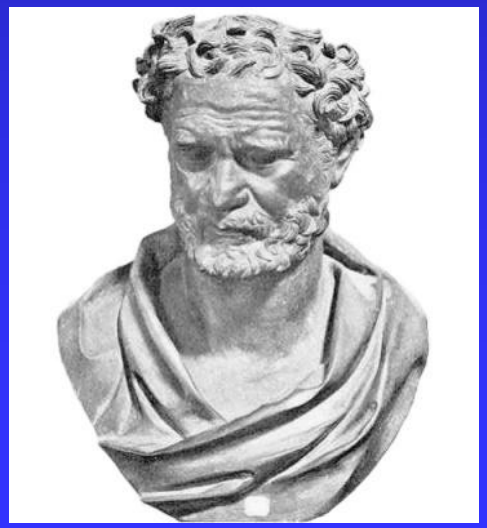
Democritus

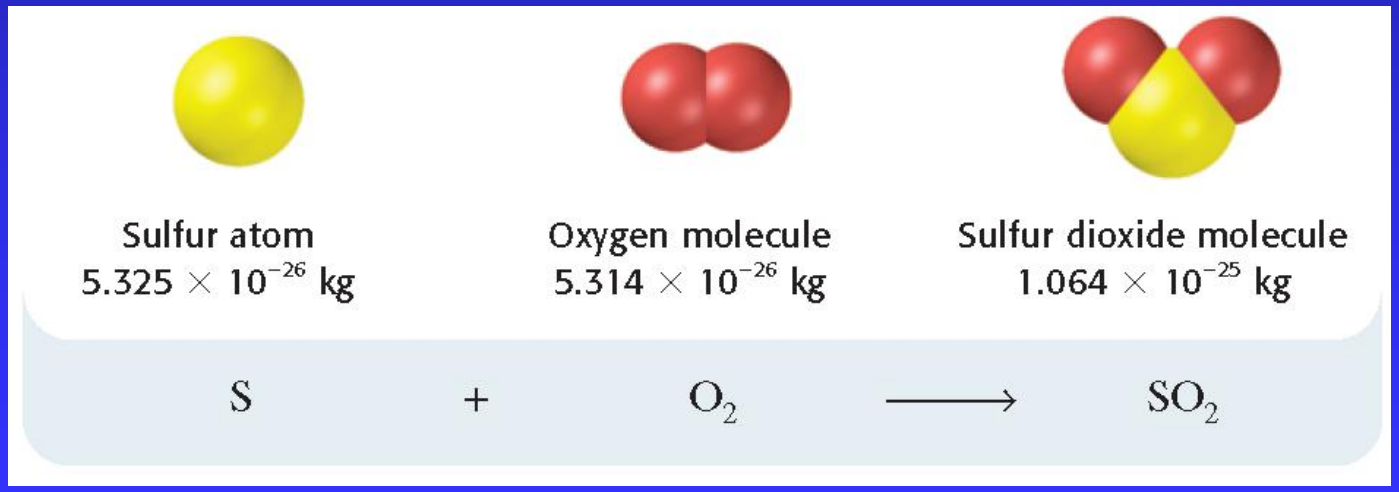
- 1st person to propose that matter was not infinitely divisible
- Called the small pieces of matter "atomos"
- These "atomos" are unique to the specific substance
- Aristotle had his own theory and because of his influence Democritus' idea was rejected



Law of Conservation of Mass

Mass can neither be created nor destroyed in a physical or chemical change

$$\text{Mass}_{\text{reactant}} = \text{Mass}_{\text{products}}$$



Law of Definite Proportions

Regardless of amount, a compound is always composed of the same elements in the same percent by mass

$$\% \text{ by mass} = \frac{\text{Mass}_{\text{element}}}{\text{Mass}_{\text{compound}}} \times 100$$

Sodium Chloride – NaCl

Sodium = 39.34%

Chlorine = 60.66%

100.0 g sample of NaCl

39.34 g Na $\frac{39.34\ g}{100.0\ g} \times 100 = 39.34\%$

60.66 g Cl $\frac{60.66\ g}{100.0\ g} \times 100 = 60.66\%$

450.0 g sample of NaCl

177.03 g Na $\frac{177.03\ g}{450.0\ g} \times 100 = 39.34\%$

272.97 g Cl $\frac{272.97\ g}{450.0\ g} \times 100 = 60.66\%$

Law of Multiple Proportions

When two different compounds are formed by a combination of the same elements, their masses combine in small whole number ratios

Name of compound	Description	As shown in figures	Formula	Mass O (*g*)	Mass N (*g*)	$\frac{\text{Mass O (g)}}{\text{Mass N (g)}}$
Nitrogen monoxide	colorless gas that reacts readily with oxygen		NO	16.00	14.01	$\frac{16.00\text{ g O}}{14.01\text{ g N}} = \frac{1.14\text{ g O}}{1\text{ g N}}$
Nitrogen dioxide	poisonous brown gas in smog		NO_2	32.00	14.01	$\frac{32.00\text{ g O}}{14.01\text{ g N}} = \frac{2.28\text{ g O}}{1\text{ g N}}$

Development of Atomic Theory

- Democritus
- John Dalton
- JJ Thomson
- Robert Millikan
- Earnest Rutherford
- James Chadwick

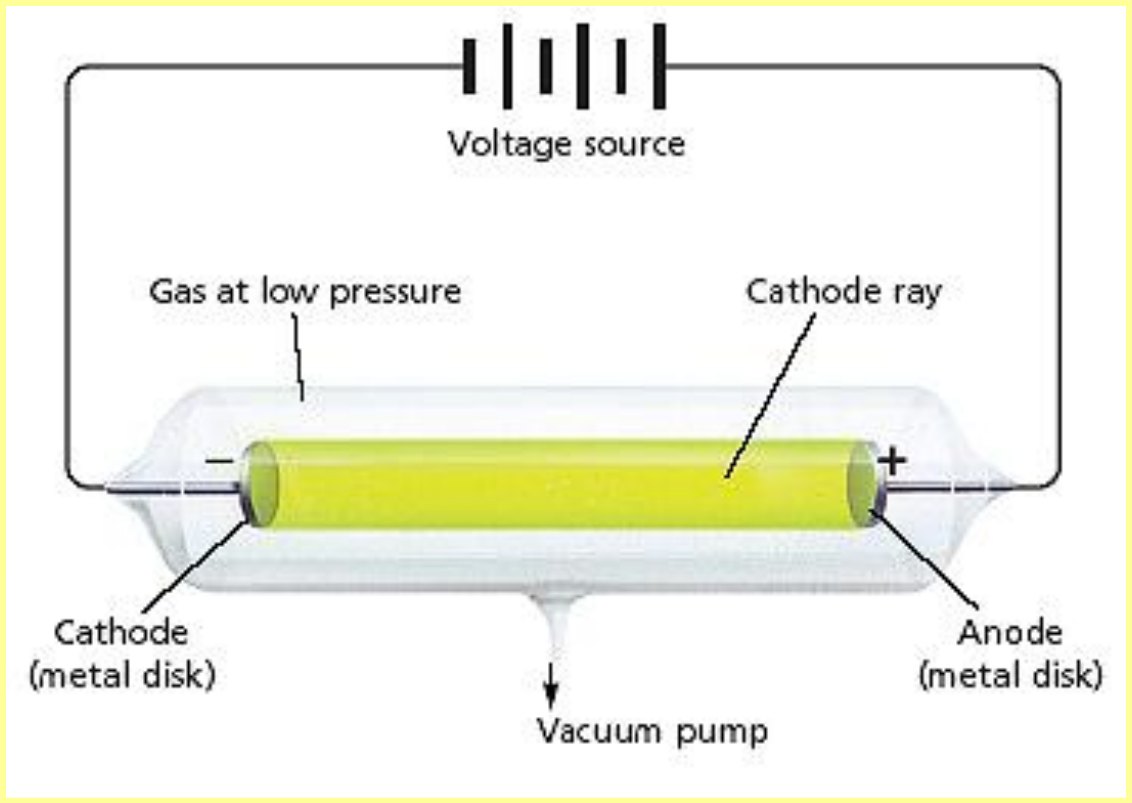
Dalton's Atomic Theory

- All matter is composed of extremely small particles called atoms (smallest particle of an element that retains the chemical prop. of that element).
- Atoms of a given element are identical in size and mass; atoms of different elements differ in size and mass
- Atoms cannot be subdivided, created, or destroyed.
- Atoms of different elements combine in simple whole-number ratios to form chemical compounds.
- In chemical reactions, atoms are combined, separated, or rearranged

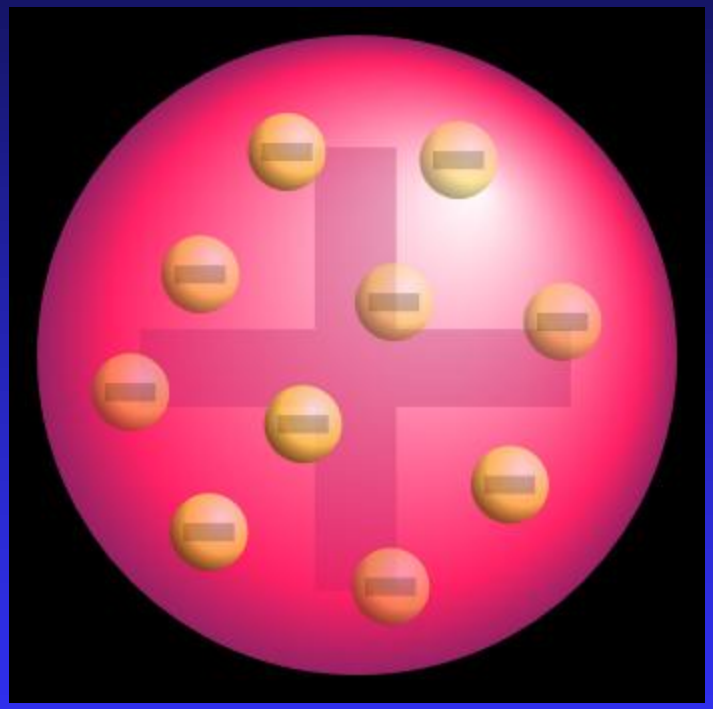
* later disproven

J.J. Thomson

- Discovered the electron from his work with cathode ray tubes.

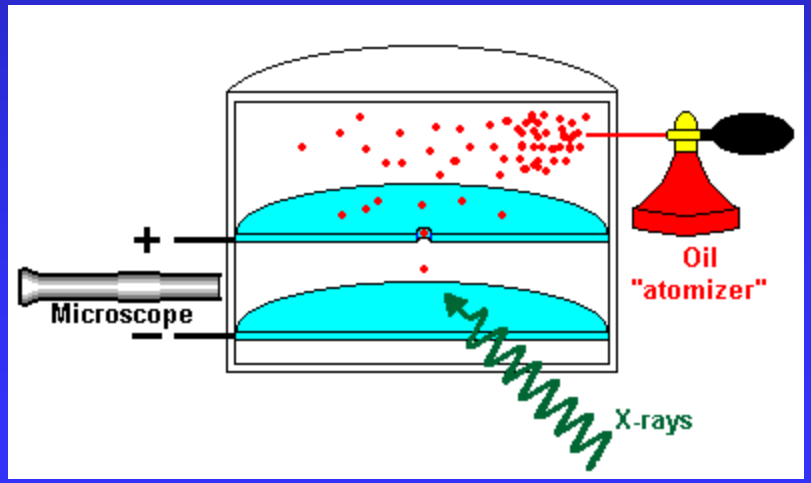


J.J Thomson's Plum-pudding Model (Chocolate Chip Cookie Model)



Robert Millikan

- Measured the charge of an electron (-1) and helped find the mass of an electron with his oil drop experiment
- Mass of an electron = 9.11×10^{-31} kg ($1/1840^{th}$ the mass of the hydrogen atom)



Ernest Rutherford Nuclear Atomic Model

- Proved Thomson's plum-pudding model incorrect
- Gold Foil Experiment - Discovered the nucleus.
- Positively charged nucleus at the center of the atom contain positively charged protons and the electrons revolve around the nucleus

James Chadwick

- Colleague of Ernest Rutherford
- Discovered the neutron.
- Neutrons help stabilize the protons in the nucleus.
- Neutrons are about the same size as protons. But have no charge

Atomic Number

- Atoms of the same element all have the same number of protons. Protons identify an element.
- The **atomic number** (Z) is the number of protons (p^+) of each atom of that element.

# of protons = # of electrons in a neutral atom

Isotopes

- **Isotopes** are atoms of the same element that have different masses.
- The isotopes of a particular element all have the same number of protons and electrons but <u>different numbers of neutrons.</u>

Mass Number

- Represents the sum of the protons and neutrons in the nucleus of a specific isotope.
- Hyphen notation: The mass number is written with a hyphen after the name of the element.

uranium-235

- Symbol: The superscript indicates the mass number and the subscript indicates the atomic number.

$$^{235}_{92}U$$

Relative Atomic Mass

- The carbon-12 atom is used as a standard by scientists to compare units of atomic mass.

  carbon-12 = 12 atomic mass units, or 12 amu.

  (each neutron and proton is approximately 1 amu)

- The atomic mass of any atom is determined by comparing it with the mass of the carbon-12 atom.

Average Atomic Mass

- A *weighted average* of all the naturally occurring isotopes of an element

Avg atomic mass = (isotope 1 mass x abundance 1) + (isotope 2 mass x abundance 2) + (isotope 3 mass x abundance 3) ...

Example: Find the atomic mass of boron using the following isotopes.

boron-10: mass = 10.01 amu, % abundance = 19.80 %

boron-11: mass = 11.01 amu, % abundance = 80.20 %

Answer :

(10.01 x .1980) + (11.01 x .8020) = 10.81 amu