CHEMISTRY NOTES - Chapter 5

Atomic Structure and the Periodic Table

Goals : To gain an understanding of :

1. Atoms and their structure.
2. The development of the atomic theory.
3. The periodic table.

NOTES:

An **atom** is the smallest part of an element that retains the properties of that element. The concept of an atom goes a long way back. It was first suggested by an ancient Greek named Democritus (the Greek word "atomos" means indivisible). Democritus theorized that if you took an object and cut it in half again and again you would eventually end up with some particle which could not be further divided.

In the early 1800's an English scientist by the name of John Dalton started relating what chemists could see to the concept of the atom. He came up with an atomic theory which could be stated as follows:

- All elements are composed of tiny indivisible particles called atoms.
- Atoms of the same element are identical and differ from atoms of other elements.
- Atoms of different elements can combine together in simple whole number ratios to form compounds.
- Chemical reactions are the rearranging of the combinations of atoms of elements in compounds. The atoms themselves remain unchanged.

Atoms are composed of three primary particles, protons, neutrons and electrons.

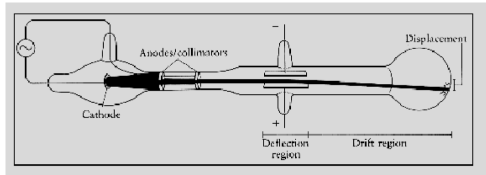
Particle	Symbol	Location	Relative electrical charge	Approximate relative mass in amu	Actual mass in grams
Electron	e^-	orbits the nucleus	1-	1/1840	9.11×10^{-28}
Proton	p^+	nucleus	1+	1	1.66×10^{-24}
Neutron	n^o	nucleus	0	1	1.66×10^{-24}

amu (atomic mass unit = 1.66×10^{-24} grams)

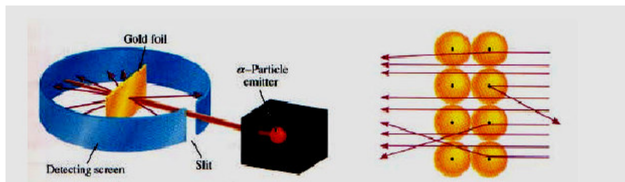
The electron was discovered by the English physicist Sir Joseph J. Thomson around 1897 with the use of a cathode ray tube. A cathode ray tube is similar to your TV. It has an anode (negative electrode) and a cathode (positive electrode). These are enclosed in an evacuated (air removed) glass container and when a charge is applied the electrons flow from anode to cathode through the open space of the glass container. Thomson observed these particles and determined that the particles:

- move at a very high speed
- have a negative charge
- have a mass of about 1/2000 of a hydrogen atom (smallest atom)
- were the same regardless of which gas was used in the container or the metal used as the electrode

The particles were eventually named "electrons." His model of the atom was the “plum-pudding” model. This discovery shattered Dalton's notion of an atom. To Dalton atoms were tiny, solid particles, not containing smaller particles.



The nucleus was discovered by Ernest Rutherford in 1911. Rutherford set up an apparatus in which he aimed alpha particles (type of radiation made up of helium nuclei) at a very thin sheet of gold foil. If the atoms were made up of evenly dispersed protons and electrons, as believed at the time, the alpha particles should go straight through unhindered. What happened was that some of the alpha particles went through; some were deflected as they passed through and some bounced back. Rutherford concluded that the positive mass of the atom (protons) must be concentrated in a very small area and most of the rest of the atom must be empty space. The area where the protons were located was called the nucleus.



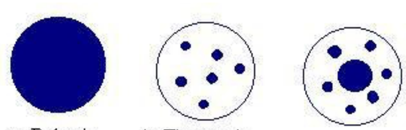
The nucleus of an atom is extremely small as compared to the entire size of the atom. It can be compared to the size of a marble in a football stadium where the marble represents the nucleus and the stadium the entire atom. Most of the space of the atom is occupied by the orbiting electrons.

Structure of the Nucleus:

1. Protons
   a. Positive charge, mass of 1.673×10^{-27} kg
   b. The number of protons in the nucleus determines the atom's identity and is called the atomic number
2. Neutrons
   a. No charge, mass of 1.675×10^{-27} kg

The different historical models are described as follows:

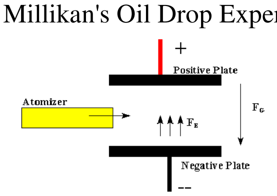
- Dalton's model of the atom - solid, tiny, indivisible particles.
- Thomson's model - often describe as the "plum pudding" model - electrons are scattered throughout the atom.
- Rutherford's model - includes the solid nucleus in the center of the atom.



a. Dalton's model
b. Thomson's model
c. Rutherford's model

Sir James Chadwick discovered the neutron in 1932.

Millikan's Oil Drop Experiment:



a. Electron is negatively charged
b. Mass is about 1/2000th of a hydrogen atom
   (a) Electron mass is 9.109×10^{-31} kg

Atomic number - indicates the number of protons and defines the element (atomic number 6 is always carbon, atomic number 7 is always nitrogen etc.).

Mass number - equals the total number of neutrons and protons in the nucleus of an atom for the most common isotope this equals the atomic mass rounded off to the nearest whole number

Atomic mass - the average mass of an atom of an element (in amu)

Calculation of the number of particles in an atom of an element :

- number of protons equals the atomic number
- number of neutrons equals the mass number minus the atomic number (remember virtually all the mass is from the neutrons and protons in the nucleus-each with an amu of 1)
- number of electrons equals the number of protons in a neutral atom

Consider the following periodic table information for carbon, nitrogen and sodium :

6	7	11
C	N	Na
12.011	14.007	22.990

Carbon's atomic number is 6, has an average mass of 12.011 amu and carbon's most common isotope has a mass number of 12 amu. Therefore, the most common type of carbon atom has 6 protons, 6 neutrons and 6 electrons.

Nitrogen's atomic number is 7, has an average mass of 14.007 amu and nitrogen's most common isotope has a mass of 14 amu. Therefore the most common type of nitrogen atom has 7 protons, 7 neutrons and 7 electrons.

Sodium's atomic number is 11, has an average mass of 22.990 amu and nitrogen's most common isotope has a mass of 23 amu. Therefore the most common type of sodium atom has 11 protons, 12 neutrons and 11 electrons.

An **isotope** is a particular form of an atom of an element. Isotopes have a different number of neutrons and therefore differ in mass (that is why there is a non-whole number atomic mass which is an average of the various isotopes). The number of protons is always the same in isotopes since they are different forms of the same element (must be same atomic number). For example, the following chart shows three isotopes of hydrogen :

Isotope	Atomic Number	Number of protons	Number of Neutrons	Number of electrons	mass (amu)
Hydrogen-1	1	1	0	1	1
Hydrogen-2 (deuterium)	1	1	1	1	2
Hydrogen-3 (tritium)	1	1	2	1	3

Note hydrogen has three isotopes, each with a whole number mass, yet hydrogen as an element has a characteristic average mass called the atomic mass (1.0079 amu).

An **atomic mass unit** (amu) is 1/12 the mass of a carbon-12 atom, or 1.66×10^{-24} grams

Mendeleev's Periodic Table (1869)

A. Organization

1. Vertical columns in atomic weight order
   a. Mendeleev placed elements in rows with similar properties
2. Horizontal rows have similar chemical properties

B. Missing Elements
1. Gaps existed in Mendeleev's table
a. Mendeleev predicted the properties of the "yet to be discovered" elements
(1) Scandium, germanium and gallium agreed with his predictions

C. Unanswered Questions
1. Why didn't some elements fit in order of increasing atomic mass?
2. Why did elements exhibit periodic behavior?

Moseley and the Modern Periodic Table (1911)

A. Protons and Atomic Number
1. The periodic table was found to be in atomic number order, not atomic mass order

B. The Periodic Law
1. The physical and chemical properties of the elements are periodic functions of their atomic numbers
2. Elements with similar properties are found at regular intervals within the periodic table

* Moseley was killed in battle in 1915, during WWI. He was only 28 years old

Organization of the Table

1. Groups or Families
   a. Vertical columns containing elements with similar chemical properties
2. Periods (series)
   a. Horizontal rows of elements
3. Metals and Nonmetals
   a. A stair-step line on the table separates the metals from the nonmetals
   b. Metalloids (Semimetals) straddle the line and have properties of both metals and nonmetals
4. Lanthanide and Actinide Series (Inner Transition Metals)
   a. Metals and man-made metal elements
5. Group 1 – Alkali metals (the most reactive metal elements) (except hydrogen (H) also in this group)
6. Group 2 – Alkaline earth metals (very reactive metal elements)
7. Group 17 – Halogens (the most reactive nonmetal elements)
8. Group 18 – Noble gases (the least reactive elements – inert and very stable)

Types of Elements

A. Metals
1. Luster
2. Good conductors of heat and electricity
3. Malleable
4. Ductile
5. High tensile strength

B. Nonmetals
1. Many nonmetals are gases at room temperature
2. Solid nonmetals tend to be brittle and non-lustrous
3. Poor conductors of heat and electricity

C. Metalloids
1. Some properties of metals and some properties of nonmetals
2. Solids at room temperature
3. Semiconductors of electricity

D. Noble Gases
1. All are gaseous members of group 18
2. Generally unreactive and stable