"We are all bold as love." *-Jimi Hendrix*

Chapter 5 Notes - Atomic Structure

HISTORY:

Atomic Theory began with Democritus – 400 BC, Greece
Atom – "un-cuttable" – first particle theory

Aristotle succeeded Democritus, did not share same view of matter, he believed it was continuous

Both ideas had no empirical backing (no experimentation)

By 1800, three fundamental laws discovered:

1) Law of conservation of mass – mass is neither created or destroyed in chemical reactions

2) Law of definite proportions – a chemical compound contains the same proportions by mass regardless of the size of the sample or the source

Ex) water is always 88% oxygen and 12 % hydrogen by mass
NaCl is always 39.34% Na and 60.66% Chlorine by mass

3) **Law of multiple proportions** – if two or more of the same compounds are composed of the same two elements, then the ratio of the masses of the second element combined with a certain mass of the first element is always a ratio of small whole numbers

For example – in CO – 1.00 g of C bonds with 1.33 g of O
In CO_2 – 1.00 g of C bonds with 2.66 g of O

Notice the 1:2 ratio of mass with the oxygens!

John Dalton, English schoolteacher took these laws and formulated an atomic theory:

DALTON'S ATOMIC THEORY (1808) (5 Main Points)

1) All matter is composed of extremely small particles called atoms.
2) Atoms of the same element are identical in size, mass and other properties.
3) Atoms cannot be subdivided, created or destroyed.
4) Atoms of different elements combine in simple, whole-number ratios to form compounds
5) In chemical reactions, atoms are combined, separated or rearranged.

This became a scientific theory because it could be tested, and Dalton himself gathered experimental data.

Parts of Dalton's theory later proven incorrect:

1) All atoms of the same element have the same mass. (It was later shown that different ***isotopes*** of the same element have slightly different masses due to different neutron amounts)
2) Atoms cannot be subdivided. (This happens naturally in the process of nuclear fission and it has been done by man in particle accelerators or atom smashers)

Finding **PERCENT COMPOSITION BY MASS** of compounds:

ex) If 102.90 g sodium bromide always contains 22.99 g of Na and 79.91 g of Br, what percentage of each element is present?

Law of multiple proportions:

If 5.00 g of X combines with14.00 g of Y to form XY_2, how many grams of X and Y would be needed to form:

XY

X_2Y_3

XY_4

5.2 Structure of the Atom

Electron discovered by use of cathode ray tube – Thomson 1897

- Experimentation proved electron has negative charge (deflected by (-) end of magnet)
- Proved electron had mass (turned paddle wheel)

In 1909 American physicist Robert Millikan found mass of **electron** to be 9.109×10^{-31} kg or about 1/1837 the mass of a hydrogen atom.

Since atoms have no net charge, it was deduced that there must be a positive charge to balance out the electron's charge.

Since electrons have such little mass compared to atoms, there must be other massive particles in the atom.

Nucleus of the atom discovered in 1911 by Rutherford.

Gold foil experiment – shot (+) alpha particles (helium nuclei) at THIN gold foil – 1 in 8000 deflected back to source

- concluded that alpha particles must have hit very small, densely packed particles of positive charge (he termed this the nucleus)

Nucleus contains **protons** – mass – $1.673\text{x } 10^{-27}$ kg which is 1836 times as massive as an electron and 1836/1837 of the mass of a hydrogen atom

In all elements after hydrogen (and some larger hydrogen isotopes) all atomic nuclei contain **neutrons** – 1.675×10^{-27} kg – slightly larger than a proton.

Strong forces hold atomic nuclei together – the nuclear strong and weak forces.

Subatomic Particle Table:

Particle	Symbol	Relative Charge	Mass number	Actual mass (amu)

5.3 Distinguishing Between Atoms

Atom type or element determined by the **proton number** or the **atomic number (Z)**

ex)

In **neutral atoms**, the proton(+) number equals the electron (-) number.

ex)

Each element may have a different percentage of **isotopes** – atoms with the same proton number but different **neutron** number

- this causes the mass of the atom to be different
- some isotopes are radioactive - meaning the nucleus is unstable
- this has no effect on the chemical reactivity, charge or identity of the element

ex) Hydrogen's isotopes:

Since neutrons are so massive, different isotopes have different masses, even though they are the same element!

mass number = number of protons and neutrons in a single atom

ex) Oxygen's Isotopes

Average atomic mass - the weighted average mass derived from all of the naturally occurring isotopes

Isotopes of elements in nature exist in certain percentages in the universe - **percent abundance**

ex) Chlorine Isotopes in Nature

mass number	**isotope mass (amu)**	**percent abundance**
35	34.97	75.77
37	36.97	24.23

ex2) Magnesium

mass number	**isotope mass (amu)**	**percent abundance**
24	23.99	78.99
25	24.99	10.00
26	25.98	11.01

Q: Why aren't AMU masses of each isotope in whole numbers like the mass numbers?
A: An amu is exactly 1/12 of the mass of a carbon-12 atom. This amount is not the same as the mass of a proton or a neutron, but very close. (since carbon 12 has 6 protons, 6 neutrons and 6 electrons)

Challenge Problem:

Cu has two stable isotopes, Copper-63 and Copper-65. If the average atomic mass of Cu is 63.55, find the percent abundance of each of its two isotopes.