Ch. 3 Notes – THE STRUCTURE OF THE ATOM

NOTE: Vocabulary terms are in **boldface and underlined**. Supporting details are in *italics.*

I. Early Ideas About Matter
  A. early "elements" – air, fire, water, earth
  B. **atom**—*the smallest particle of an element retaining the properties of that element*
  C. early theories and ideas, pro and con
    1) *Democritus* of Abdera (460-370 B.C.)
      a) first proposal of finitely divisible matter (could not test his ideas)
      b) "*atoma" / "atomos*"—indivisible, indestructible particles in matter
    2) *Aristotle* (384-322 B.C.): did not believe in atoms
      a) "hyle"— continuous state of all matter
      b) His idea was widely accepted until the 17th century!
    3) Sir Isaac Newton (1642-1727) worked without proof to support atomic theory (Laws of physics, gravitation....)
    4) Robert Boyle (1627-1691) also worked to support atomic theory (gas laws, structured the scientific method, a founder of chem.)
    5) Antoine Lavoisier (1743-1794) – "father of Modern Chemistry"
    6) **atomic theory**—matter is made up of atoms
  D. Atomic Model Development
    1) *John Dalton* (1766-1844); his model (pub.1807) stated that atoms are indivisible
    2) J.J. Thomson; (1856-1940); work begun in 1897
      a) adapted model with subatomic particles: protons and electrons
      b) *"plum pudding" model—electrons stuck in a proton lump*
    3) E. Rutherford (1871-1937); model in 1911
      a) *nucleus as the dense center with* p^+ *and* n^0*;* e^- *outside it*
      b) the atom is mostly space (gold foil experiment)
    4) Niels Bohr (1885-1962); model proposed in 1913
      a) *nucleus as the center, composed of* p^+ *and* n^0
      b) *e- orbit the nucleus; similar to planetary motion*
      c) e^- in an orbit have a fixed energy level
      d) lowest energy levels are closest to the nucleus
      e) *quantum*—a bundle of energy needed to make an electron "jump" to a higher level, which is a *quantum leap*
    5) quantum mechanical model
      a) Erwin Schrödinger (1887-1961); model proposed 1926
      b) based on probability of e^- location, not exact path
      c) **electron cloud model**; "boundary surface diagram"
  E. Atomic theory, conservation of matter, and recycling
    1) natural cycles: nitrogen, carbon, phosphorus, sulfur, water
    2) Laws of Conservation of Mass and Energy apply
  F. Dalton's atomic theory (see next page)
    1) "Father of Atomic Theory" – John Dalton (1766-1844)
    2) Dalton's Atomic Theory (1803)

DALTON'S ATOMIC THEORY
1) All elements are composed of submicroscopic particles called atoms. 2) Atoms are indivisible and indestructible. (He didn't know about subatomic particles and how to split an atom.) 3) Atoms of the same element are identical. (Not really true, as we'll see later.) Atoms of different elements are different. 4) Atoms of elements can physically mix or form compounds by chemically combining in whole-number ratios. (Law of Multiple Proportions) 5) Chemical reactions involve the separation, joining, or rearranging of atoms. Atoms of an element are never changed into atoms of another element in a chemical reaction. (He didn't know about nuclear reactions.)

II. Defining the Atom

A. **electrons** (e^-)— negatively charged subatomic particles

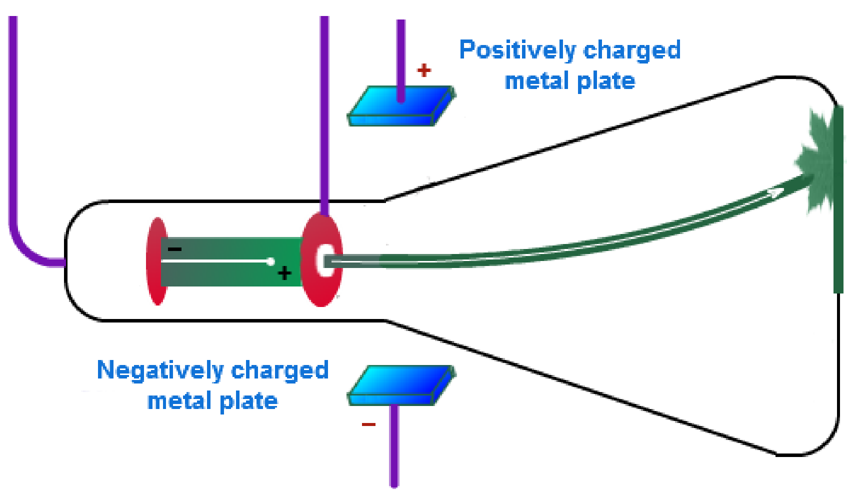
1) characteristics
   a) *fixed charge of -1*
   b) very light mass (9.11×10^{-28} g); 1/1840 of a proton or neutron
   c) they orbit the center: electrons are kept in motion so they don't fall into the positively-charged nucleus

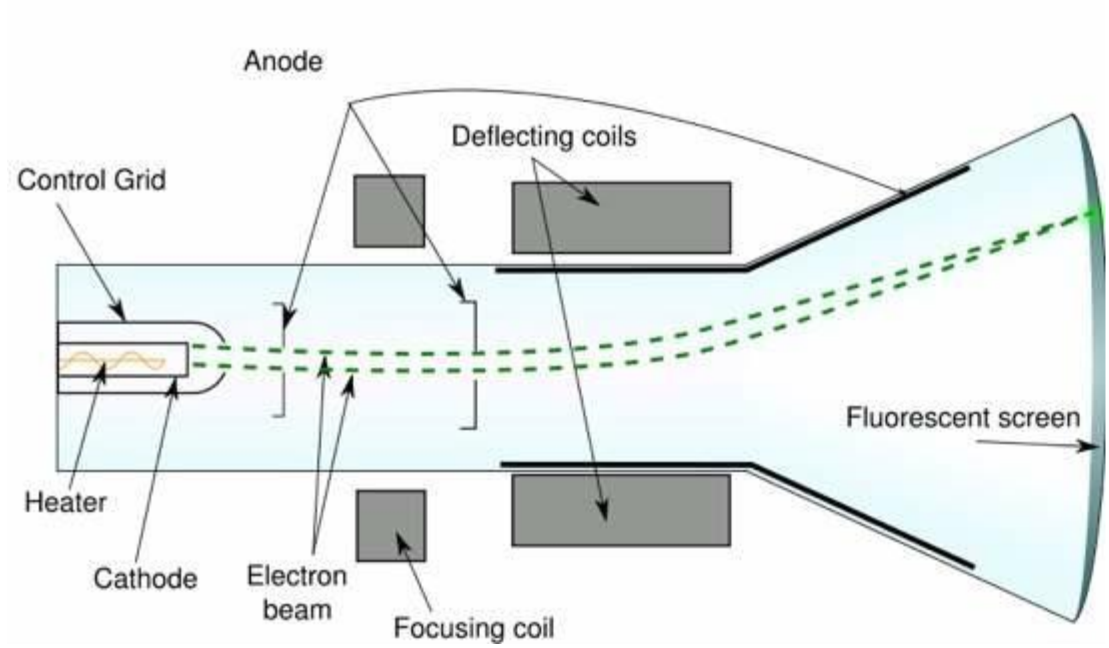
2) Sir William Crookes (1832-1919) discovered cathode rays in a CRT
   a) *CRT (cathode ray tube)*—a closed glass tube with metal electrodes at the ends, containing low-density gases at low pressure, subjected to high voltage.
   b) **cathode ray**— *glowing light beam* arising from the cathode (-) and traveling to the anode (+); *composed of electrons*

3) Sir Joseph John "J.J." Thomson (1856-1940) discovered e^- by CRT experiments

4) Robert Millikan (1868-1953) – oil drop experiments on e^- charge & mass

CATHODE RAY TUBE (CRT) from tutorcircle.com





B. **protons** *(p+)— positively charged subatomic particles*
(a hydrogen atom stripped of its electron is a "raw proton")
1) characteristics
a) *fixed charge of +1*
b) same mass as a neutron (1.67 x 10^{-24} g)
c) located in the center of an atom
2) canal rays—positive CRT beam attracted to the cathode (found by Eugene Goldstein 1850-1930)

C. **neutrons** *(n^0)— neutral subatomic particles*
1) characteristics
a) *fixed charge of 0*
b) same mass as a proton (1.67 x 10^{-24} g)
c) located in the nucleus
2) Sir James Chadwick (1891-1974) discovered the neutron

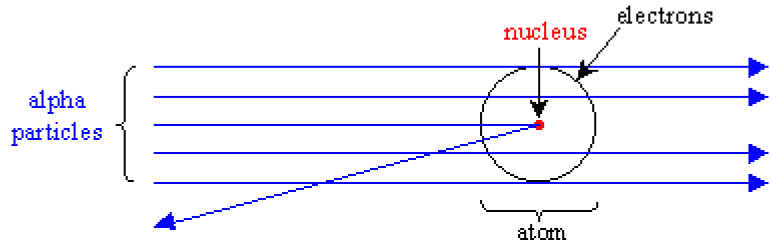
D. other subatomic particles—(hundreds)
1) leptons: muon, tau, neutrino
2) baryons, composed of quark triplets & mesons etc.

E. the nuclear model of the atom
1) **nucleus**—*central core of an atom containing p^+ and n^0*
2) very dense as compared to the rest of the atom
3) the nucleus has an overall positive charge

G. *Rutherford's gold foil experiment*
1) Ernest Rutherford (1871-1937)
2) shot a stream of alpha (α) particles at a sheet of gold foil
3) most of the particles went straight through (because the atoms are mostly empty space)
4) a few particles were deflected (those that grazed a nucleus)
5) even (~1/8000) fewer bounced directly back (those that hit a nucleus head-on)

THE ATOM IS MOSTLY EMPTY SPACE!
If an atom were the size of an average professional football stadium, the nucleus would be the size of a marble.

RUTHERFORD'S EXPERIMENT from www.visionlearning.com :



III. How atoms differ: Atomic numbers vs. atomic masses

A. **atomic number**—*number of protons in the nucleus of an atom*

1) characteristics
   a) the atomic number is the unique I.D. number of an element
   b) each element only has one atomic number

2) examples
   E1) What is the atomic number of the following elements?
   O (8)   I (53)   Cl (17)   Au (79)

B. *atomic neutrality*

1) *atoms are electrically neutral*

number of protons = number of electrons in an atom

2) examples
   E2) How many electrons does Cu have? (29)
   E3) How many electrons does Rn have? (86)

C. **mass number**—*the total number of protons and neutrons in the nucleus*

MASS NUMBER = PROTONS + NEUTRONS
 # OF NEUTRONS = MASS NUMBER - ATOMIC NUMBER

Symbols can be written two ways:

mass number	**12**		
SYMBOL	**C**	OR	**C - 12**
atomic number	**6**		

1) mass number is the total mass of the nucleus
2) Mass number is *not* the decimal number on the periodic table! (that's atomic mass)
3) examples

   E4) How many p^+, n^0 and e^- are in an atom of S-34?

   S = sulfur, which is #16. S has $\boxed{16\ p^+}$ and because # p^+ = # e^-, S has $\boxed{16\ e^-}$.
   Mass number = 34. # n^0 = mass # - atomic # = 34-16 = $\boxed{18\ n^0}$.

E5) How many p^+, n^0 and e^- are in an atom of $^{41}_{19}K$?

K = potassium, which is #19. K has $\boxed{19\ p^+}$ and because # p^+ = # e^-, K has $\boxed{19\ e^-}$.
Mass number = 41. # n^0 = mass # - atomic # = 41-19 = $\boxed{22\ n^0}$.

D. **Isotopes**—*atoms of the same element that contain different numbers of neutrons*
1) same number of p^+
2) different mass numbers
3) different atomic masses
4) in nature, most elements occur as a mix of two or more isotopes
5) examples of oxygen:

ISOTOPE	MASS #	ATOMIC #	p^+	n^0	e^-
O-16	16	8	8	8	8
O-17	17	8	8	9	8
O-18	18	8	8	10	8

Remember, # n^0 = mass number – atomic number.

E. **atomic mass**—*a weighted average based on mass and relative abundance of all naturally occurring isotopes of an element*

ATOMIC MASS =
(MASS x RELATIVE ABUNDANCE) of natural isotope #1 +
(MASS x RELATIVE ABUNDANCE) of natural isotope #2 +
(MASS x RELATIVE ABUNDANCE) of natural isotope #3 … etc.

1) unit is amu = atomic mass unit
2) synthetic isotopes (made in lab, not found in nature) are not considered

E6) Magnesium has three isotopes: Mg-24, Mg-25, and Mg-26:

ISOTOPE	ABUNDANCE	ATOMIC MASS
Mg-24	78.70%	23.985
Mg-25	10.13%	24.986
Mg-26	11.17%	25.983

The atomic mass of Mg:
ATOMIC MASS = (MASS x RELATIVE ABUNDANCE)
(23.985)(0.7870) + (24.986)(0.1013) + (25.983)(0.1117) = $\boxed{24.31\text{ amu}}$

IV. Radioactive decay (overlap with Chapter 23)

A. Nuclear reactions
1) **nuclear reactions**—*chemical reactions converting matter to energy*
2) violate the Conservation Laws
3) involves **transmutation**—*the changing of one element into another element*

B. Radioisotopes
1) **radioisotopes** (*radioactive isotopes* or *radionuclides*)—*radioactive forms of an element*
2) unstable isotopes which spontaneously release particles

C. **half life**—*the time it takes for half the amount of a radioisotope to decay* (from a fraction of a second to thousands of years)

D. **radiation**—*emissions from a radioactive material*; can be rays and/or particles

TYPES OF EMISSIONS:

1) **alpha particle (α)**
   a) characteristics: made of He nuclei (2 protons, 2 neutrons); *positively charged*
   b) low **penetrating power** (can't pass through matter easily)
   c) alpha decay example:

$$^{241}_{95}\mathbf{Am} \rightarrow {}^{237}_{93}\mathbf{Np} + {}^{4}_{2}\mathbf{He}$$

2) **beta particle (β)**
   a) characteristics: made of electrons; *negatively charged*
   b) intermediate penetrating power (smaller size than alphas)
   c) beta decay example:

$$^{3}_{1}\mathbf{H} \rightarrow {}^{3}_{2}\mathbf{He} + {}^{0}_{-1}\mathbf{e}$$

3) **gamma radiation (γ)**
   a) characteristics: made of electromagnetic (em) radiation, no charge
   b) high penetrating power (no charge, no mass)
   b) gamma decay example:

$$^{3}_{2}\mathbf{He} \rightarrow {}^{3}_{2}\mathbf{He} + \gamma$$