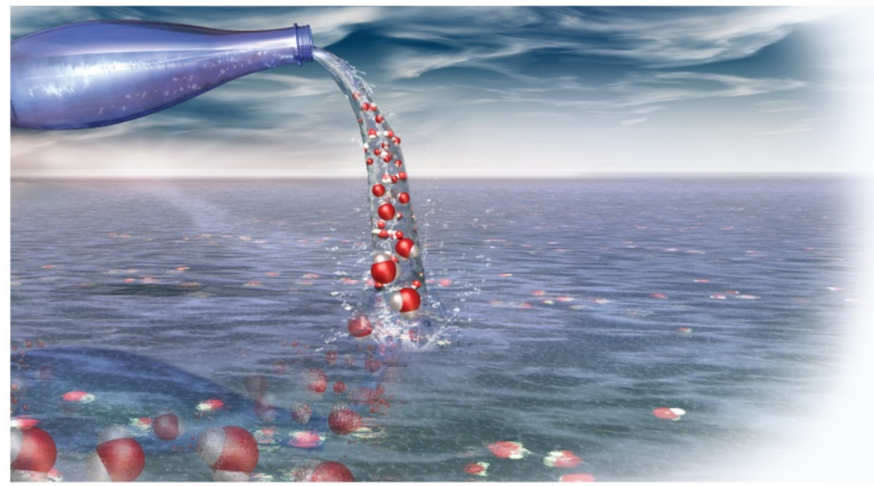
Chapter 1

Atoms



Matter from the Particulate Point of View

- **Matter is composed of particles.**
  - **Example: subatomic particles such as neutrons, protons, and electrons, atoms, and molecules.**

- **How the particles come together dictates the physical properties of matter.**

- **Matter is defined as anything that has mass and occupies space (e.g., has volume).**

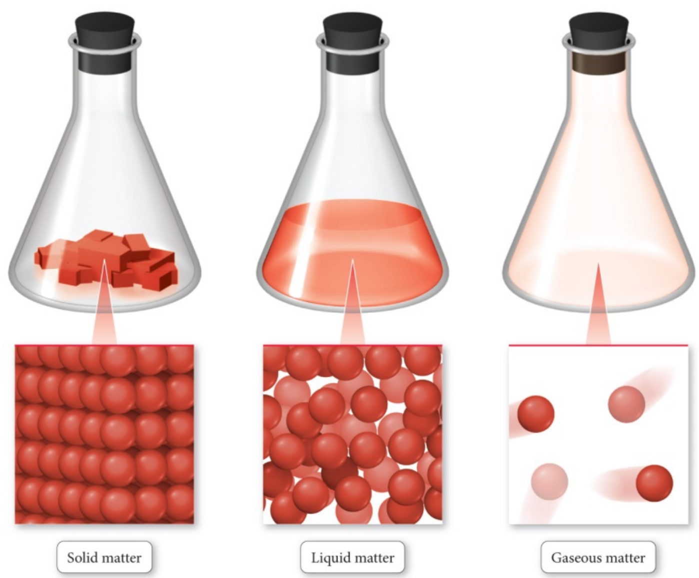
Elements, Molecules, and Mixtures: The Types of Matter

- **Atoms**:
  - Basic submicroscopic particles that constitute the fundamental building blocks of ordinary matter
- **Molecules:**
  - Substances formed when two or more atoms come together (bond) in specific geometric arrangements
- Atoms and molecules determine how matter behaves
- **Chemistry** is a discipline that seeks to understand matter and its properties, and the transformations that matter undergoes- particularly between molecules.

The Classification of Matter

- Matter can be classified according to
  - its **state**—its physical form (i.e., solid, liquid, or gas) based on what properties it exhibits;
  - its **composition** or the types of particles.

- The state of matter changes from solid to liquid to gas with increasing temperature.



Solid Matter

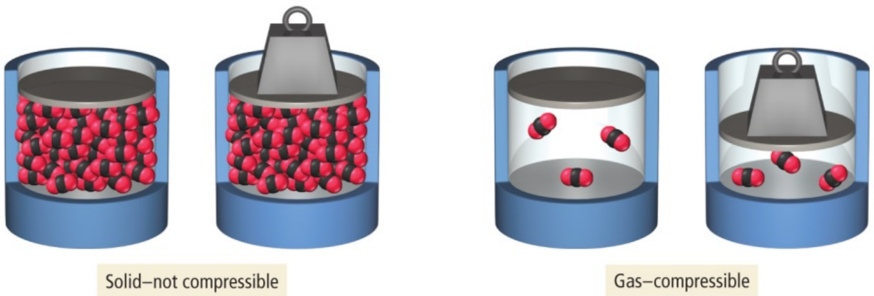
- In *solid matter*, atoms or molecules pack close to each other in fixed locations.
- Although the atoms and molecules in a solid vibrate, they do not move around or past each other.
- Consequently, a solid has a fixed volume and rigid shape.
  - Ice, aluminum, and diamond are good examples of solids.

Liquid Matter

- In *liquid matter*, atoms or molecules pack about as closely as they do in solid matter, but they are free to move relative to each other.
- Liquids have fixed volume but not a fixed shape.
- Liquids' ability to flow makes them assume the shape of their container.
  - Water, alcohol, and gasoline are all substances that are liquids at room temperature.

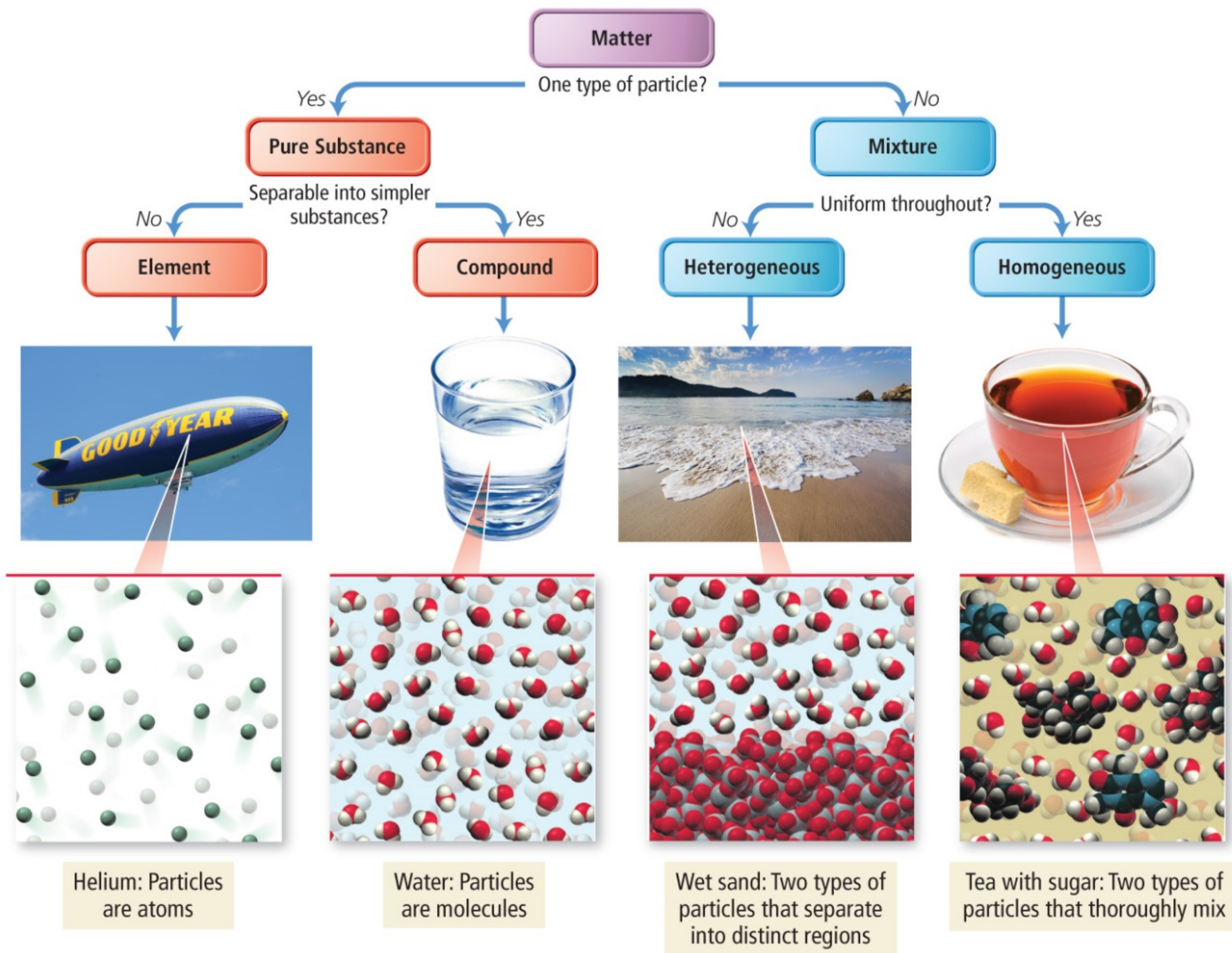
Gaseous Matter

- In *gaseous matter*, atoms or molecules have a lot of space between them.
- They are free to move relative to one another.
- Fill available space
- These qualities make gases *compressible*.



Classification of Matter by Components

- Matter can be classified according to its composition: elements, compounds, and mixtures.



Classification of Matter by Components

- The first division in the classification of matter is between a *pure substance* and a *mixture*.
- A **pure substance** is made up of only one component, and its composition is invariant.
- A **mixture**, by contrast, is a substance composed of two or more components in proportions that can vary from one sample to another.

Classification of Pure Substances

- Pure substances can be categorized into two types:
  - Elements
  - Compounds
- This categorization depends on whether or not they can be broken down (or decomposed) into simpler substances.

Classification of Pure Substances

- An **element** is a substance that cannot be chemically broken down into simpler substances.
  - Basic building blocks of matter
  - Composed of single type of atom, like helium
- A **compound** is a substance composed of two or more elements in fixed definite proportions.
- Most elements are chemically reactive and combine with other elements to form compounds.

Classification of Mixtures

- **Mixtures** can be categorized into two types:
  - **Heterogeneous mixtures**
  - **Homogeneous mixtures**
- This categorization of mixtures depends on how *uniformly* the substances within them mix.

Heterogeneous Mixture

- A **heterogeneous mixture** is one in which the composition varies from one region of the mixture to another.
  - Made of multiple substances, whose presence can be seen (example: a salt and sand mixture)
    - Portions of a sample of a heterogeneous mixture have different composition and properties.

Homogeneous Mixture

- A **homogeneous mixture** is one made of multiple substances, but it appears to be one substance.
- All portions of a sample have the same composition and properties (like sweetened tea).
- Homogeneous mixtures have uniform compositions because the atoms or molecules that compose them mix uniformly.

The Scientific Approach to Knowledge

- The approach to scientific knowledge is empirical.
  - It is based on ***observation*** and ***experimentation***.

- The scientific method is a process for understanding nature by observing nature and its behavior through experimentation.

- Key characteristics of the scientific method:
  - **Observations**
  - Formulation of **hypotheses**
  - **Experimentation**
  - Formulation of **laws and theories**

Observations

- Observations:
  - They are also known as **data**.
  - They are the descriptions about the characteristics or behavior of nature.
- Observations, verification of observations, and experimentation can lead scientists to formulate a **hypothesis**.

The Scientific Approach to Knowledge

- Scientists try to understand the universe through empirical knowledge gained through observation and experiment

Gathering Empirical Knowledge — Observation

- Some observations are descriptions of the characteristics or behavior of nature — **qualitative**

- Some observations compare a characteristic to a standard numerical scale — **quantitative**

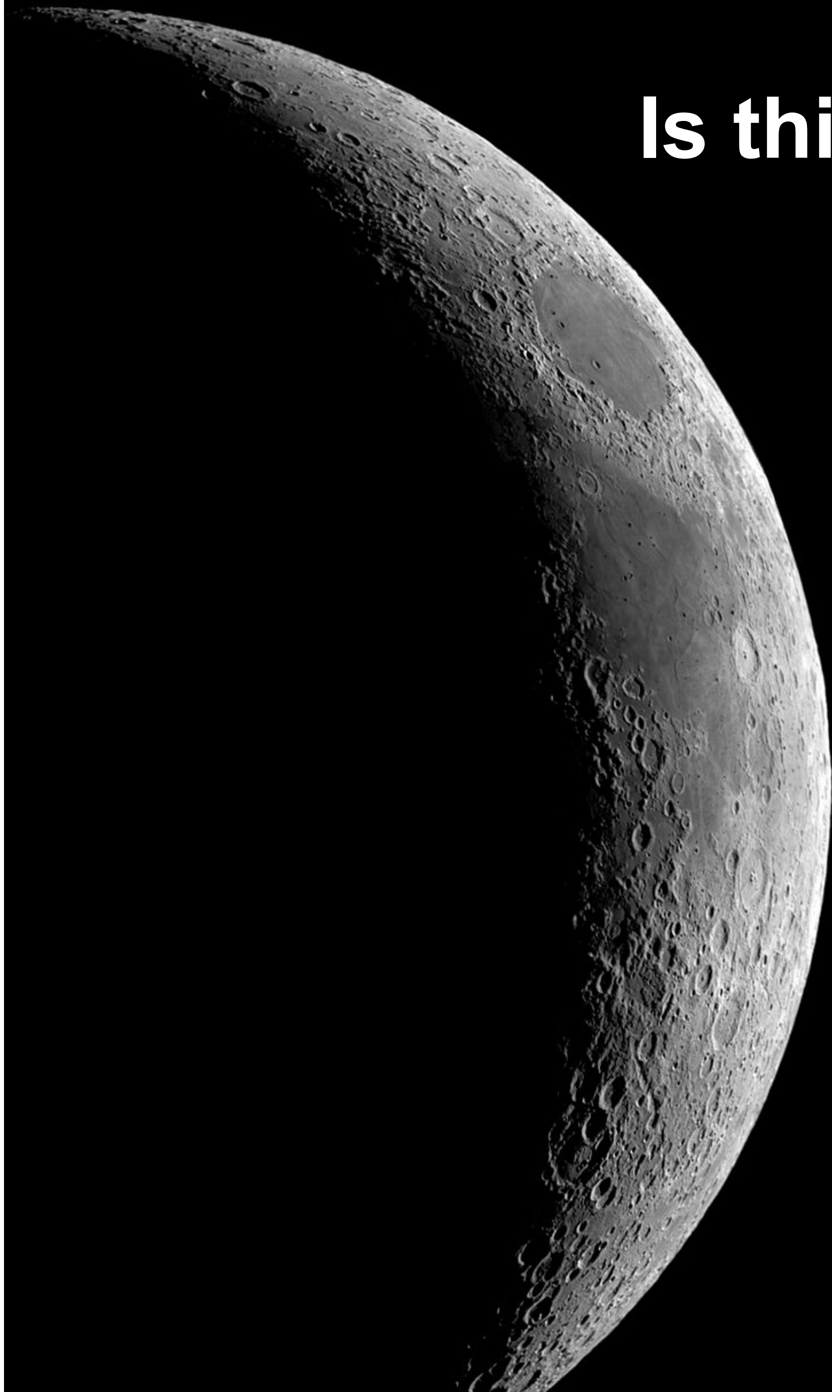
From Observation to Understanding

- **Hypothesis** – a tentative interpretation or explanation for an observation

- A good hypothesis is one that can be tested to be proved wrong!

Testing Ideas

- Ideas in science are tested with experiments
- An **experiment** is a set of highly controlled procedures designed to test whether an idea about nature is valid
- The experiment generates observations that will either validate or invalidate the idea



Is this a good hypothesis?

- Yesterday, the phase of the moon was a crescent. I observed that it rained a lot yesterday.
- It also rained during the previous crescent moon.
- *I hypothesize* that crescent moons cause rain showers

From Specific to General Understanding

- A **hypothesis** is a potential explanation for a single or small number of observations
- A **scientific theory** is a general explanation for why things in nature are the way they are and behave the way they do

Think about it

- Given what you just learned about the term “Theory”, what’s wrong with this statement:

  “The theory of gravity is just a theory and therefore probably wrong”

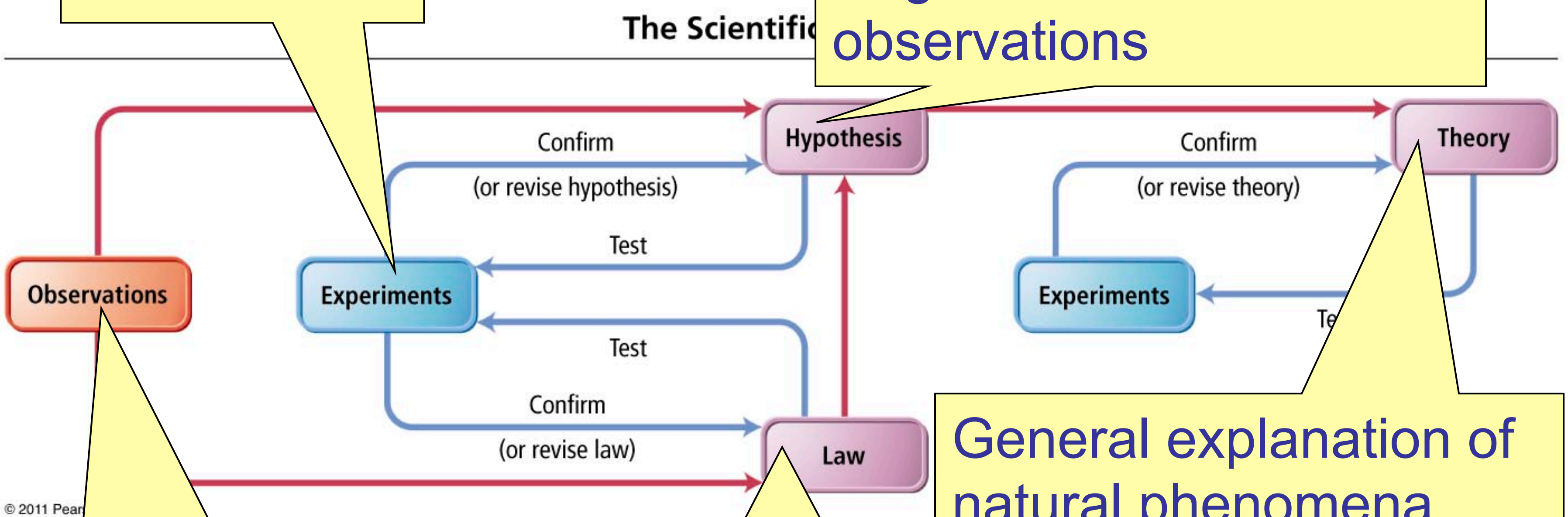
From Specific to General Observations

- A **scientific law** is a statement that summarizes all past observations and predicts future observations
  - ✓ **Law of Conservation of Mass** – "In a chemical reaction matter is neither created nor destroyed."
- A scientific law allows you to predict future observations

Scientific Method

Procedure designed to test an idea

Tentative explanation of a single or small number of observations



General explanation of natural phenomena

Careful noting and recording of natural phenomena

Generally observed occurence in nature

Relationships Between Pieces of the Scientific Method

	Applies to single or **small number** of events	Applies to **all** events
Describes ***what*** happens	observation	law
Explains ***why*** things happen	hypothesis	theory

The Scientific Approach to Knowledge

It is necessary to be careful with the information presented by an experimentalist who lacks theoretical principles… [he] gathers at random several facts and presents them as proofs… scientific knowledge without reasoning [theory] does not exist.”

J. le R. d’Alembert (1717-1783), from *Nouvelles Experiences sur la Resistance des Fluids*. Jombert, Paris (1997)

The Scientific Approach to Knowledge

D'Alembert was always surrounded by controversy. ….Unfortunately he carried [his]... pugnacity into his scientific research and once he had entered a controversy, he argued his cause with vigour and stubbornness. He closed his mind to the possibility that he might be wrong...

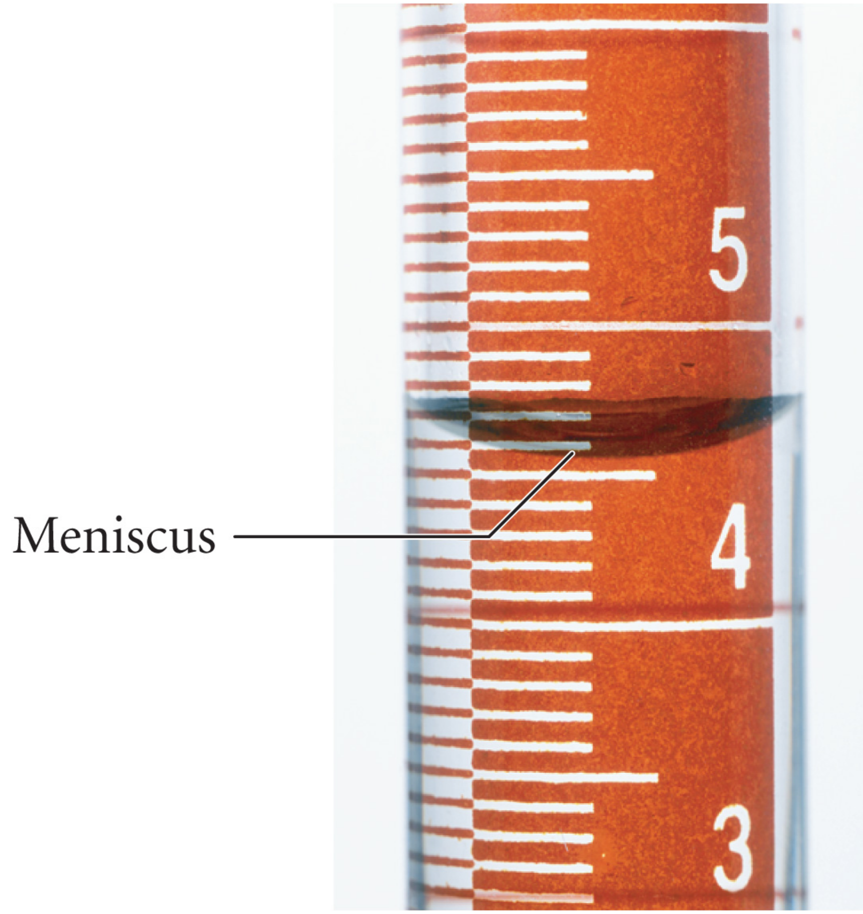
Thomas L. Hankins (1990). *Jean D'Alembert: Science and the Enlightenment*. Taylor & Francis. p. 236. ISBN 2881243991

Why Is Scientific Measurement Important?

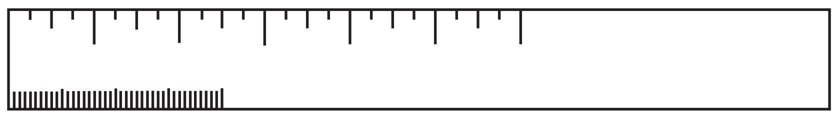
- Scientific data can be either qualifiable or quantifiable.
  - Qualifiable data are observational.
    - Subjective in nature
    - Examples: color, shape
  - Quantifiable data are measurable (empirical).
    - Objective in nature
    - Uses equipment (e.g., glassware, balance, instrumentation) capable of generating empirical data with standardized **UNITS**.
      - English system (e.g., inch, feet, etc.)
      - International System of Units (SI)
        - Metric system

What Is a Measurement?

- quantitative observation
- comparison to an agreed- upon standard
- every measurement has a number and a unit



Significant Figures



No measurement can be more accurate than half the smallest division

Note: Tro says 1/10 of smallest division

Often write $\pm$ to indicate this: 11$\pm$1 mm, 1.1 $\pm$0.3 cm

No error range given? Then assume $\pm$ half the least Significant Figure

Significant Figures

- Any digit that is not zero is significant

    1.234 kg   4 significant figures

- Zeros between nonzero digits are significant

    606 m   3 significant figures

- Zeros to the left of the first nonzero digit are **not** significant

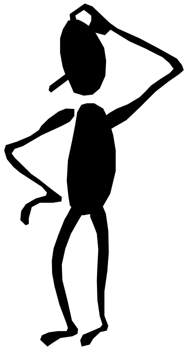
    0.08 L   1 significant figure

- If a number is greater than 1, then all zeros to the right of the decimal point are significant

    2.0 mg   2 significant figures

- If a number is less than 1, then only the zeros that are at the end and in the middle of the number are significant

    0.004020 g   4 significant figures



How many significant figures are in each of the following measurements?

24 mL	2 significant figures
3001 g	4 significant figures
0.0320 m^3	3 significant figures
6.4 x 10^4 molecules	2 significant figures
560 kg	2 significant figures

Rounding Rules for Significant Figures

Digit being rounded >5		Add 1 to least significant digit
Result: 4.16 with 2 Sig. Fig.	→	4.2
Digit being rounded <5		Leave least Sig. Fig. unchanged
Result 4.14 with 2 Sig. Fig.	→	4.1

ROUND AT THE END OF THE PROBLEM

Significant Figures

Addition or Subtraction

The answer cannot have more digits to the right of the decimal point than any of the original numbers.

$$\begin{array}{r} 89.332 \\ +1.1 \\ \hline 90.432 \end{array}$$

+1.1 ← one significant figure after decimal point

90.432 ← round off to 90.4

$$\begin{array}{r} 3.70 \\ -2.9133 \\ \hline 0.7867 \end{array}$$

3.70 ← two significant figures after decimal point

0.7867 ← round off to 0.79

Significant Figures

Multiplication or Division

The number of significant figures in the result is set by the original number that has the ***smallest*** number of significant figures

$$4.51 \times 3.6666 = 16.536366 \quad = 16.5$$

3 sig figs

round to
3 sig figs

$$6.8 \div 112.04 = 0.0606926 \quad = 0.061$$

2 sig figs

round to
2 sig figs

Significant Figures

Exact Numbers

Numbers from definitions or numbers of objects are considered to have an infinite number of significant figures

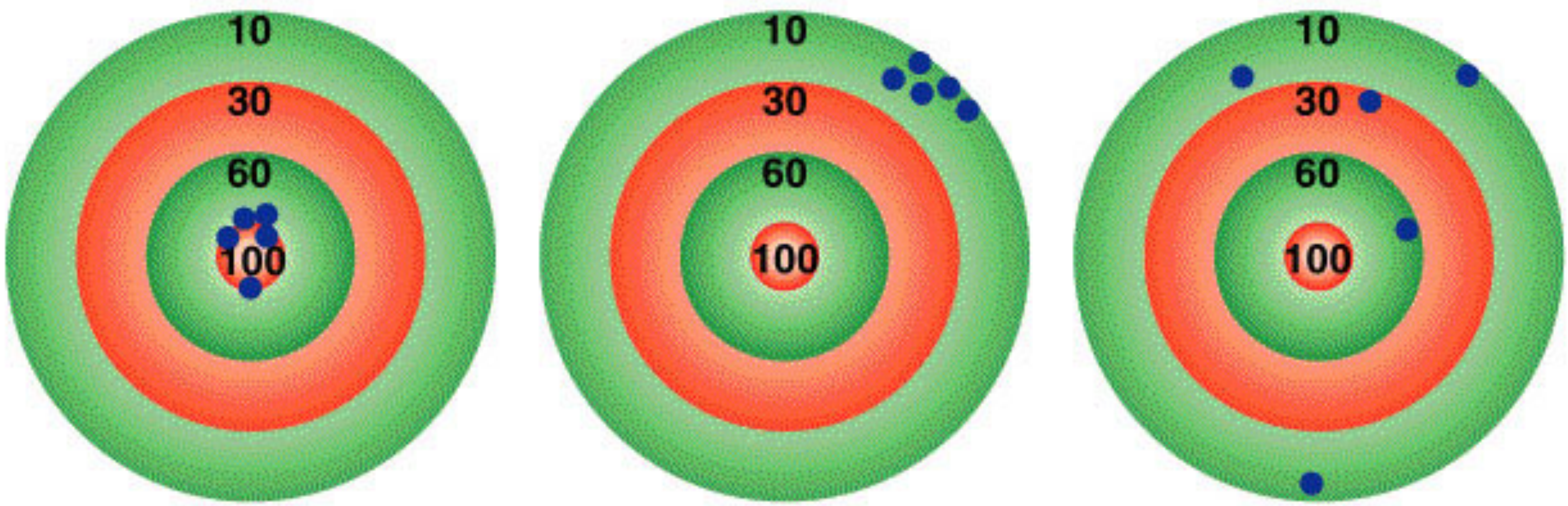
The average of three measured lengths; 6.64, 6.68 and 6.70?

$$\frac{6.64 + 6.68 + 6.70}{3} = 6.67333 = 6.67 \quad = \cancel{7}$$

Because 3 is an exact number

Accuracy – how close a measurement is to the *true* value

Precision – how close a set of measurements are to each other



accurate
&
precise

precise
but
not accurate

not accurate
&
not precise

Our Current Understanding of the Structure of Matter

Early Ideas about the Building Blocks of Matter

- ***Leucippus*** (fifth century B.C.) and his student ***Democritus*** (460–370 B.C.) were the first to propose that matter was composed of small, indestructible particles.
  - Democritus wrote, "Nothing exists except atoms and empty space; everything else is opinion."
    - He proposed that atoms existed in different shapes and sizes and moved randomly through empty space.

- ***Plato and Aristotle*** did not embrace the atomic ideas of Leucippus and Democritus.
  - They held that matter had no smallest parts.
  - They proposed that different substances were composed of various proportions of fire, air, earth, and water.

Early Building Blocks of Matter Ideas

- An English chemist, ***John Dalton*** (1766–1844) offered convincing evidence that supported the early atomic ideas of Leucippus and Democritus.
  - Dalton's atomic theory of matter

- Modern Atomic Theory and Its Laws:
  - Law of conservation of mass
  - Law of definite proportions
  - Law of multiple proportions

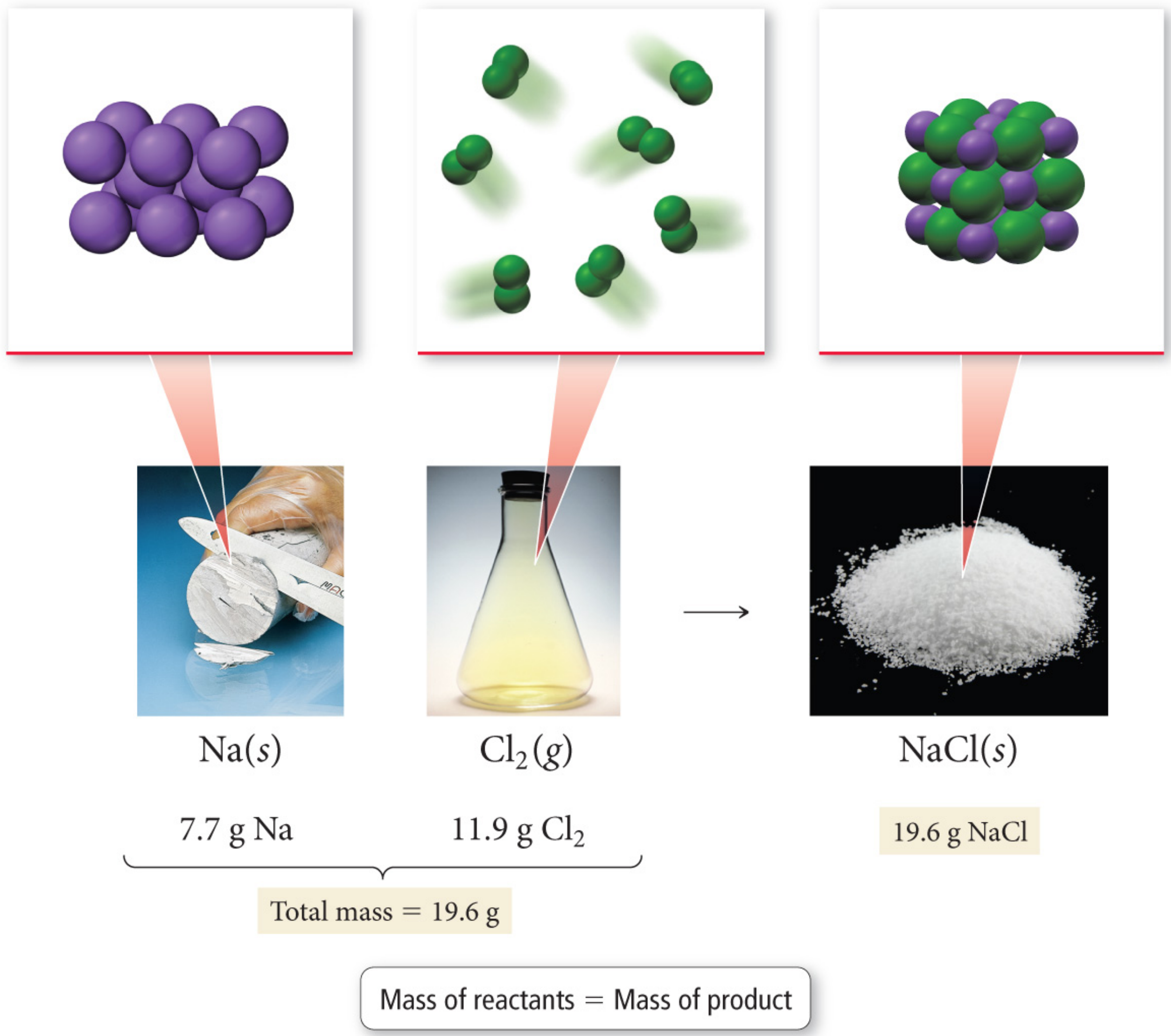
John Dalton and the Atomic Theory

- Dalton's **atomic theory** explained the laws as follows:
  1. Each element is composed of tiny, indestructible particles called atoms.
  2. All atoms of a given element have the same mass and other properties that distinguish them from the atoms of other elements.
  3. Atoms combine in simple, whole-number ratios to form compounds.
  4. Atoms of one element cannot change into atoms of another element. In a chemical reaction, atoms change only the way that they are bound together with other atoms.

The Law of Conservation of Mass

- Antoine Lavoisier formulated the **law of conservation of mass**, which states the following:
  - ***In a chemical reaction, matter is neither created nor destroyed.***
- When a chemical reaction occurs, the total mass of the substances involved in the reaction does not change.
  - This law is consistent with the idea that matter is composed of small, indestructible particles.

The Law of Conservation of Mass (Matter)



The Law of Definite Proportions

- In 1797, a French chemist named Joseph Proust made observations on the composition of compounds.

- He summarized his observations in the **law of definite proportions**:
  - ***All samples of a given compound, regardless of their source or how they were prepared, have the same proportions of their constituent elements.***

  - The law of definite proportions is sometimes called the law of constant composition.

An Example of the Law of Definite Proportions

The decomposition of 18.0 g of water (H_2O) results in 16.0 g of oxygen (O_2) and 2.0 g of hydrogen (H_2),

or

an oxygen-to-hydrogen mass ratio of 8:1.

$$\text{Mass ratio} = \frac{16.0\ \text{g O}}{2.0\ \text{g H}} = 8.0 \text{ or } 8:1$$

Problem Solving: The Law of Definite Proportions

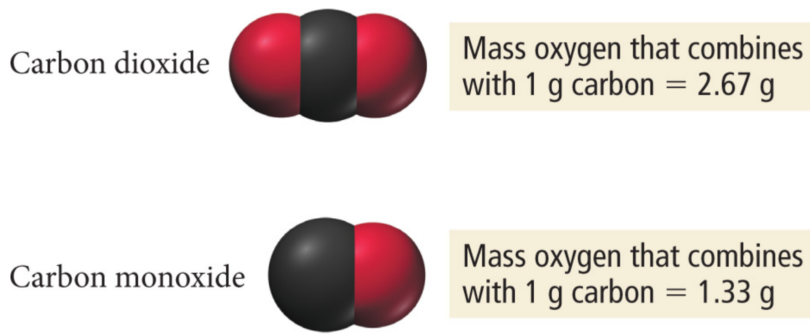
Two samples of carbon dioxide decompose into their constituent elements. One produces 25.6 g oxygen and 9.60 g carbon, the other 21.6 g oxygen, and 8.10 g carbon. Show that these results are consistent with the law of definite proportions.

The Law of Multiple Proportions

- In 1804, John Dalton published his **law of multiple proportions**:
  - ***When two elements (call them A and B) form two different compounds, the masses of element B that combine with 1 g of element A can be expressed as a ratio of small whole numbers.***
    - When an atom of A combines with either one, two, three, or more atoms of B, the following molecular compounds are possible: AB_1, AB_2, AB_3, etc.

The Law of Multiple Proportions

- Carbon monoxide and carbon dioxide are two compounds composed of the same two elements: carbon and oxygen.



  – The mass ratio of oxygen to carbon in carbon dioxide is 2.67:1; therefore, 2.67 g of oxygen reacts with 1 g of carbon.
  – In carbon monoxide, however, the mass ratio of oxygen to carbon is 1.33:1, or 1.33 g of oxygen to every 1 g of carbon.

- The ratio of these two masses is itself a small whole number.

$$\frac{\text{Mass oxygen to 1 g carbon in carbon dioxide}}{\text{Mass oxygen to 1 g carbon in carbon monoxide}} = \frac{2.67}{1.33} = 2$$

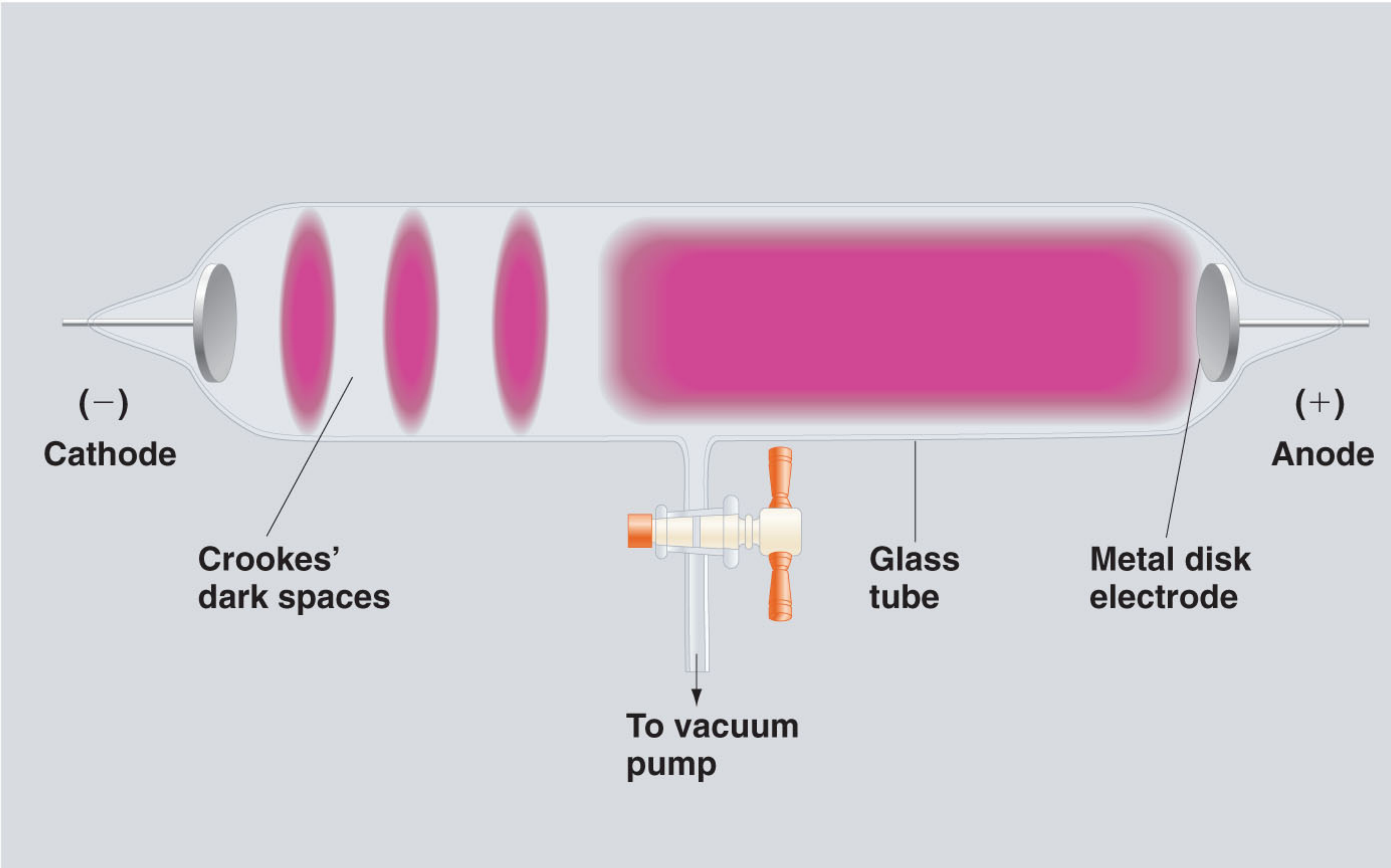
Problem Solving: The Law of Multiple Proportions

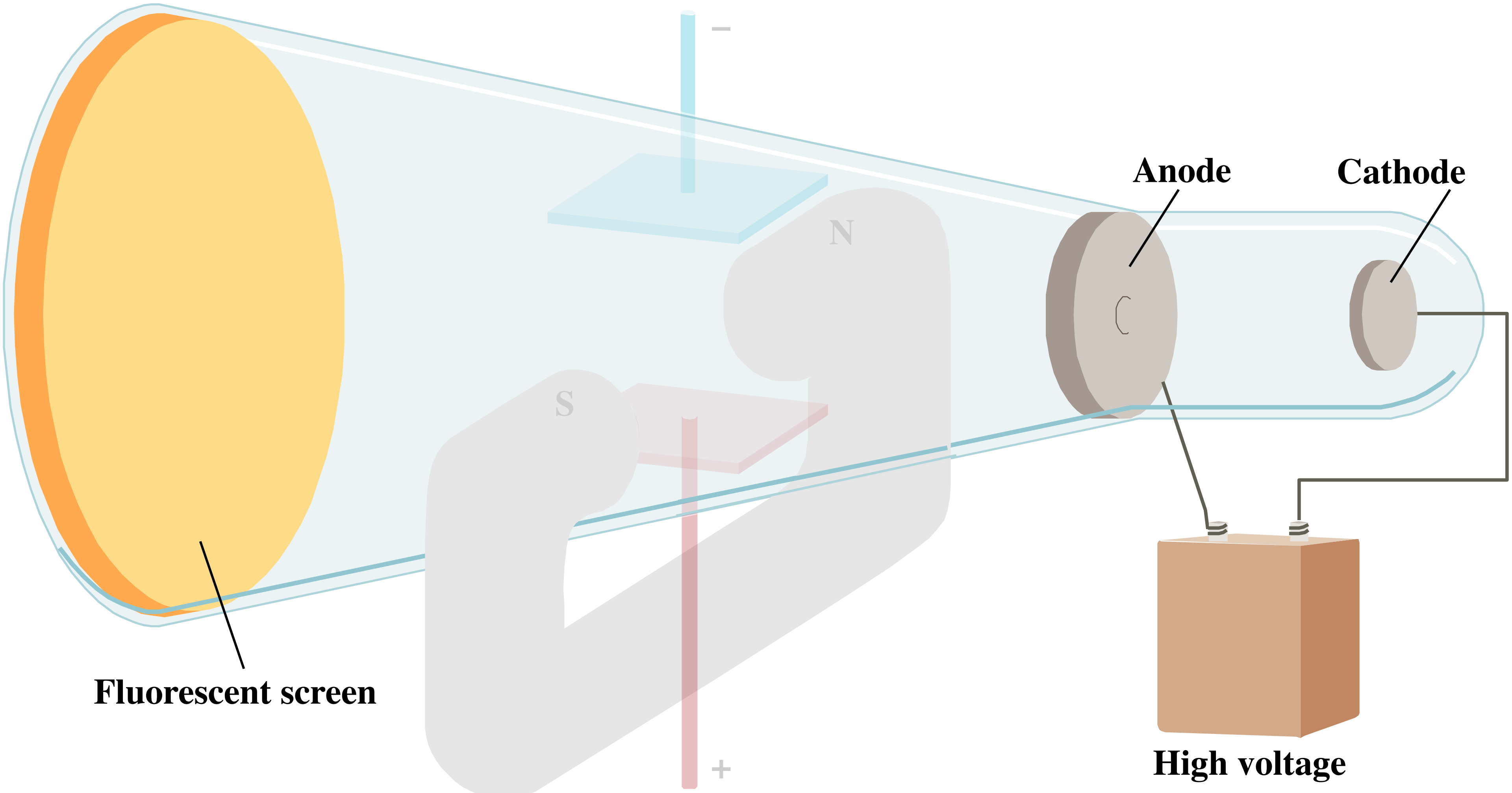
Nitrogen forms a number of compounds with oxygen. Measurements of the masses of nitrogen and oxygen that form upon decomposing these compounds show that nitrogen dioxide contains 2.28 g oxygen to every 1.00 g of nitrogen, while dinitrogen monoxide contains 0.570 g oxygen to every 1.00 g nitrogen. Show that the results are consistent with the law of multiple proportions:

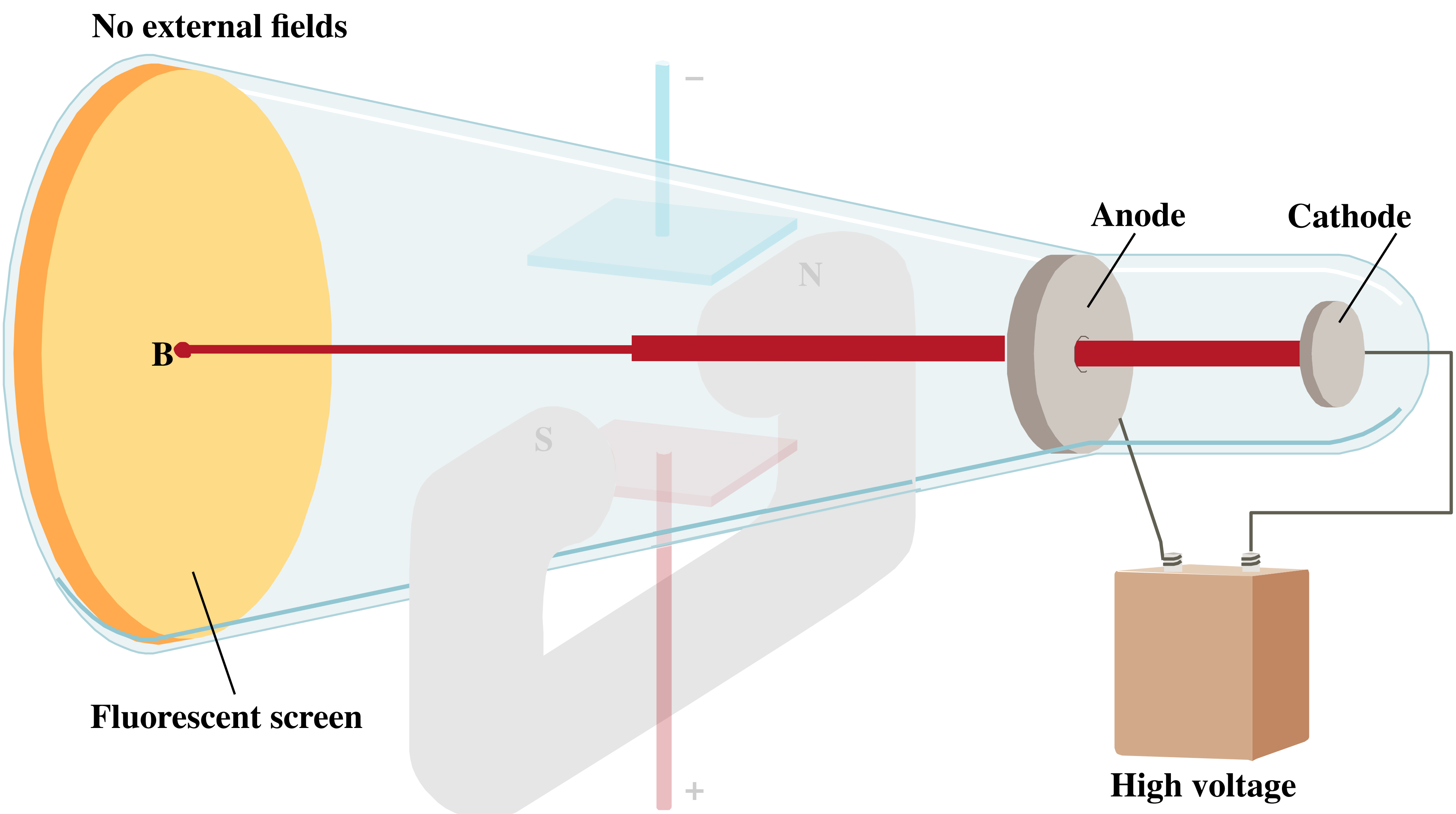
The Structure of the Atom

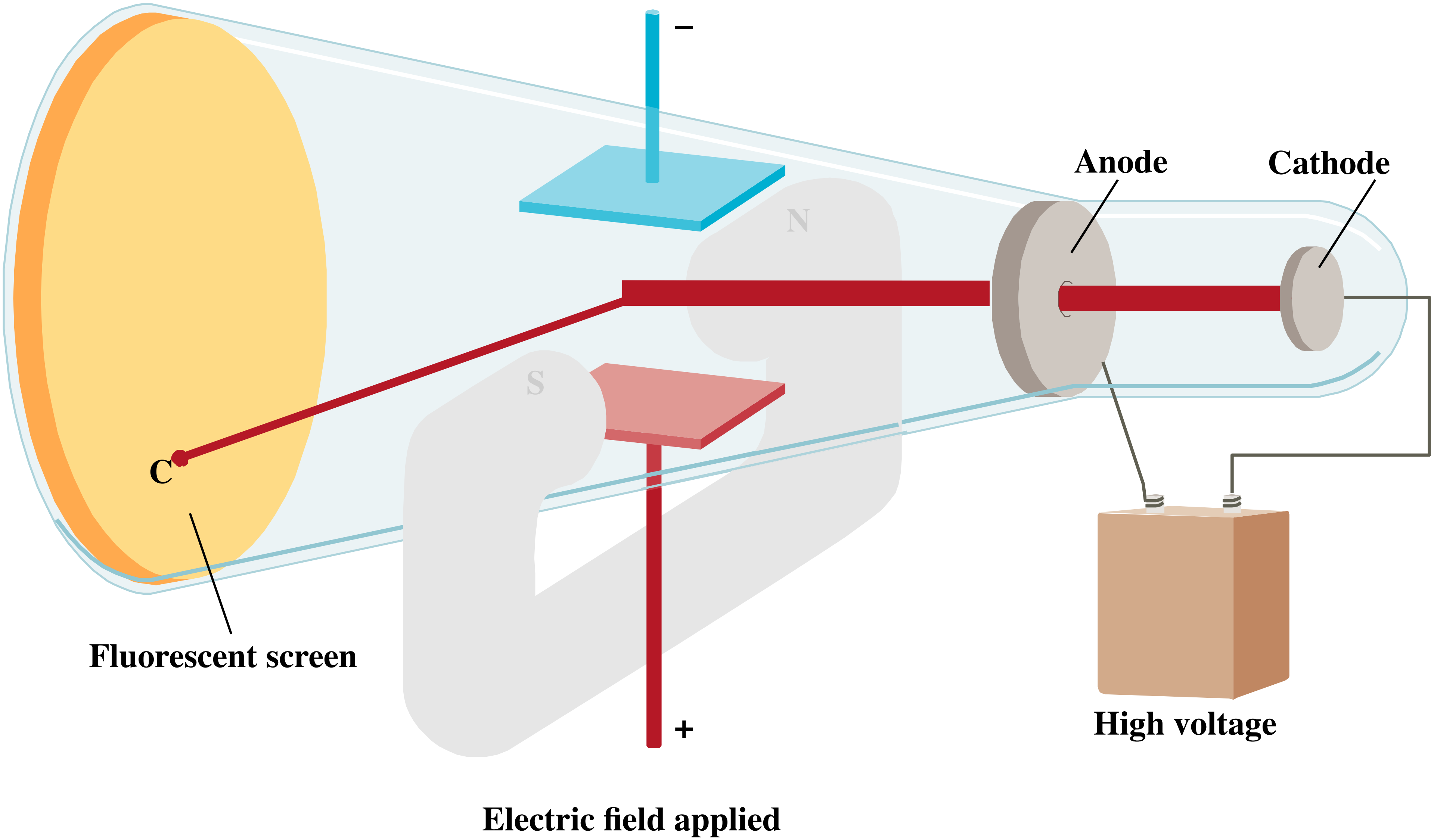
The Discovery of the Electron

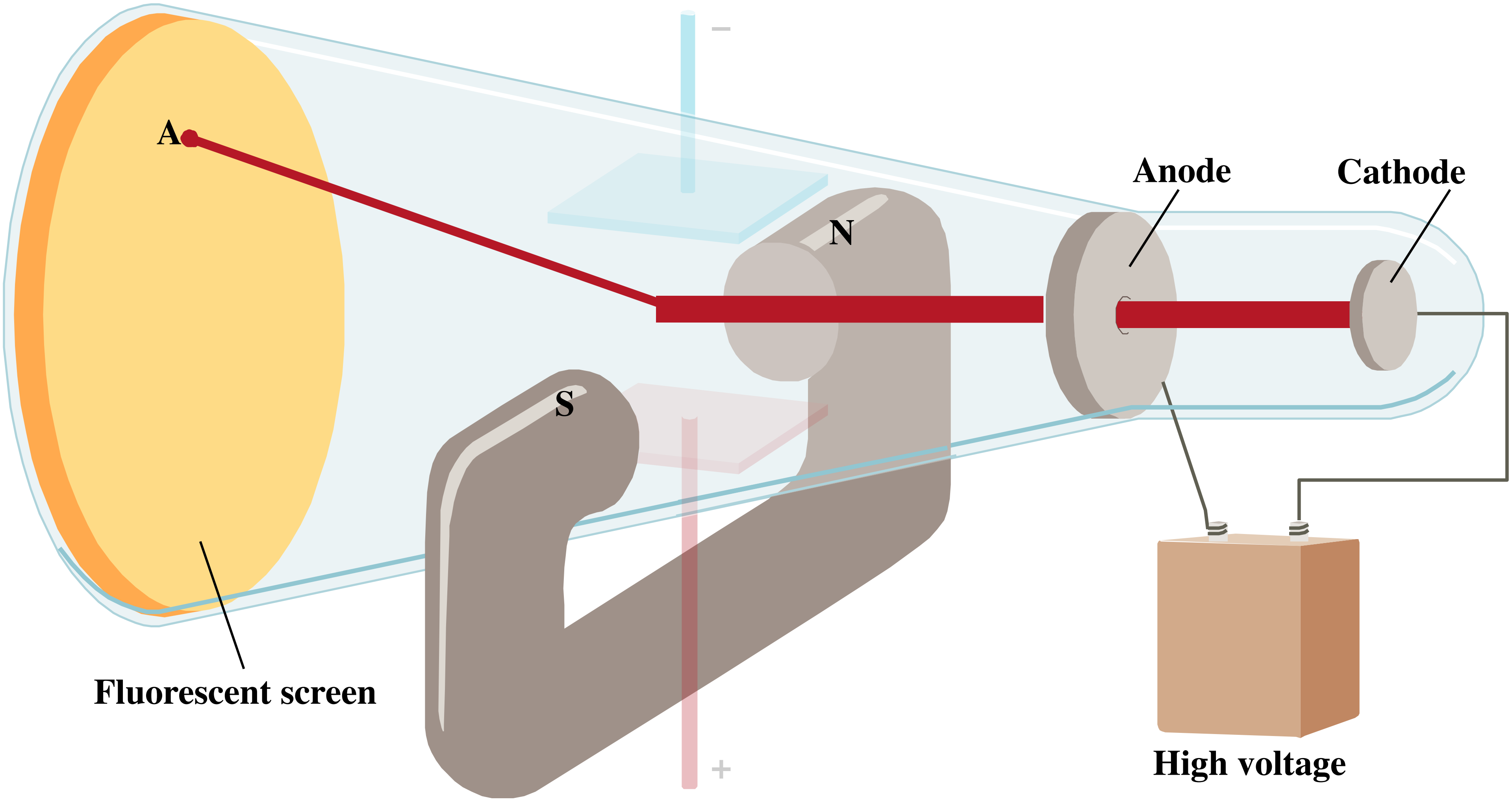
- From J. J. Thomson's (1856–1940) **cathode rays** experiments, it was observed that
  - a beam of particles, called cathode rays, traveled from the negatively charged electrode (called the cathode) to the positively charged one (called the anode);
  - the particles that compose the cathode ray have the following properties:
    - They travel in straight lines.
    - They are independent of the composition of the material from which they originate (the cathode).
    - They carry a negative **electrical charge**.
- J.J. Thomson was able to measure the **charge-to-mass** ratio of the cathode ray particles by deflecting them using electric and magnetic fields, as shown in the figure on the bottom of the next slide.
  - The value he measured was -1.76×10^3 coulombs (C) per gram.


(−)
Cathode
(+)
Anode
Crookes'
dark spaces
Glass
tube
Metal disk
electrode
To vacuum
pump

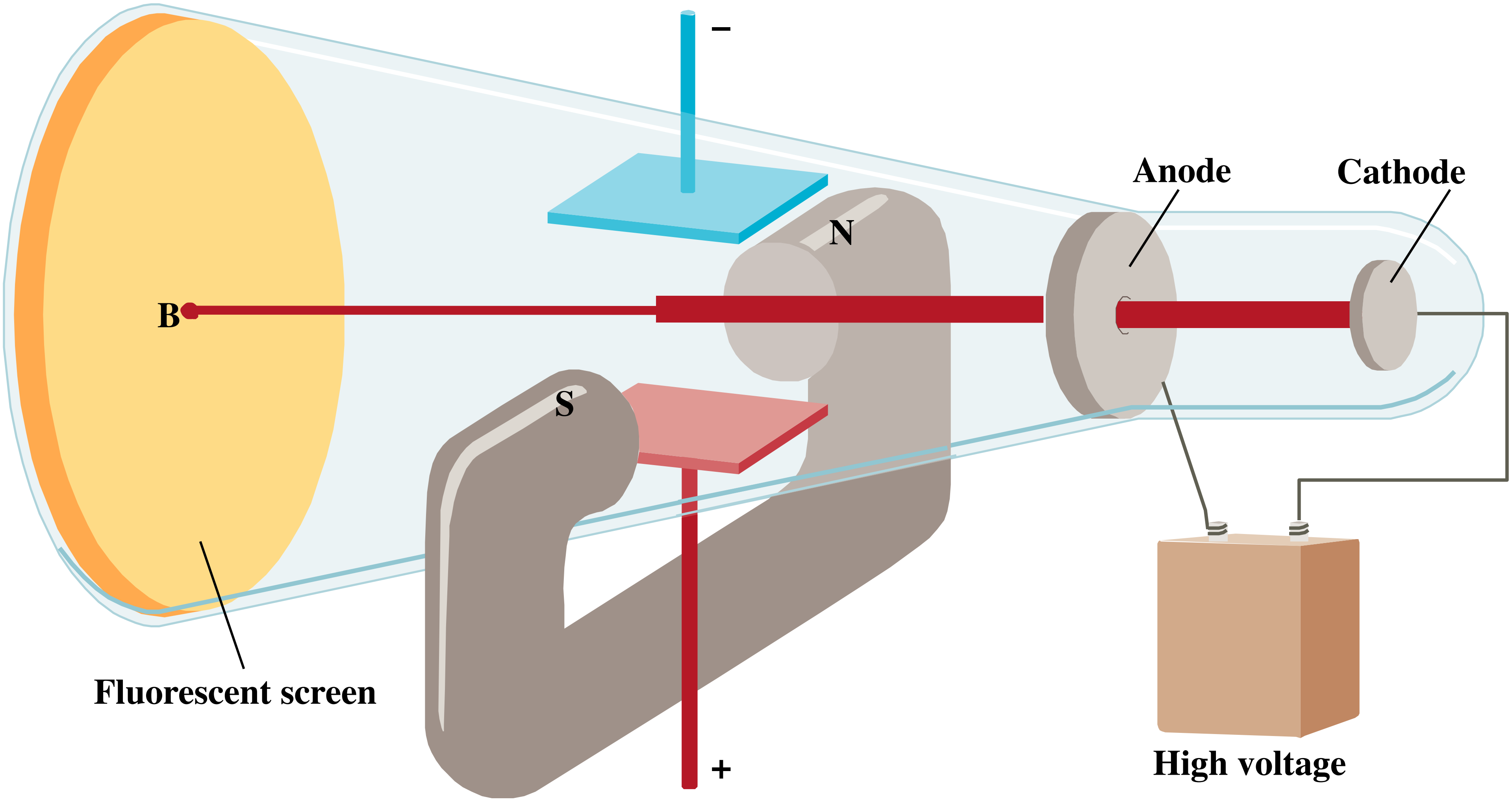

–
N
S
+
Anode
Cathode
Fluorescent screen
High voltage


No external fields
–
N
S
+
B
Anode
Cathode
Fluorescent screen
High voltage

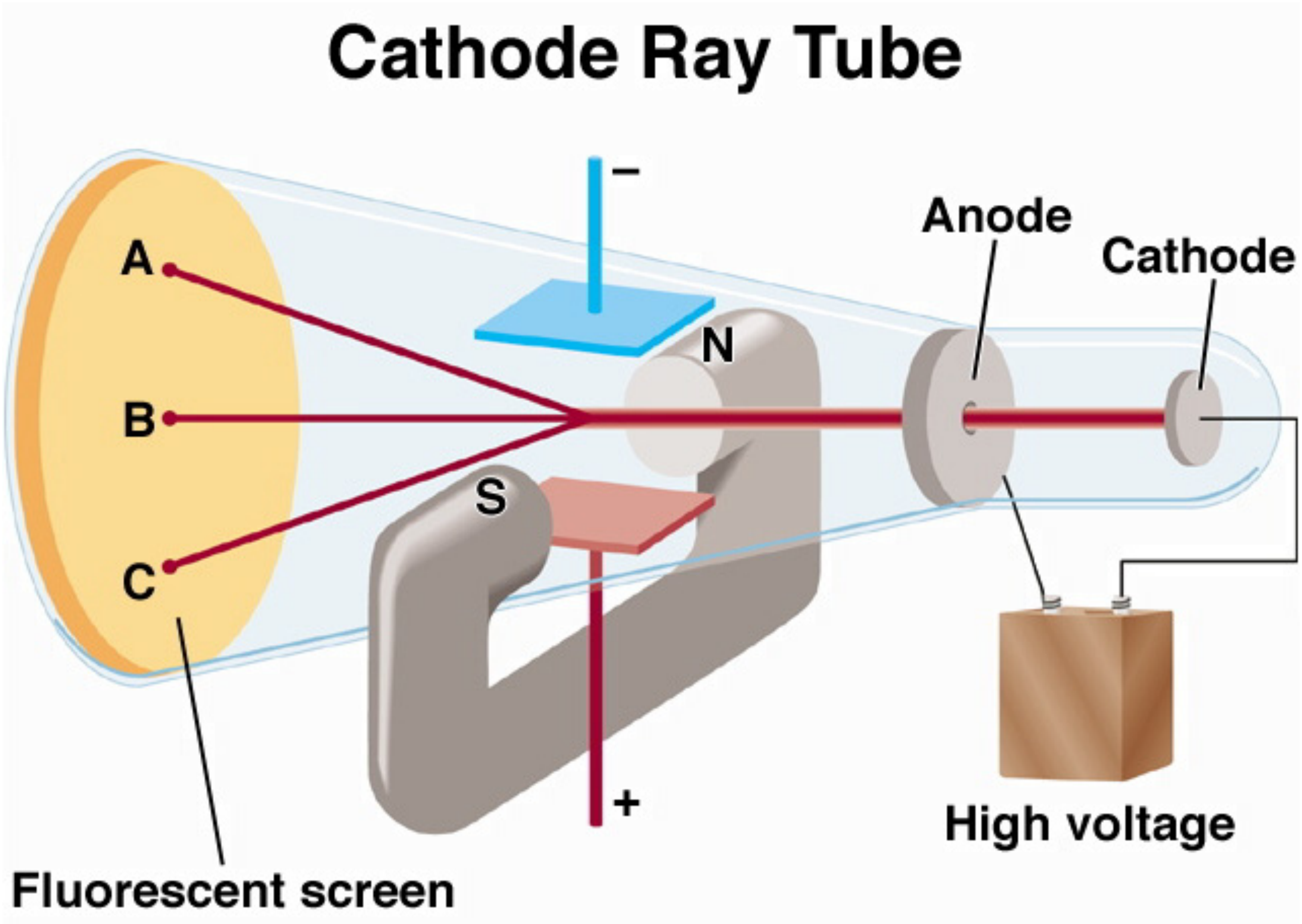

−
Anode
Cathode
N
S
C
Fluorescent screen
+
High voltage
Electric field applied



Magnetic field applied



Effects of electric field and magnetic field cancel



J.J. Thomson, measured mass/charge of e^-

(1906 Nobel Prize in Physics)

Thomson:

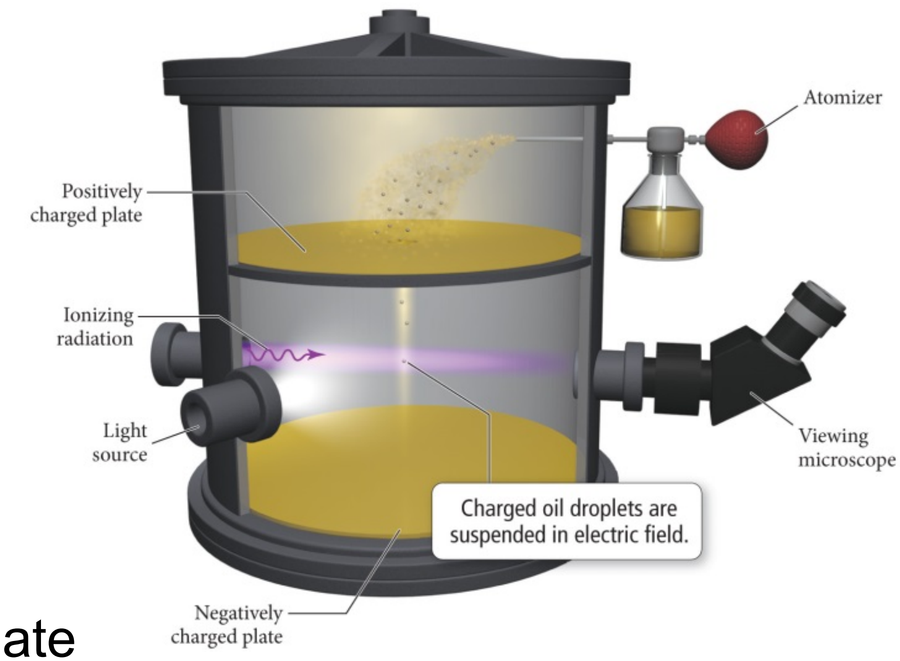
$q/m = -1.7588 \times 10^{8}$ C g^{-1} for electron

For Hydrogen it was 1836 times smaller (and of opposite sign)

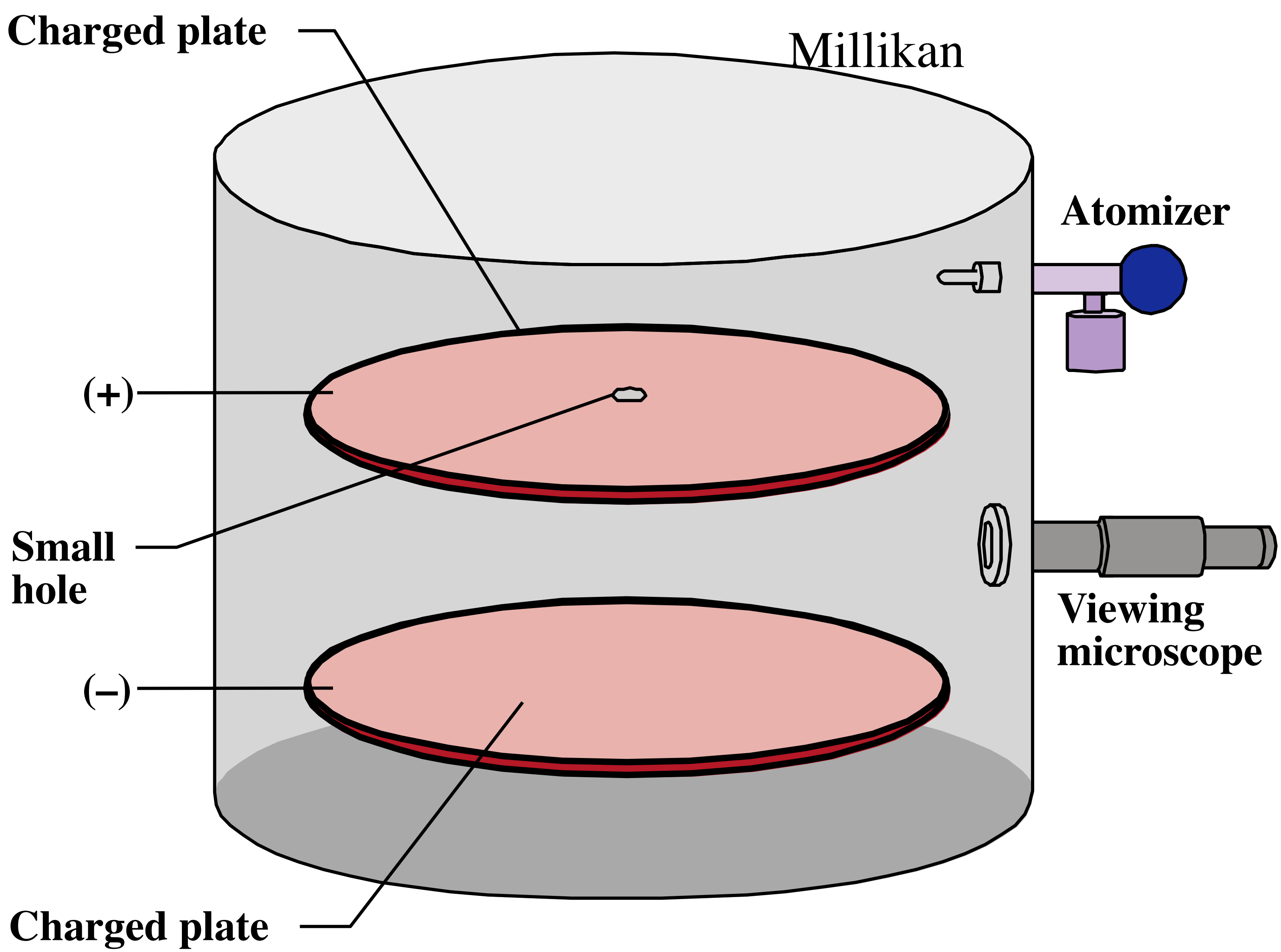
Millikan's Oil Drop Experiment: Determining the Charge of an Electron

- Data from his oil drop experiment led physicist Robert Millikan (1868–1953) to deduce the charge of a single electron.

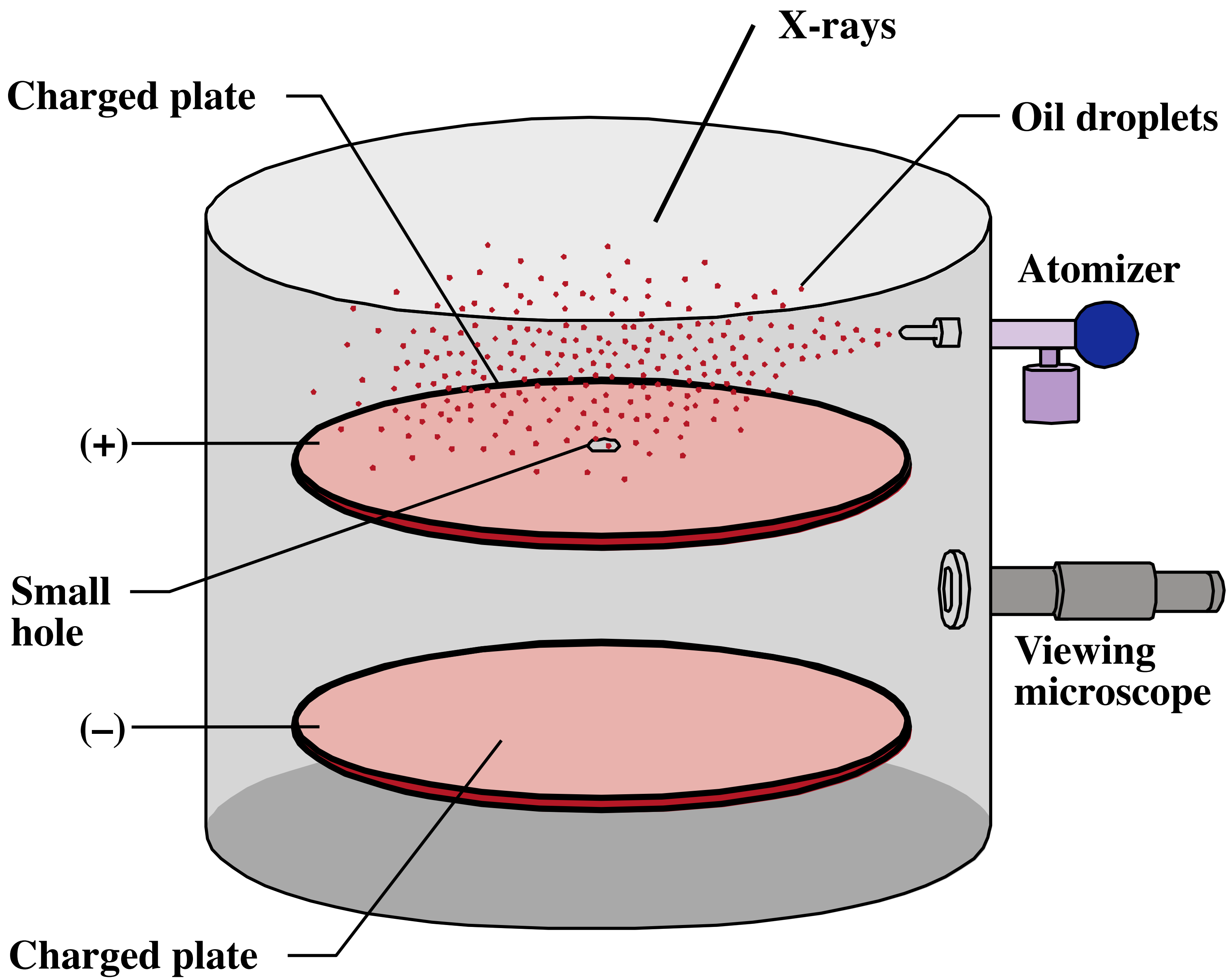
- The experiment
  - measured the strength of the electric field required to halt the free fall of the oil drops;
  - determined the mass of oil drops from their radii (size) and density.

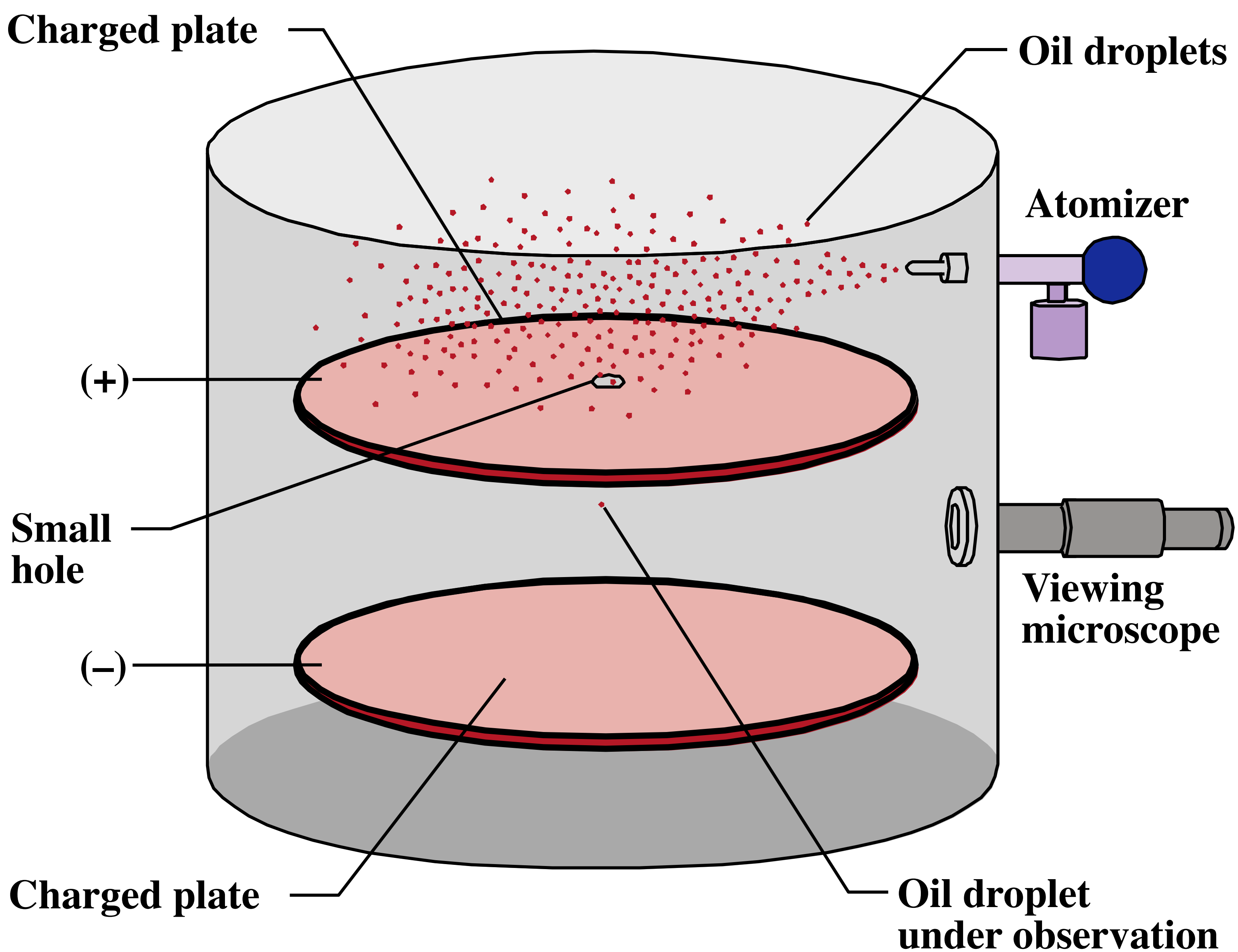


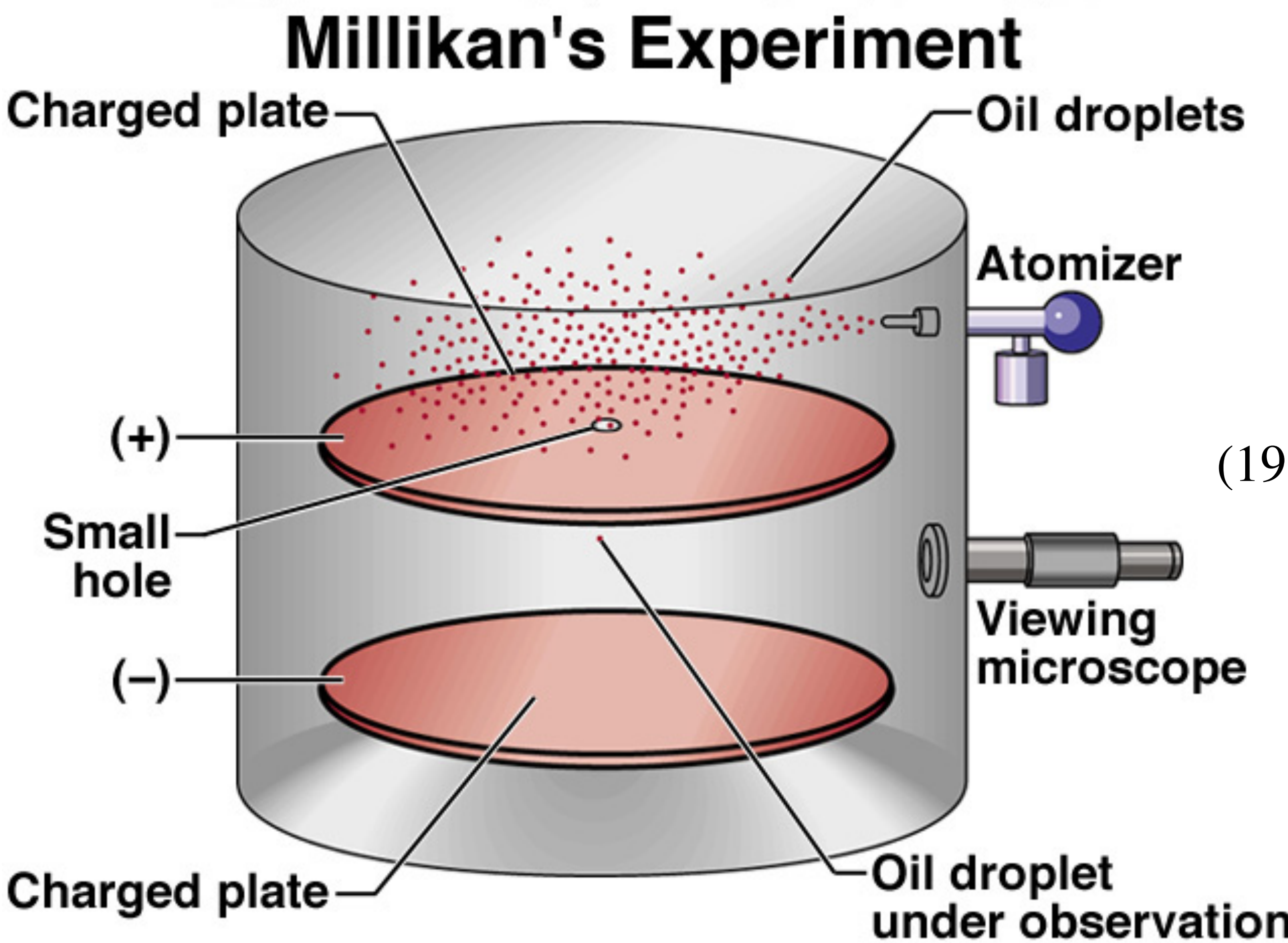
- The data allowed Millikan to calculate the charge of each drop.
  - Value determined was (-1.60×10^{-19} C) per drop of oil.



Charged plate
Millikan
Atomizer
(+)
Small hole
Viewing microscope
(–)
Charged plate


X-rays
Charged plate
Oil droplets
Atomizer
(+)
Small hole
Viewing microscope
(–)
Charged plate


Charged plate
Oil droplets
Atomizer
(+)
Small hole
Viewing microscope
(–)
Charged plate
Oil droplet under observation



Measured mass of e^-

(1923 Nobel Prize in Physics)

e^- charge = -1.60×10^{-19} C

Thomson's charge/mass of e^- = -1.76×10^{8} C/g

e^- mass = 9.10×10^{-28} g

H^+ mass = 1.67×10^{-24} g

Millikan's Oil Drop Experiment: The Charge-to-Mass Ratio for an Electron

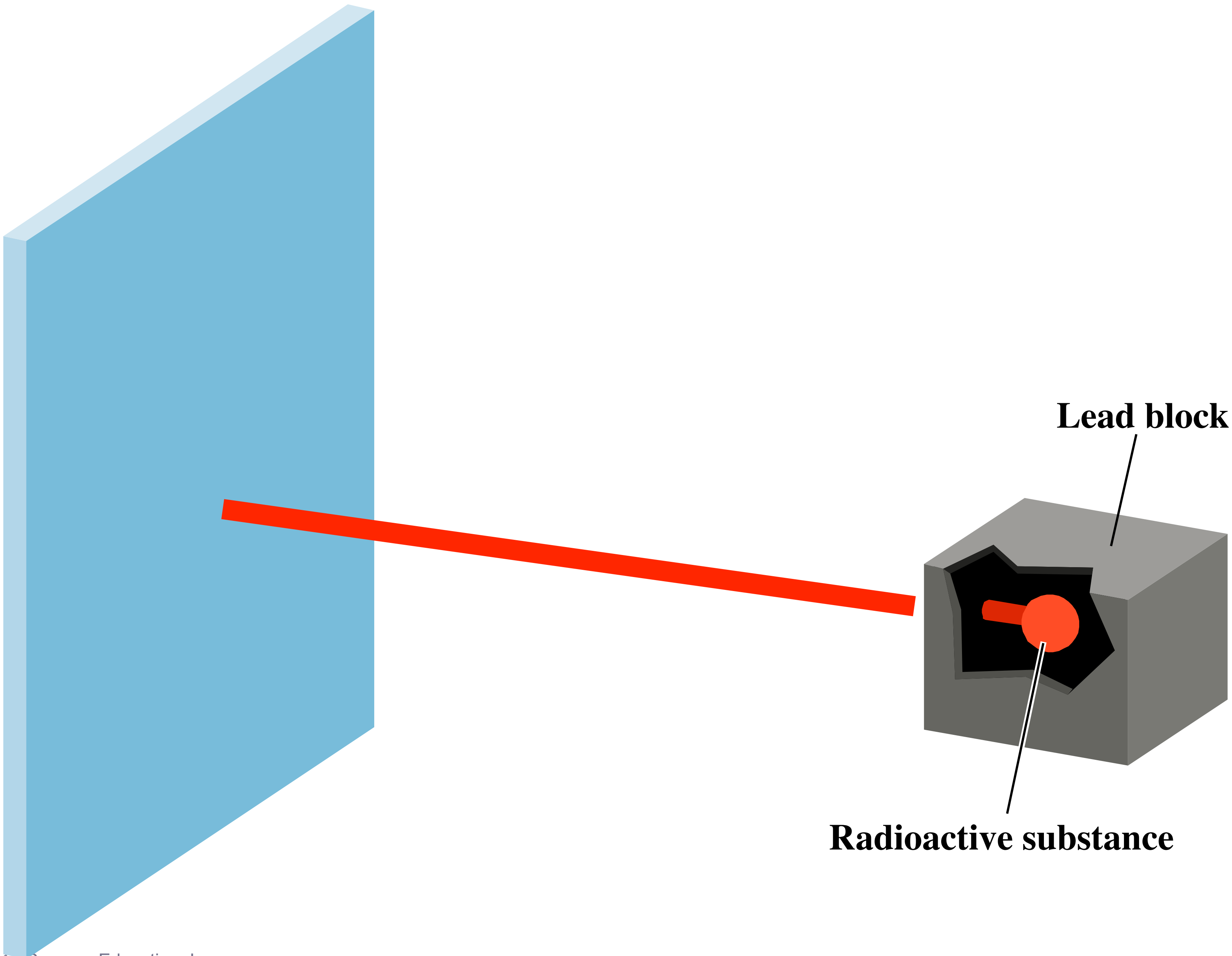
Using data from

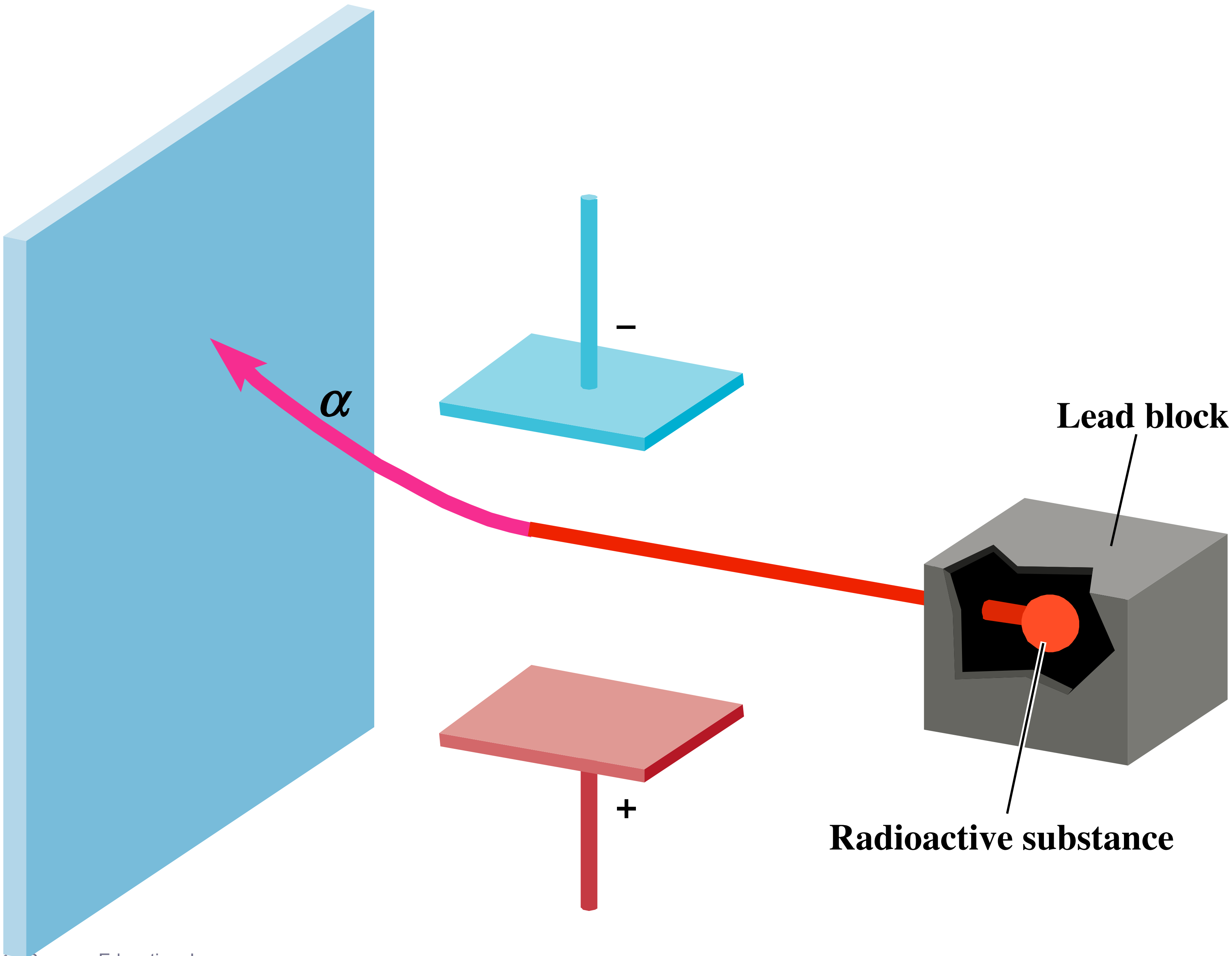
- Millikan's experiment (–1.60 × 10^{-19} C/electron);
- Thomson's mass-to-charge ratio for electrons, it can be deducted that the mass of an electron is as follows:

$$\cancel{\text{Charge}} \times \frac{\text{mass}}{\cancel{\text{charge}}} = \text{mass}$$

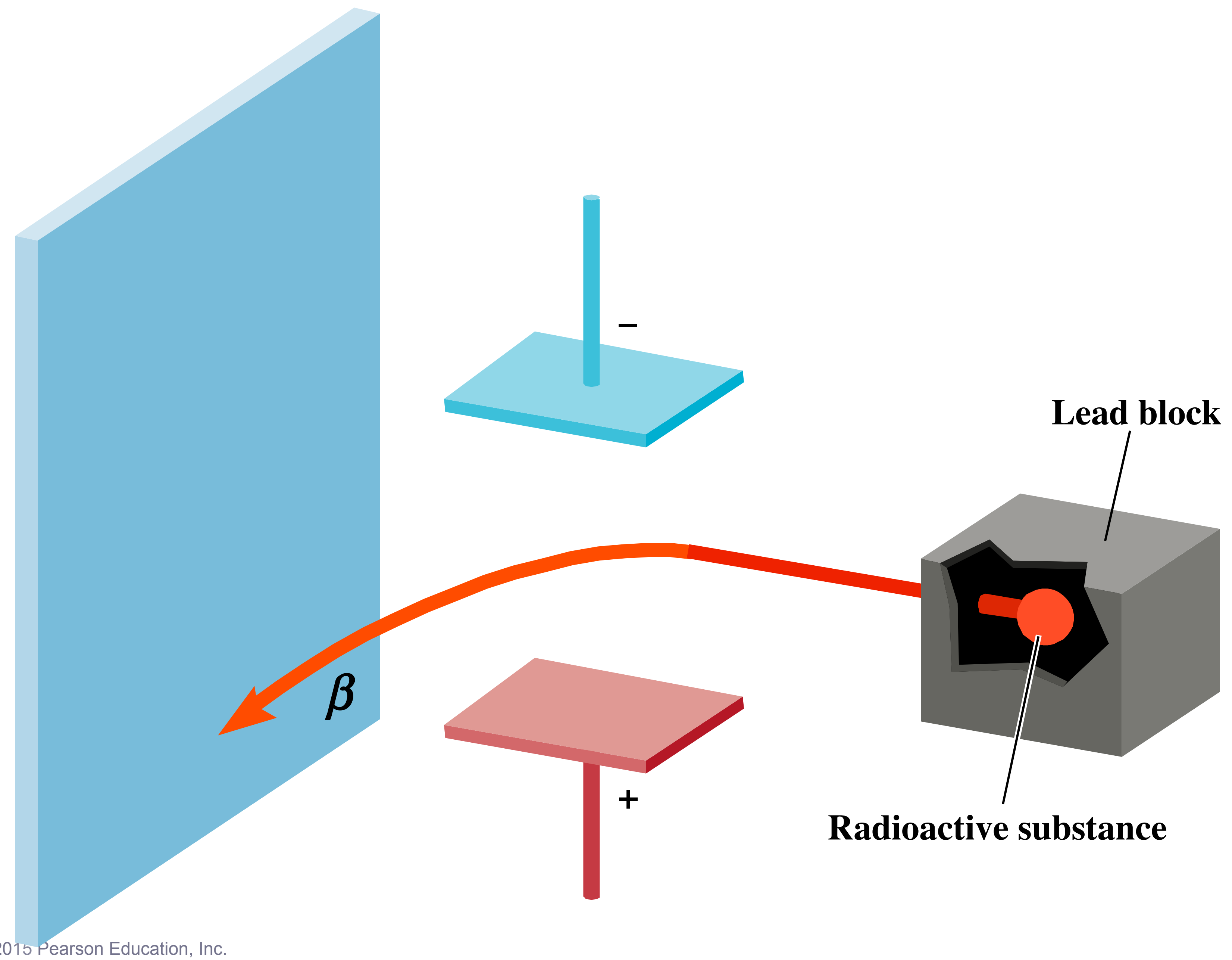
Radioactivity

Some elements: Uranium, Radium, Thorium
emit high energy radiation:
alpha, beta, gamma rays

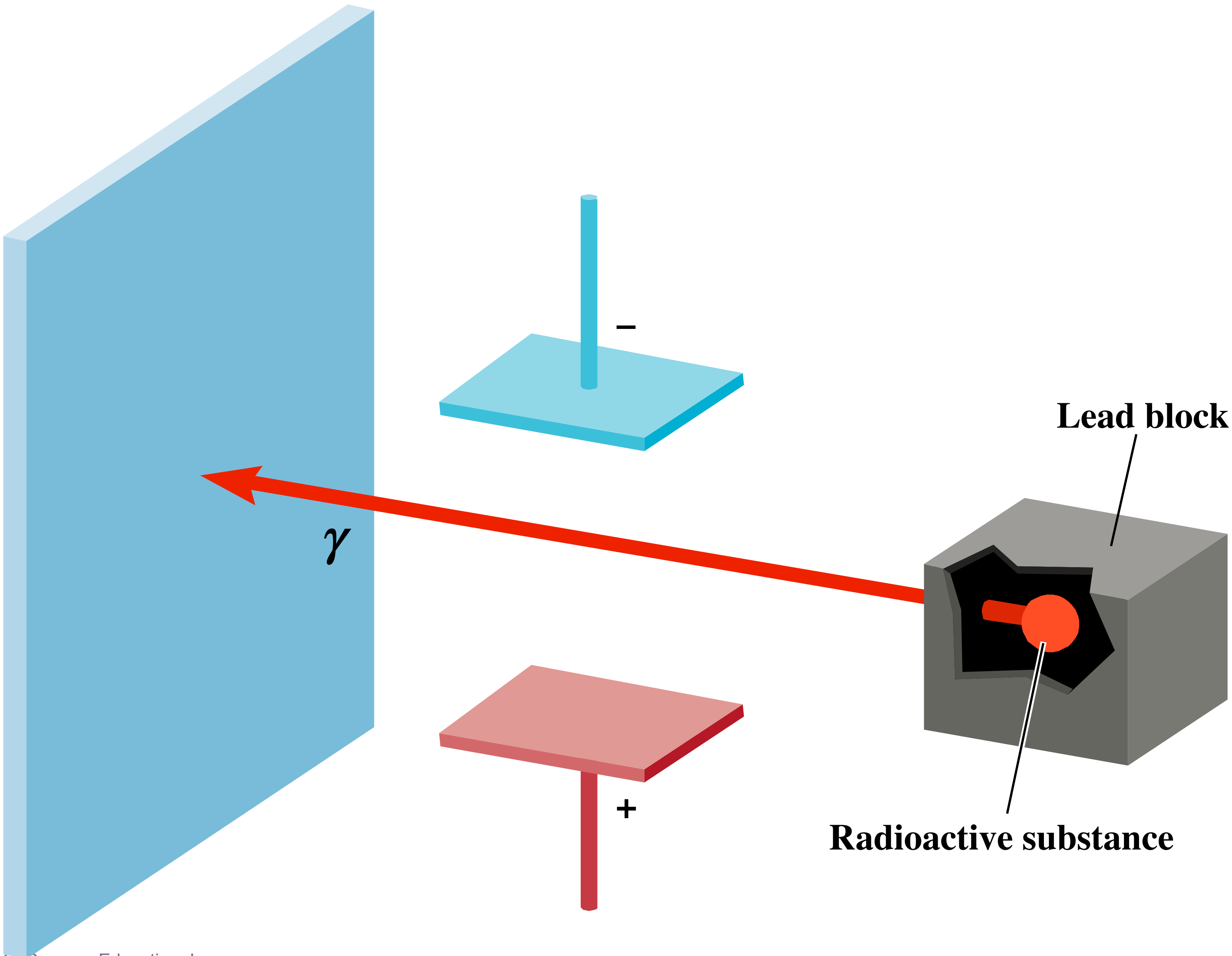

Lead block
Radioactive substance



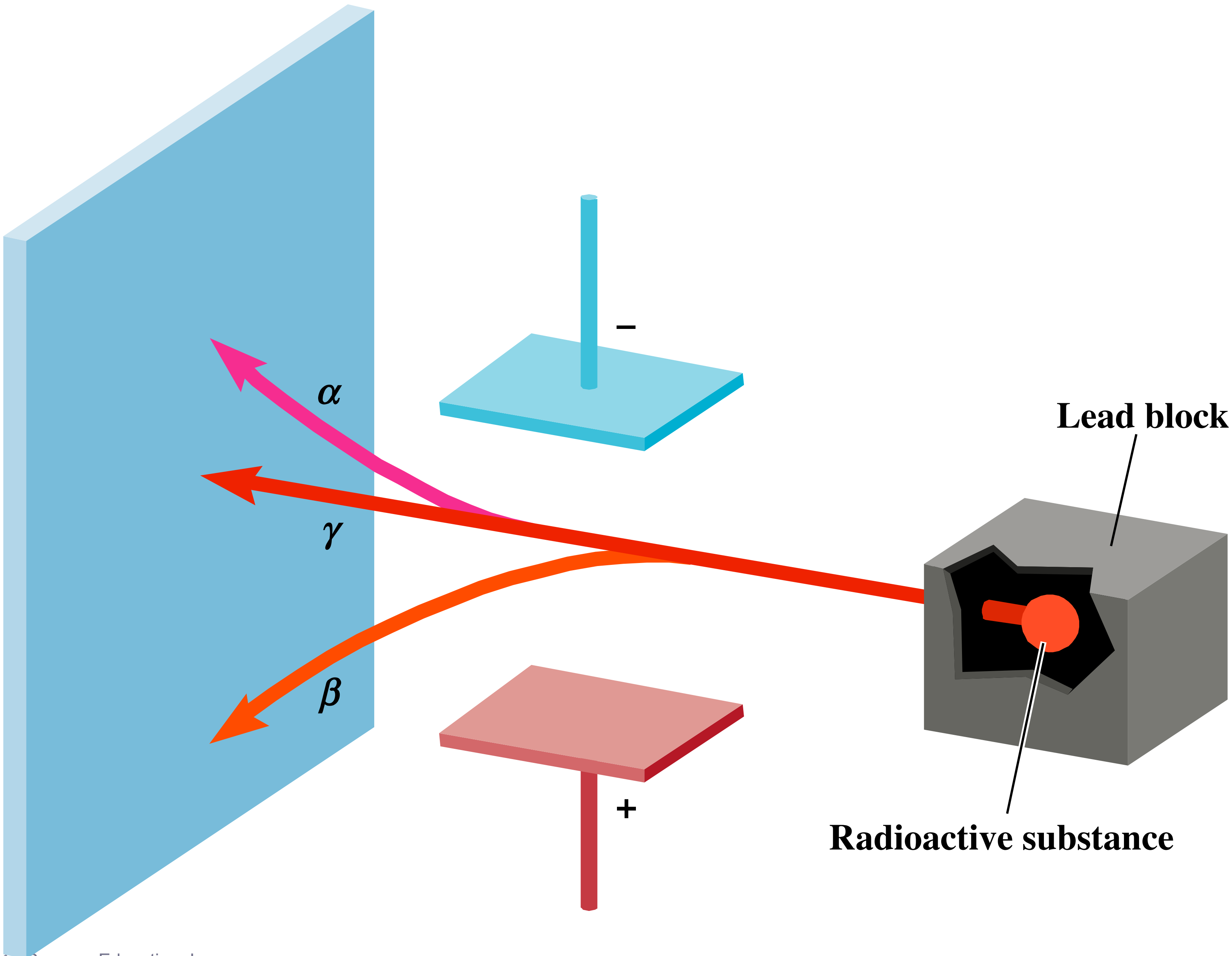
α
−
Lead block
+
Radioactive substance



−
Lead block
β
+
Radioactive substance



Lead block
−
γ
+
Radioactive substance



Lead block
−
α
γ
β
+
Radioactive substance

Putting the pieces together:

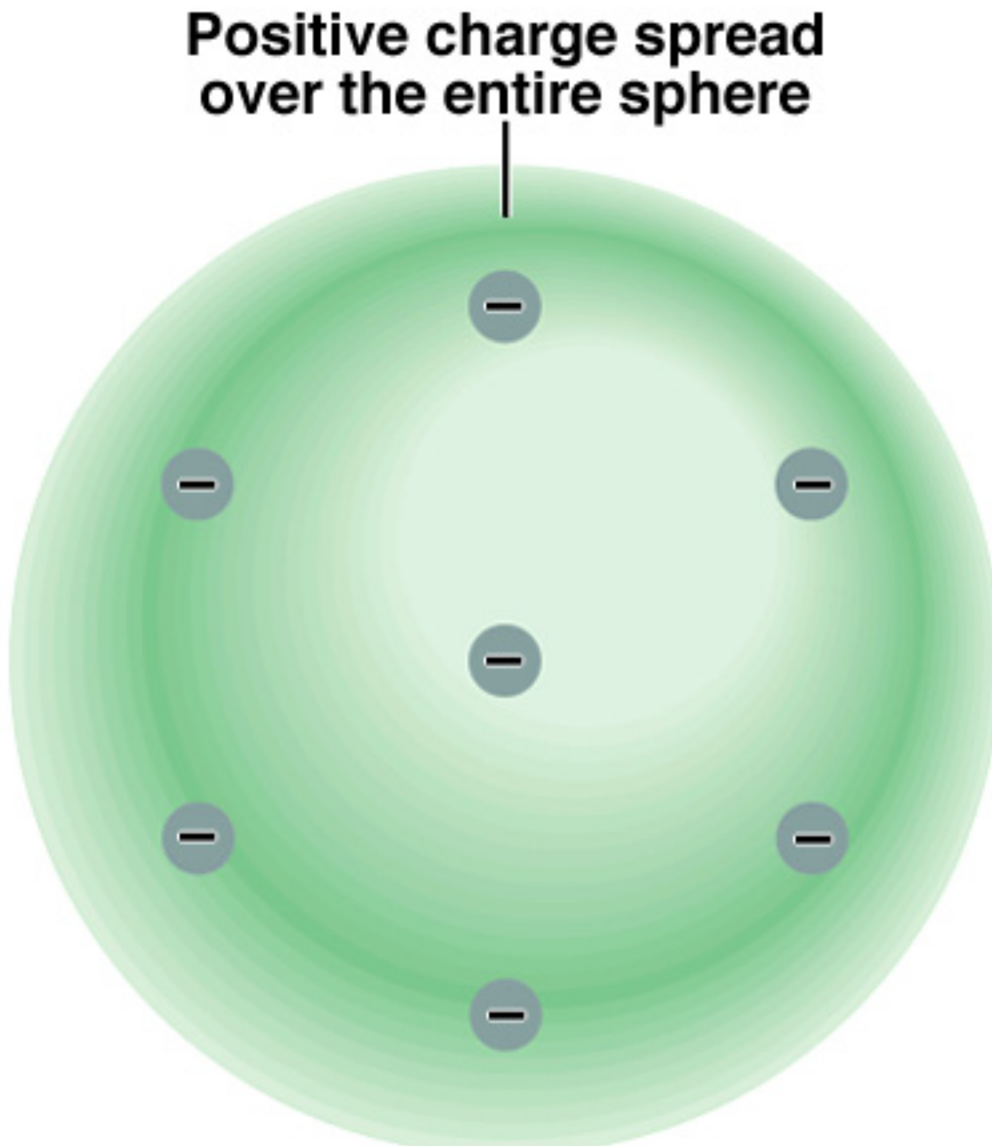
Beta particles: negatively charged = electrons

Alpha particles: positively charged
Ernest Rutherford showed that these are chemically part of helium atoms

Gamma Rays: like light and X-rays, only more energy

So where are the Electrons and the Positive Charge in an atom?

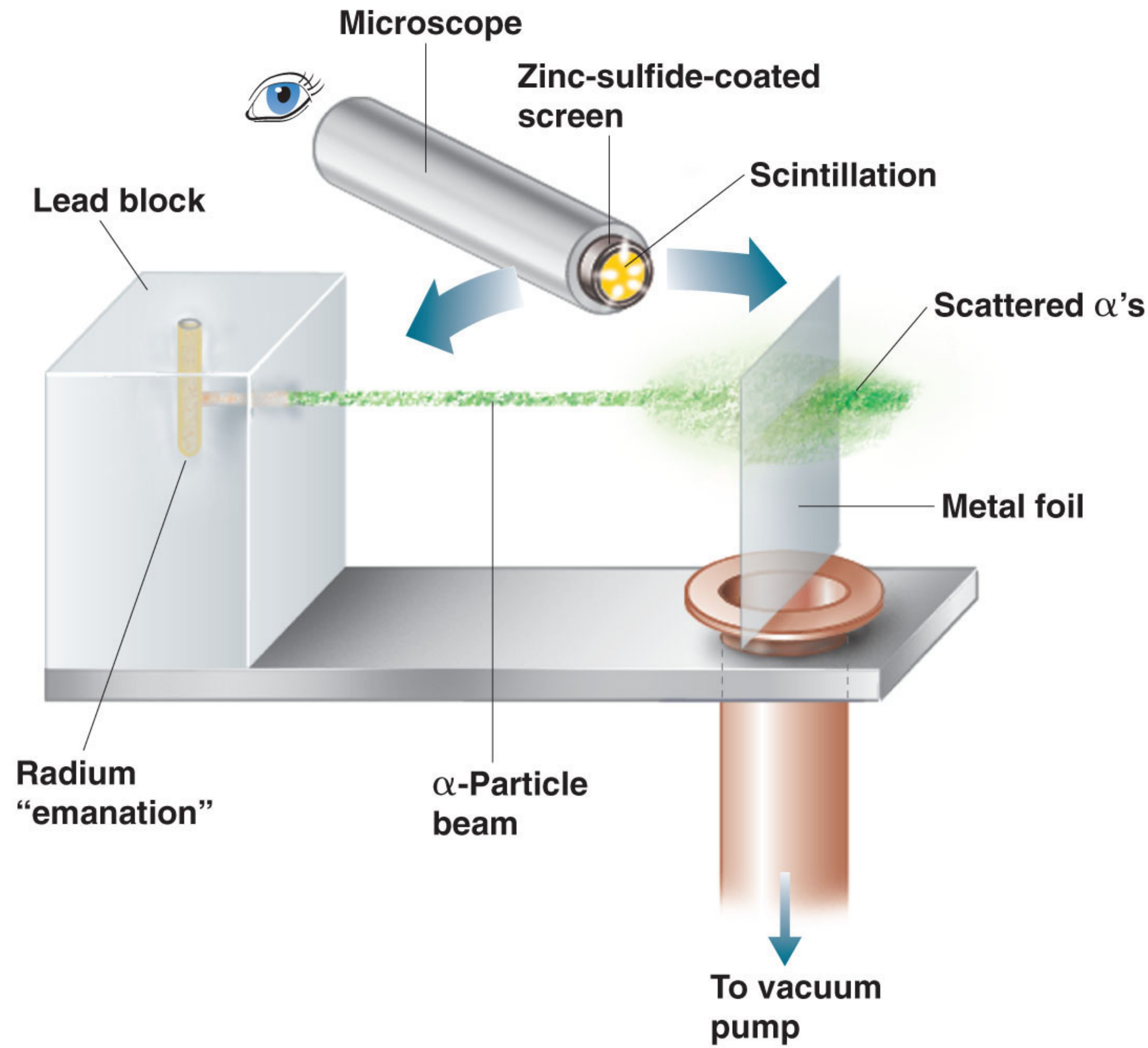
Thomson's Model of the Atom

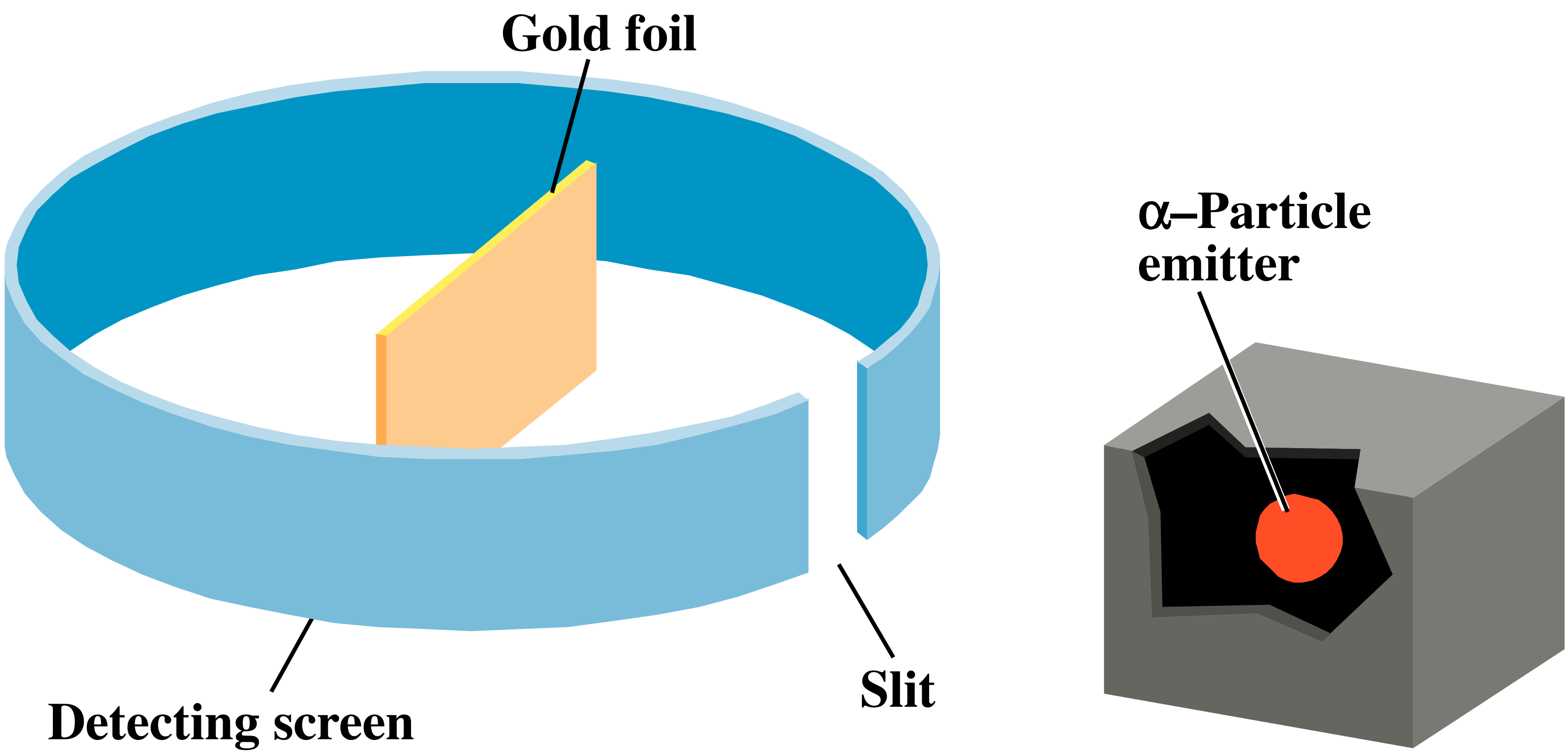


The Ultimate Undergraduate Research Project

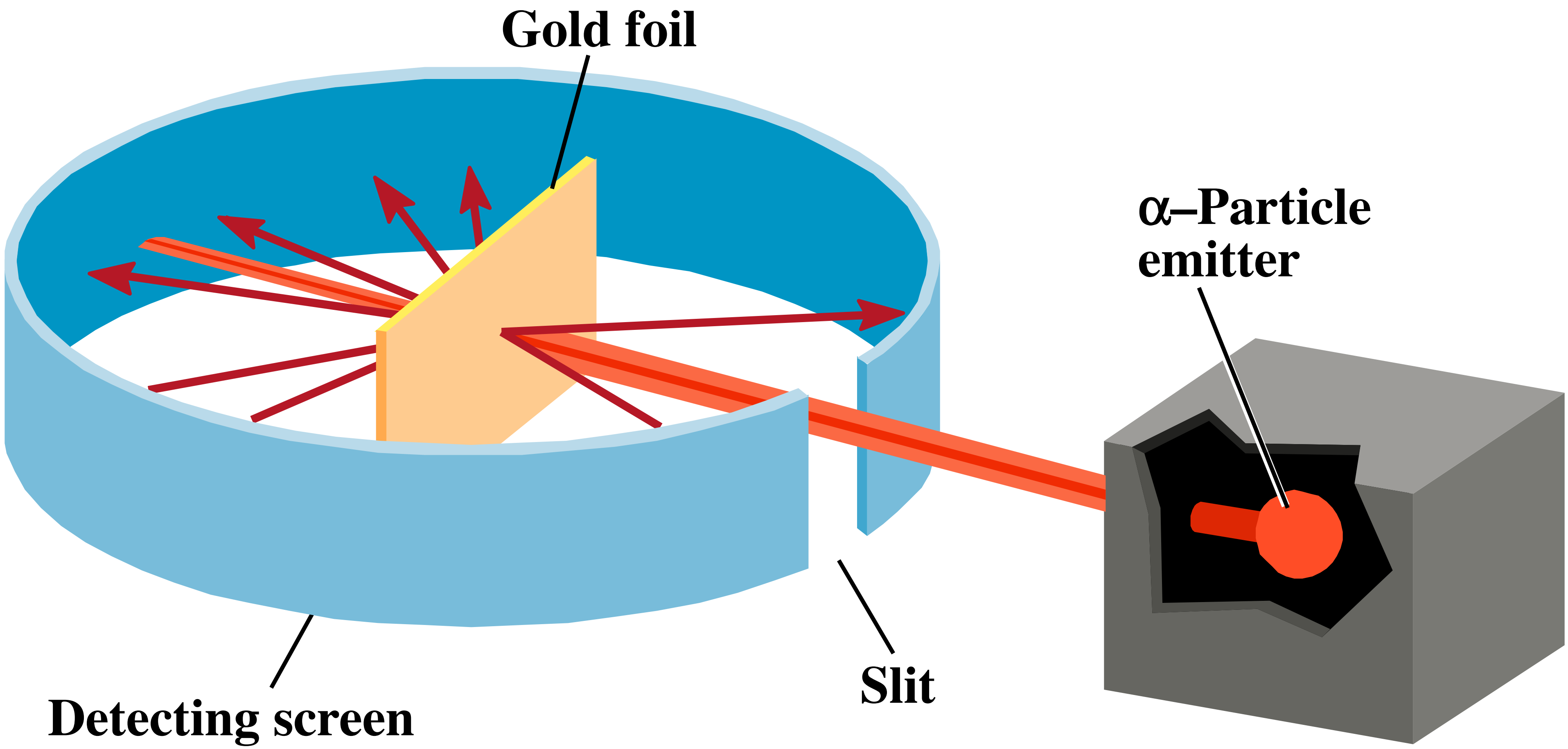
(as carried out by Ernest Marsden under the direction of Johannes Gieger and Ernest Rutherford)

Rutherford's Experimental Design:


Microscope
Zinc-sulfide-coated
screen
Scintillation
Lead block
Scattered α's
Metal foil
Radium
"emanation"
α-Particle
beam
To vacuum
pump



(a)



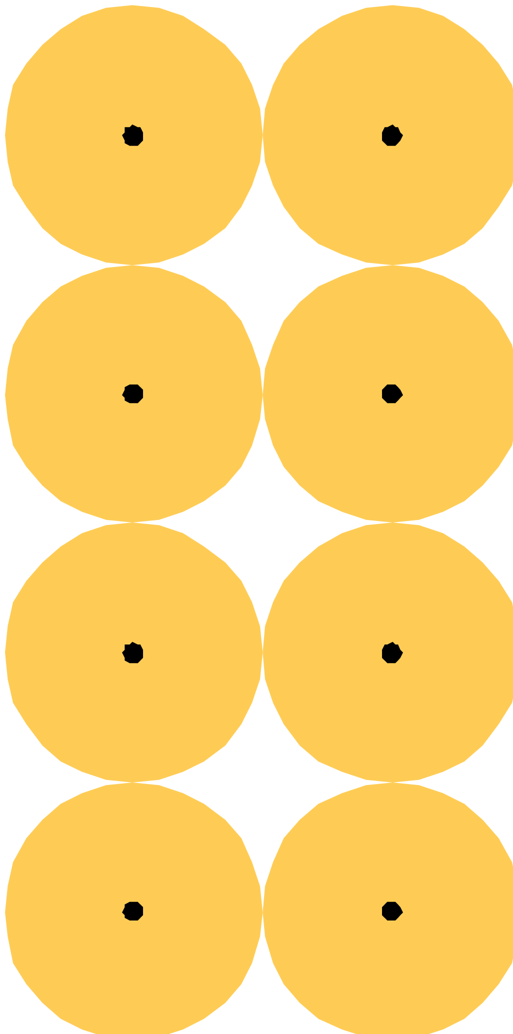
(a)

Thomson's Model Predicts:

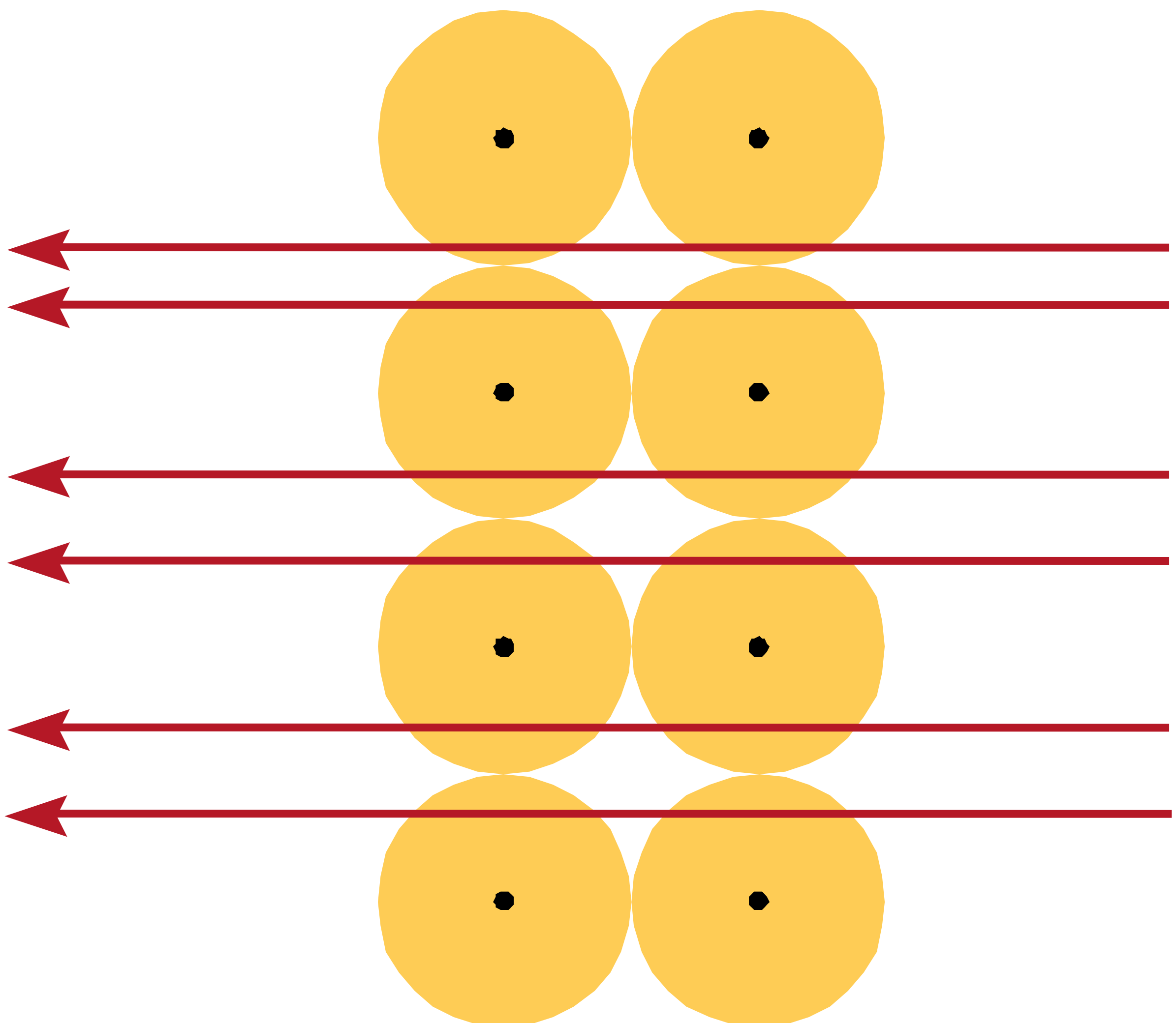
Multiple collisions, small deflections

Rutherford on the large deflections:

'.. *About as credible as if you had fired a 15-inch [artillery] shell at a piece of paper and it came back and hit you*.'



(b)



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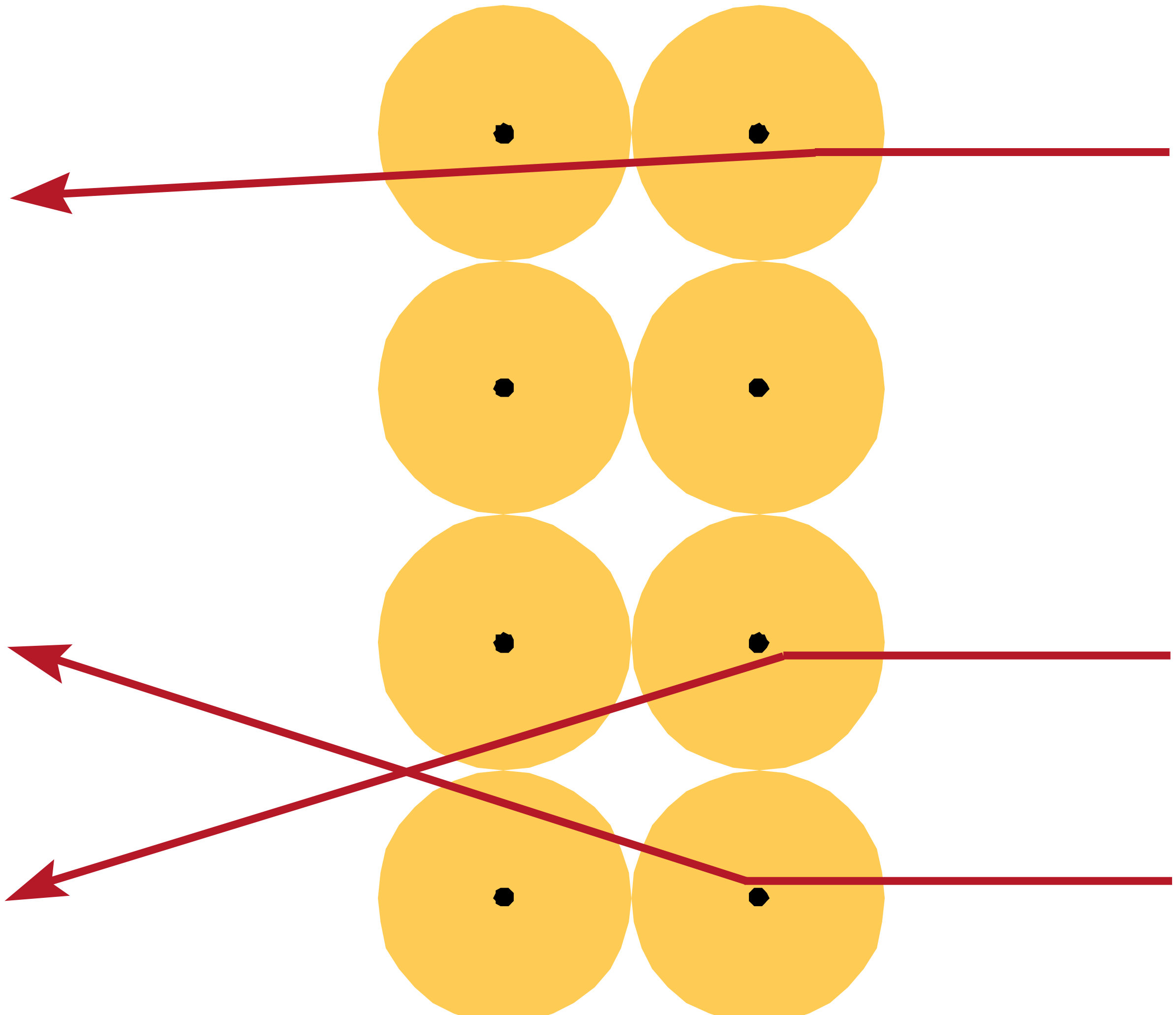

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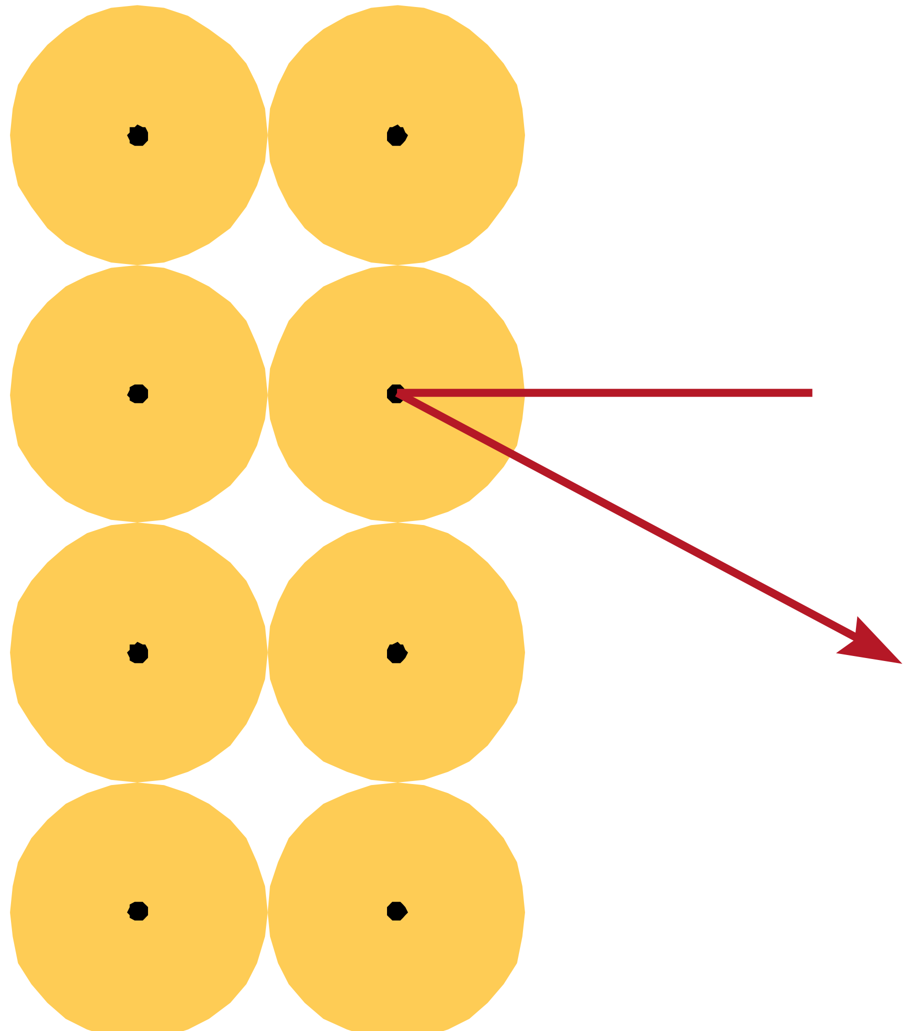
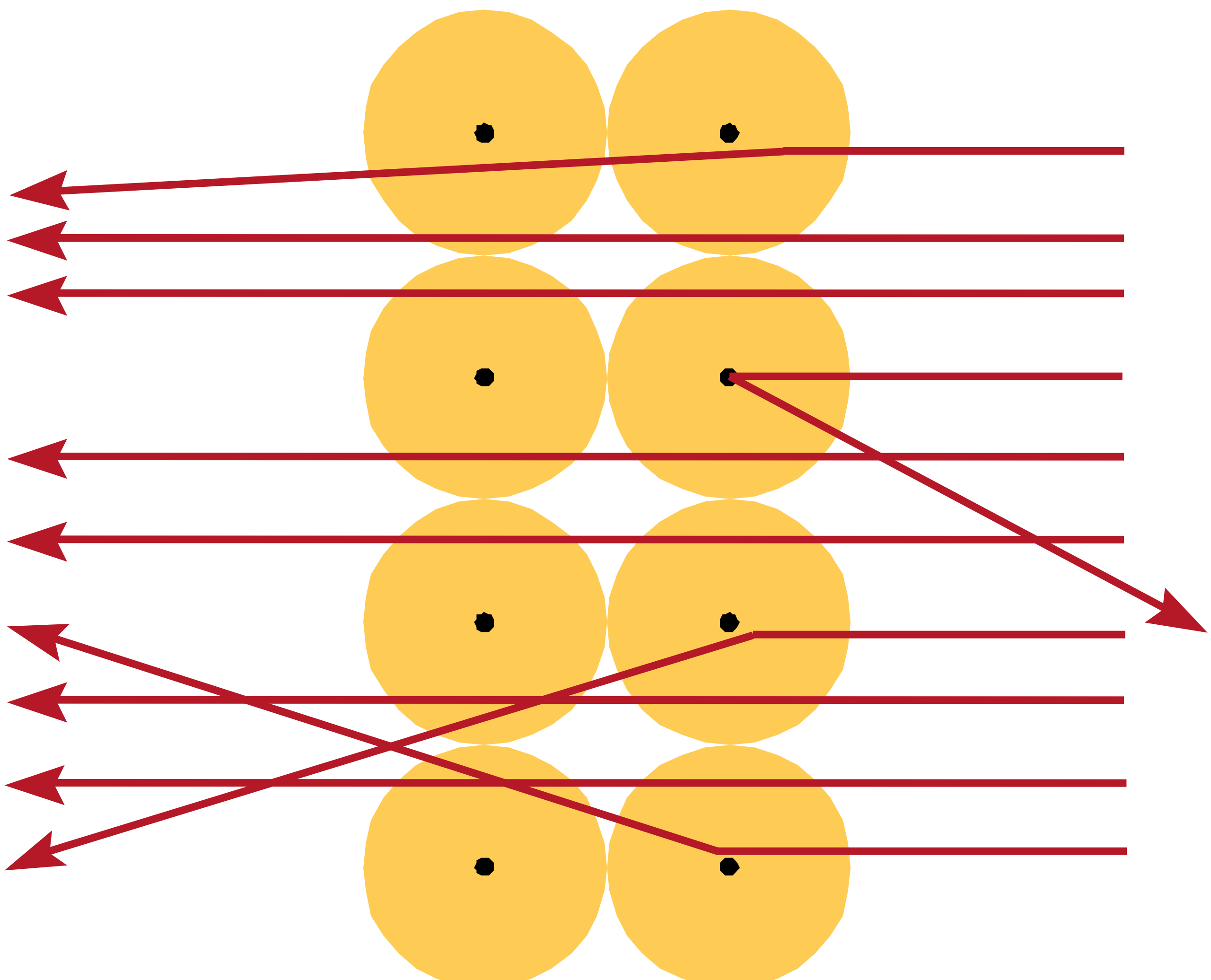
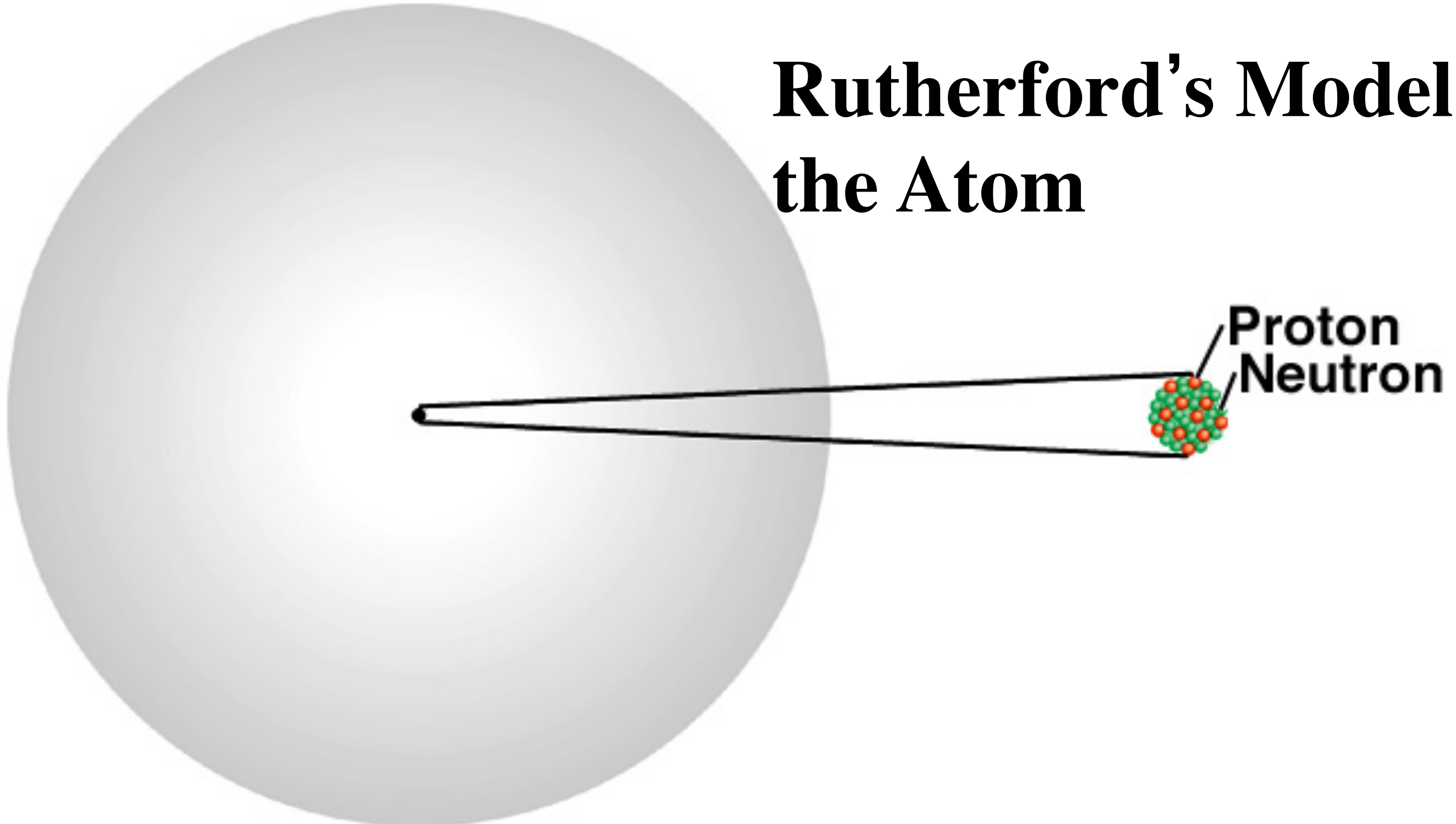


(b)



(b)

Rutherford's Model of the Atom



atomic radius ~ 100 pm = 1×10^{-10} m

nuclear radius ~ 5×10^{-3} pm = 5×10^{-15} m

Rutherford's Model of the Atom

If the atom is Bryant Denny Stadium, then the nucleus is a marble on the 50-yard line.

Ernest Rutherford, Baron Nelson, BSc, DSc, Canterbury College, University of New Zealand

Lecture Theatre
The University of Canterbury, Old Site

Building on the Rutherford Atomic Model: The Nuclear Atom Model

- The **nuclear theory** of the atom has three basic parts:
  1. Most of the atom's mass and all of its positive charge are contained in a small core called a **nucleus**.
  2. Most of the volume of the atom is empty space, throughout which tiny, negatively charged electrons are dispersed.
  3. There are as many negatively charged electrons outside the nucleus as there are positively charged particles (named **protons**) within the nucleus, so that the atom is electrically neutral.

Protons (and Neutrons)

Had bare proton (nucleus of hydrogen)
Charge of alpha particle x 2 proton

Mass of alpha particle x 4 proton

So must be something else in the nucleus:

Neutron

Chadwick's Experiment (1932)

H atoms - 1 p; He atoms - 2 p

mass He/mass H should = 2

measured mass He/mass H = 4

$$\alpha + {}^{9}Be \longrightarrow {}^{1}n + {}^{12}C + \text{energy}$$

neutron (n) is neutral (charge = 0)

n mass ~ p mass = 1.67×10^{-24} g

The Atom's Subatomic Particles

- All atoms are composed of the same subatomic particles:
  - Protons
  - Neutrons
  - Electrons

- Protons and neutrons have nearly identical masses.
  - The mass of the proton is 1.67262×10^{-27} kg.
  - The mass of the neutron is 1.67493×10^{-27} kg.
  - The mass of the electron is 9.1×10^{-31} kg.

- The charge of the proton and the charge of the electron are equal in magnitude but opposite in sign. The neutron has no charge.

Subatomic Particles

TABLE 1.2 Subatomic Particles

	Mass (kg)	Mass (amu)	Charge (relative)	Charge (C)
Proton	1.67262×10^{-27}	1.00727	1+	$+1.60218 \times 10^{-19}$
Neutron	1.67493×10^{-27}	1.00866	0	0
Electron	0.00091×10^{-27}	0.00055	1−	-1.60218×10^{-19}

Elements: Defined by Their Numbers of Protons

- The most important number to the identity of an atom is the number of protons in its nucleus.
- *The number of protons defines the element*.
- The number of protons in an atom's nucleus is its **atomic number** and is given the symbol ***Z***.

Isotopes: Elements with Varied Number of Neutrons

- All atoms of a given element have the same number of protons; however, they do not necessarily have the same number of neutrons.
  - Example:
    - All neon atoms contain 10 protons, but they may contain 10, 11, or 12 neutrons.
    - All three types of neon atoms exist, and each has a slightly different mass.

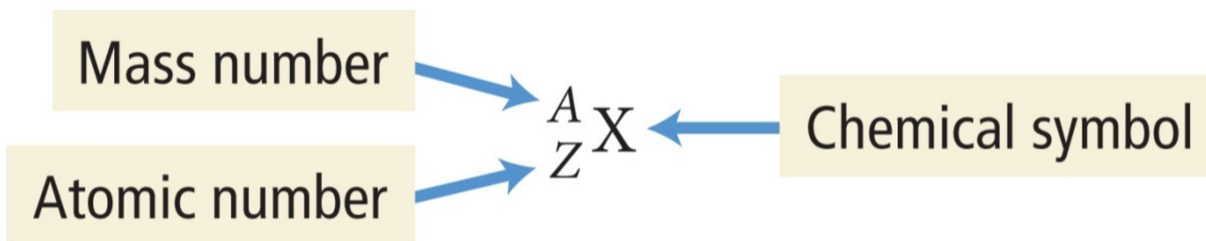
$$^{20}_{10}\mathrm{Ne} \quad ^{21}_{10}\mathrm{Ne} \quad ^{22}_{10}\mathrm{Ne}$$

- Atoms with the same number of protons but a different number of neutrons are called **isotopes**.

Isotopes: Representation

- The sum of the number of neutrons and protons in an atom is its **mass number** and is represented by the symbol ***A***.

***A* = number of protons (p^+) + number of neutrons (n)**



- X is the chemical symbol, *A* is the mass number, and *Z* is the atomic number.

Isotopes: Representation

- A second common notation for isotopes is the chemical symbol (or chemical name) followed by a dash and the mass number of the isotope.

Chemical symbol or name → $X\text{-}A$ ← Mass number

Ne-20	Ne-21	Ne-22
neon-20	neon-21	neon-22

Isotopes: Varied Number of Neutrons

- The relative amount of each different isotope in a naturally occurring sample of a given element is roughly constant.

- These percentages are called the **natural abundance** of the isotopes.

Symbol	Number of Protons	Number of Neutrons	*A* (Mass Number)	Natural Abundance (%)
Ne−20 or $^{20}_{10}\mathrm{Ne}$	10	10	20	90.48
Ne−21 or $^{21}_{10}\mathrm{Ne}$	10	11	21	0.27
Ne−22 or $^{22}_{10}\mathrm{Ne}$	10	12	22	9.25

Ions: Charged Atoms
Losing and Gaining Electrons

- The number of electrons in a neutral atom is equal to the number of protons in its nucleus (designated by its atomic number *Z*).

- In chemical changes, however, atoms can lose or gain electrons and become charged particles called **ions**.

  – **Positively charged ions** are called **cations**.
    - Metal elements, such as Na^+, form cations.

  – **Negatively charged ions** are called **anions**.
    - Nonmetal elements, such as F^-, form anions.

Atomic Mass: The Average Mass of an Element's Atoms

- Atomic mass is sometimes called *atomic weight* or *standard atomic weight.*
- The atomic mass of each element is directly beneath the element's symbol in the periodic table.
- The atomic mass of an element represents the **average mass of the isotopes** that compose that element.
  - It is a *weighted value based on the element's natural abundance of each isotope*.

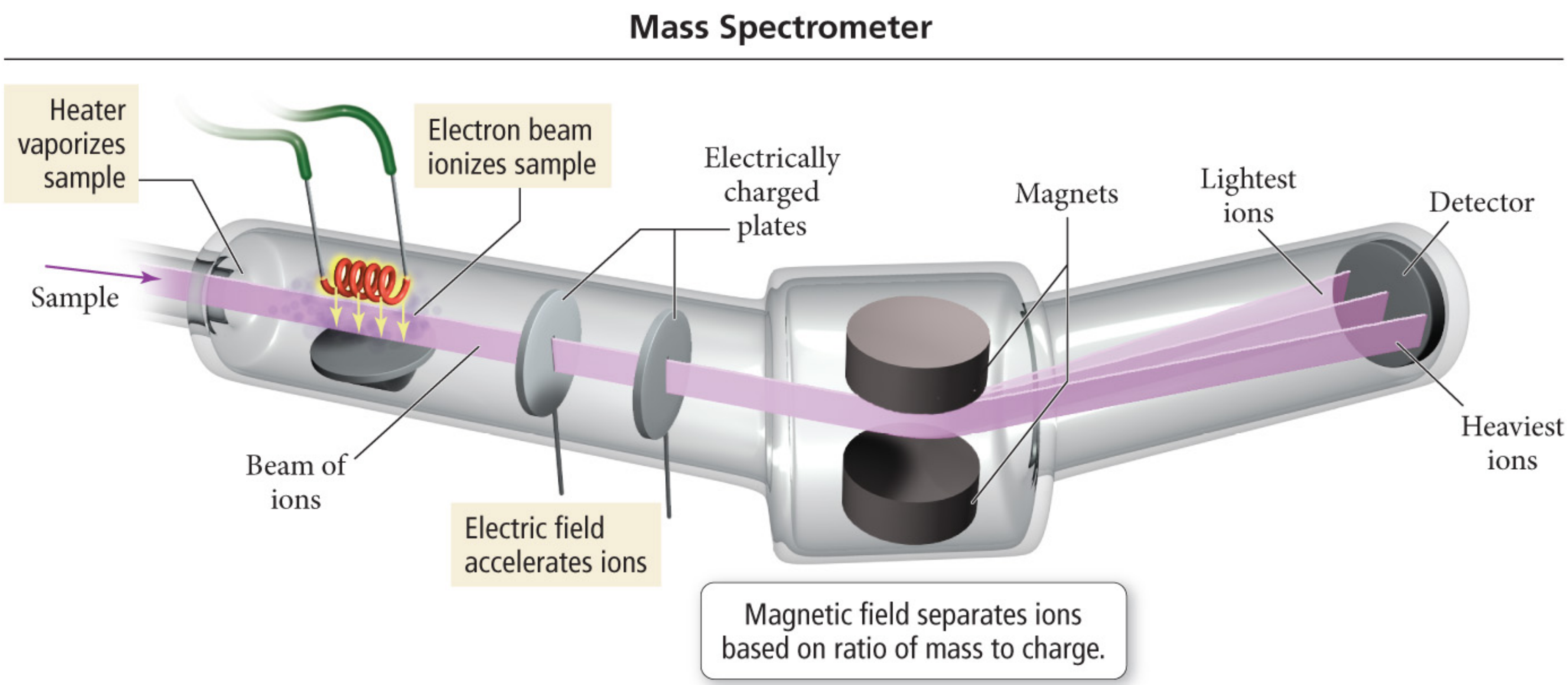
Atomic Mass

17
Cl
35.45
chlorine

- In general, we calculate the atomic mass with the following equation:

$$\begin{aligned}
\text{Atomic mass} &= \sum_{n} (\text{fraction of isotope } n) \times (\text{mass of isotope } n) \\
&= (\text{fraction of isotope 1} \times \text{mass of isotope 1}) \\
&+ (\text{fraction of isotope 2} \times \text{mass of isotope 2}) \\
&+ (\text{fraction of isotope 3} \times \text{mass of isotope 3}) \ + \ \ldots
\end{aligned}$$

Mass Spectrometry



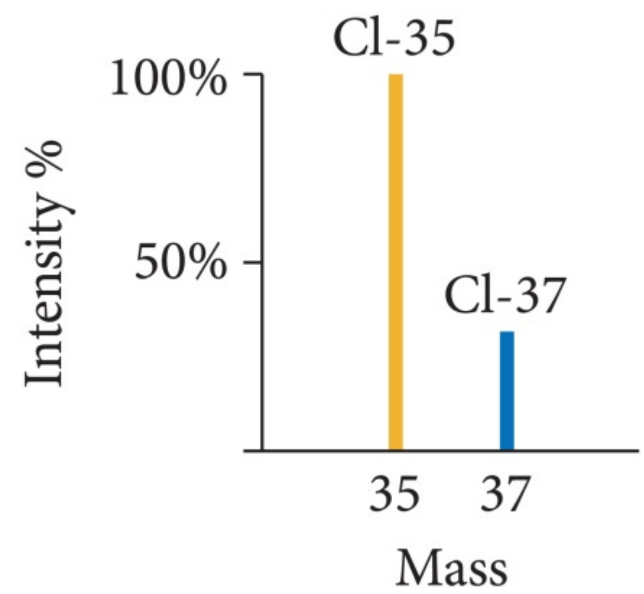
$$KE = 1/2 \times m \times v^2$$

$$v = (2 \times KE/m)^{1/2}$$

$$F = q \times v \times B$$

Mass Spectrometry: Measuring the Mass of Atoms and Molecules

- The masses of atoms and the percent abundances of isotopes of elements are measured using **mass spectrometry**—a technique that separates particles according to their mass.



Atomic Mass: Problem

- Naturally occurring chlorine consists of 75.77% chlorine-35 atoms (mass 34.97 amu) and 24.23% chlorine-37 atoms (mass 36.97 amu).
- Calculate chlorine's atomic mass.