ATOMIC MODEL

- All forms of matter are made up of PARTICLES
- These PARTICLES may be:
  - atoms (Carbon atoms in diamond)
  - molecules (H_2O molecules in water)
  - ions (Na^+ and Cl^- ions in sodium chloride)
- The particulate nature of matter as we know it today evolved over a very long period of time.

I. EARLY THOUGHTS (ancient Greek philosophers)

- Empedocles (440 B.C.) stated that all matter was composed of "four elements": earth, air, water, and fire.
- Democritus (about 470-370 B.C.) thought that all forms of matter were finitely divisible into very small particles which cannot be divided further. He called these particles atoms.(atomos = indivisible in Greek)
- Democritus idea was a speculative hypothesis which was not based on scientific observations.
- Aristotle (384-322 B.C.) opposed the theory of Democritus, and the existence of atoms was no longer considered until the 17th century

II. DALTON'S ATOMIC THEORY

- John Dalton revived the concept of atoms and proposed an atomic theory based on indirect experimental evidence (1803-1810)
- Dalton's Atomic Theory:
  - explains the difference between an element and a compound
  - explains two scientific laws, and
  - predicts a new scientific law.

DALTON'S ATOMIC THEORY

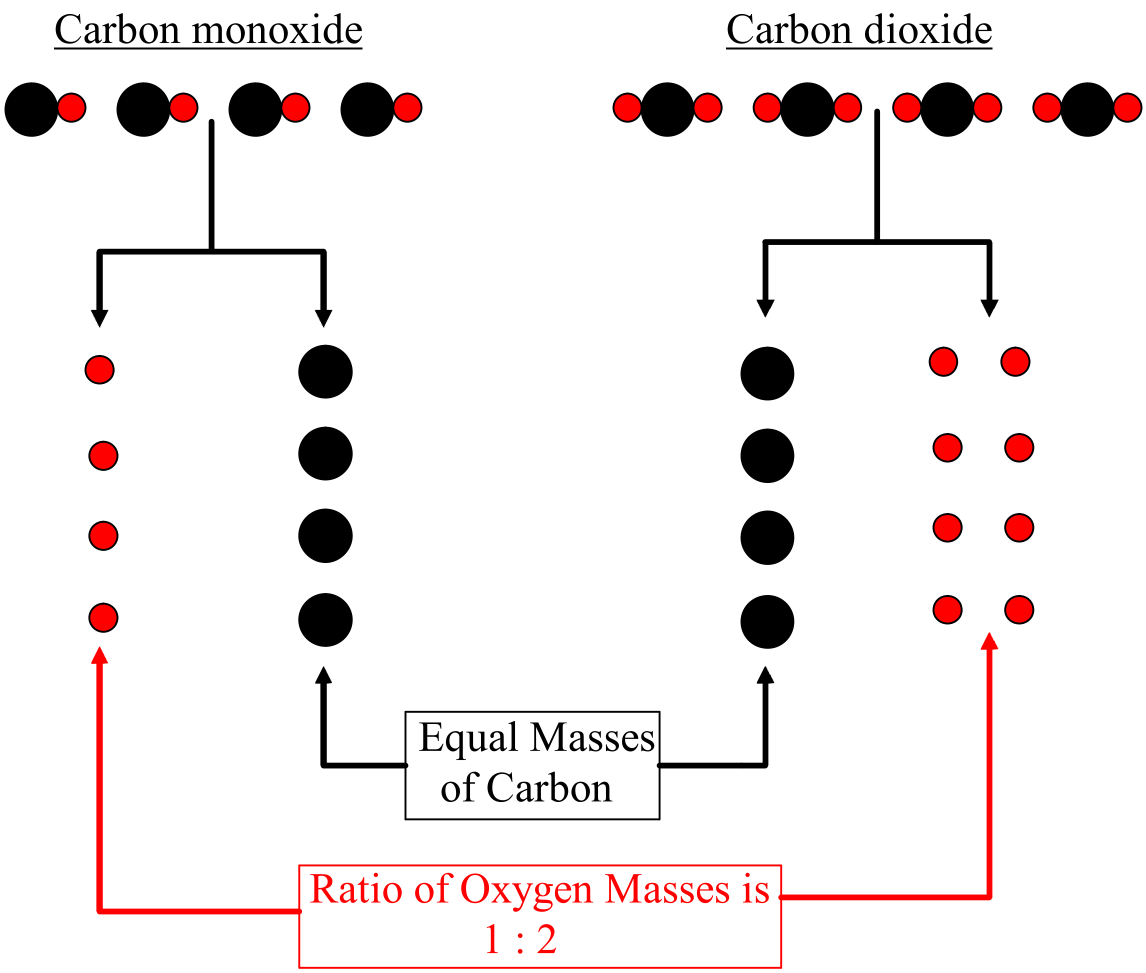
POSTULATES	DEDUCTIONS
• All matter is composed of indivisible particles called atoms. • An atom is a very small particle of matter that retains its identity during a chemical change.	
• An element is a type of matter composed of only one kind of atom. • All atoms of a particular element are identical (e.g. they have the same mass)	• Gives a more precise definition for an element.
• A compound is a type of matter composed of atoms of two or more elements chemically combined in fixed proportions.	• Give a more precise definition for a compound. • Supports the "Law of Definite Composition". • Predicts the "Law of Multiple Proportions.
• A chemical reaction consists of the rearrangement of atoms from one set of combination to another. • Atoms are not created or destroyed, but are rearranged.	• Supports the "Law of Conservation of Mass"

THE LAW OF MULTIPLE PROPORTIONS

- When two elements form more than one compound, the masses of one element in these compounds for a fixed ratio mass of the other element are in ratios of small whole numbers.

Example: CARBON and OXYGEN form two compounds: CO and CO_2

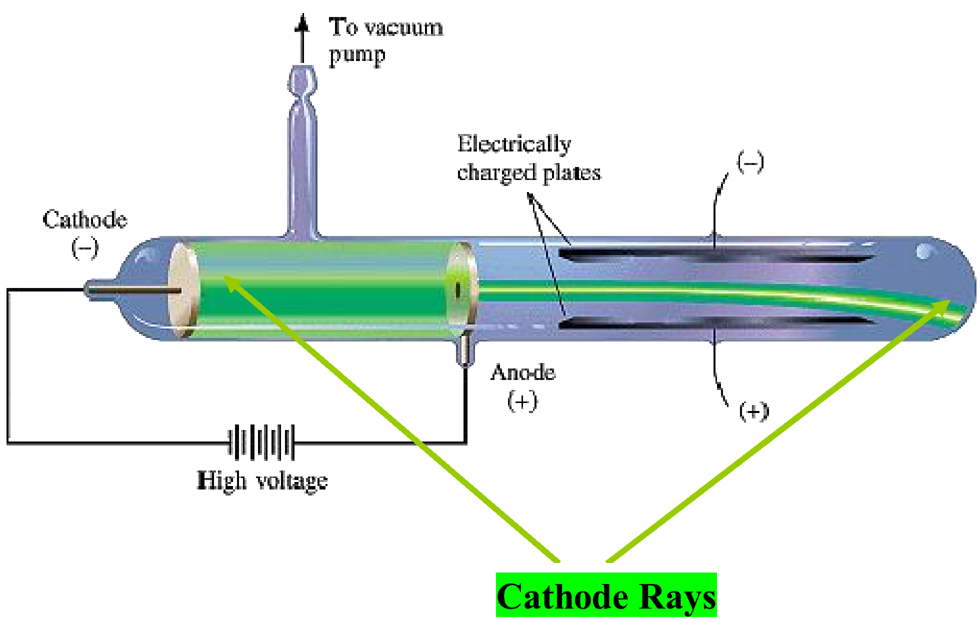
Meaning: The masses of OXYGEN, in these two compounds, for the same mass of CARBON are in a ratio of small whole numbers.



ATOMIC MODEL

III. THE DISCOVERY OF THE ELECTRON

- J. J. Thomson showed that the atoms of any element can be made to emit tiny negatively charged particles.
- He investigated the conductance of gases through glass tubes called "gas discharge tubes"



J.J. Thomson investigated the properties of "CATHODE RAYS" and concluded:

- Cathode Rays are a stream of very small and very light particles emitted the cathode.
- These particles are negatively charged, since they travel from the Cathode (-) toward the Anode (+), and they are repelled by the negative plate of an electric field.
- These particles are present in any form of matter (No matter what metal was used as a Cathode, each emitted Cathode Rays).
- J.J. Thomson called these particles "electrons"

Conclusions:

The atom is not the smallest particle of matter.
Atoms contain smaller particles, such as the electron.

ATOMIC MODEL

IV. The Discovery of the Proton

- Protons were first observed by E. Goldstein in 1886
- William Thomson (better known as Lord Kelvin, and no relation to J.J. Thomson) investigated the properties of the proton.
- He concluded:

    1. The proton is a positively charged particle, whose electrical charge is equal in magnitude but of opposite sign to that of the electron
    2. The mass of the proton is about the same as the mass of the H atom (1840 times larger than the mass of the electron)

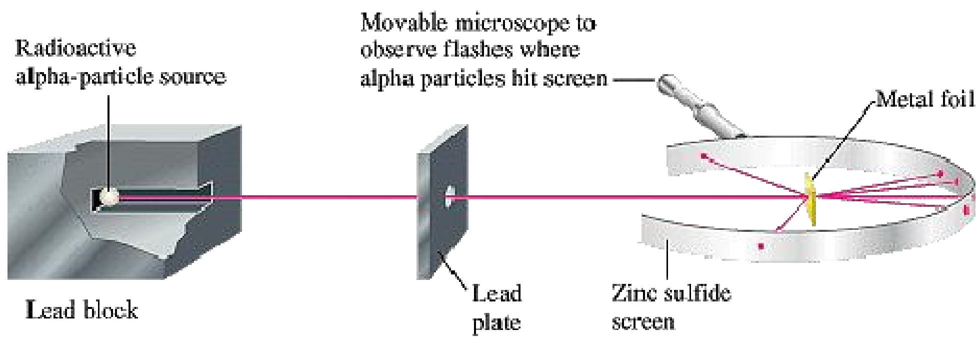
Conclusion:

The atoms contained both negatively and positively charged particles.

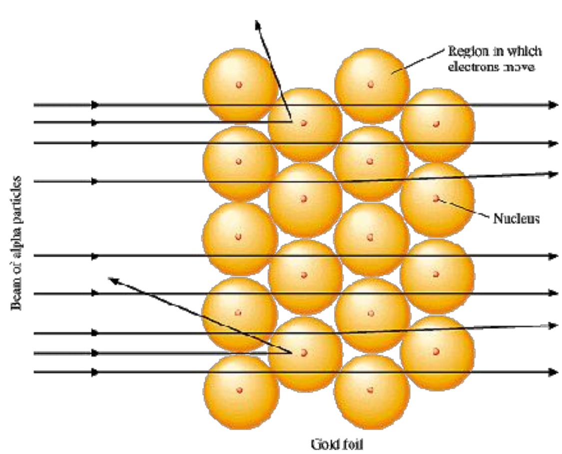
ATOMIC MODEL

V. The Discovery of the Nucleus

- Rutherford investigated the penetrating power of a particular type of radioactive rays (Alpha rays) through thin metallic foils
- Alpha (α) rays consist of a stream of particles (α particles) which:
    - are positively charged
    - are relatively heavy (four times as heavy as the H atom)



OBSERVATIONS	INTERPRETATIOINS
• Most of the α-particles pass through the gold foil.	• The gold atom is mostly occupied by lightweight electrons that cannot deflect the α-partilces.
• A few α-particles (1 in 8000) were scattered at large angles and sometimes deflected back.	• A very small part of the atom is dense (contains most of its mass) and is positively charged.



ATOMIC MODEL

Conclusion:

- The **nuclear model** of the atom contains a nucleus and electrons surrounding them.

NUCLEUS	ELECTRONS
• Located at the center of the atom • Contains all the protons • Is positively charged • Contains most of the mass of the atom • Is very small • Is very dense	• Are located outside the nucleus • Are negatively charged • Have negligible mass • Occupy most of the volume of the atom

VI. The Discovery of the Neutron

- James Chadwick discovered in 1932 that the atom contains a third subatomic particle called neutron.
- The **neutron:**
  - is located in the nucleus
  - is electrically neutral
  - has about the same mass as the proton

Properties of Subatomic Particles

Particle	Relative Mass	Electrical Charge
Electrons	0	–1
Proton	1	+1
Neutron	1	0

- The nucleus has a positive charge (contains protons and neutrons)
- The positive charge of the nucleus is referred to as **NUCLEAR CHARGE** and is abbreviated **Z**.
- The NUCLEAR CHARGE can be determined experimentally and is equal to the number of protons and the number of electrons (in neutral atoms)
- The mass of an atom is the sum of the masses of the protons and of the neutrons

A = Mass Number = Number of protons + Number of neutrons

ATOMIC STRUCTURE

- Any nucleus is characterized by:

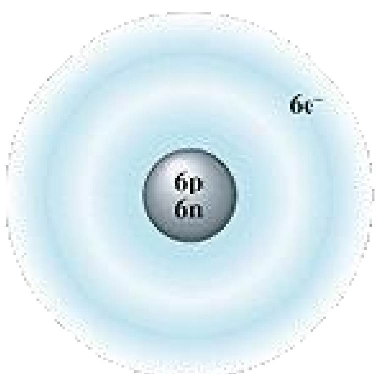
Z = Nuclear Charge = number of protons
A = Mass Number = number of protons + number of neutrons

- A NUCLIDE, is an atom, characterized by:
    - a definite nuclear charge, Z (atomic number)
    - a definite Mass Number, A

- Symbolic Representation of a Nuclide:

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

For example: A Carbon (C) nuclide: $^{12}_{6}C$ (Z = 6 A = 12)



Examples:
Determine the subatomic particles for the nuclides indicated below:

1. $^{208}_{82}Pb$ p^+ = n^0 = e^- =

2. $^{137}_{56}Ba$ p^+ = n^0 = e^- =

ISOTOPES

- Atoms with the same number of protons (same Atomic Number, Z) but different number of neutrons (different Mass Number, A) are called **isotopes**.
- Most of the elements occur in nature as mixtures of isotopes.
- Naturally occurring copper is found to contain 2 isotopes:

Isotope	$^{63}_{29}\mathbf{Cu}$	$^{65}_{29}\mathbf{Cu}$
Nuclear Particles	29 protons 34 neutrons	29 protons 36 neutrons
Relative Abundance	69.09 %	30.91 %
Relative Atomic Mass	62.9298 amu	64.9278 amu
Average Atomic Mass	63.55 amu	

Isotopes & Atomic Mass:

- The mass of an atom is measured relative to the mass of a chosen standard (carbon-12 atom), and is expressed in atomic mass units (amu).
- The average atomic mass of an element is the **weighted average** for the mass of all the naturally occurring isotopes of that element expressed in atomic mass units (amu).
- Therefore the atomic mass of an element is closest to the mass of its most abundant isotope.
- As an example, the two naturally occurring isotopes of copper are listed below:

Isotope	copper-63	copper-65
Relative Abundance	69.09 %	30.91 %
Relative Atomic Mass	62.9298 amu	64.9278 amu

- The average atomic mass of copper can be calculated as follows:

$$\begin{Bmatrix}\text{Atomic mass}\\ \text{of an element}\end{Bmatrix} = \left\{\begin{pmatrix}\text{Abundance}\\ \text{of isotope 1}\end{pmatrix} \text{x} \begin{pmatrix}\text{Mass of}\\ \text{isotope 1}\end{pmatrix}\right\} + \left\{\begin{pmatrix}\text{Abundance}\\ \text{of isotope 2}\end{pmatrix} \text{x} \begin{pmatrix}\text{Mass of}\\ \text{isotope 2}\end{pmatrix}\right\}$$

$$\text{Avg atomic mass} = (62.9298\text{ amu})\left(\frac{69.09}{100}\right) + (64.9278\text{ amu})\left(\frac{30.91}{100}\right)$$

$$= 43.48\text{ amu} + 20.07\text{ amu}$$

$$= 63.55\text{ amu}$$

Examples:

1. Which of the structures shown below represent isotopes?

$$^{207}_{82}X \qquad ^{207}_{83}Y \qquad ^{209}_{84}Z \qquad ^{209}_{82}R$$

2. An element has three naturally occurring isotopes with the following masses and abundances. Determine the average atomic mass and identity of this element.

Isotope Mass (amu)	**Abundance (%)**
27.977	92.21
28.976	4.70
29.974	3.09

3. Two samples of different compounds of sulfur and oxygen have the following composition. Show that the compound follows the law of multiple proportions. What is the ratio of oxygen in the two compounds for a fixed amount of sulfur?

	Amount S	**Amount O**
Compound A	1.210 g	1.811 g
Compound B	1.783 g	1.779 g