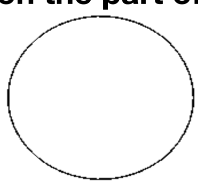
Chemistry Test on Topic III-1: Basic Atomic Structure

I. Basic Atomic Structure - pp. 73; 78-79; 100; 81-85; 206-207

A. Early Concepts: Greek & Roman **[4.1a(stop at Bohr)]**

- **Democritus** and Leucippus (400 B.C) – The idea of small, invisible particles being the building blocks of matter can be traced back more than 2,000 years to the Greek philosopher Democritus and Leucippus. These particles, considered to be so small and indestructible that they could not be divided into smaller particles were called **"atomos"** the Greek word for indivisible. The English word **atoms** comes from this Greek word**. This early concept of atoms was not based upon experimental evidence but simply a result of thinking and reasoning on the part of the philosophers.**



- The Greeks were the first to try to explain why chemical changes occur. By about 400 B.C., they had proposed that **all matter was composed of four fundamental substances: fire, earth, water, and air**.
- **Alchemy** dominated the next 2000 years of chemical history. Some alchemists were fakes who were obsessed with the idea of turning cheap metals into gold. However, many alchemists were sincere scientists and this period saw important events: the elements mercury, sulfur, and antimony were discovered and alchemists learned how to prepare acids.

B. Beginnings of The Modern Model

1. Mathematical nature of matter – Real Chemistry from this point on
    - The 1st scientist to recognize the importance of careful **measurements** was the Irishman **Robert Boyle**. He insisted that science should be firmly grounded in experiments. His definition of ***element*** was based on experiments: a substance was an element unless it could be broken down into two or more simpler substances (by chemical means). He believed that metals were not true elements and eventually one would be able to change one metal into another (belief of alchemists).
    - As scientists of the eighteenth century studied the nature of materials, several things became clear:

1. Most natural materials are mixtures of pure substances.
2. Pure substances are either elements or combinations of elements called compounds.
3. A given compound always contains the same proportions (by mass) of the elements. For example, water *always* contains 8 g of oxygen for every 1 g of hydrogen, and carbon dioxide *always* contains 2.7 g of oxygen for every 1 g of carbon. This principle became known as the **law of constant composition.** It means that a given compound always has the same composition, regardless of where it comes from.

2. Dalton's atomic model **(1803)–** Billiard Ball Theory

- **Proposed the 1st Atomic Theory in 1803**
- It was not until the 18th century that experimental evidence in favor of the atomic hypothesis began to accumulate. Finally, around 1805, John Dalton proposed some basic assumptions about atoms based on what was known through scientific experimentation and observation at the time. These assumptions are very closely related to what scientists presently know about atoms. For this reason, Dalton is often referred to as **the father of modern atomic theory**.
- Some of these basic ideas were: (**some are not true today)**
    1) All matter is made of very small, discrete particles called atoms. (No internal structure, just solid matter)
    2) All atoms of an element are alike in weight, and this weight is different from that of any other kind of atom.
        - **Not true today –** isotopes: have different numbers of neutrons
    3) Atoms cannot be subdivided, created, or destroyed.
        - **Not true today** – can split atoms, they are not indivisible
    4) Atoms of different elements combine in simple whole-number ratios to form chemical compounds
    5) In chemical reactions, atoms are combined, separated or rearranged. (bonds broken and formed)
    - (An atom is **neutral)**

Dalton's Atomic Theory

1. Elements are made of tiny particles called **atoms.**
2. All atoms of a given element are identical.
3. The atoms of a given element are different from those of any other element.
4. Atoms of one element can combine with atoms of other elements to form compounds. A given compound always has the same relative numbers and types of atoms.
5. Atoms are indivisible in chemical processes. That is, atoms are not created or destroyed in chemical reactions. A chemical reaction simply changes the way the atoms are grouped together.

- Dalton used his model to predict how a given pair of elements might combine to form more than one compound. Since he correctly predicted the formation of multiple compounds between two elements, his atomic theory became widely accepted.

- By the second half of the 1800s, many scientists believed that all major discoveries related to the elements had been made. The only thing left for young scientists to do was refine what was already known. This came to a surprising halt when **J.J. Thomson discovered the electron beam in a cathode ray tube in 1897.** Soon afterward, Henry Becquerel announced his work with radioactivity, and Marie Curie and her husband, Pierre, set about trying to isolate the source of radioactivity in her laboratory in France.

C. Composition of The Atom

1. **Matter's electrical connection**
    - From around the beginning of the 20^{th} Century, scientists have been gathering evidence about the structure of atoms and fitting the information into a model of the atomic structure.

    a. anode/cathode rays; photoelectric effect

    <u>JJ Thomson - Anode/Cathode Rays</u>

The discovery of the electron (means beam) as the 1^{st} subatomic particle is credited to **J.J. Thomson** (England, 1897). He used an evacuated tube connected to a spark coil as shown in the figure below. As the voltage across the tube was increased, a beam became visible. This was referred to as a cathode ray. Thomson found that the beam was deflected by both electrical and magnetic fields. Therefore, he concluded that cathode rays are made up of very small, negatively charged particles, which became known as **electrons**.

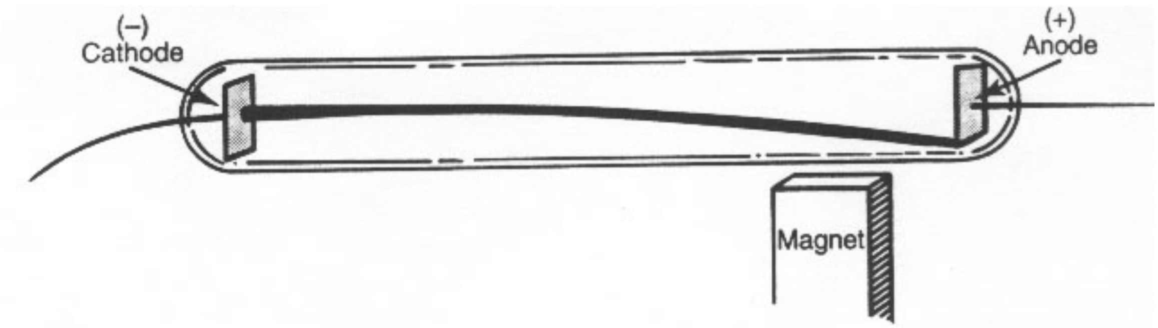
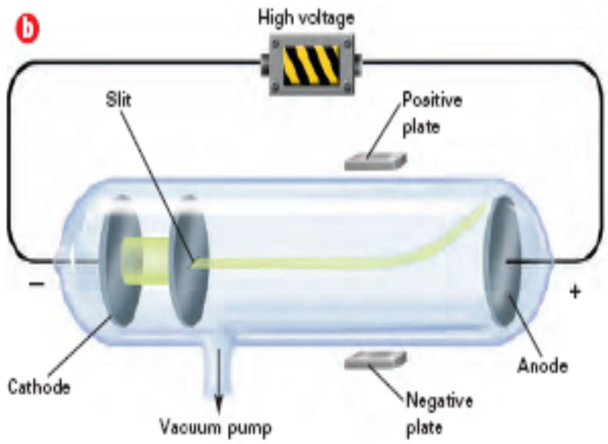
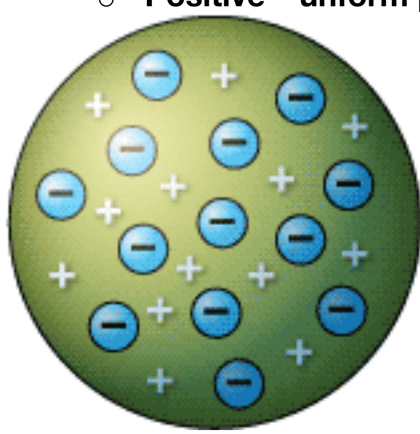


FIGURE 3. *Electron Experiment*



- **He concluded that atoms must also contain positive particles besides negative particles, giving the atom a zero (neutral) overall charge.**
- **Plum Pudding Model (Chocolate Chip Cooke)**
    - **Negative – plums/chips**
    - **Positive – unform pudding/cookie**



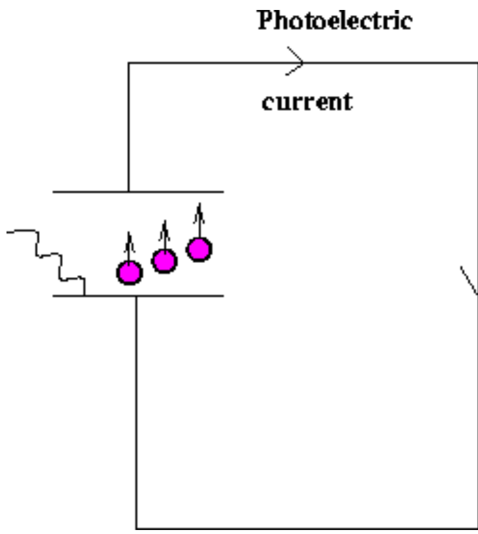
- Further experimentation led Thomson to find the ratio of the electrical charge of the electron to its mass (**charge to mass ratio – is much greater for electrons than protons**). This was a major step toward understanding the nature of the particle.
- “Charge” Unit is a **Coulomb (c)**

    - + or –
        - + - attract
        - ++ **and** -- repel
- electron “e-“ is what moves to create electricity

<u>Photoelectric Effect – for more see packet</u>

- An experiment which provides compelling proof for the photon nature of light (also called the particle nature of light) is the **photoelectric effect**. In this effect light is shone at a metal plate, and it is found that electrons are ejected. These electrons then get accelerated to a nearby plate by an external potential difference, and a photoelectric current is established, as below.
- The explanation, which was first given by Einstein and which won him the Nobel Prize, is as follows. Each photon hits an electron in the metal, giving up its energy to the electron. This can sometimes be enough energy to free the electron from the attractive forces holding it in the metal. The electron is then accelerated (or moves quickly) towards the other side, causing a flow of charges and hence a current. This effect, which arises in devices such as burglar alarms, cannot be explained using a “wave” picture of light. For example, it is found experimentally that the photoelectric current depends critically on the frequency of the light being used. If the frequency of the light used is too low, then no current is observed. Red light is the lowest frequency light and violent is the highest frequency light in the visible spectrum.
- Important conclusions from the work of Einstein and Planck:
    1. Energy is quantized. It can occur in only discrete amounts of quantum. (quantum mechanics)
    2. Light has characteristics of both waves and particles (called photons). The phenomenon is sometimes referred to as the **dual nature of light.**

- **Einstein used monochromatic light**
- **A minimum frequency is required in order to produce a current**
    - **Current = flow of electrons**
- **Einstein – light travels in bundles/packets of energy called photon**
- **Energy=constant time frequency of light**
- **Einstein single handedly advanced science 50 years at the time**

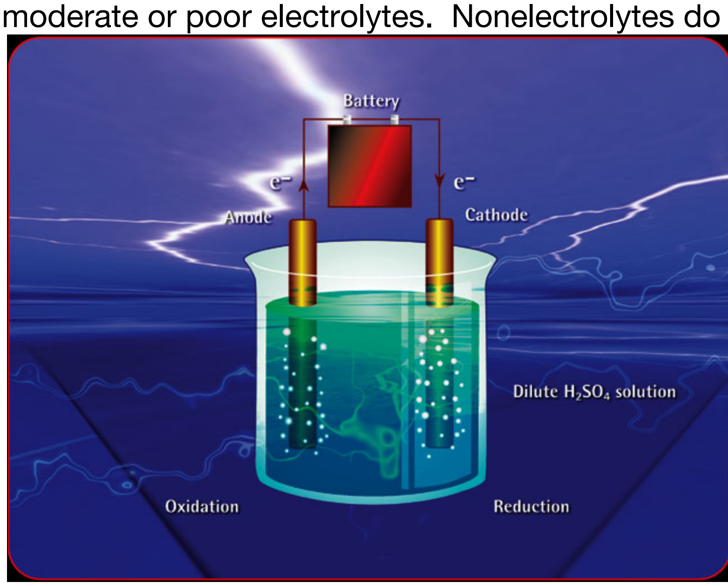


Flame Test (Lab)

- Electrons in **ground state** are at the **lowest energy state**
- **Excited state**
- When come down to ground state, unique color is produced
- On test: If gave ROYGBIV – Where is the longest wave length? – Red, Shortest - Violet

b. **Faraday and electrolysis**

In the early 1830s, Michael Faraday discovered that water solutions of certain substances conduct an electric current. He called these substances **electrolytes**. Our definition today of an electrolyte is much the same. It is a substance that dissolves in water to form a solution that will conduct an electric current. The usual apparatus to test for this conductivity is a light bulb placed in a series of two prongs that are immersed in the solution tested, as shown below. Solutions can be good, moderate or poor electrolytes. Nonelectrolytes do not conduct at all.



2. Electrons: Thompson and Millikan – see above and below for more

Thomson

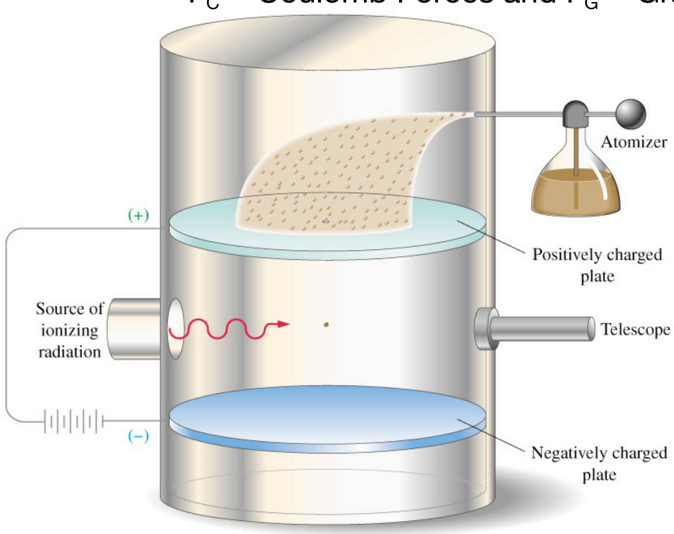
a. discovering the electron & its charge

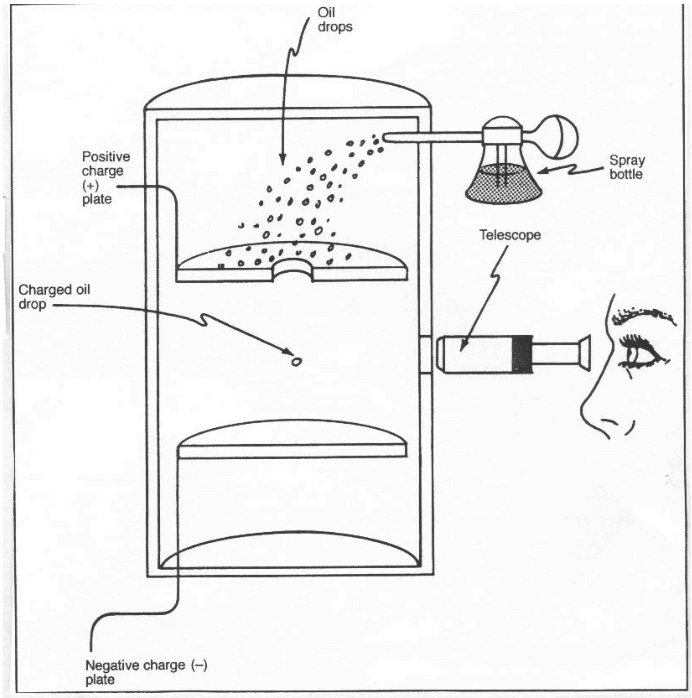
- **Discovered the electron, which is negatively charged**
- See more info above

b. Thompson's atomic model – The Plum Pudding Model – see above

Millikan

- It was an American scientist, **Robert Millikan**, who in 1909 was able to measure the charge on an using the apparatus pictured below.
- Oil droplets were sprayed into the chamber and, in the process, became randomly charged by gaining or losing electrons. The electric field was adjusted so that a negatively charged drop would move slowly upward in front of the grid in the telescope. Knowing the rate at which the drop was rising, the strength of the field, and the mass of the drop, Millikan was able to calculate the charge on the drop. Combining the information with the results of Thomson, he could calculate a value for the mass of a single electron. Eventually this number was found to be **9.11×10^{-28} gram.** (He initially determined the magnitude of the electron to be 1.60×10^{-19} Coulombs which led him to the mass.)
- F_C = Coulomb Forces and F_G = Gravity --- They are equal.



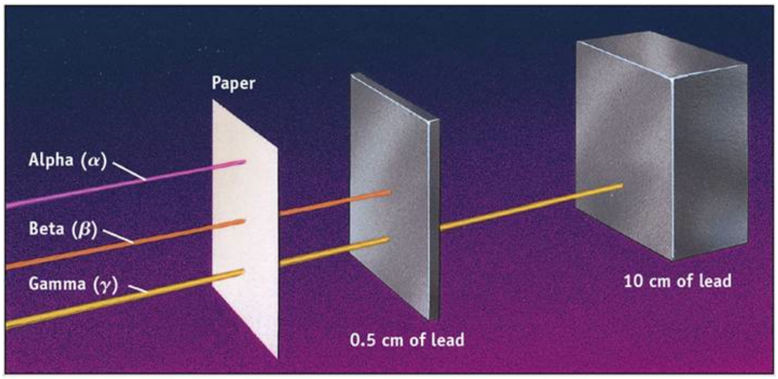


3. The Nucleus: Rutherford – Gold Foil Experiment
   a. discovery of radioactivity; types of radiation **[4.2a]**
   - **Origin of radiation = nucleus**
   - **Light = radiation**
   - **3 types of radioactive emissions**
     - α – alpha particles: charge 2x that of electron, opposite charge
     - β – beta particles: high-speed electron
     - ɣ - gamma rays: high energy "light"

Radioactivity

In the late nineteenth century, scientists discovered that certain elements produce high-energy radiation. For example, in 1896 the French scientist Antoine Henri Becquerel accidentally found that the image of a piece of mineral containing uranium could be produced on a photographic plate in the absence of light. He attributed this phenomenon to a spontaneous emission of radiation by the uranium, which his student, Marie Curie, called **radioactivity.** Studies in the early twentieth century demonstrated three types of radioactive emission: gamma (γ) rays, beta (β) particles, and alpha (α) particles. A γ ray is high-energy "light"; a β particle is a high-speed electron; and an α particle has a 2+ charge—that is, a charge twice that of the electron and with the opposite sign. The mass of an α particle is 7300 times that of

an electron.



b. **Rutherford's experiment**

- **Ernest Rutherford** (England, 1911) performed a gold foil experiment (see below) that had tremendous implications for atomic structure.
- Alpha particles (helium nuclei) passed through the foil with few deflections. However, some deflections (1 per 8,000) were almost directly back toward the source. This was unexpected and suggested an atomic model with mostly empty space between a **nucleus**, in which most of the mass of the atom was located and which was positively charged, and the electrons that defined the volume of the atom. After two years of studying the results, Rutherford finally came up with an explanation. He reasoned that the rebounded alpha particles must have experienced some powerful force within the atom. And he assumed this force must occupy a very small amount of space, because so few alpha particles had been deflected. He concluded that the force must be a densely packed bundle of matter with a positive charge. He called this positive bundle the nucleus. He further discovered that the volume of a nucleus was very small compared with the total volume of an atom. If the nucleus were the size of a marble, then the atom would be about the size of a football field. The electrons, he suggested, surrounded the positively charged nucleus like planets around the sun, even though he could not explain their motion.
- Further experiments showed that the nucleus was made up of still smaller particles called **protons**. Rutherford realized, however, that protons by themselves could not account for the entire mass of the nucleus. He predicted the existence of a new nuclear particle that would be neutral and would account for the missing mass. In 1932, James Chadwick (England) discovered this particle, the **neutron**. Today the number of subatomic particles indentified and named as discrete units has risen to well over 90.

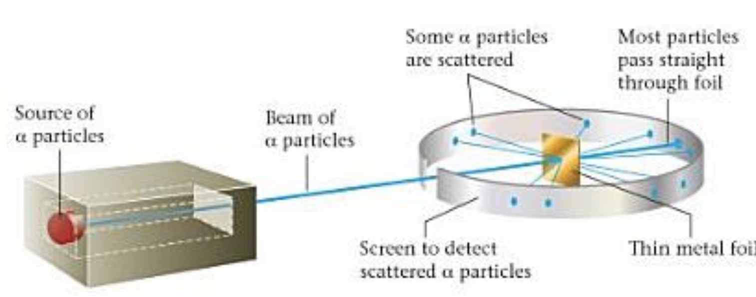


Figure 4.5
Rutherford's experiment on α-particle bombardment of metal foil.

(radium is source)

c. Rutherford's atomic model

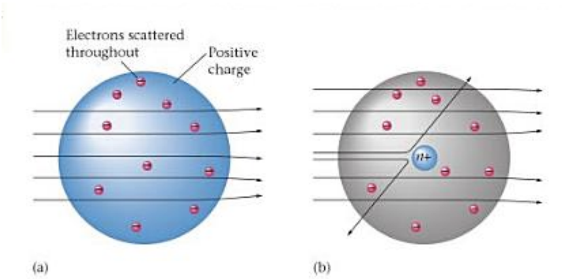
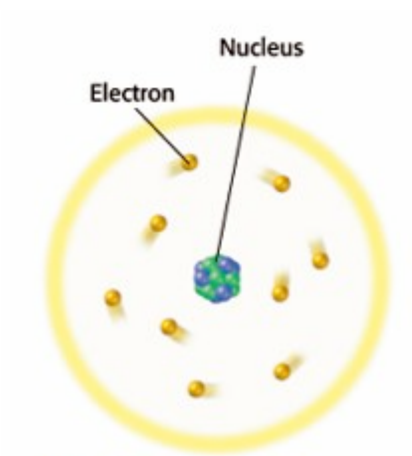


Figure 4.6
(a) The results that the metal foil experiment would have yielded if the plum pudding model had been correct. (b) Actual results.



D. Typical Atomic Properties

- Protons – positively charged
- Neutrons – neutrally (no) charge
- Electrons – negatively charged

1. Atomic sizes

Name	Charge	Symbol	Actual Mass (g)	Relative Mass compared to proton	Location
Electron	- (e^-)	${}^{0}_{-1}e$	9.109×10^{-28}	1/1,837	Outside Nucleus
Proton	+ (P^+)	${}^{1}_{1}H$	1.673×10^{-24}	1	Nucleus
Neutron	0 (n^0)	${}^{1}_{0}n$	1.675×10^{-24}	1 (slightly larger than proton)	Nucleus

- Protons are much bigger than electrons
- The charge to mass ratio is much greater for electrons and much less for protons (Electrons are packed with charge)

2. Atomic number; Moseley
   - Atomic # - The number of protons in the nucleus of an atom determines the **atomic number**.
   - All atoms of the same element have the same number of protons and therefore the same atomic number; atoms of different elements have different atomic numbers. Thus, the atomic number indentifies the element. An English scientist, **Henry Moseley**, first determined the atomic numbers of the elements through the use of x-rays.

3. Atomic masses

   - Since the actual masses of subatomic particles and atoms themselves are very small numbers when expressed in grams, scientists use atomic mass units (amu) instead. An **atomic mass unit** is defined as ½ the mass of a carbon atom. Thus, the mass of any atom is expressed relative to the mass of one atom of carbon-12, which is sometimes called the **unfied atomic mass unit** (u) or the **dalton.**
   - The sum of the # of protons and the # of neutrons in the nucleus is called the **mass number.**

   a. relative mass scale – see above
   b. isotopes
   - **Isotope** - atoms of the same elements that have different masses, meaning a different number of neutrons and the same number of protons and electrons.

- To specify which of the isotopes of an element we are talking about, we use the symbol:

$$^{A}_{Z}\text{X}$$

- Where
  - X = the symbol of the element
  - A = the mass number (number of protons and neutrons)
  - Z = the atomic number (number of protons)
- Mass number = number of protons + number of neutrons or A = Z + # of neutrons, so A – Z = # of neutrons

- **Average atomic mass –** the weighed average of the atomic masses of the naturally occurring isotopes of an element.

c. **<u>average isotopic (atomic) mass; calculations</u> [4.1b]**

- **The Average atomic mass can be calculated by multiplying the atomic mass of each isotope by its relative abundance (expressed in decimal form) and adding the results.**
- Size of one carbon atom = 12 amu