Atomic Theory & Isotopes

Chapter 2

Theories of Matter Composition

Democritus

(5^{th} – 4^{th} century BC)

ATOMISM

Aristotle

(4^{th} – 5^{th} century BC)

CONTINUOUS MATTER

FOUR ELEMENTS – Earth, Air, Fire, Water

Boyle

(17^{th} century)

Reintroduced ATOMISM in modern times.

Dalton

(19^{th} century)

Atomic Theory to explain results of EXPERIMENTS.

Development of the Modern Atomic Theory

In 1803, John Dalton proposed an atomic theory that is still the basis for many of our theories about the atom.

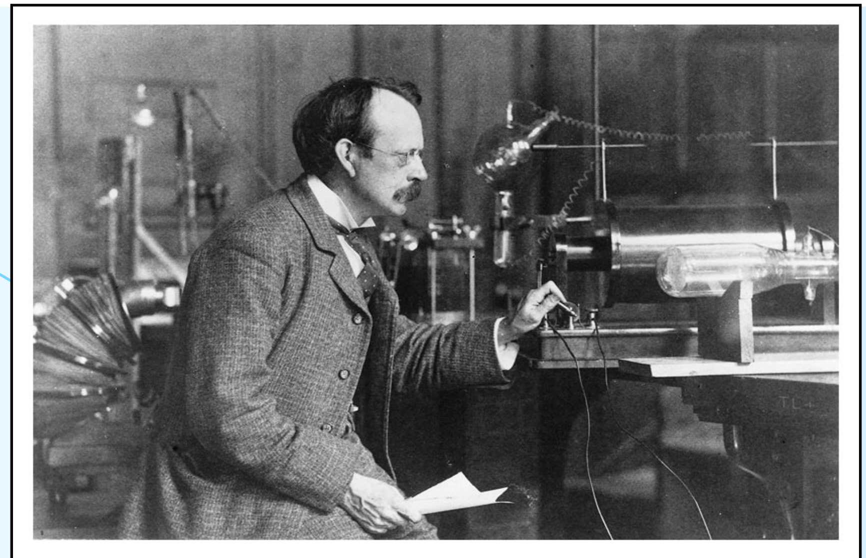
1. All matter is composed of atoms, which are tiny, indivisible particles.

2. A chemical reaction is a rearrangement of atoms to form different compounds. Atoms are neither created nor destroyed in a chemical reaction (the law of conservation of mass).

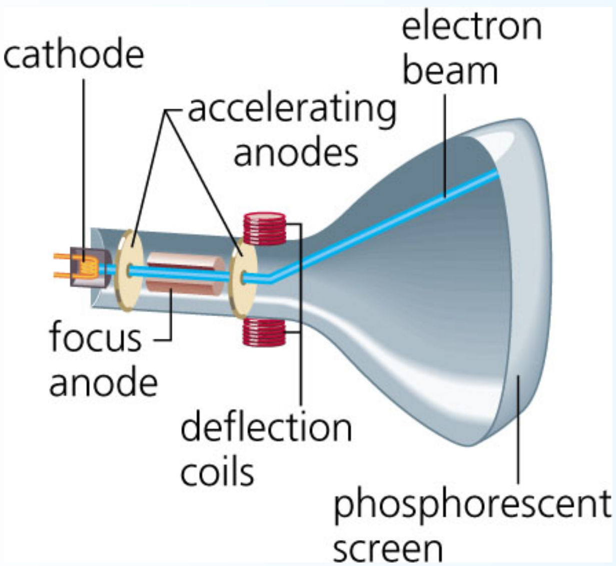
3. Atoms of one element cannot be converted into another element. Atoms of an element are identical in mass and other properties, and are different from every other element.

4. A compound is a combination of atoms of two or more elements in specific ratios (the law of definite composition).

J.J. Thompson and the Discovery of the Electron.

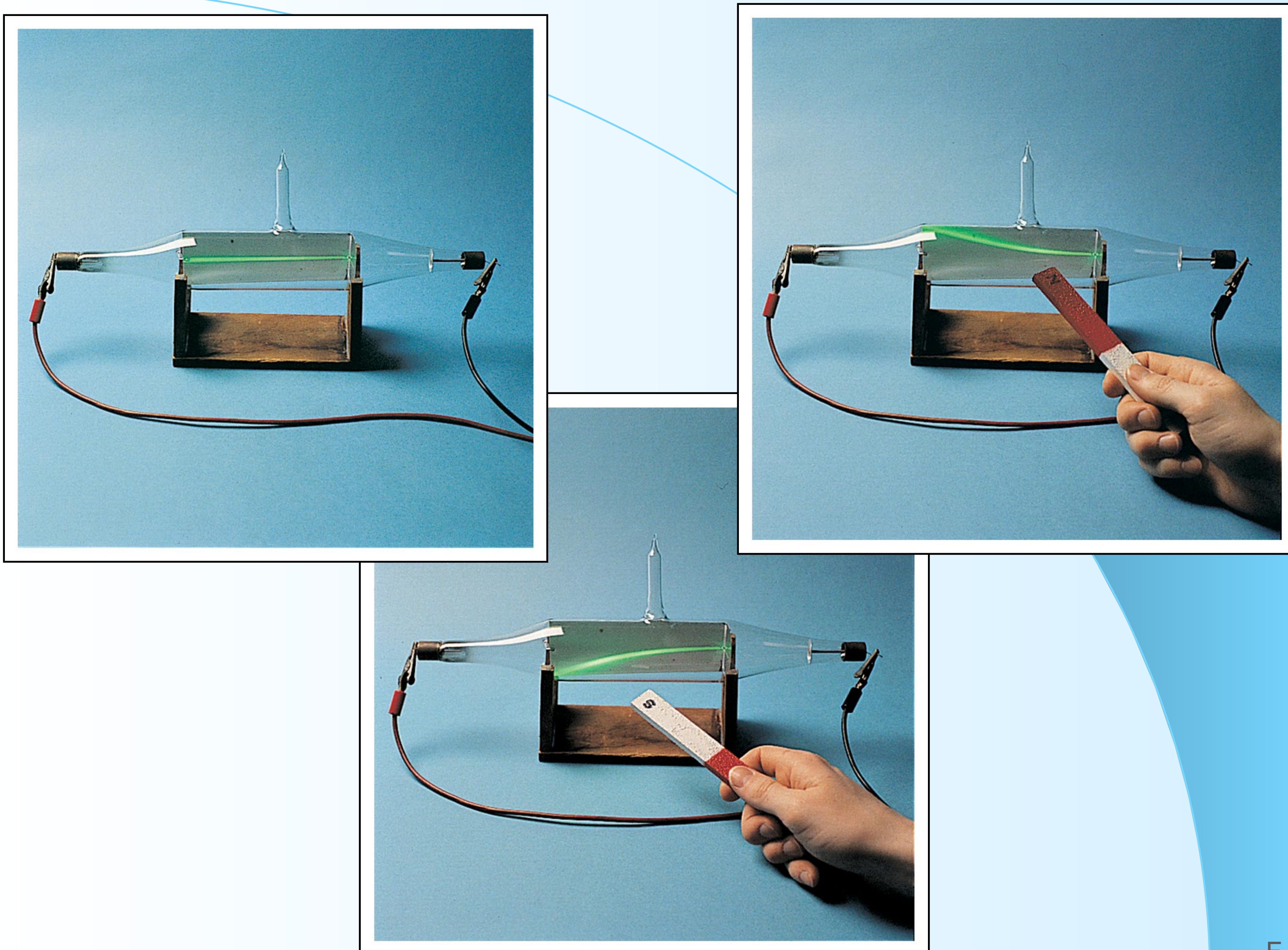


Cathode Ray Tube:



Electron: Very small, subatomic, negatively charged particle.

From his experiments, Thompson calculated the ratio of the electron's mass, $\boldsymbol{m_e}$, to its electric charge, **e**.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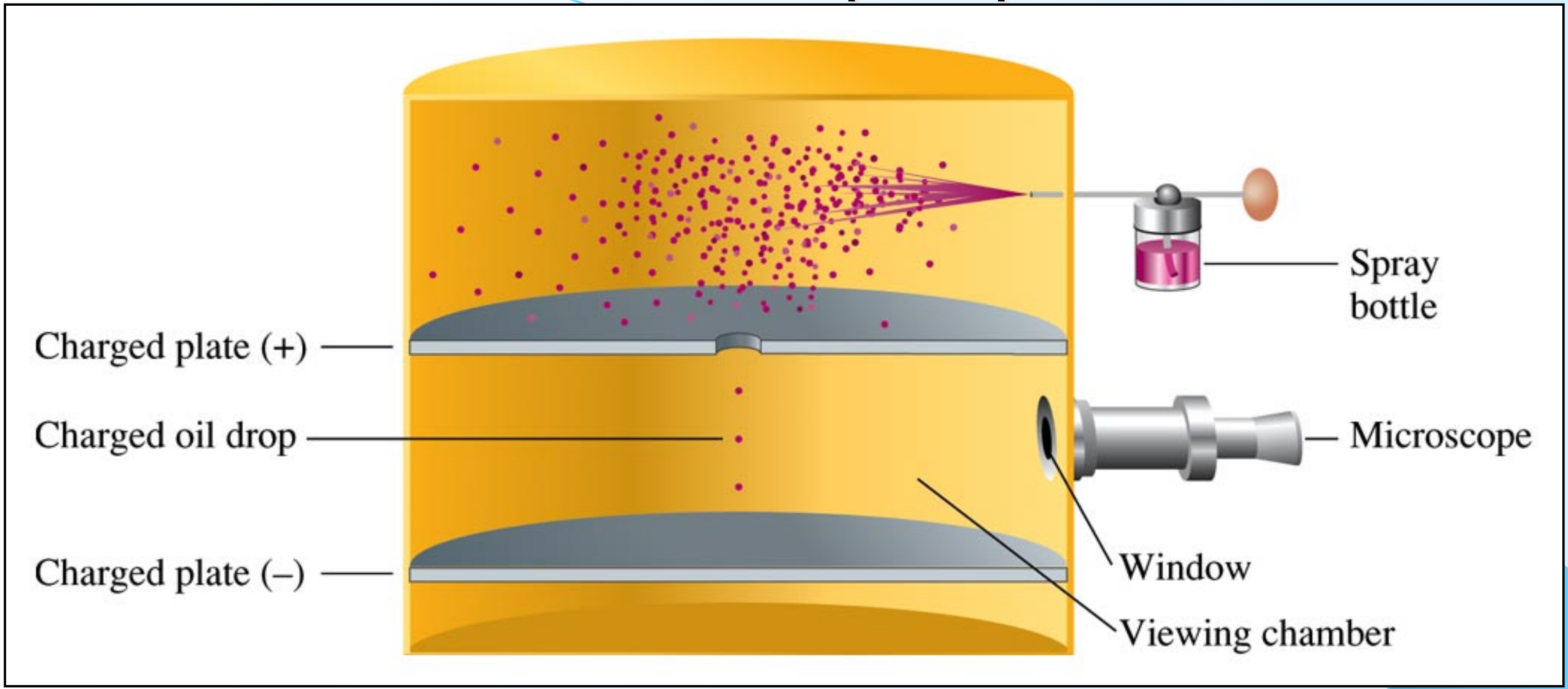
Millikan's Oil-drop Experiment

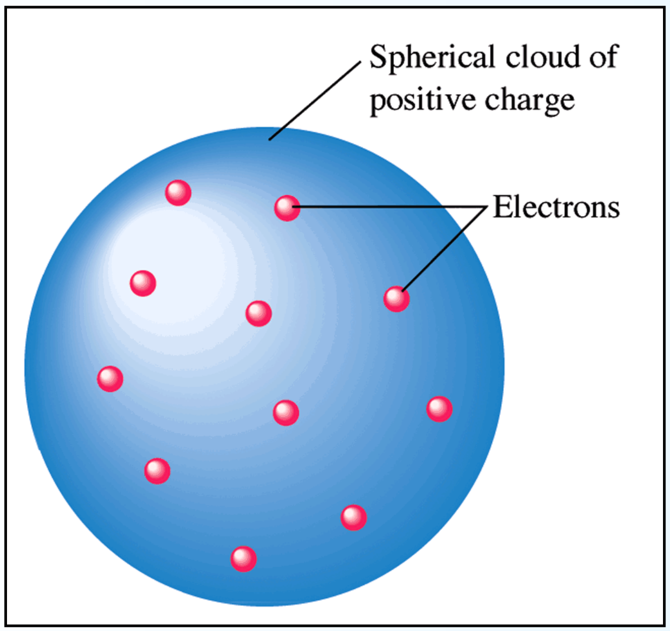


The work of Thomson and Millikan combined provided us with the electron’s mass of **9.109 x 10-31 kg**, which is more than 1800 times smaller than the mass of the lightest atom (hydrogen).

These experiments also showed that the electron is a subatomic particle.

Thomson's Atomic Theory:

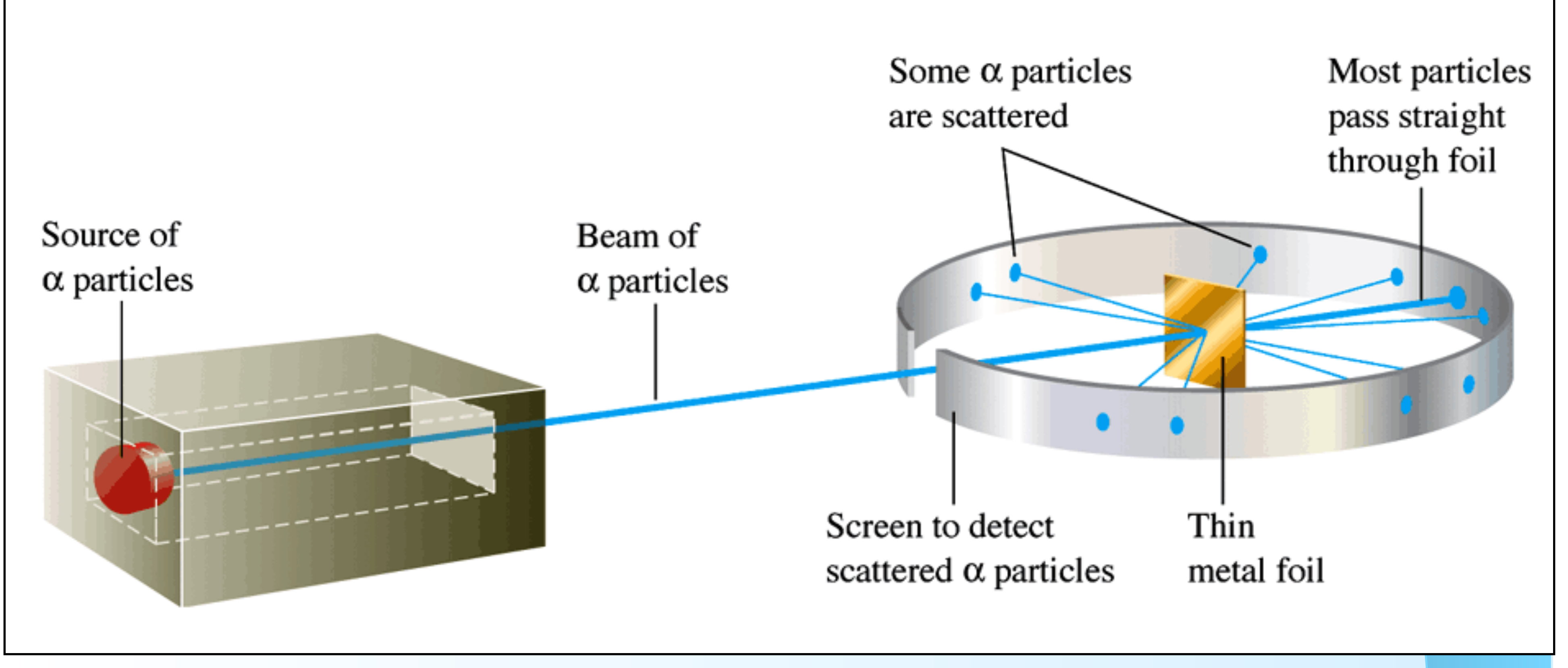
- JJ Thomson favored the ***plum pudding model***. He believed that the atom is a large mass of positive charge, with tiny electrons moving around in circles on the surface.



- Ernest Rutherford, a student of Thomson, was working with alpha particles, large positively charged particles.
- He and his students (Geiger & Morrison) found that the behavior of alpha particles was not consistent with Thomson's model.

Rutherford's Gold Foil Experiment

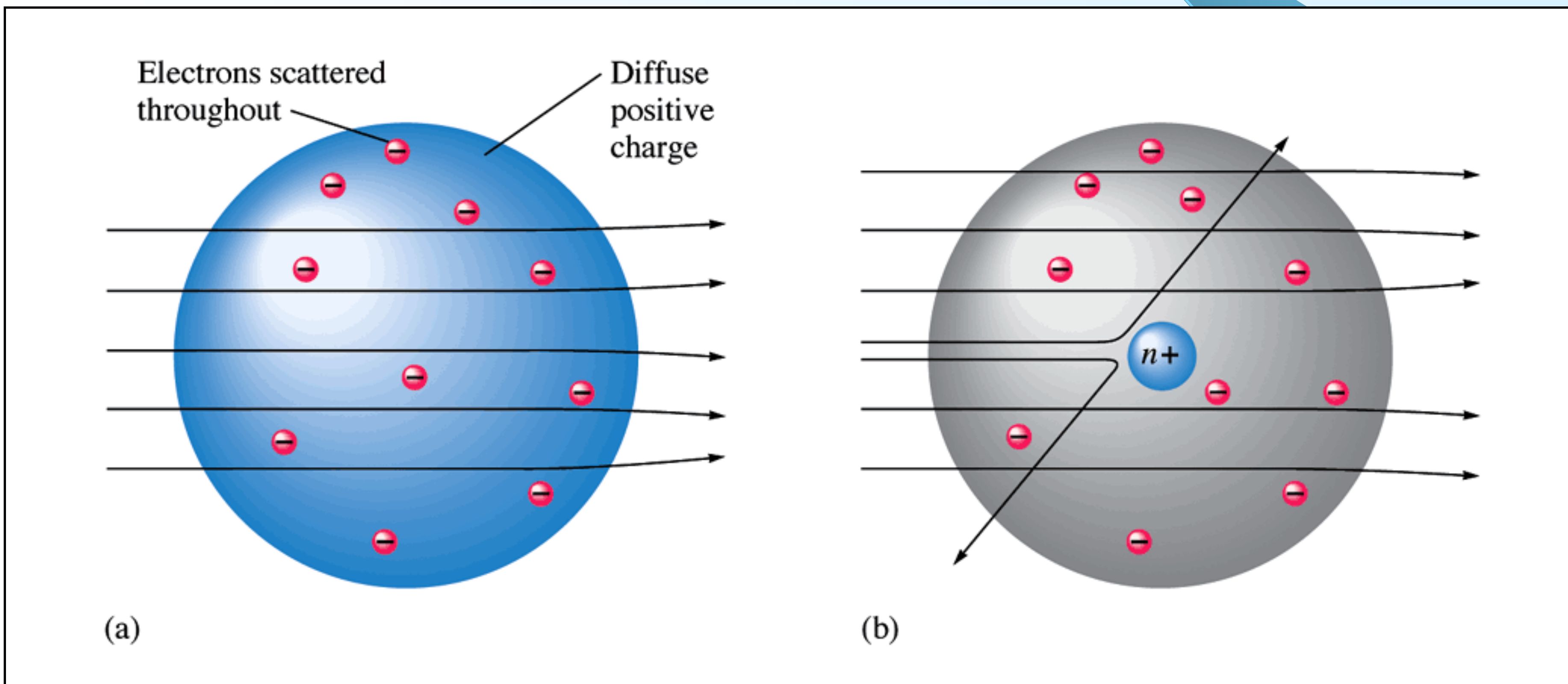
- Rutherford used alpha particles to further study the atomic model.
- He directed alpha particles from a source at thin pieces of metal foil that were a relatively small number of atoms thick.



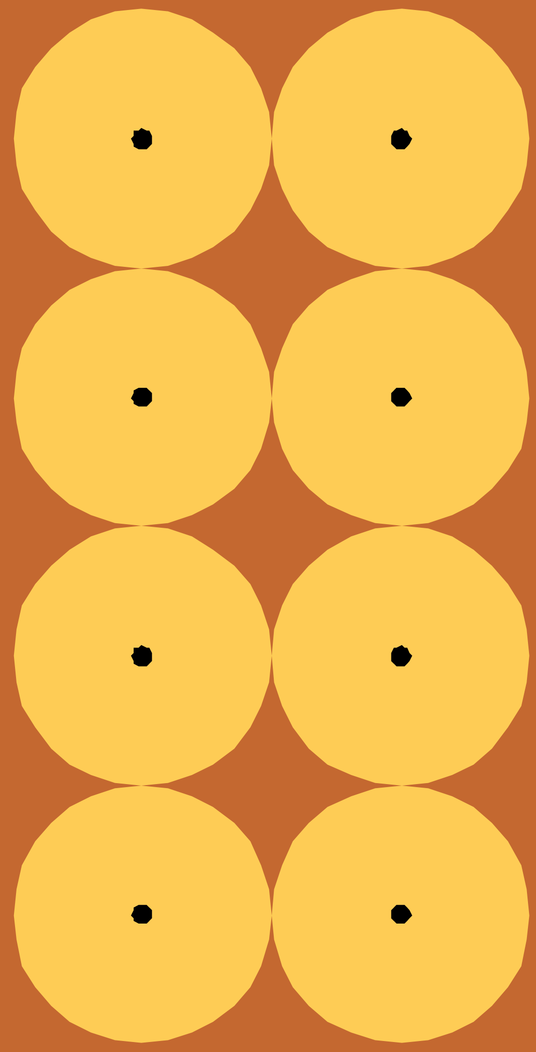
Rutherford's Experiment

(a) The expected results of the metal foil experiment if the diffuse model were correct.

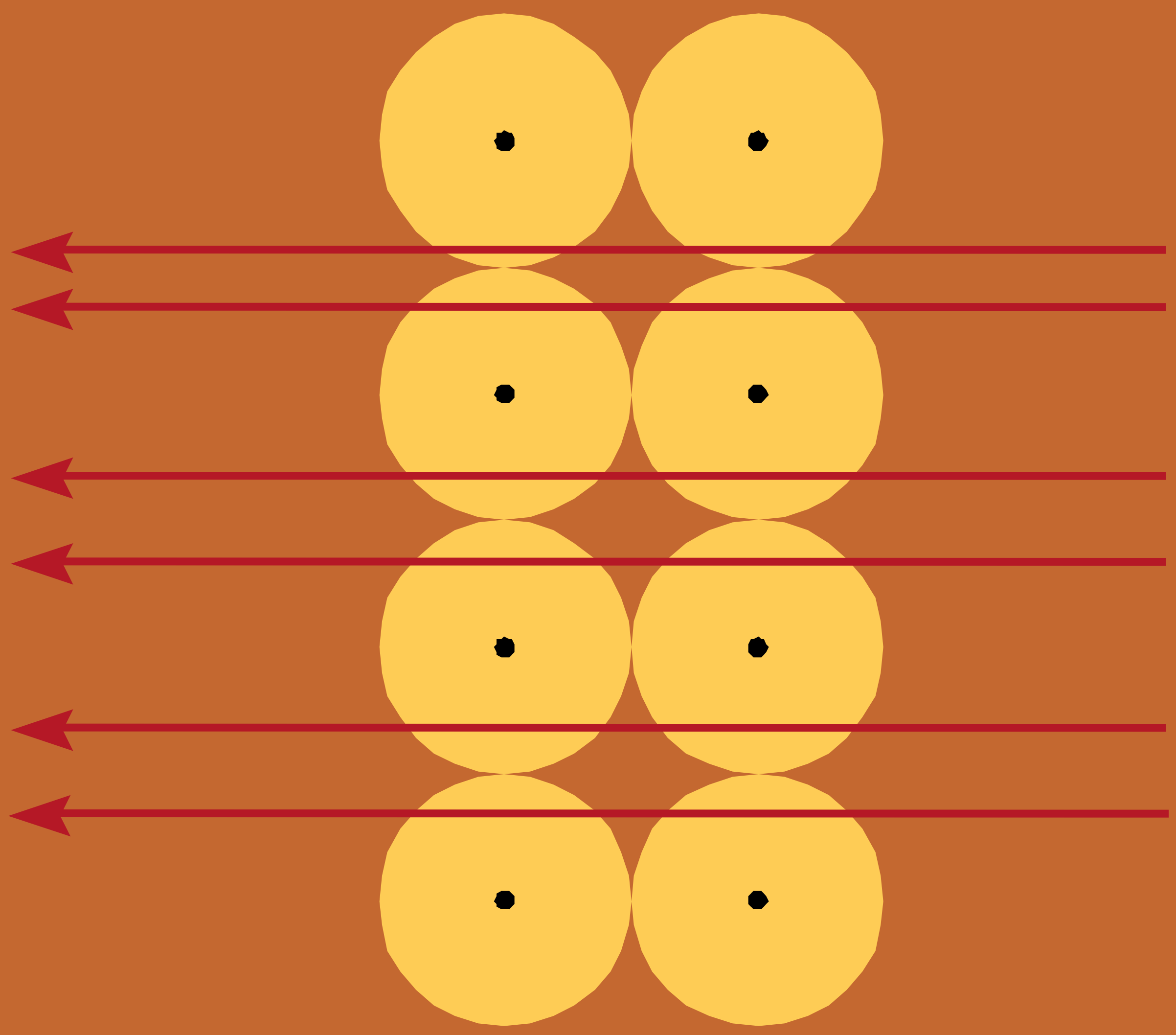
(b) Actual results.



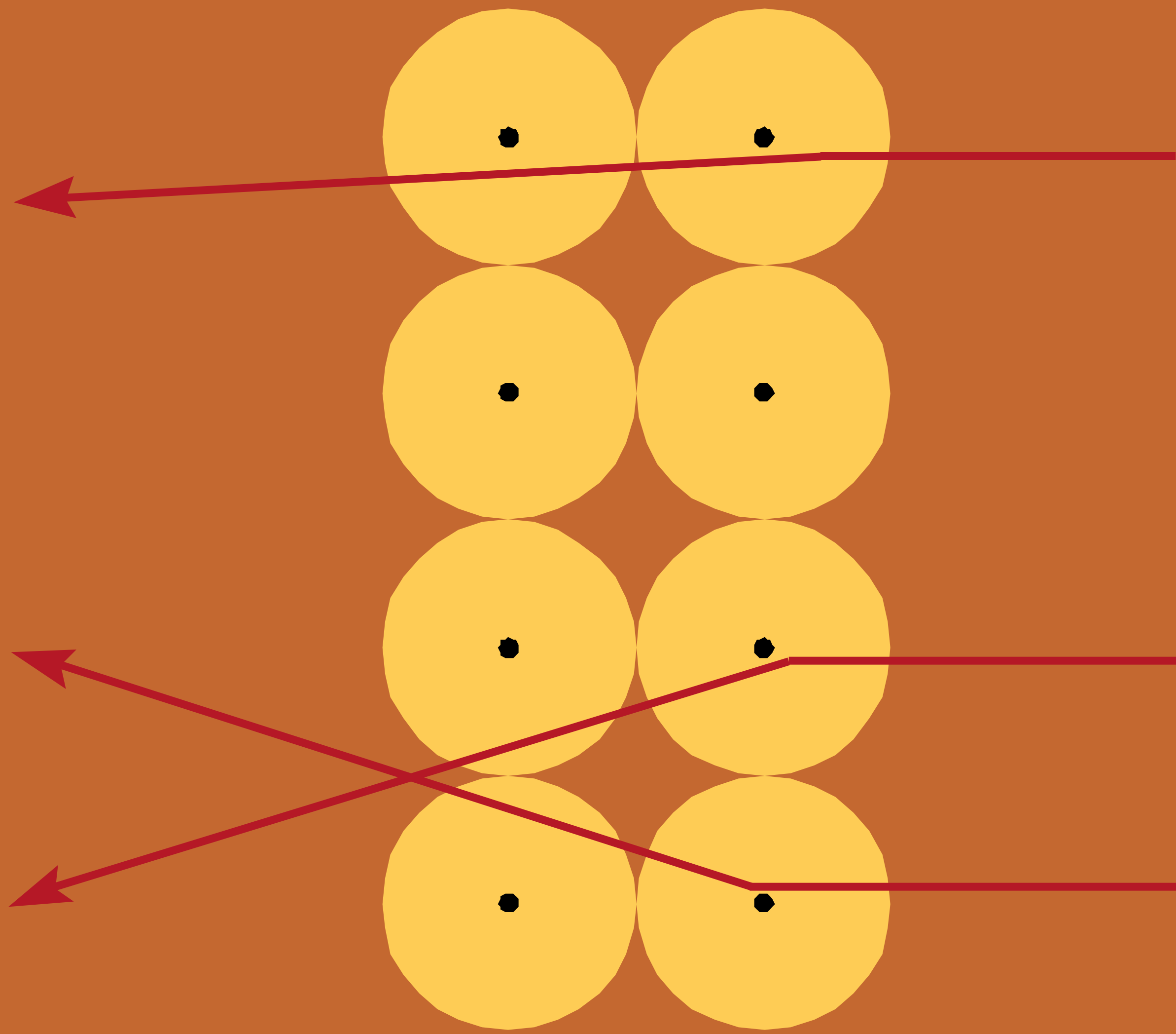
Rutherford’s Experimental Design



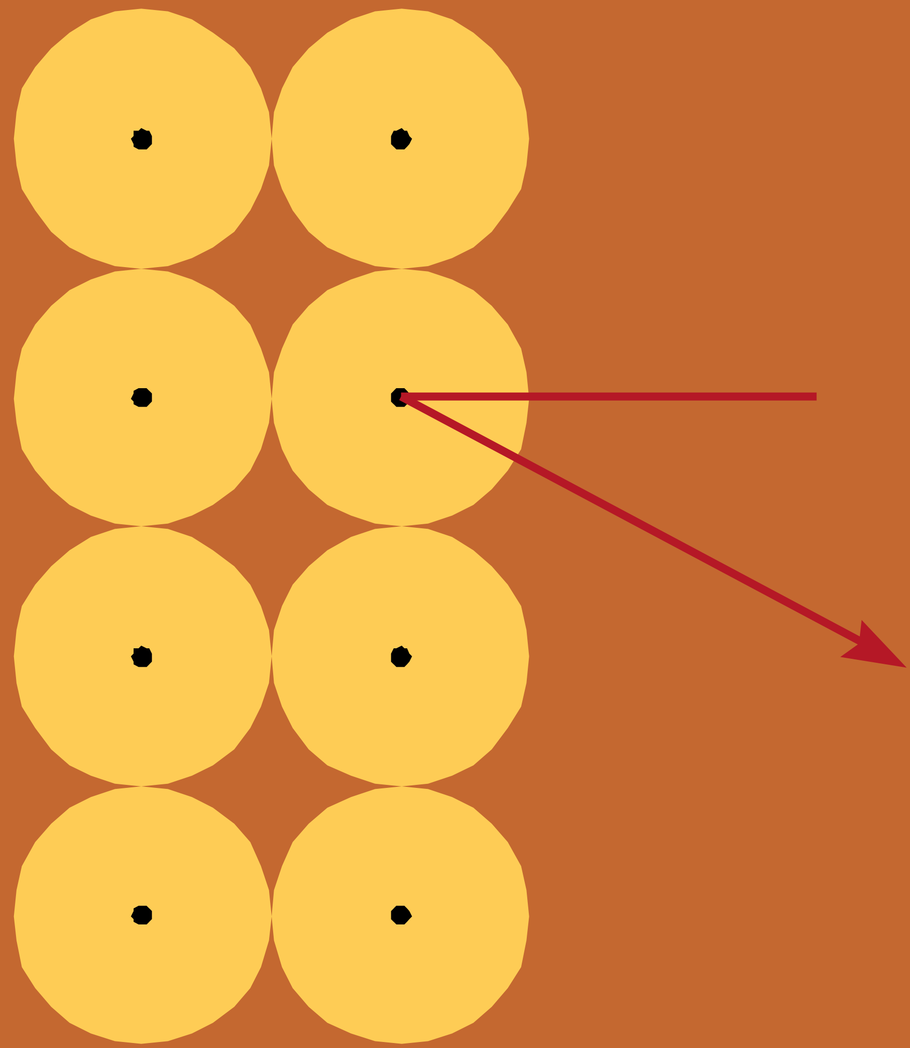
Rutherford's Experimental Design



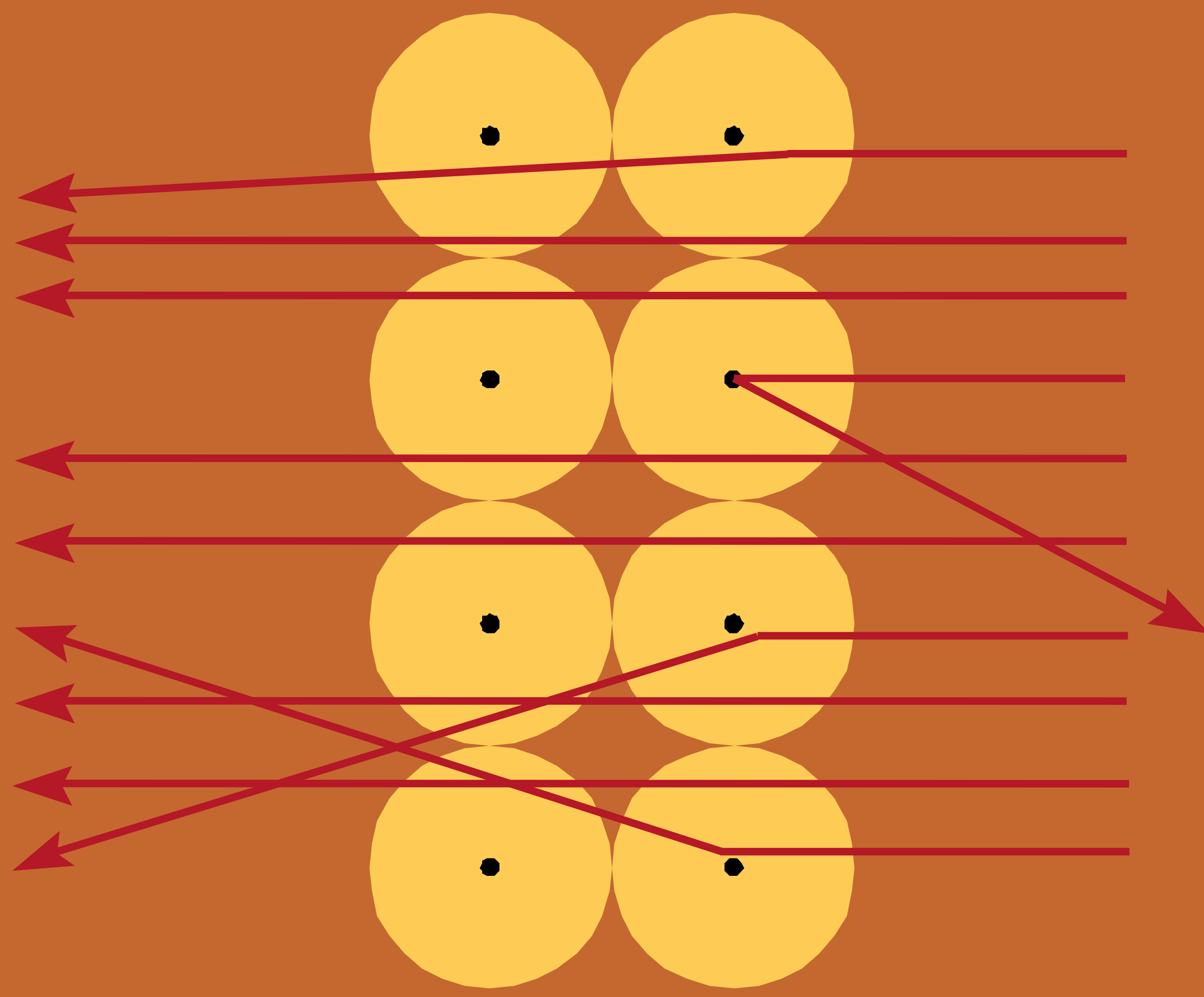
Rutherford’s Experimental Design



Rutherford's Experimental Design



Rutherford's Experimental Design



Rutherford's model of the atom:

An atom is made up of a small, central, positively charged nucleus, surrounded by even smaller, negatively charged electrons that are moving around the atom.

NUCLEUS:

1) Subatomic particle: Core of the atom.
2) Positively charged.
3) Takes up very little space in the atom.
4) Contains most of the mass of the atom.

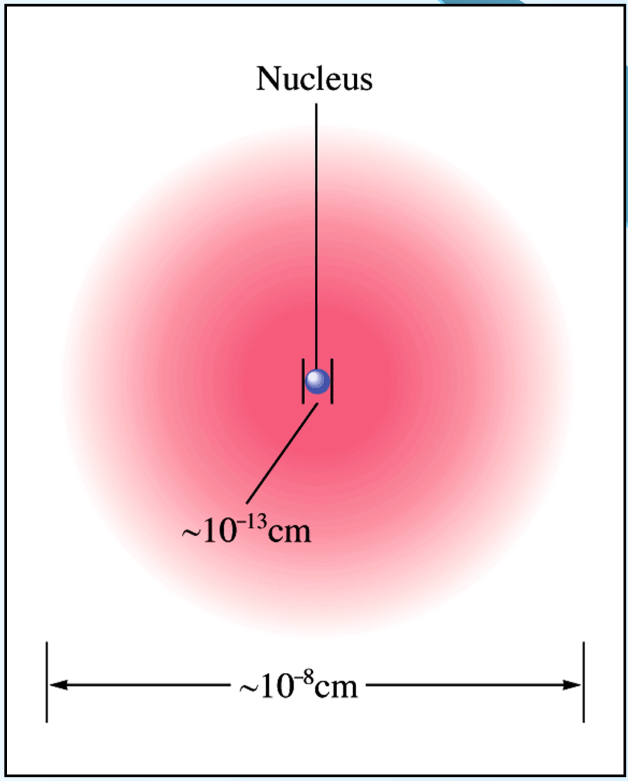


Figure:

A nuclear atom viewed in cross section.

(Note that this drawing is not to scale.)

The Modern Model of the Atom

Atoms are made up of three subatomic particles:

Proton (p^+): Positively charged subatomic particle found in the nucleus of the atom.

Neutron (n^0): Neutral subatomic particle found in the nucleus of the atom.

Electron (e^-): Negatively charged subatomic particle found outside the nucleus.

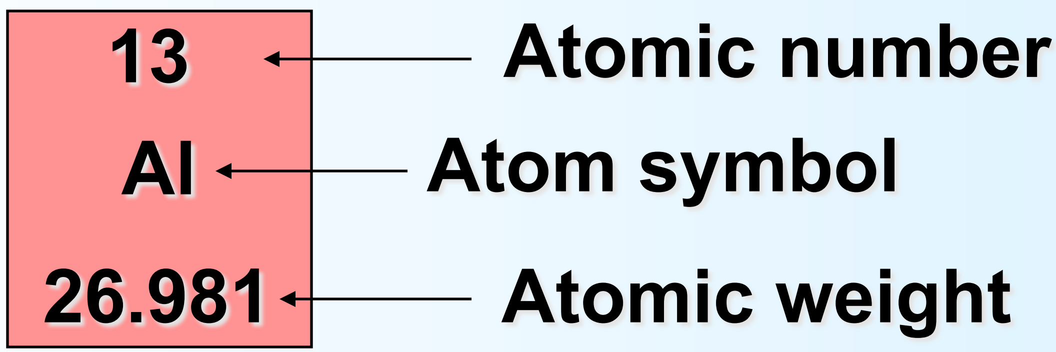
Particle	Mass	Charge*
Electron	9.11×10^{-31} kg	1−
Proton	1.67×10^{-27} kg	1+
Neutron	1.67×10^{-27} kg	None

Elements

- ***The identity of an element is determined by the number of protons in the nucleus.***
- For example, any atom with 6 protons in the nucleus is a Carbon atom.
- Number of protons in the nucleus = ***atomic number (Z)***
- **In a neutral atom, *# electrons = # protons.***
- The symbol for an element is simply its 1, 2, or 3 letter abbreviation from the periodic table.
- Elements are arranged in the periodic table by their ***atomic number (Z)***

Elements and Atomic Number, Z

- All atoms of the same element have the same number of protons in the nucleus, Z.
- Elements are arranged on the periodic table by their atomic number.

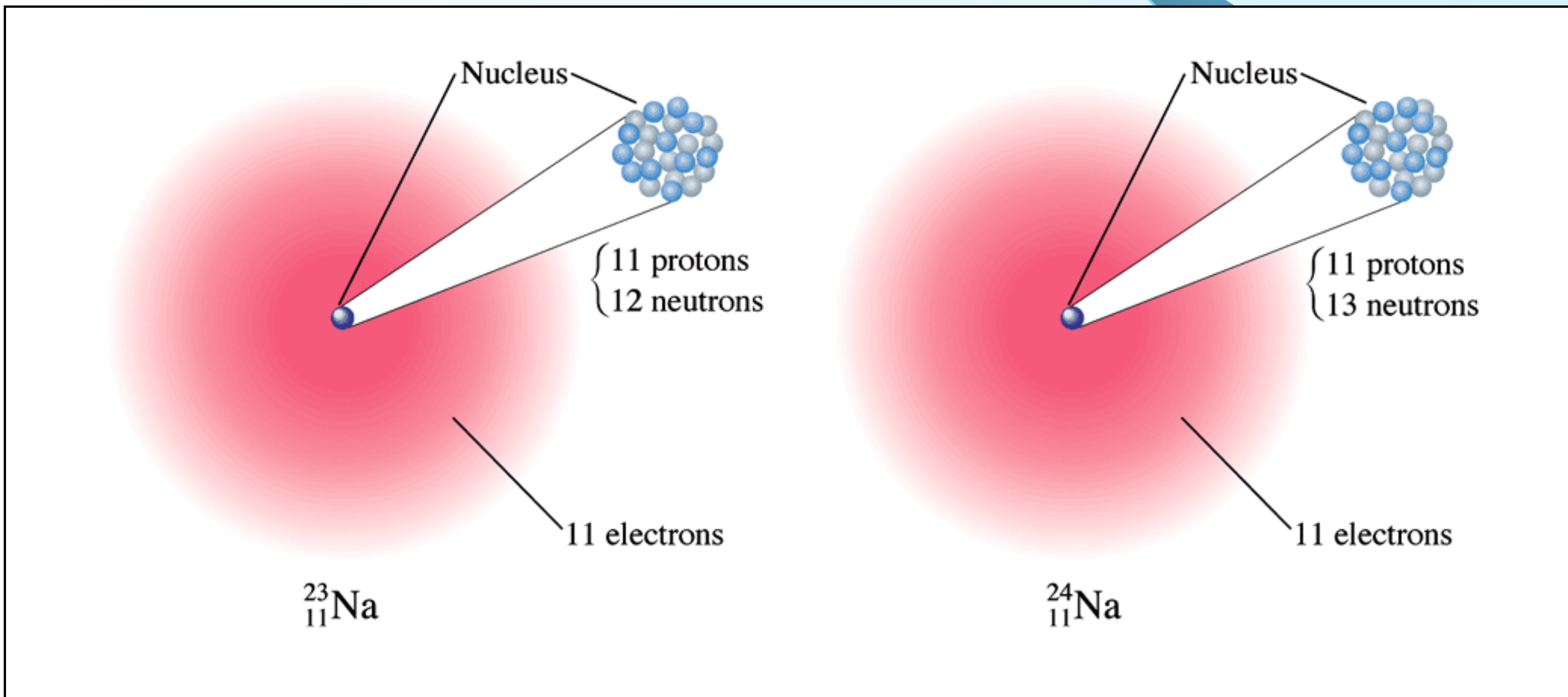


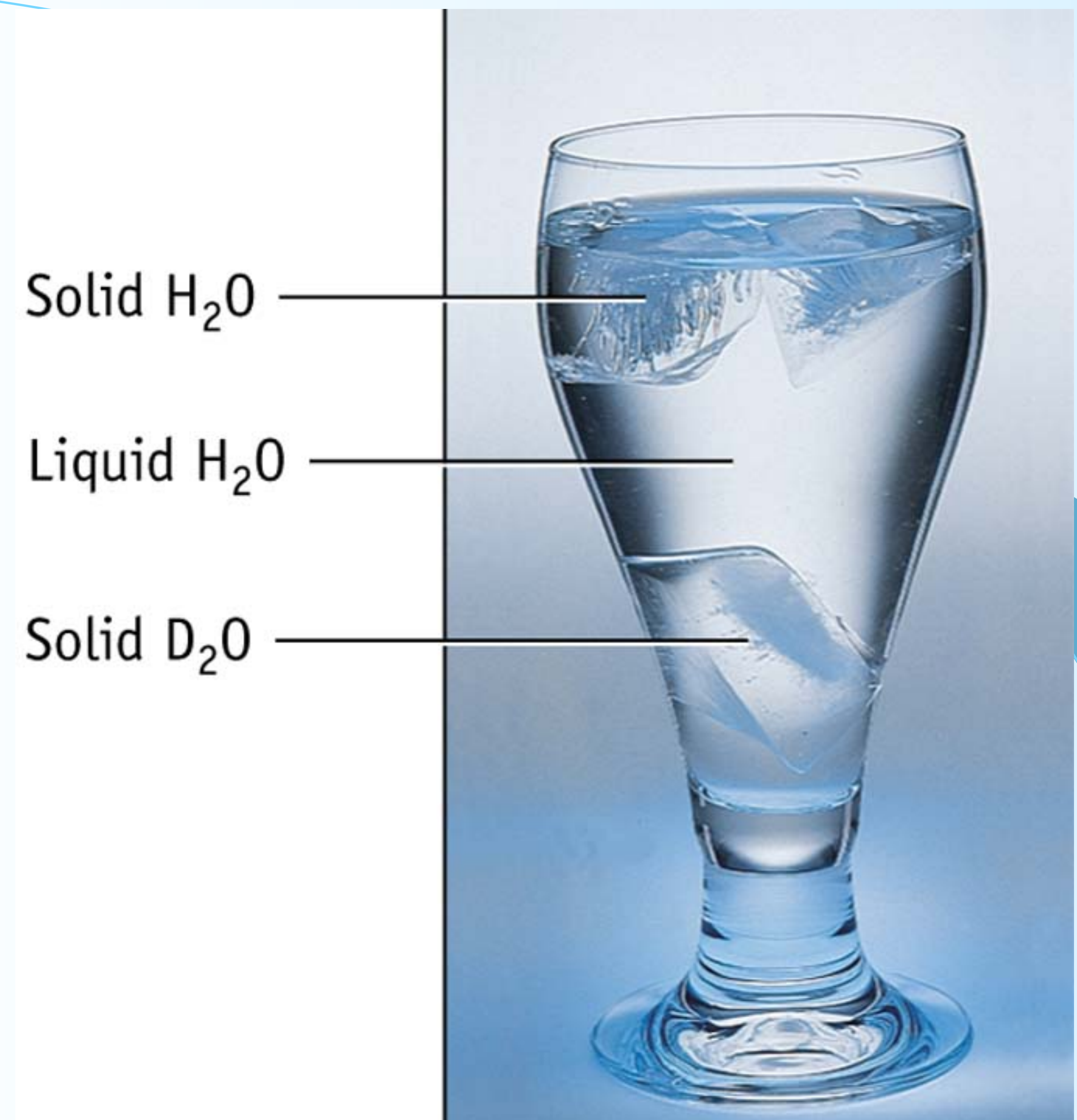
Isotopes

- ***Isotopes*** are atoms of the same element with different numbers of neutrons.
- Most elements have more than one isotope.
- Isotopes of an element are have nearly identical properties.
- The number of protons and electrons, which are the same in all isotopes of an element, has much more to do with the chemical and physical properties of an element.

Isotopes

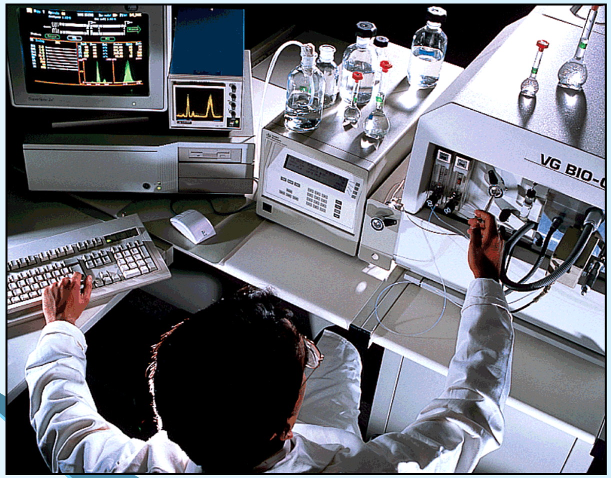
Atoms of the same element (Z) but different mass number (A).



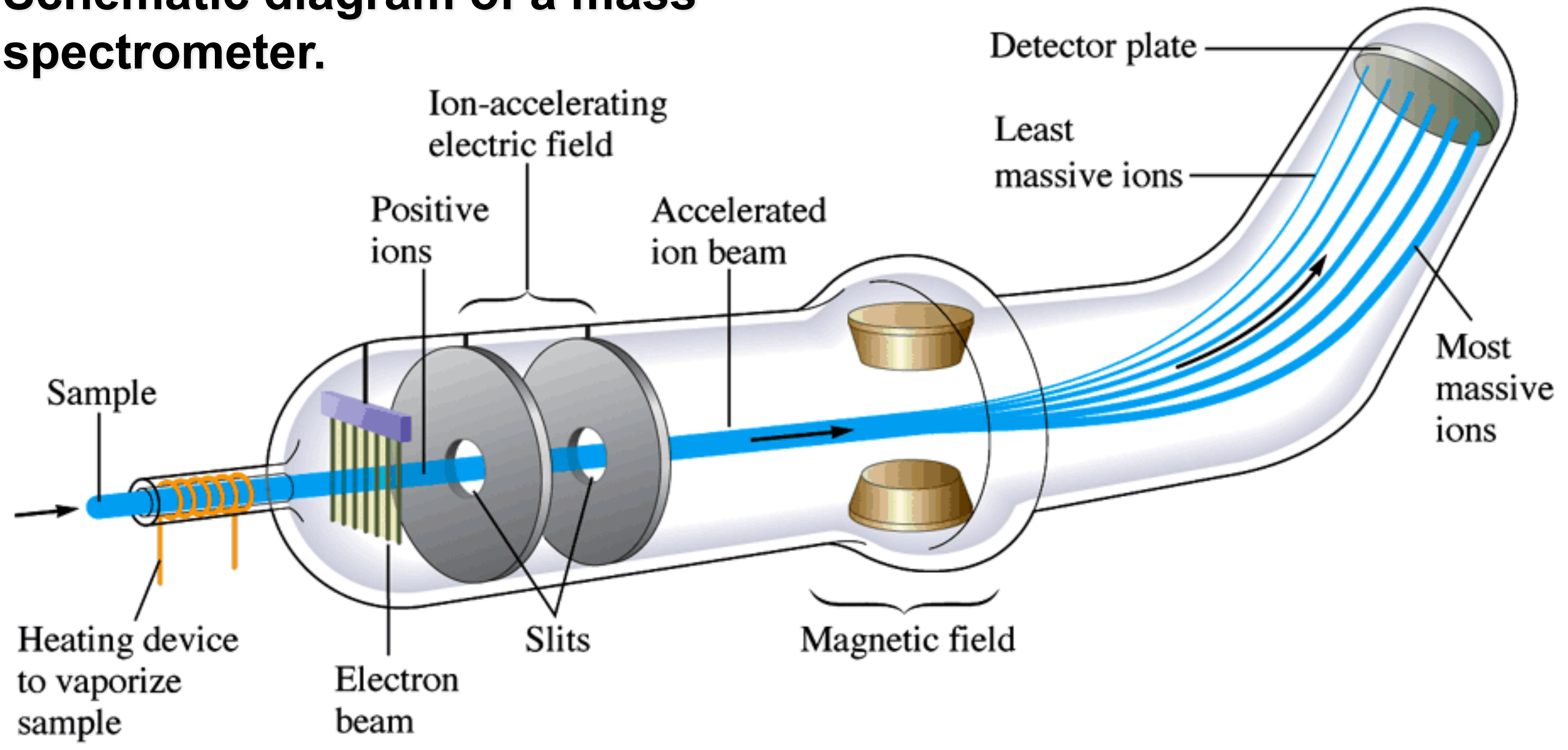

Solid H_2O
Liquid H_2O
Solid D_2O

Mass Spectrometry

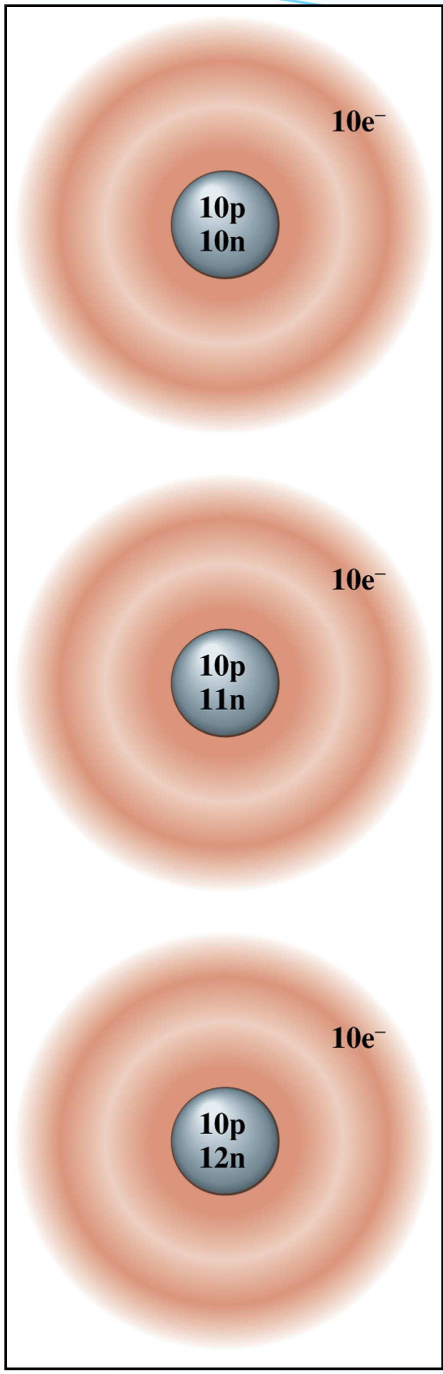
- Can be used to separate isotopes of an element.



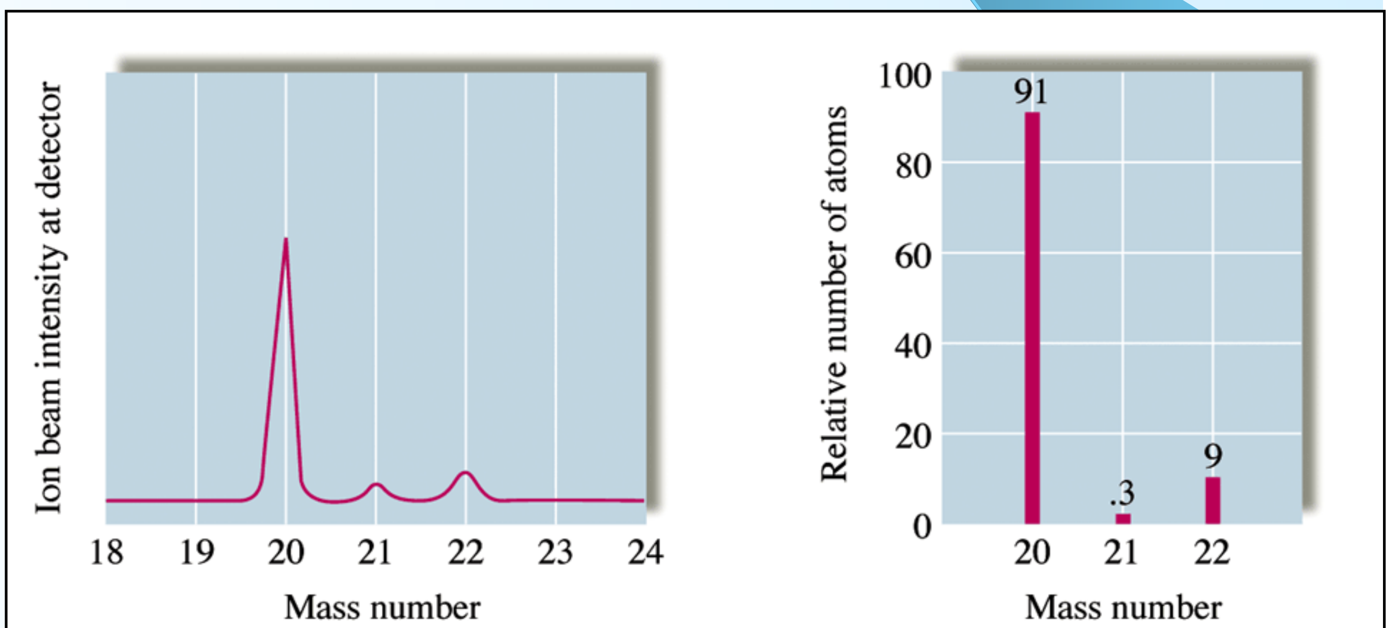
Schematic diagram of a mass spectrometer.



Representations of the three Naturally Occurring Isotopes of Neon.



Mass spectrograph of Neon



Mass Number

- The ***mass number* (*A*)** is the sum of the protons and the neutrons in the nucleus of an atom.

mass number = # protons + # neutrons

- The name for an isotope is the element name followed by the mass number.

- The symbol for an isotope is its element symbol along with its mass number (A) and atomic number (Z).

Mass Number (cont.)

- Consider the isotope Phosphorus-31
  symbol: $^{31}_{15}P$

A = 31 (mass number = the sum of protons and neutrons)

Z = 15 (atomic number = protons)

- In order to determine how many neutrons are in the nucleus of an atom simply subtract:
- **# neutrons = mass number – atomic number**

  = 31 – 15 = **16 neutrons**

Isotope Name	Isotope Symbol	Atomic Number (Z)	Mass Number (A)	protons	electrons	neutrons
Rhenium - 187						
		80	201			
			27		13	
	$^{237}_{93}Np$					

Atomic Mass

- ***Atoms are very small.***

- For example, 18 mL (~ 4 teaspoons) of water contains ~ 602,200,000,000,000,000,000,000 molecules of H_2O.

- A single water molecule has a mass of 2.99×10^{-23} grams.

- Working in grams to describe the mass of single atoms or molecules is not convenient.

- Instead, we typically express the mass of atom in terms of ***atomic mass units***, ***amu***, or simply ***u***.

Atomic Mass (cont.)

- The atomic mass unit is defined as $^{1}/_{12}$ of the mass of an atom of the carbon-12 isotope.

  1 amu = 1/12 mass of C-12

- Both the neutron and the proton have a mass of approximately 1 amu:

  1 amu ≈ 1 proton ≈ 1 neutron

- Electrons are very small:

  1 electron ≈ 0.00055 amu

- Because the mass number is the sum of protons and neutrons, the mass number is a whole number ***approximation*** of the atomic mass of an isotope.

Average Atomic Mass

- Most elements are a mixture of two or more isotopes.
- The percentage of an isotope in a naturally occurring sample of an element is called the ***isotopic abundance*** (or *percentage abundance*) of that that isotope.
- The ***isotopic mass*** is the mass of a single atom of an isotope.
- The ***average atomic mass*** is the *weighted average* of the masses of isotopes of an element.

Calculating Average Atomic Mass from Isotope Data

$$\sum(\text{Fractional Abundance} \times \text{Isotope Mass})$$

Consider the element ***Magnesium****, composed of three isotopes of the following percentage abundances:*

Isotope	% Abundance	Isotope Mass (amu)
Mg-24	79.0 %	23.985
Mg-25	10.0 %	24.986
Mg-26	11.0 %	25.982

Calculating Percent Abundances from Mass Data

Consider the two isotopes of Boron:

Isotope	Isotope Mass (amu)	%Abundance
Boron-10	10.0129	
Boron-11	11.0093	

What is the percent abundance of each isotope?