The History of the Atom

How did we learn about the atom?

The Atomic Theory of Matter

- All matter is made up of fundamental particles.

- *What does "fundamental" mean?*

The Greek Philosophers, 400 B.C.

- Democritus coined the term, “atom”
- Greek word *atomos*, means “indivisible.”

- Atoms are the smallest particle (indivisible and indestructible).

- *Are atoms indivisible?*

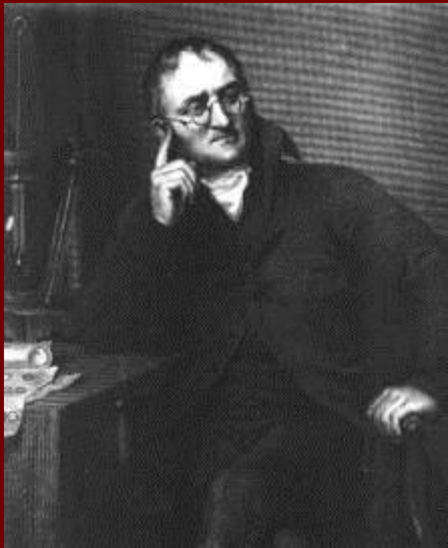
The 1st Scientific Theory, 1803

- **Dalton's Atomic Theory**
  1. All matter is made up of atoms.
  2. Atoms are indivisible and indestructible.
  3. A) Atoms of the same element are identical.

     B) Atoms of one element are different from atoms of another element.
  4. Atoms can physically mix together or can chemically combine in whole-number ratios to form compounds.
  5. Chemical reactions occur when atoms are separated, joined, or rearranged.

Still True Today?

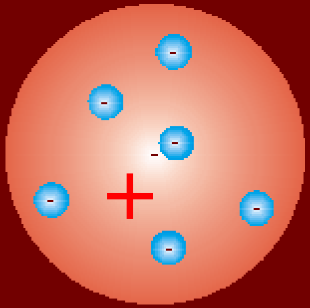
1. Yes
2. No: subatomic particles
3. A) No: isotopes
   B) Yes: differ by # protons
4. Yes
5. Yes



1st Discovery of Subatomic Particles, 1897

- J. J. Thomson
  - Discovered the electron and the proton.

Plum Pudding Model of the Atom: Tiny particles of negative charge embedded in a ball of positive charge.



How did Thomson discover the electron?

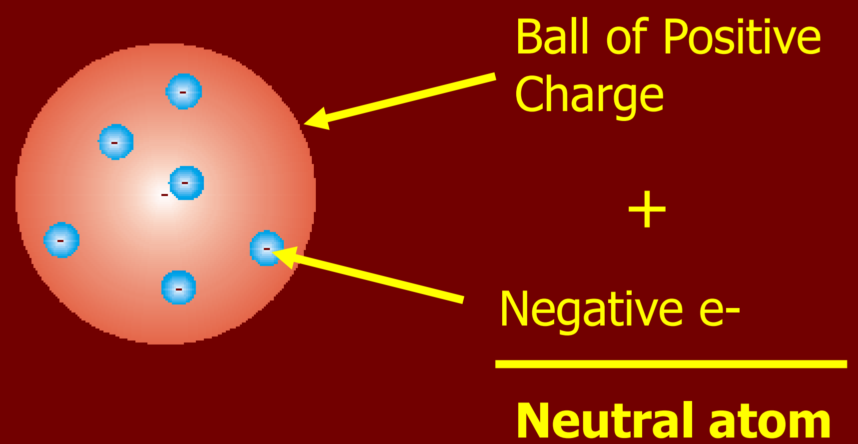
- Passed an electric current through gases in a cathode-ray tube producing a glowing beam.

- *If electrically charged plates are placed near the cathode tube, the cathode ray will be attracted towards the positive plate and repelled from the negative plate.*
- *Therefore, the ray must be made of negatively-charged particles.*

How did Thomson know that there were protons?

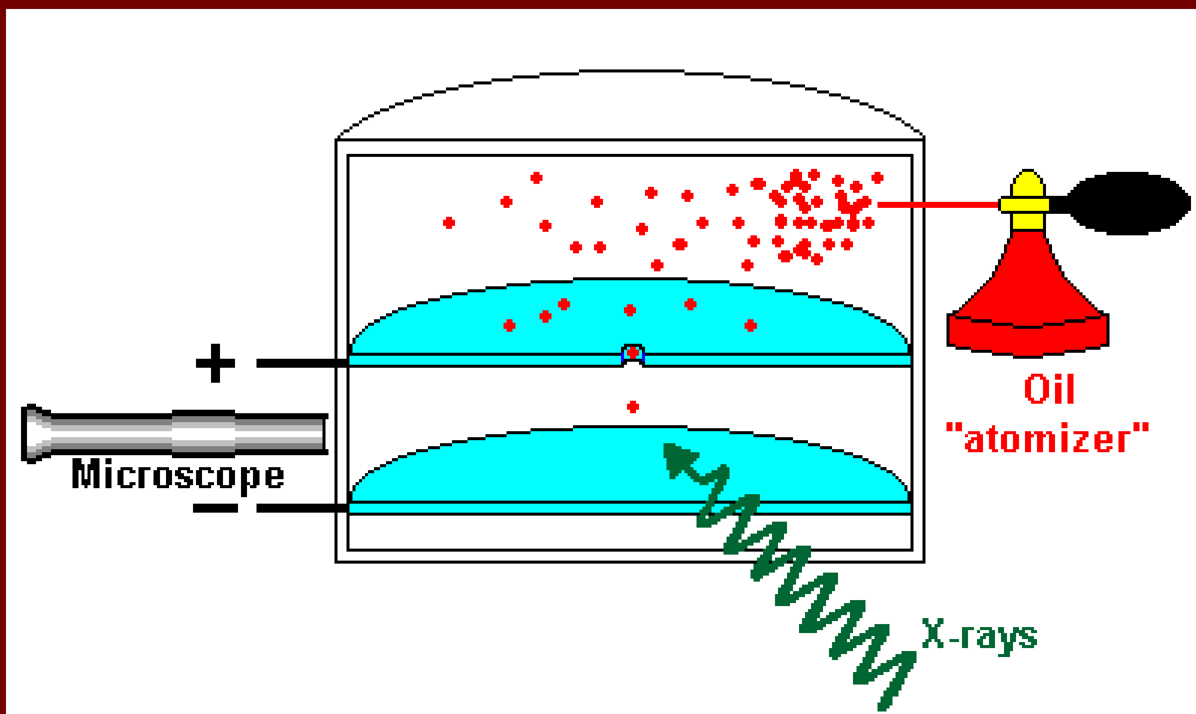
- Thomson inferred the existence of positively charged particles since matter is neutral!

  You do not get a shock every time you touch matter!



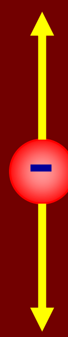
Millikan's Oil Drop Experiment, 1916

- Found the mass and charge of the electron.



Force from the negatively charged plate.

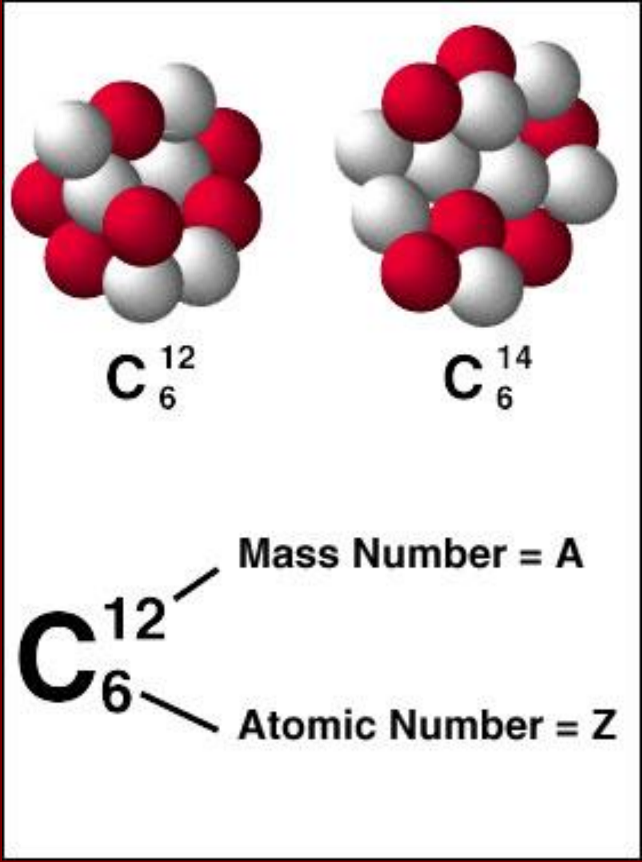
(up)



Force of Gravity (down)

Here’s what Millikan’s Apparatus really looked like…

Chadwick, 1932



- Discovered the <u>neutron</u>.
- <u>Isotopes</u>: Atoms of the same element that have different masses.
  - Same number of protons and electrons, different number of neutrons!
  - Ex. Carbon-14 and Carbon-12

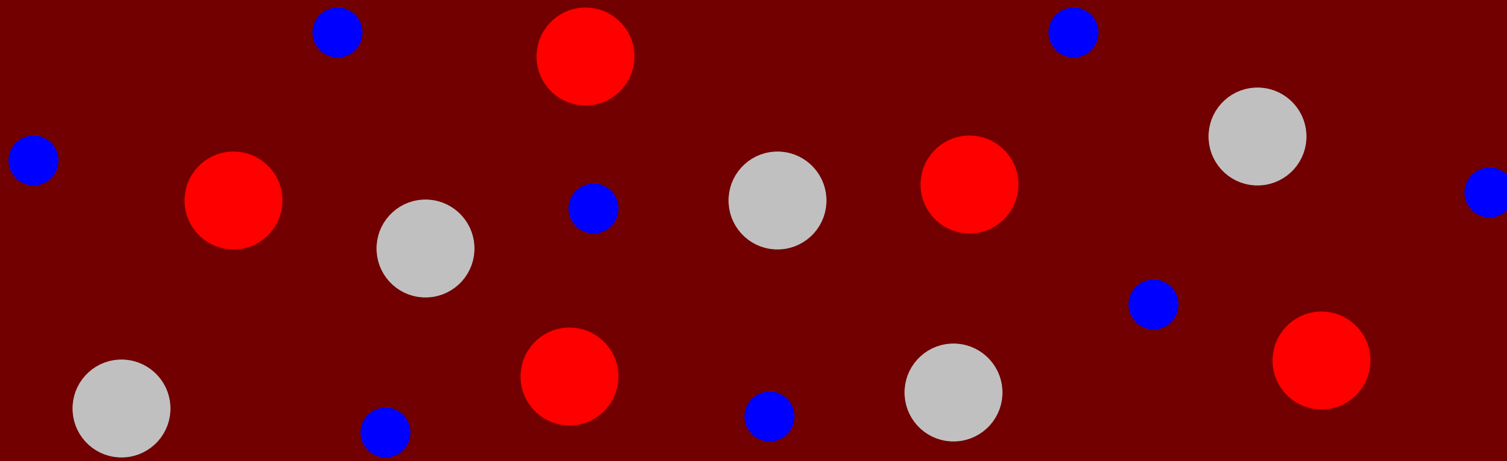
Subatomic Particles

Particle	Symbol	Relative Charge	Relative Mass	Actual Mass (g)
Proton	p^+			
Neutron	n^o			
Electron	e^-			

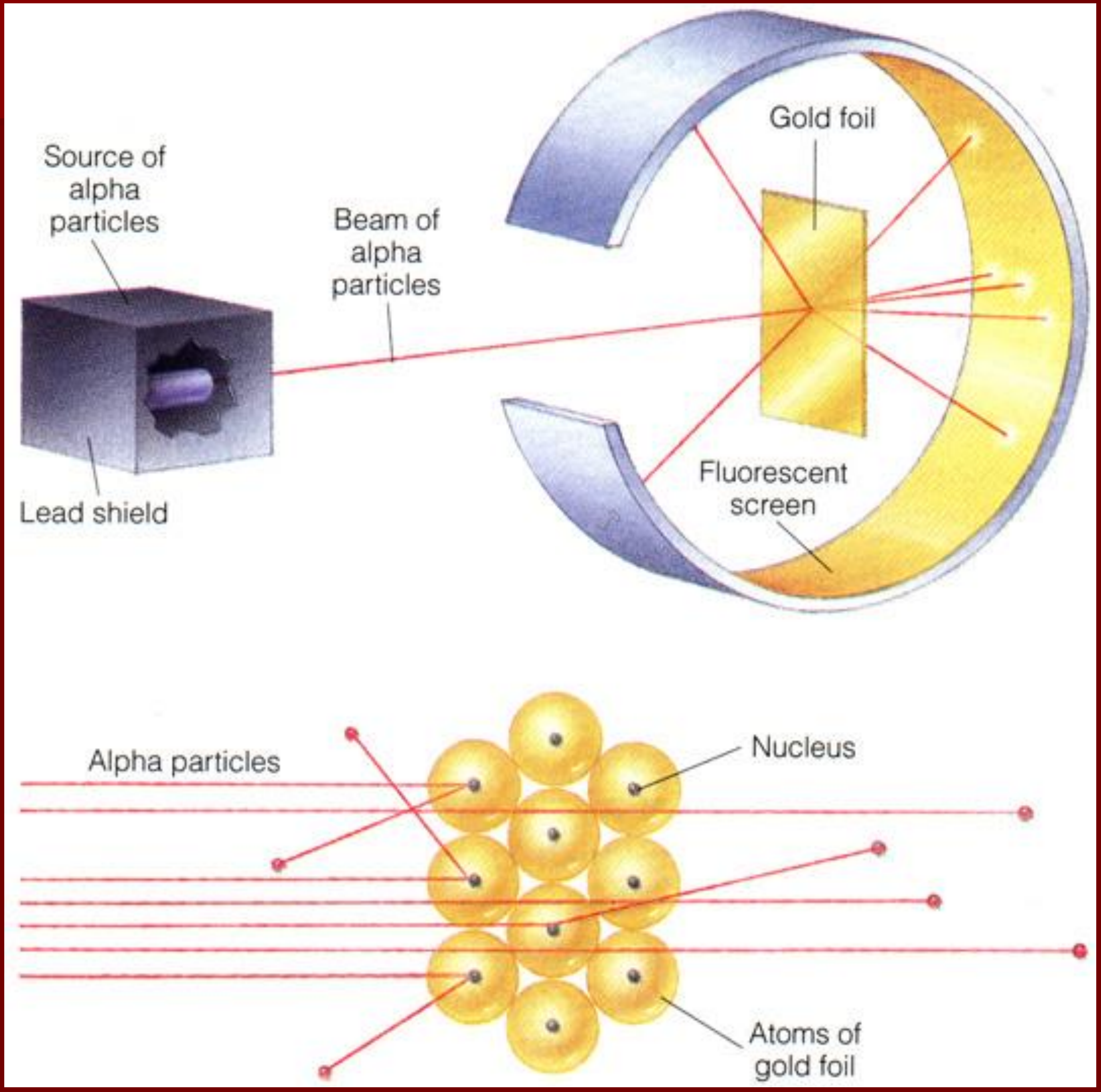
Which particle(s) make up most of the mass of the atom?

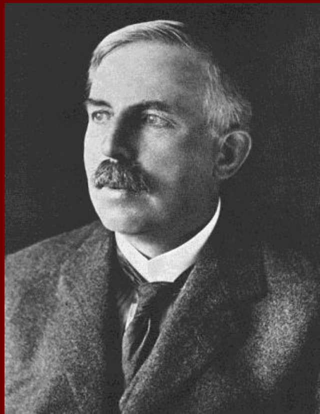
Okay, so now we have subatomic particles…

How are the particles arranged in the atom???

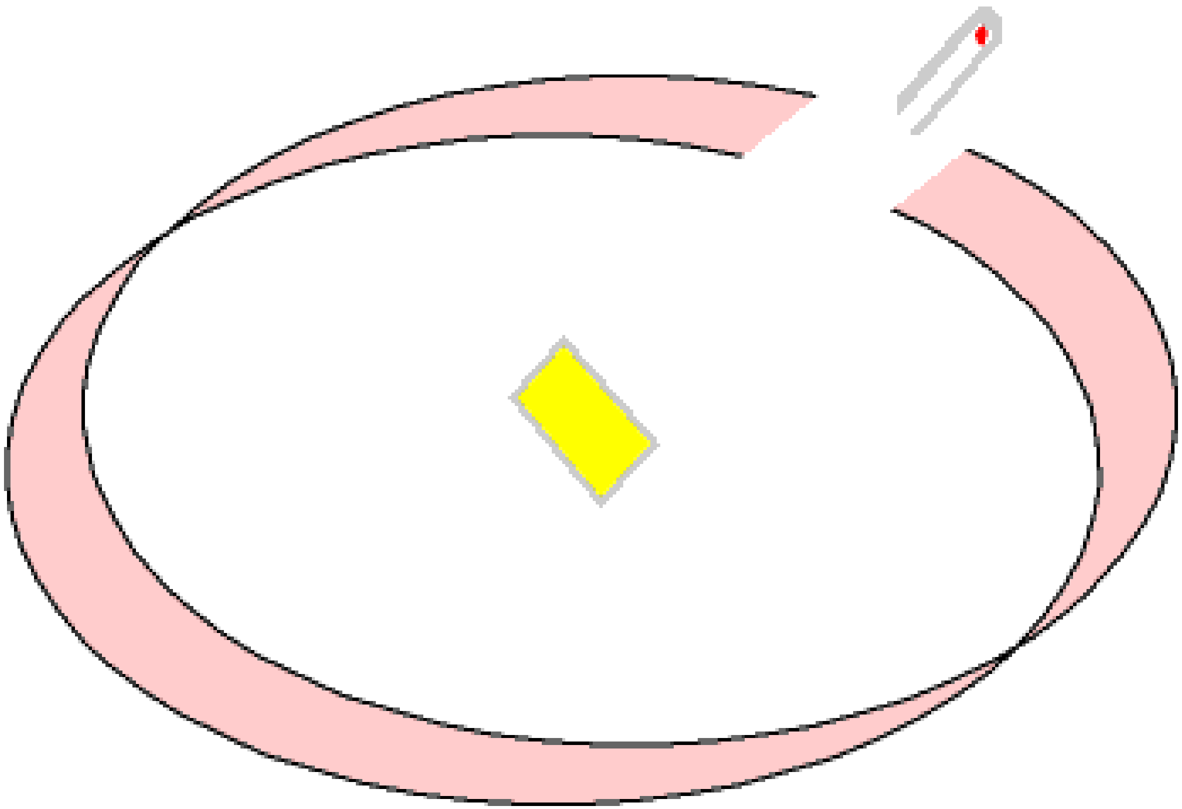


Rutherford's Gold Foil Experiment, 1911



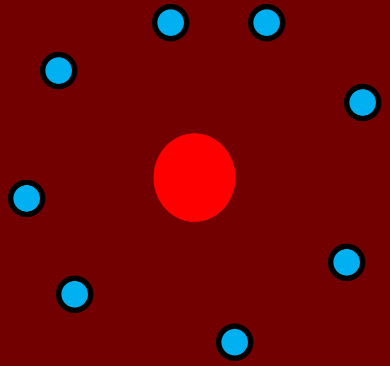


The Rutherford gold foil experiment used an alpha particle gun.



Here's what you need to remember:

- Rutherford shot alpha particles (2+ charge) at gold foil.

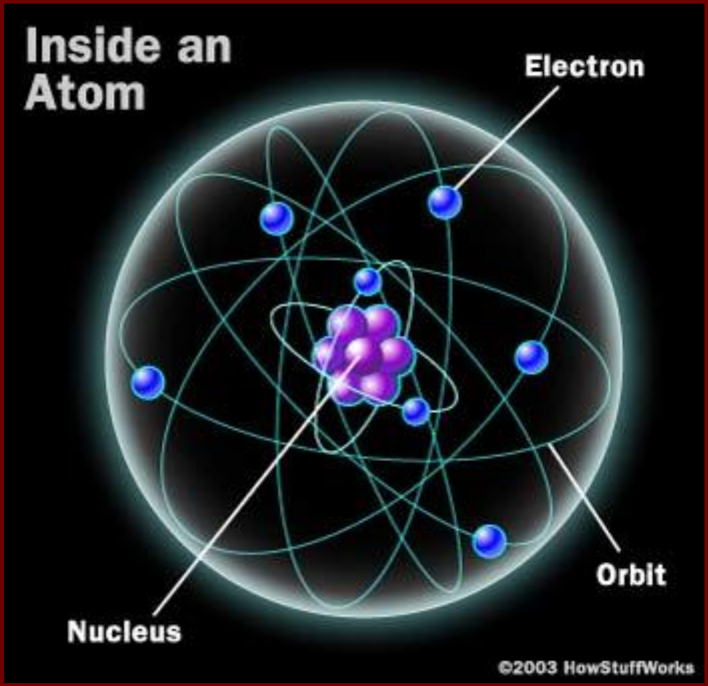


1) Most of the alpha particles went straight through…
   => Atoms are mostly empty space.
2) A few bounced back…
   => Small, dense positively-charged nucleus.

- Nucleus: small, dense core containing protons and neutrons.
- Electrons surround the nucleus.

Bohr's Planetary Model, 1913

- Electrons orbit the nucleus like planets orbit the sun.



Bohr: Electrons in Orbits

- *Bohr studied how atoms emit or absorb light.*

- Given more energy, e- will move to an orbit farther away from the nucleus.

- Electrons can occupy only certain orbits (energy levels)

Energy Levels

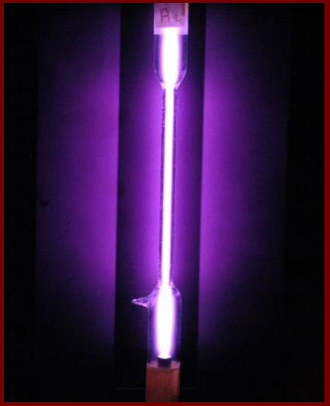
- Regions of space in which e- can move about the nucleus of an atom.

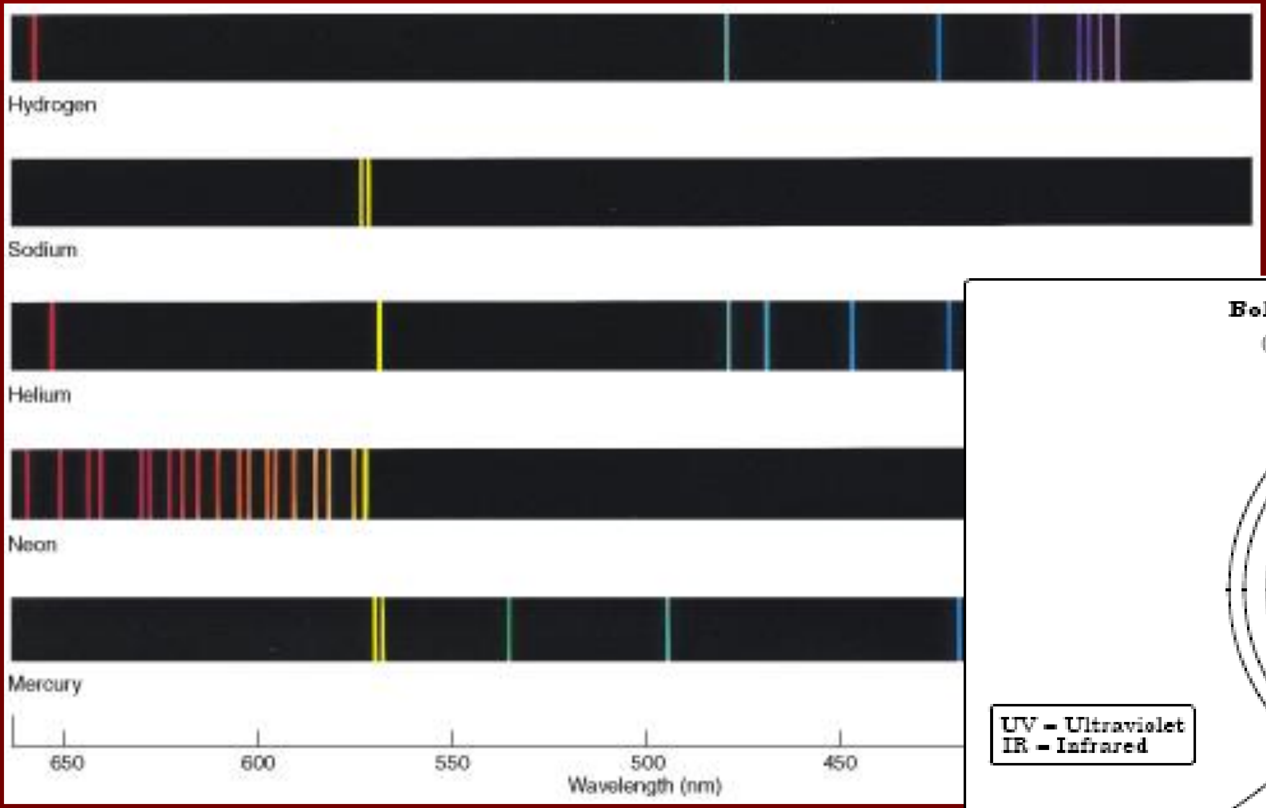
- Evidence for Energy Levels: Emission Spectrum of Atoms!

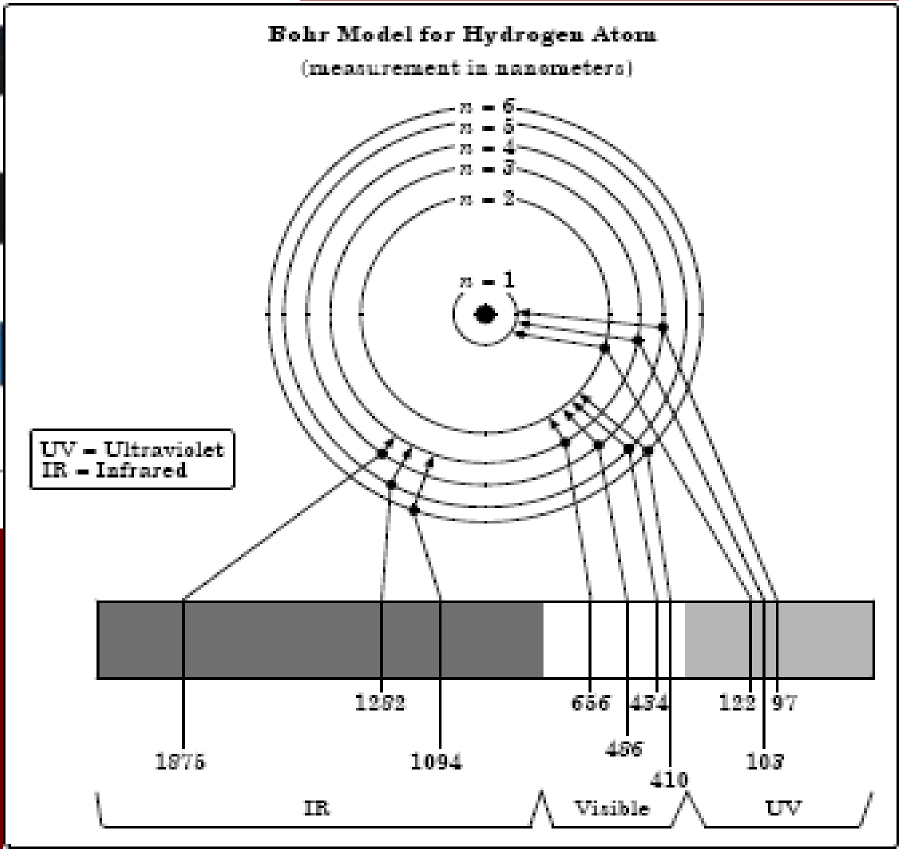
(details later…)

Emission Spectra

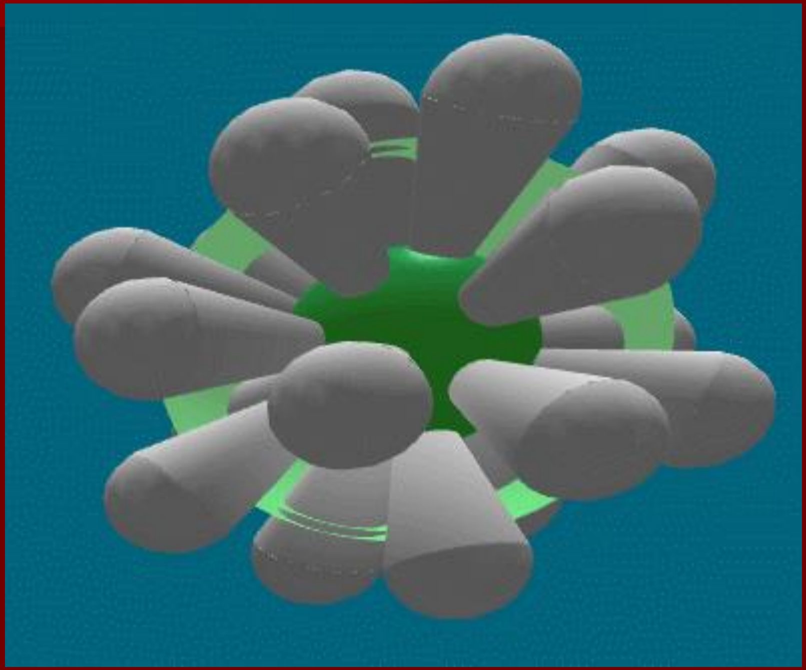
(Argon)

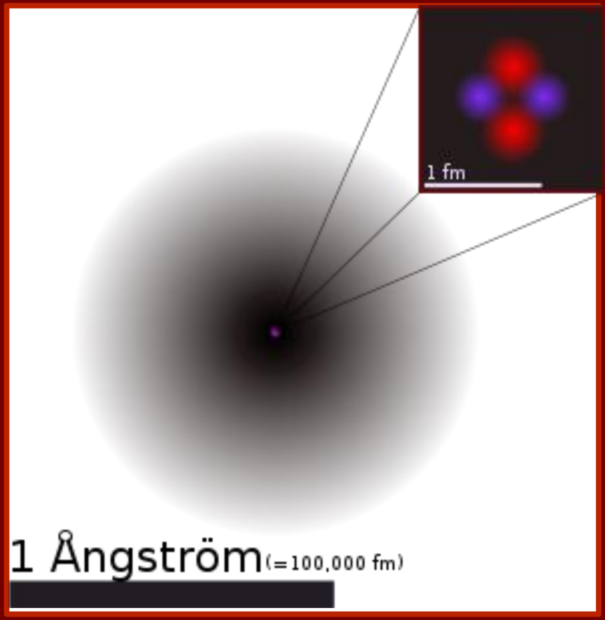






Electron Cloud Model - Schrodinger, 1926





a.k.a. Quantum Mechanical Model, Wave Model

Why is it called the electron cloud model?

- Electrons do not orbit.
- Electrons can only have certain energies.
- Cannot pinpoint the exact location of electrons (only probability).
- Probability of finding an e- is represented by an electron cloud.

- Heisenberg's Uncertainty Principle: It is impossible to know the velocity and the position of a particle at the same time.

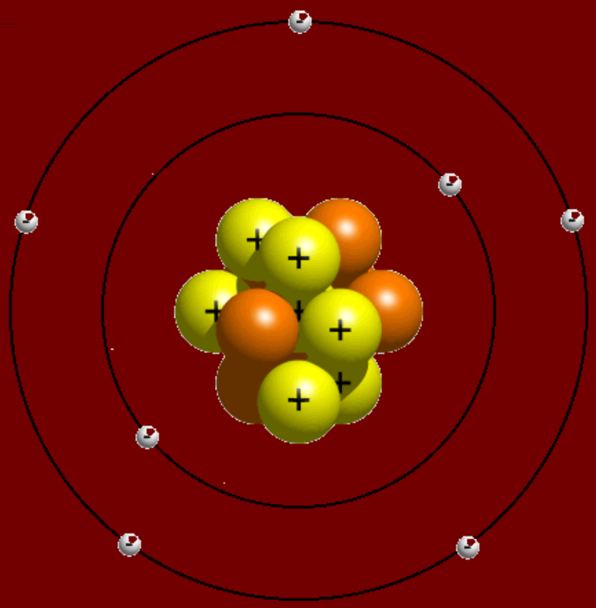
DISTINGUISHING BETWEEN ATOMS

Atomic Number

- # of protons
- Each element has a different number of protons.
- In a neutral <u>atom</u>,
  # of protons = # of electrons
- *Where is it found on the periodic table?*

Mass Number

- # of protons + # of neutrons



- Most of the mass of the atom is concentrated in the nucleus.

To find the number of neutrons in an atom…

- # neutrons = Mass Number – Atomic Number
- Examples: $^{9}_{4}Be$ $^{2}_{1}H$

Be-9 H-2

How many protons and neutrons are found in the nucleus of these atoms?

1. $^{23}_{11}\text{Na}$
2. $^{32}_{16}\text{S}$
3. $^{108}_{47}\text{Ag}$
4. $^{207}_{82}\text{Pb}$
5. Carbon-14
6. Fluorine-19
7. Chromium-52

Atomic Structure Worksheet

- *Atomic Mass – on the periodic table*
  - *If you are not given the mass number, round the atomic mass to a whole number (to get the most common isotope).*
- # of Electrons
  - No charge (neutral): $\#e^- = \#p^+$
  - Negative charge: gained electrons
  - Positive charge: lost electrons

Isotopes

- Atoms that have the same number of protons but a different number of neutrons.
- Isotopes have different mass numbers because they have a different # of neutrons.

Isotope Examples

- Example: Ne-20, Ne-21, Ne-22
  (ALL neon atoms have 10 protons and 10 electrons)

- Example:
  hydrogen-1 (hydrogen),
  hydrogen-2 (deuterium),
  hydrogen-3 (tritium)

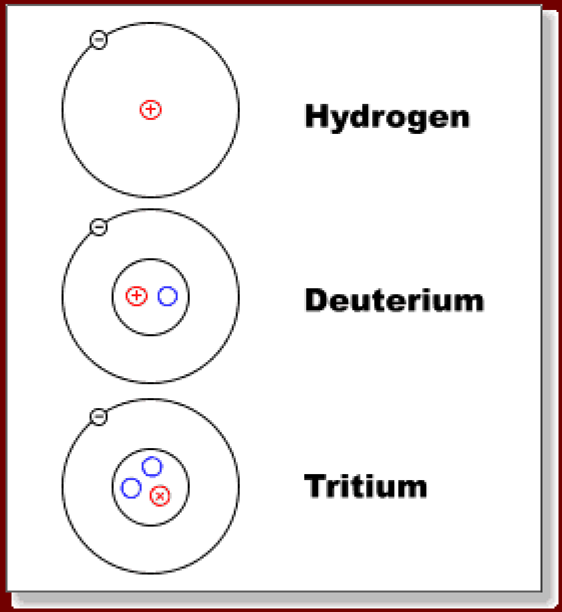
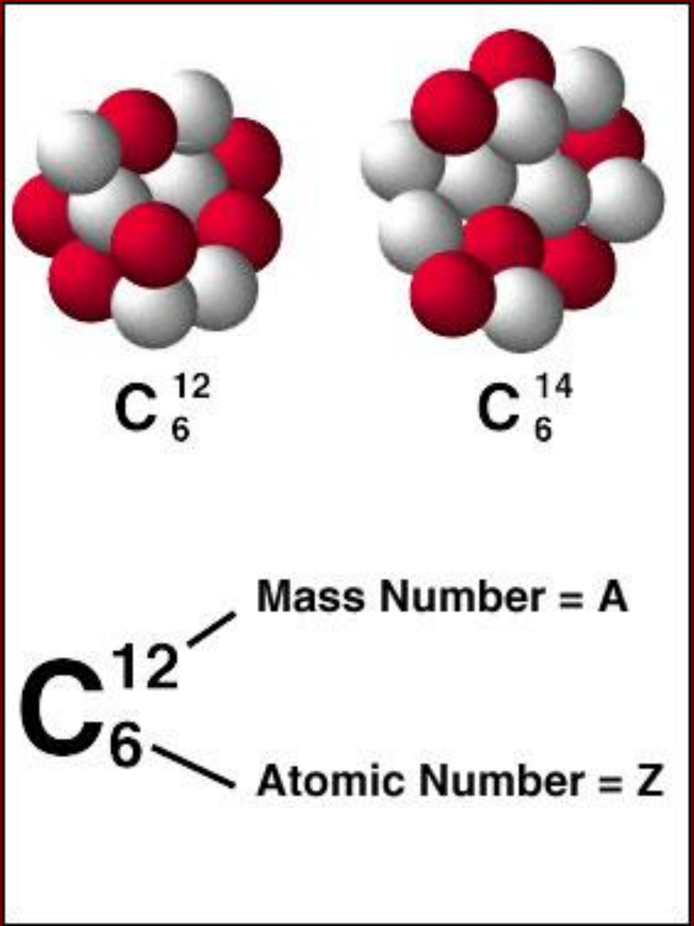
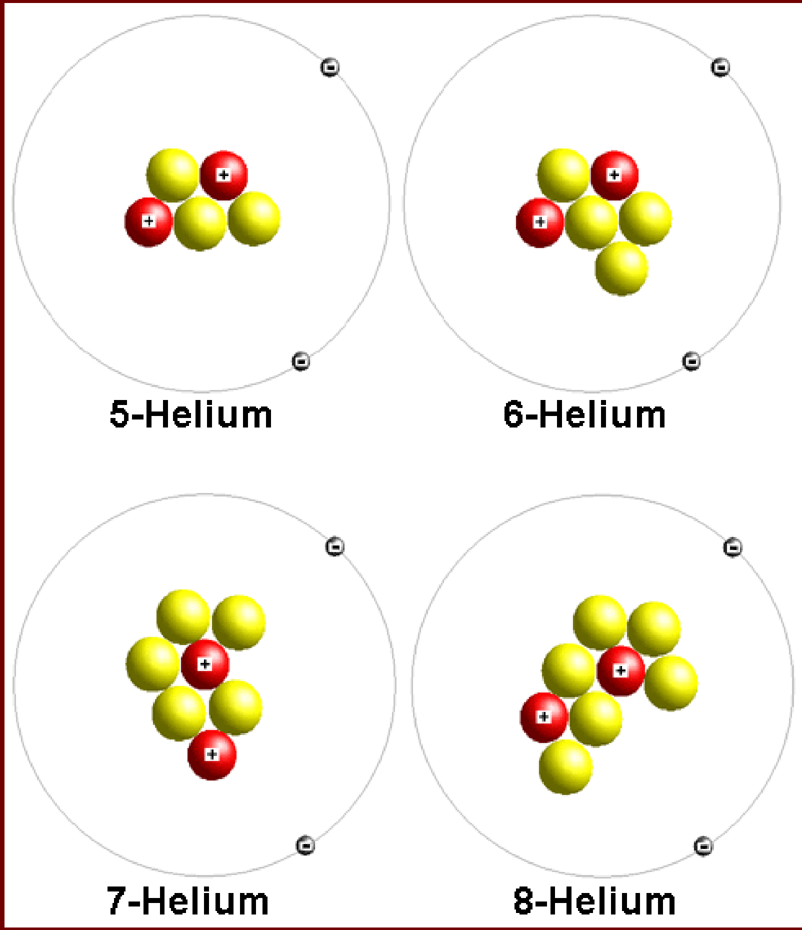


Figure 3

Isotope Examples





Atomic Mass

- Atomic Mass: sum of the masses of all the protons and neutrons
  - Measured in grams or atomic mass units (amu).
  - An atomic mass unit (amu) is defined as one twelfth of the mass of a carbon-12 atom.

Average Atomic Mass

- Average atomic mass (on the periodic table): the weighted average of all the naturally occurring isotopes of that element.

Isotope	# p^+	$\#n^o$	$\#e^-$	Mass #	Atomic Mass	% abundance
Chlorine-35						80%
Chlorine-37						20%

Average Atomic Mass		

- – Mass of a proton = 1.001 amu
- – Mass of a neutron = 1.001 amu

Preview to the Periodic Table

- Organized by increasing atomic number.
- Period = horizontal row
- Group or Family = vertical column

Terms

- Atomic Number
- Mass Number
- Isotopes
- Atomic Mass
- Average Atomic Mass