Chapter 4 – Atoms

4.1 Dalton and the Atomic Theory

In 1808 John Dalton came up with the first atomic theory which consisted of five ideas:

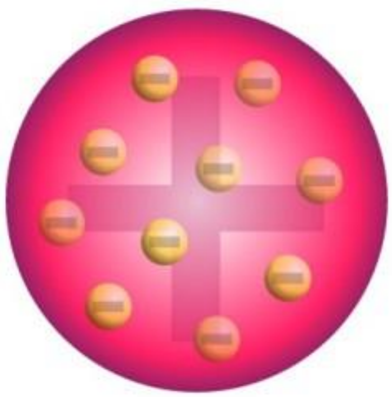
1. An element is composed of tiny, indivisible, indestructible particles called atoms.
2. All atoms of an element are identical and have the same properties.
3. Atoms of different elements combine to form compounds.
4. Compounds contain atoms in small whole number ratios.
   e.g. Each H_2O molecule consists of one O and two H atoms
5. Atoms can combine to form different compounds.
   e.g. carbon and oxygen can form CO_2 or CO

Note: Dalton's first two ideas were later found to be incorrect. Do you know why?

4.2 Thomson and the Electron

J.J. Thomson was given credit for discovery of electron in 1897. He determined that cathode rays are deflected by electric fields & magnetic fields, so cathode rays are composed of tiny, negatively charged subatomic particles called electrons **(e^-).**

In 1904 J.J. Thomson proposed a model where an atom is a positively charged sphere with electrons scattered about like raisins in plum pudding. This was called the plum pudding model of the atom and it was very short lived.



[YouTube Thomson and the Cathode Ray Tube]

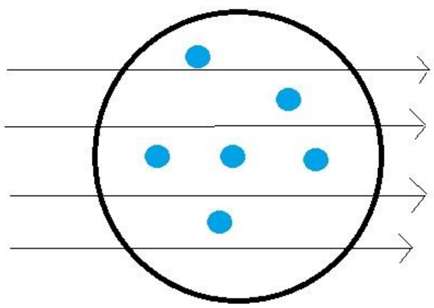
4.3 Rutherford and the Nucleus

In 1911 Ernest Rutherford is credited with discovering the nucleus through his alpha scattering experiment. He shot alpha particles (α) at gold foil surrounded by a detector. Most of the α particles went straight through the foil, but a few actually bounced sideways or even backwards! This was astonishing! It was as surprising as a fired bullet bouncing off a Kleenex!

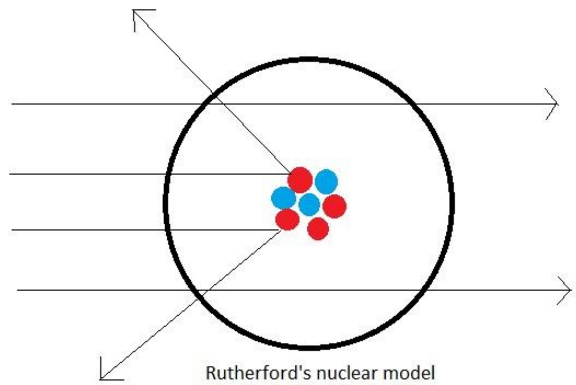
[YouTube Rutherford's experiment]

Rutherford's interpretations:

1. Most alpha (α) particles passed through foil
   $\Rightarrow$ The atom is mostly empty space with electrons moving around the space.
2. Some α particles were deflected or bounced back
   - Atoms contains a *small dense region* and when particles strike this region they recoil.
   - The dense region = **atomic nucleus** (contains atom's protons and neutrons)



Thomson's model - Plum Pudding



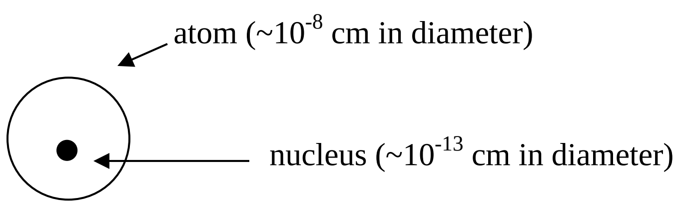
Rutherford's nuclear model

Rutherford's Planetary Model of the Atom

$\Rightarrow$ *Negatively charged* **e^-'s** move around the *positively charged* nucleus

In 1917 Rutherford is also generally credited with discovering that the positive charge of an atom is in a particle called the proton. He was working with hydrogen nuclei.

Rutherford also estimated the size of the atom and its nucleus:

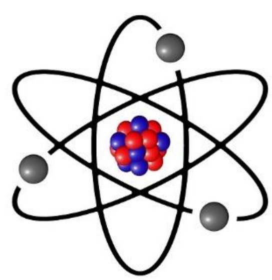


If nucleus = size of a small marble, then the atom is the size of the Cardinal's stadium!

James Chadwick discovered neutrons (1932): he won the Nobel Prize for this in 1935.

$\Rightarrow$ **(n)** = *neutral* subatomic particle

So at this point we have a model with protons (red) and neutrons (blue) inside a center dense nucleus with electrons (gray) orbitting around the nucleus. This is the nuclear model of the atom.



4.4 Subatomic Particles – electrons, protons, neutrons

Particle	Symbol	Location	Charge	Relative Mass (amu)
electron	e^-	outside nucleus	–1	1/1836 ≈ 0
proton	p^+	inside nucleus	+1	1
neutron	n	inside nucleus	0	1

Note the mass of the electron is 1836 times smaller than the mass of a proton or neutron. Thus comparatively speaking it is basically zero.

4.5 Atomic Notation, Atomic Number, Mass Number

Every atom of an element has the same number of protons. The number of protons never changes for an atom. An atom is defined by the number of protons. 6 protons = Carbon always and Carbon always has 6 protons.

The Atomic Notation is shorthand for keeping track of the protons and neutrons in the nucleus:

$$^{\text{Mass Number = A}}_{\text{Atomic Number = Z}}\text{Sy} \text{ - element symbol}$$

Atomic number Z: total number of protons

Mass number A: total number of protons and neutrons in an atom's nucleus

Example: $^{17}_{8}\text{O}$ means 8 protons and 17 is the mass = protons + neutrons

So how can you calculate the number of neutrons in an atom? Well if the mass is protons and neutrons, and the atomic number is protons only, you just subtract. So for the above 17-8 = 9 neutrons.

For neutral atoms: total number of p^+ = total number of e^-

4.6 Isotopes

Isotopes: Atoms of an element that have a different number of neutrons. Isotopes of an element have the same atomic number or number of protons, but a different mass number because of the different neutrons.

⇒ We often refer to a specific isotope of an element by giving the name of the element followed by the mass number.

e.g. carbon-12 (^{12}C), carbon-13 (^{13}C) and carbon-14 (^{14}C) are isotopes of carbon

Masses of atoms are so small that we define the atomic mass unit (amu) to scale up the numbers.

Carbon-12 was chosen as the **reference** and given a mass value of exactly **12.000 amu**. The mass of all other atoms are scaled relative to mass of Carbon-12.

The Atomic Mass of an Element in the Periodic Table is the ***weighted average*** of all naturally occurring isotopes for that element. So that is why there are decimal places and not whole numbers. The mass in the periodic table is an average of all the isotopes.

For example, there are two main naturally occurring isotopes of carbon: ^{12}C and ^{13}C

More carbon exists as carbon-12 (98.89%) compared with carbon-13 (1.11%), so the atomic mass reported for carbon (12.01 amu) is closer to carbon-12 than carbon-13. Thus we call carbon-12 the **most abundant isotope**.

Some elements are *radioactive* and *unstable*. Radioactive means the atoms is unstable because the nucleus emits ionizing radiation such as alpha particles, beta particles, or gamma rays.

⇒ distinguished on the Periodic Table with parentheses around the ***mass number (****not the reported atomic mass****)*** of the most common isotope for the radioactive element.

e.g. Element # 96, curium, Cm, has a mass number of (247).

4.7 Chemical Formulas

Chemical formula:

- expresses the type and number of atoms present in a compound, but not their bonding order
- number of atoms is indicated by a subscript following the element's symbol
  - (If there is no subscript, only *one* atom of that element is in the compound.)

Example: water = H_2O which means 2 H atoms and 1 O atom = 3 atoms total

Parentheses are used when there is more than one polyatomic group present in the compound.

Example: $(NH_4)_2CO_3$ $(NH_4)_2$ = 2 NH_4's = 2 N + 2×4 H = 2 N + 8 H

⇒ TOTAL: 2 N, 8 H, 1 C and 3 O = 14 atoms total

Law of Definite (or Constant) Composition: Compounds always contain the same elements in the same proportion by mass. Water always contains 11.2% hydrogen and 88.8% oxygen by mass.

Practice Problems

1. Write the atomic notation for sodium-23. How many neutrons are in a sodium-23 atom?

2. Write the atomic notation for chlorine-37. How many neutrons are in a chlorine-37 atom?

3. Indicate the mass number, number of protons, neutrons, and electrons in this table:

Isotope	**mass #**	**# of protons**	**# of neutrons**	**# of electrons**
carbon-12				
carbon-13				
carbon-14				
calcium-41				
chlorine-35				

4. Use the atomic weight reported on the Periodic Table to determine the most abundant naturally occurring isotope for each of the following:
   a. lithium-6 or lithium-7
   b. chlorine-35 or chlorine-37

5. Which scientist thought that electrons were scattered about an atom like raisons in plum pudding?

6. Which subatomic particle has a +1 charge?
7. Which scientist decided that atoms has a dense center nucleus?
8. How many atoms are in $Al(NO_3)_3$ total?

Answers to Practice Problems

1. $^{23}_{11}Na$ 23-11=12 neutrons

2. $^{37}_{17}Cl$ 37-17=20 neutrons

Isotope of carbon	**mass #**	**# of protons**	**# of neutrons**	**# of electrons**
carbon-12	**12**	**6**	**6**	**6**
carbon-13	**13**	**6**	**7**	**6**

carbon-14	**14**	**6**	**8**	**6**
Calcium-41	**41**	**20**	**21**	**20**
Chlorine-35	**35**	**17**	**18**	**17**

3.
4. a. Li – 7 b. Cl-35
5. Thompson
6. Proton
7. Rutherford
8. 13