
- For example, sodium's valence configuration is just "$3s^1$".
- The valence configuration of bromine is "$4s^2 4d^5$". We leave out the "$3d^{10}$" electrons because they are not in the outermost principal energy level.

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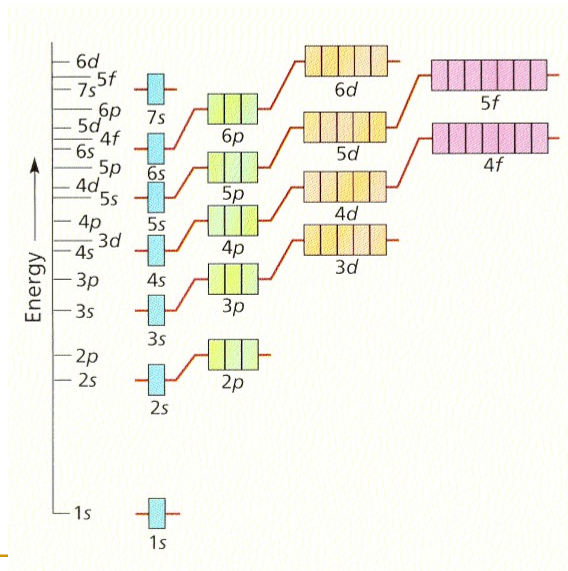
Orbital Energies in Multielectron Atoms



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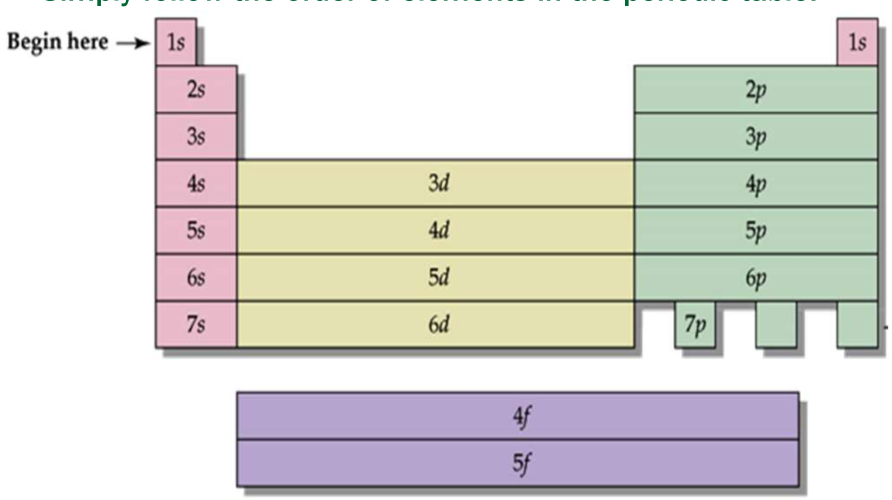
The Order of Filling of Orbitals

Simply follow the order of elements in the periodic table!



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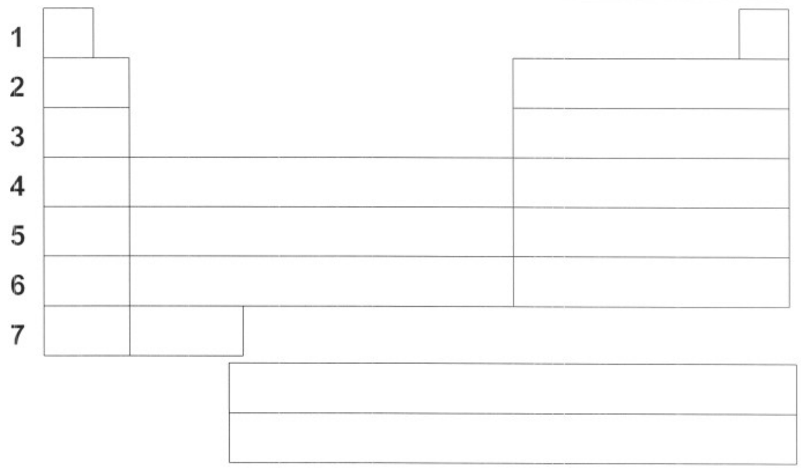
Sublevel Blocks



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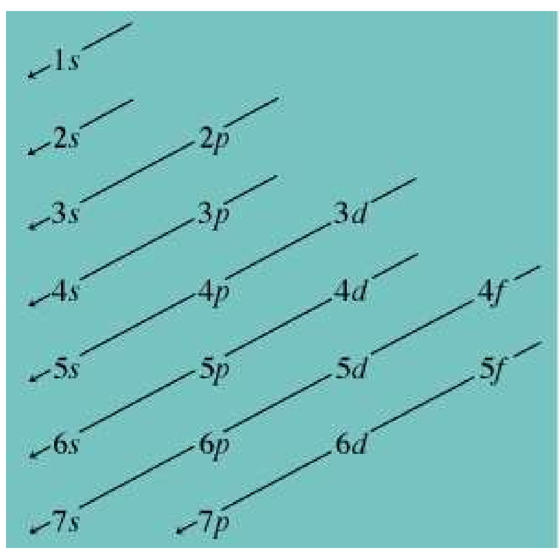
Another useful device...

Start at the top and follow the arrows:

1s,
2s,
2p, 3s,
3p, 4s,
3d, 4p, 5s…



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Examples of Electron Configurations

- Helium has 2 electrons in the 1s orbital, **He: $1s^2$**
- Carbon has 6 electrons, **C: $1s^2\ 2s^2\ 2p^2$**
- Calcium has 20 electrons, **Ca: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$**
- Calcium cation, **Ca^{2+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$**
- Sulfide anion, **S^{2-}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$**

- Notice that the superscripts add up to the total number of electrons on the atom or ion! Each of these examples represents the "ground state" of the atom/ion because the electrons are in the lowest possible energy levels.

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Quick Check

Indicate if each configuration is correct or incorrect for potassium. Explain why or why not?

A. $1s^2 2s^2 2p^6 3s^1$

B. $1s^2 2s^2 2p^6 3s^2 3p^6$

C. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ √

D. $1s^2 2p^8 3s^1$

E. $1s^2 2s^2 2p^6 3s^2 3p^7$

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For phosphorus, indicate if each configuration is correct or incorrect. Explain why or why not.

A. 2, 2, 8, 5

B. 2, 8, 3

C. 2, 8, 5 √

D. 2, 6, 7

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Using the periodic table, write the complete electronic configuration for each:

A. Cl

B. Sr

C. I

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Using the periodic table, write the complete electronic configuration for each:

A. Cl

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

B. Sr

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2$ OR

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

C. I

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^5$ OR

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

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Learning Check

A. The final two notations for Co are

1) $3p^64s^2$
2) $4s^24d^7$
3) $4s^23d^7$

B. The final three notations for Sn are

1) $5s^25p^24d^{10}$
2) $5s^24d^{10}5p^2$
3) $5s^25d^{10}5p^2$

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Solution

A. The final two notations for Co are

3) $4s^2\ 3d^7$

B. The final three notations for Sn are

2) $5s^2\ 4d^{10}\ 5p^2$

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Learning Check

A. Number of electrons in a p orbital

1) 1e 2) 1e or 2e 3) 3e

B. Number of orbitals in a p subshell

1) 1 2) 2 3) 3

C. Number of orbitals in 4d subshell

1) 1 2) 3 3) 5

D. Number of electrons (maximum) in a 3d orbital

1) 2e 2) 5e 3) 10e

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The Noble Gases

- The noble gases are noted for their chemical stability and existence as monatomic
  molecules.
- Except for helium, they share a common electron configuration that is very stable.
- This configuration has 8 valence-shell electrons.

	valence e^-
He	2
Ne	8
Ar	8
Kr	8
Xe	8
Rn	8

- When cations are formed, they take on the electron configuration resembling that of the nearest noble gas.
- The Na atom loses an electron to become an Na cation.
- Its electron configuration becomes the same as Ne.

The electron configuration for:

- Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$
- Na+: $1s^2$ **$2s^2\ 2p^6$**
- Ne: $1s^2$ **$2s^2\ 2p^6$** **Octet**

Magnesium atoms lose both valence electrons when they become cations. They take on the electron configuration of Ne, the closest noble gas:

- Mg: $1s^2\ 2s^2\ 2p^6\ 3s^2$
- Mg^{2+} : $1s^2$ **$2s^2\ 2p^6$**
- Ne: $1s^2$ **$2s^2\ 2p^6$** **Octet**

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Pseudo-Noble Gas Configurations

- Many ions do not have noble gas configurations.
- Their ions are exceptions to the octet rule.
- These atoms (most often transition metals) lose enough electrons to have a more stable electron configuration.
- By filling a *d* orbital, for example. . .

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This is a pseudo-noble gas configuration.

Example:

Ag: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^1\ 4d^{10}$

- Addition of 7 e- would be necessary to achieve the noble gas configuration of Xe.
- Loss of 11 e- would give Ag the configuration of Kr.
- Loss of the $5s^1$ e- gives 18 e- in the outer energy level, a more stable pseudo-noble gas configuration.

Learning Check . . .

- How many valence electrons does each of the following atoms have?

K	C	Mg	O
1	**4**	**2**	**6**

- Write the electron configurations for the unipositive ions of Cu and Au and the dipositive ions of Zn and Hg.

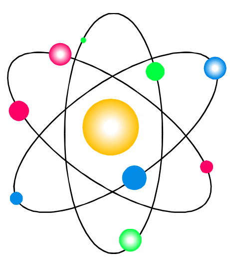
Cu^{+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$

Au^{+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

$5s^2\ 4d^{10}\ 5p^6\ 4f^{14}\ 5d^{10}$

Zn^{2+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$

Hg^{2+}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6 5s^2\ 4d^{10}$
$5p^6\ 4f^{14}\ 5d^{10}$



40S Chemistry

Lewis/Electron Dot Structures

Valence Electrons

- Determine the chemical properties of an element.
- The same for each element in a periodic table group.
- Usually the only electrons involved the formation of chemical bonds.
- Are best depicted using Lewis dot structures.

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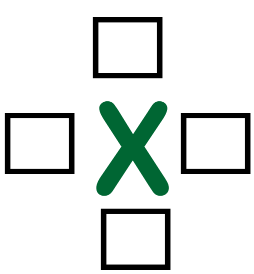
Lewis Symbols: Basic Rules



- Draw the atomic symbol.
- Treat each side as a box that can hold up to two electrons.
- Count the electrons in the valence shell.
- Start filling box - don't make pairs unless you need to.

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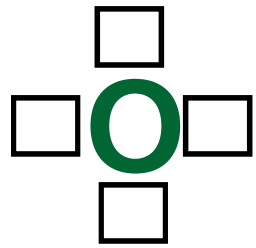
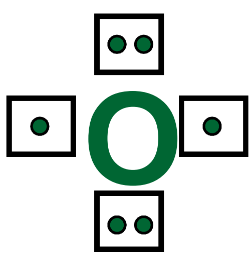
- Oxygen has 6 electrons in its valence - VIA.
- Start putting them in the boxes.

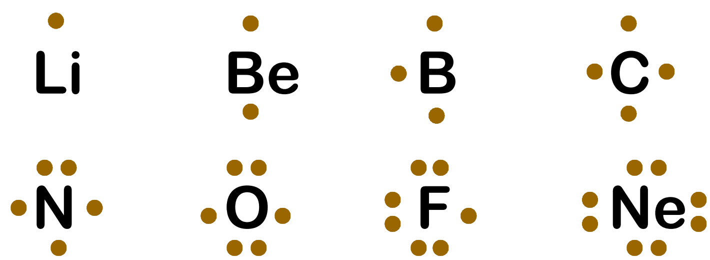




- The Lewis symbol for oxygen contains two non bonding electron pairs and two unpaired bonding electrons

This is the Lewis symbol for oxygen

Lewis Dot Structures of Second Period Elements



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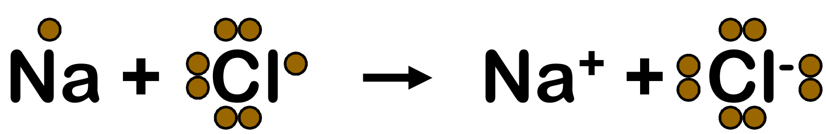
Lewis Dot Structres and the Formation of NaCl



The electron from Na moves over to the Cl. Now both satisfy the octet rule.

- Na becomes Na^{+} - a cation
- Cl becomes Cl^{-} - an anion

The + and - charges attract each other and form an ionic bond.

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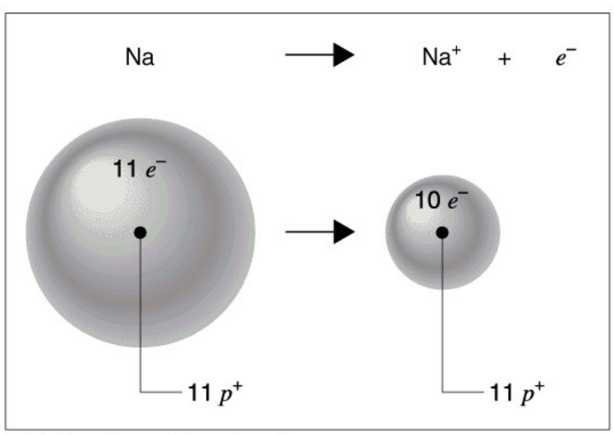
Stable Cation Electron Configuration

- It is the nature of matter to adjust to achieve the lowest possible energy.
- Noble gas electron configurations are stable and therefore not chemically reactive.
- Other elements have atoms that have more energy and unstable electron configurations.

- Ionic radius is smaller than atoms of the element.



- Compounds form to allow atoms to achieve the lowest possible energy.
- Atoms are most stable if they have a filled or empty outer layer of electrons.
- Except for H and He, a filled layer contains 8 electrons - an octet.
- Atoms will

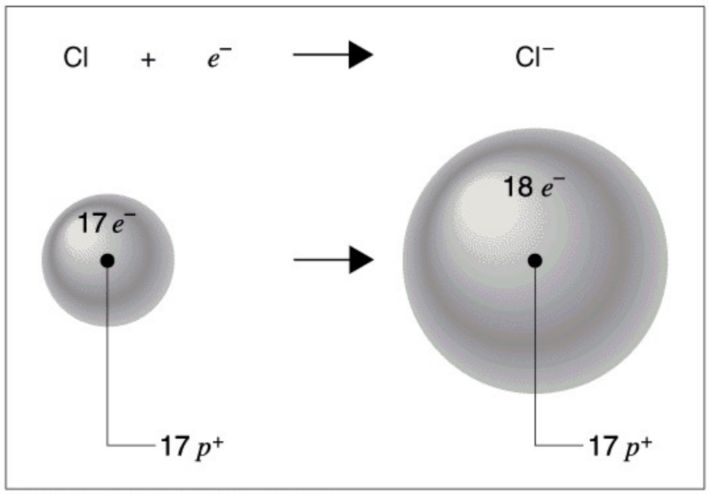
  gain or lose (ionic compounds)

  share (covalent compounds)

electrons to make a filled outer layer.

Stable Electron Configuration for Anions

- An ion or group of ions with a negative charge.
- Attain a stable electron configuration by gaining e-.
- Ionic Radius is larger than the atom of the same element.



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- Halide ions form when members of the halogen family gain e- to become anions.

$$Cl + 1\ e^- \longrightarrow Cl^-$$

- An oxide ion forms when an oxygen atom, with 6 valence e-, gains 2 e-.

$$O + 2\ e^- \longrightarrow O^{2-}$$

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Learning Check . . .

- How many e- will the following elements gain or lose in forming an ion?

Ca	F	Al	O
lose 2	**gain 1**	**lose 3**	**gain 2**

What will the ionic formula be when each of the following become ions?

S	Na	F	Ba
S^{2-}	Na^{+}	F^{-}	Ba^{2+}