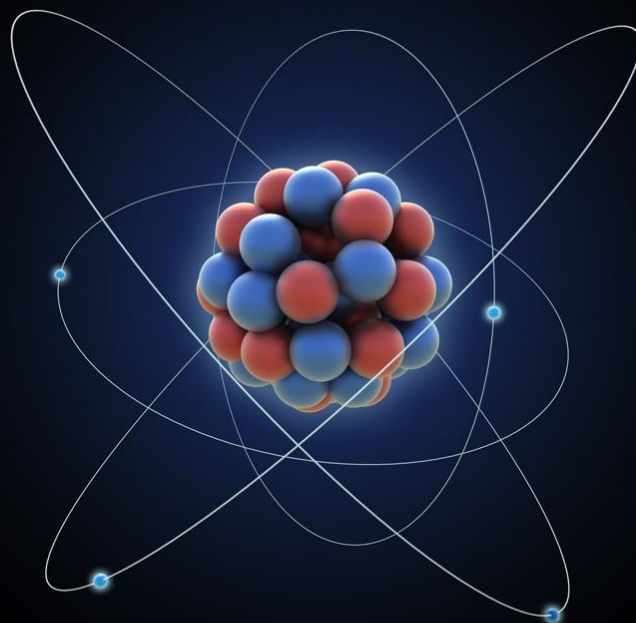




Chapter 2. Atoms, Molecules and Ions

General Chemistry I, CHEM 103

Lecture 3: Atoms, Molecules and History of Atomic Theory

Register for Achieve: <https://achieve.macmillanlearning.com/courses/4txh35>.

Course ID: 4txh35

Achieve HW #1 due this Wednesday 9/9 at 5 pm

Learning Objectives for Chapter Units 2.1 – 2.4

- Describe the basic properties of each physical state of matter: solid, liquid, and gas.
- Apply the law of conservation of matter.
- Classify matter as an element, compound, homogeneous mixture, or heterogeneous mixture with regard to its physical state and composition.
- Define and give examples of atoms and molecules.
- Identify properties of and changes in matter as physical or chemical.
- Outline milestones and conclusions in the development of modern atomic theory.
- Summarize and interpret the results of the experiments of Thomson, Millikan, Rutherford, and Chadwick.
- Understand that elements and compounds are pure substances. Explain the difference between elements and compounds in terms of composition.
- Explain the structure of an atom.
- Calculate the number of protons, neutrons and electrons in an atom or an ion when provided appropriate resources as the ones introduced in class (e.g., periodic table, atom shorthand notation, isotope notation, charge...)
- Calculate atomic weights of elements when given isotope distribution or mass spectrometry data. Solve problems involving isotope distribution and atomic masses.

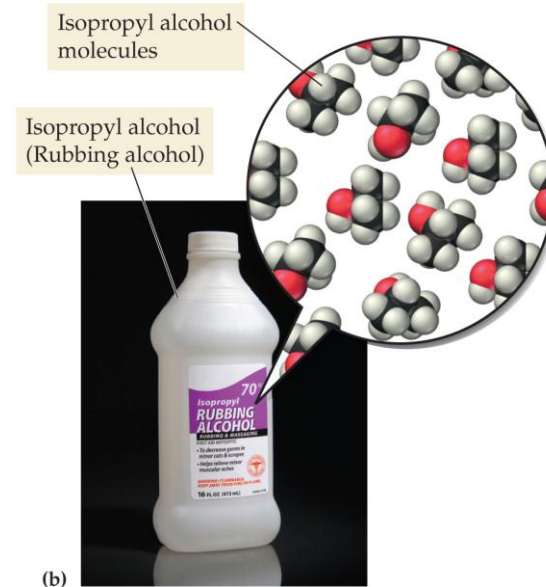
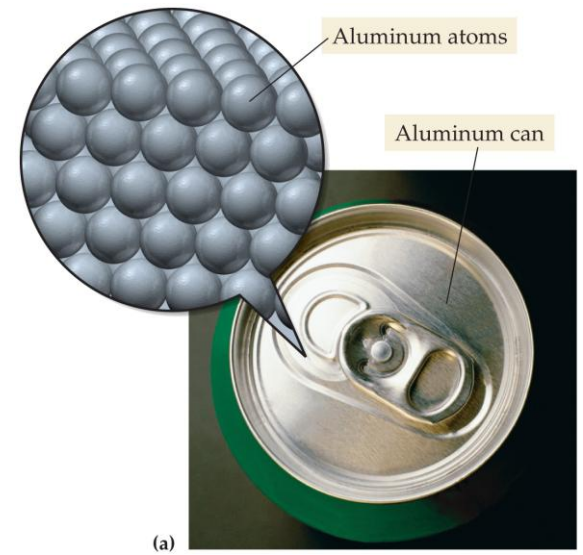
Matter

Matter is defined as anything that occupies space and has mass.

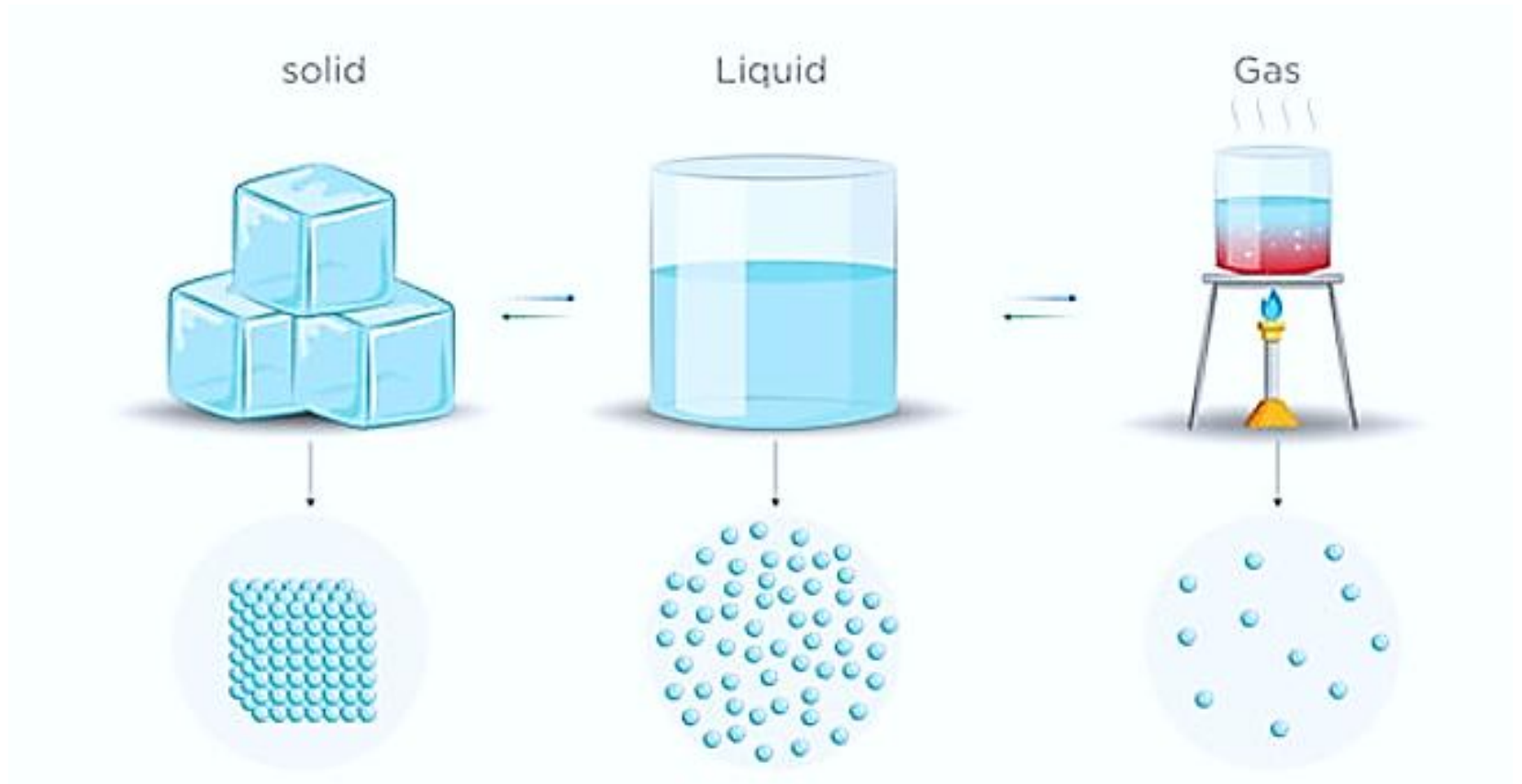
Matter can be visible or microscopic (ex: air or dust)

Atoms: fundamental building blocks of matter.

Molecules: **two or more atoms** joined to one another in specific geometric arrangements.



Three States of Matter



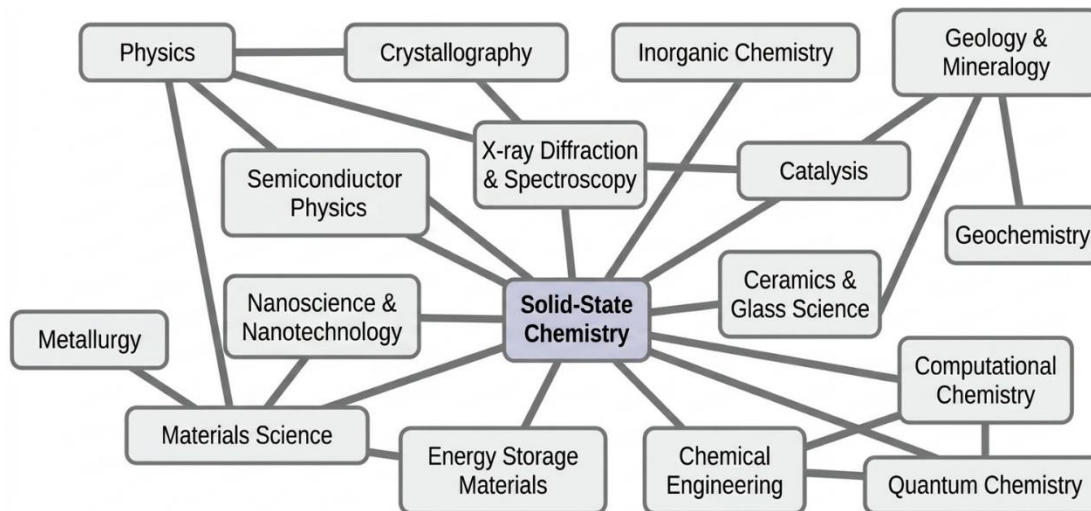
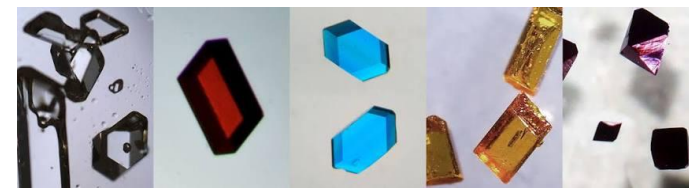
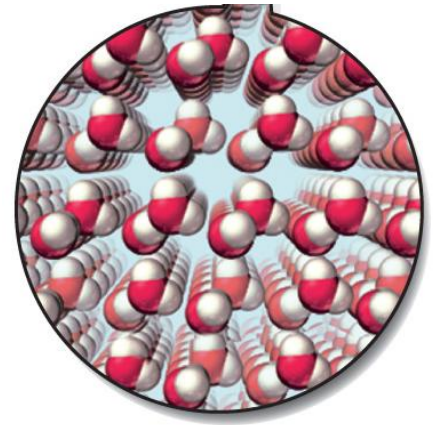
States of Matter: Solid

Atoms or molecules **pack close** to each other in fixed locations.

Neighboring atoms or molecules in a solid may **vibrate or oscillate**, but they do not move around each other.

Solids have **fixed volume** and rigid shape.

Ex: Ice, diamond, quartz, iron



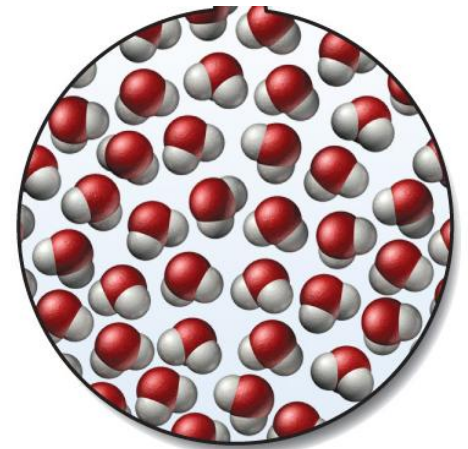
States of Matter: Liquid

Atoms or molecules are close to each other but are free to move around.

Liquids have a fixed volume because their atoms or molecules are in close contact.

Liquids assume the shape of their containers because the atoms or molecules are free to move around each other.

Ex: Water, gasoline, alcohol, mercury.



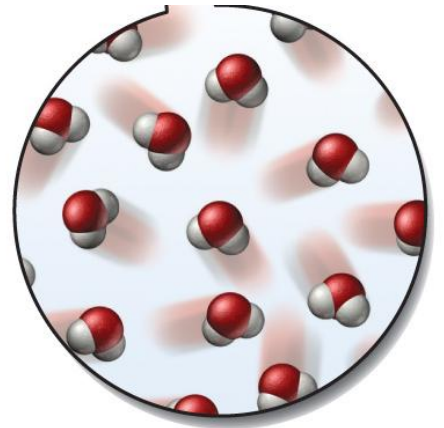
States of Matter: Gas

Atoms or molecules are separated by large distances and are free to move relative to one another.

Atoms or molecules that compose gases are not in contact with one another, so gases are compressible.

Gases always assume the shape and volume of their container.

Ex: Oxygen, helium, carbon dioxide

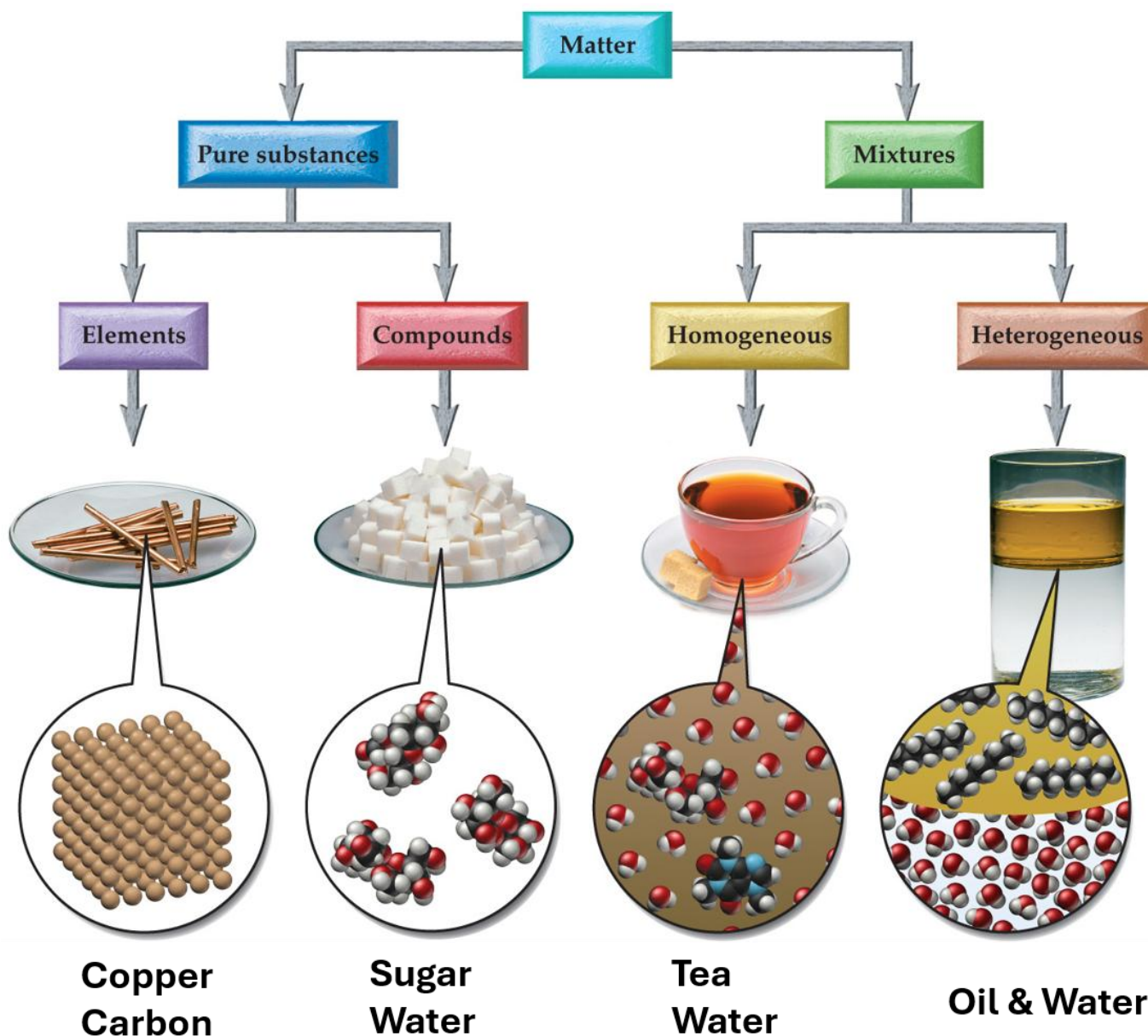


Properties of Solids, Liquids, and Gases

State	Atomic/ Molecular Motion	Atomic/ Molecular Spacing	Shape	Volume	Compressibility
Solid	Oscillation/ vibration about fixed point	Close together	Definite	Definite	Incompressible
Liquid	Free to move relative to one another	Close together	Indefinite	Definite	Incompressible
Gas	Free to move relative to one another	Far apart	Indefinite	Indefinite	Compressible

Classifying matter: elements, compounds and mixtures

But what is everything fundamentally made of?



Properties of Matter

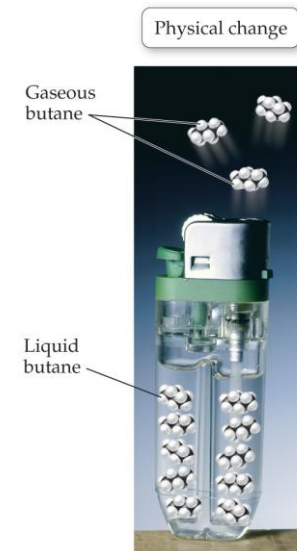
- Physical change

Matter changes appearance
Composition stays the same
Can (often) be reversed
Phase change

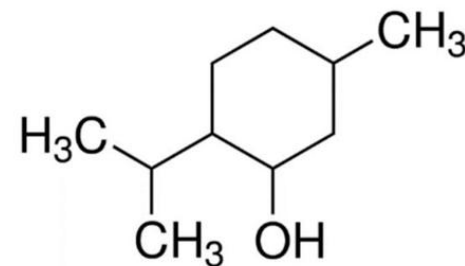
melting ice (solid water) to form liquid water

boiling liquid water to form steam (gaseous water)

tearing a piece of paper



Menthol



Properties of Matter

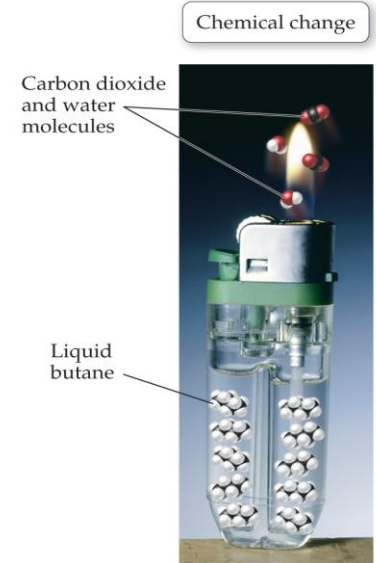
Chemical change

Change in composition of matter

Chemical reaction occurs

Not reversible without another reaction

Color change, production of light, heat, gas

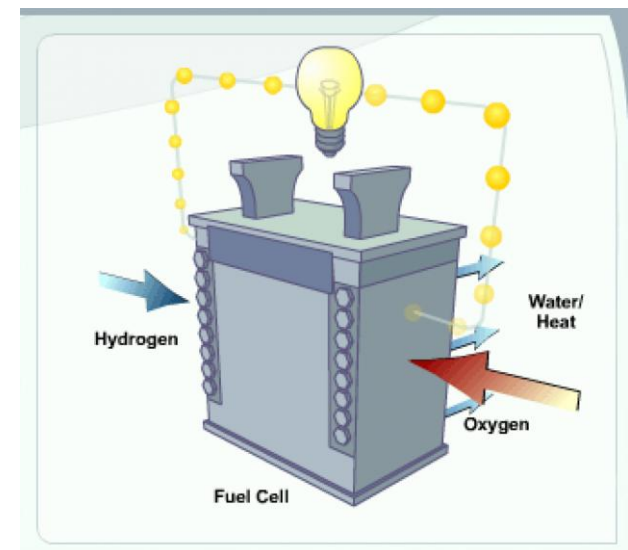


Burning a piece of paper

Metabolizing an apple for energy

Hydrogen fuel cell:

Oxygen and hydrogen **combining** to form water



Participation Questions

1. Evaporation of alcohol
2. Burning of coal
3. Dissolving sugar into coffee
4. Breaking a beaker made of glass
5. Rusting of iron

Click:

https://docs.google.com/forms/d/e/1FAIpQLSdPmLt8PjgS_c0go58ERDoGNFJmagncpBdKpOw0qzXGvKMhzhg/viewform?usp=header



Law of Conservation of Mass

- **There is no new matter.**
- Matter is neither created nor destroyed in a **chemical reaction**. In **chemical reactions**, the changes in mass are so small that they can be ignored.
- In a **nuclear reaction**, significant changes in mass can occur.
- During physical and chemical changes, the total amount of matter remains constant.

Nuclear Reactions	Chemical Reactions
Proceed by redistribution of nuclear particles (nucleons)	Proceed by redistribution of extra-nuclear particles (electrons)
One element may be converted into another	No new element can be produced
Often accompanied by release or absorption of enormous amount of energy	Accompanied by release or absorption of relatively small amount of energy
Rate of reaction is unaffected by external factors such as concentration, temperature, pressure and catalyst	Rate of reaction influenced by external factors

Law of Conservation of Mass

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Why Are Nuclear Reactions So Much More Explosive?

- The energy stored in a nuclear bond is on the order of 1 MeV. A typical chemical bond, on the other hand, stores energy on the order of 1eV, approximately one million times less.

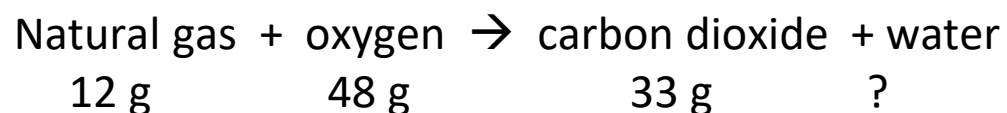
[Is nuclear energy good or bad?](#)

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Example: Conservation of Mass



Example: 12 g of natural gas combines with 48 g of oxygen in a flame. The chemical reaction produces 33 g of carbon dioxide and how many grams of water?



Early Ideas About Atoms

Around 400 BCE: Ancient Greek philosophers Leucippus and Democritus:
All matter is made of irreducible “building blocks” that couldn’t be divided further.

“a-tomos” = “in-divisible”

Aristotle believed matter was composed of four “elements:”



Atomic Theory

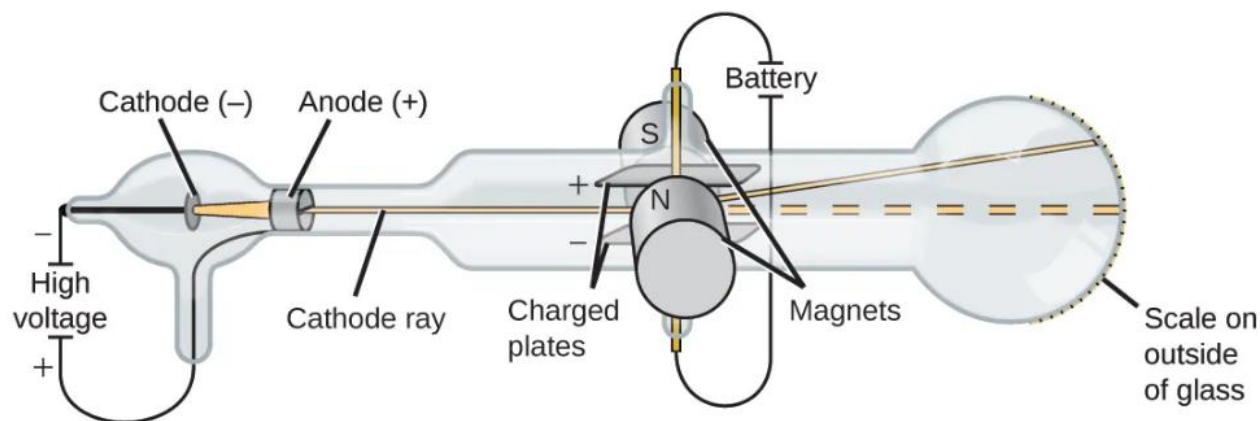
1808 CE: John Dalton's Atomic Theory expanded on the centuries-old Greek philosophies

1. **All matter** is composed of **atoms**.
2. For a given **element**, all **atoms** have the **same mass and other properties**.
3. **Different elements** have atoms with **different properties**.
4. A **compound** consists of atoms of **two or more elements** combined in **whole-number ratios**.
Example: Water H_2O has 2 atoms of hydrogen for every 1 atom of water
5. **Atoms are not created or destroyed** in a chemical change. They rearrange to make new compounds (with new ratios of atoms).

Revising Dalton's Atomic Theory

Dalton was wrong about indivisibility, but right that elements are unique types of matter.

1897 CE: J.J. Thompson built on (and challenged) Dalton's atomic theory



Electricity applied in a vacuum generates a beam of negatively charged particles.

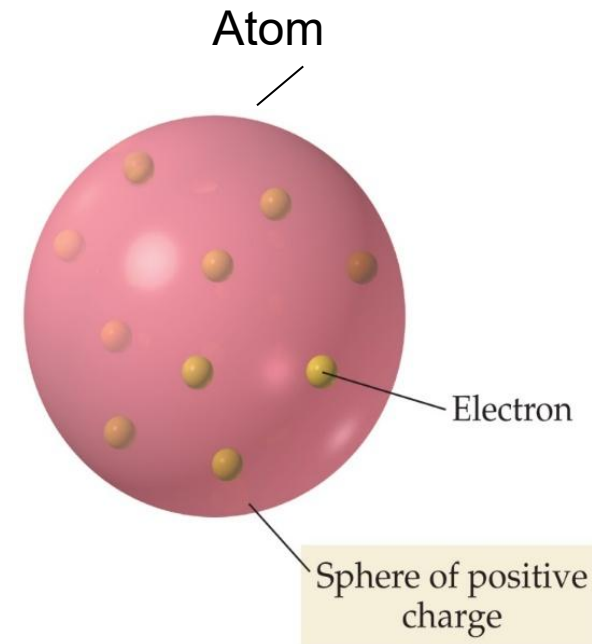
Experimentation allowed calculation of the mass-to-charge ratio of these particles.

They were negatively charged, and much lighter than any atoms that had been observed!

Atomic theory was revised to account for new observations.

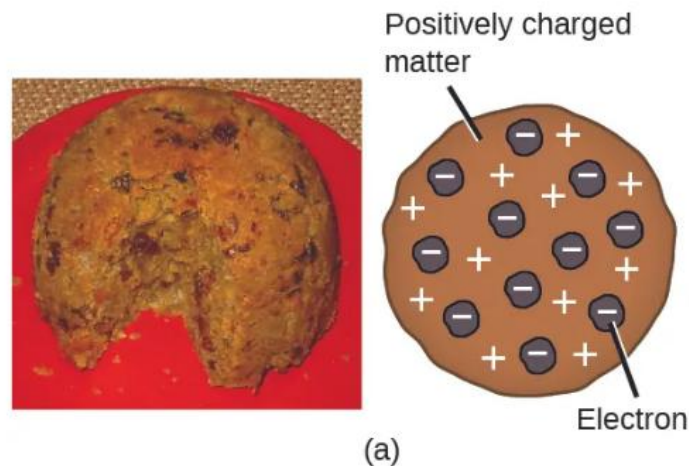
J.J. Thompson's Plum Pudding Model of the Atom

Atoms were known to be neutral so there must be positive charges to balance these negatively charged “electrons”



1904 CE: “Plum Pudding Model”

The atom is a sphere of positive charge with negative particles (electrons) mixed in

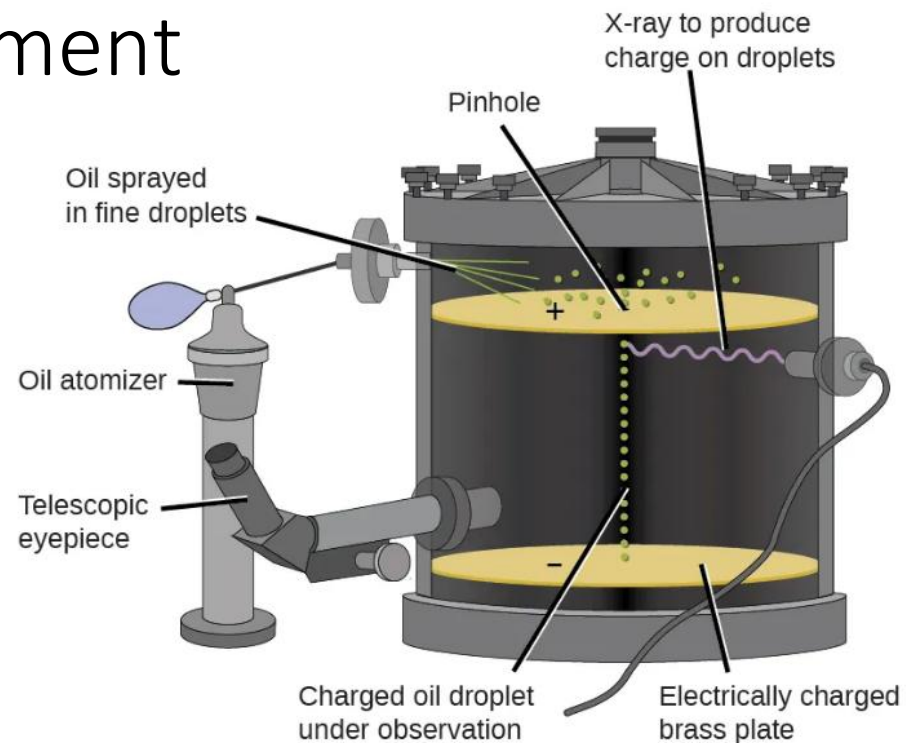


Millikan's Oil Drop Experiment

1909 CE: Further investigation into negatively charged electrons.

Experiments produced tiny oil droplets with slight negative charges

Observed that charge on droplets was always a multiple of the same number!



Oil drop	Charge in coulombs (C)
A	$4.8 \times 10^{-19} \text{ C}$
B	$3.2 \times 10^{-19} \text{ C}$
C	$6.4 \times 10^{-19} \text{ C}$
D	$1.6 \times 10^{-19} \text{ C}$
E	$4.8 \times 10^{-19} \text{ C}$

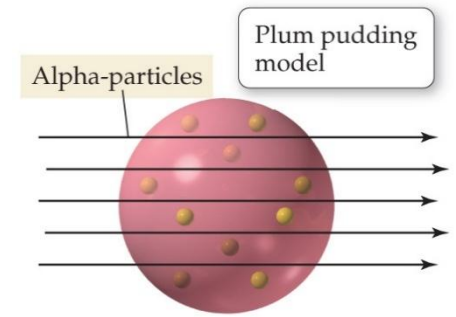
Determined the **smallest unit of negative charge must be $1.6 \times 10^{-19} \text{ C}$**

Combined with Thompson's data, allowed **calculation of mass of one electron**

Rutherford's Gold Foil Experiment

ca. 1909 CE: Ernest Rutherford and co. tried to confirm Thompson's Plum Pudding Model

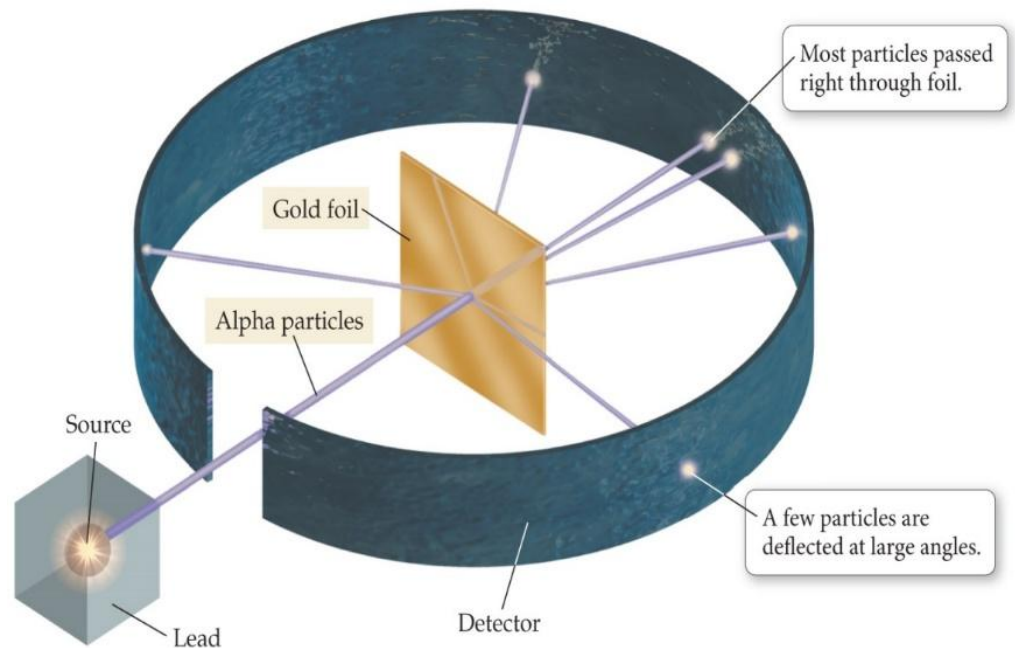
Tiny **alpha-particles** were directed at a thin sheet of gold foil.



(a) Rutherford's expected result

Most of the particles passed directly through the foil.

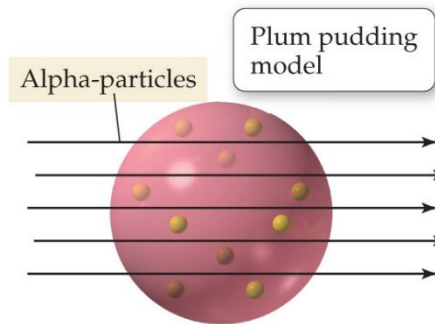
A few were deflected, some of them at sharp angles.



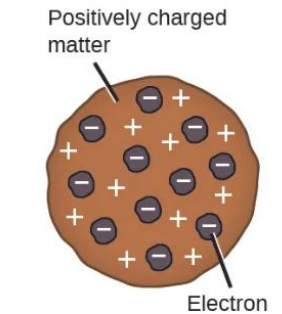
Nuclear Theory of the Atom (1911 CE)

a) Rutherford's Expected result:
Alpha-particles should pass through gold foil with minimal deflection.

Deflections should be roughly uniform.



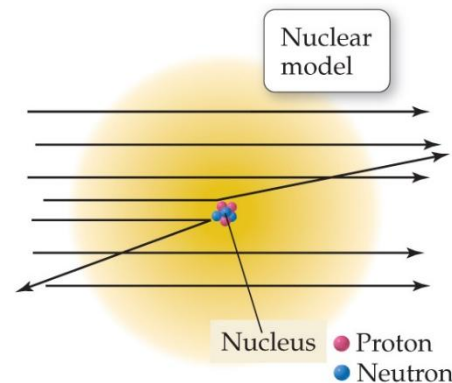
(a) Rutherford's expected result



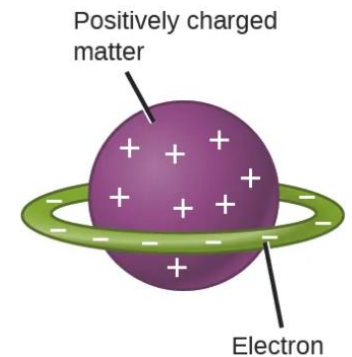
"Plum Pudding" model

(b) Actual result:
A small number of alpha-particles were deflected or bounced back.

Supported Hantaro Nagaoka's
"planetary" model of the atom
(1903)



(b) Rutherford's actual result



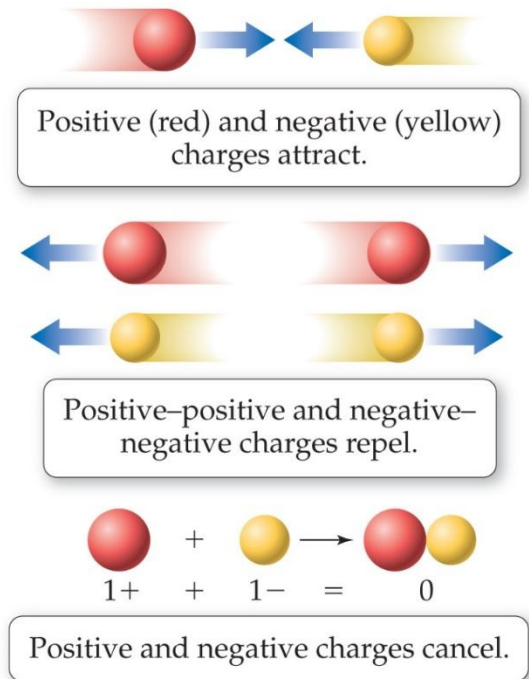
"planetary" model

Nuclear Theory of the Atom (1911 CE)

1. Most of **atom's mass and positive** charge is contained in a small central core called the **nucleus**
2. Atom is **mostly empty space**, with small negatively charged electrons interspersed.
3. Amount of **positive charge = amount of negative charge** (atom is neutral)

Experiments showed that nuclei of **all elements** contained Hydrogen nuclei as “building blocks”

Rutherford named this building block of nuclear mass and positive charge the “**proton**.”



James Chadwick Discovering the Neutron

Experiments to this point had determined:

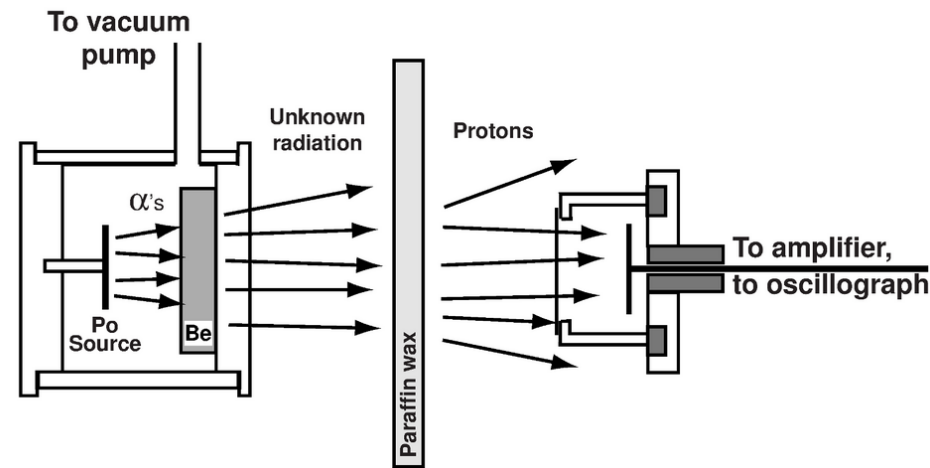
Atoms contain **negatively (-) charged e^-** , which have very little mass.

Most of an **atom's mass is in the nucleus**

The nucleus contains **positive (+) charged protons**

An atom of Helium is 4x as heavy as a proton, but the nucleus only has 2x as much (+) charge.

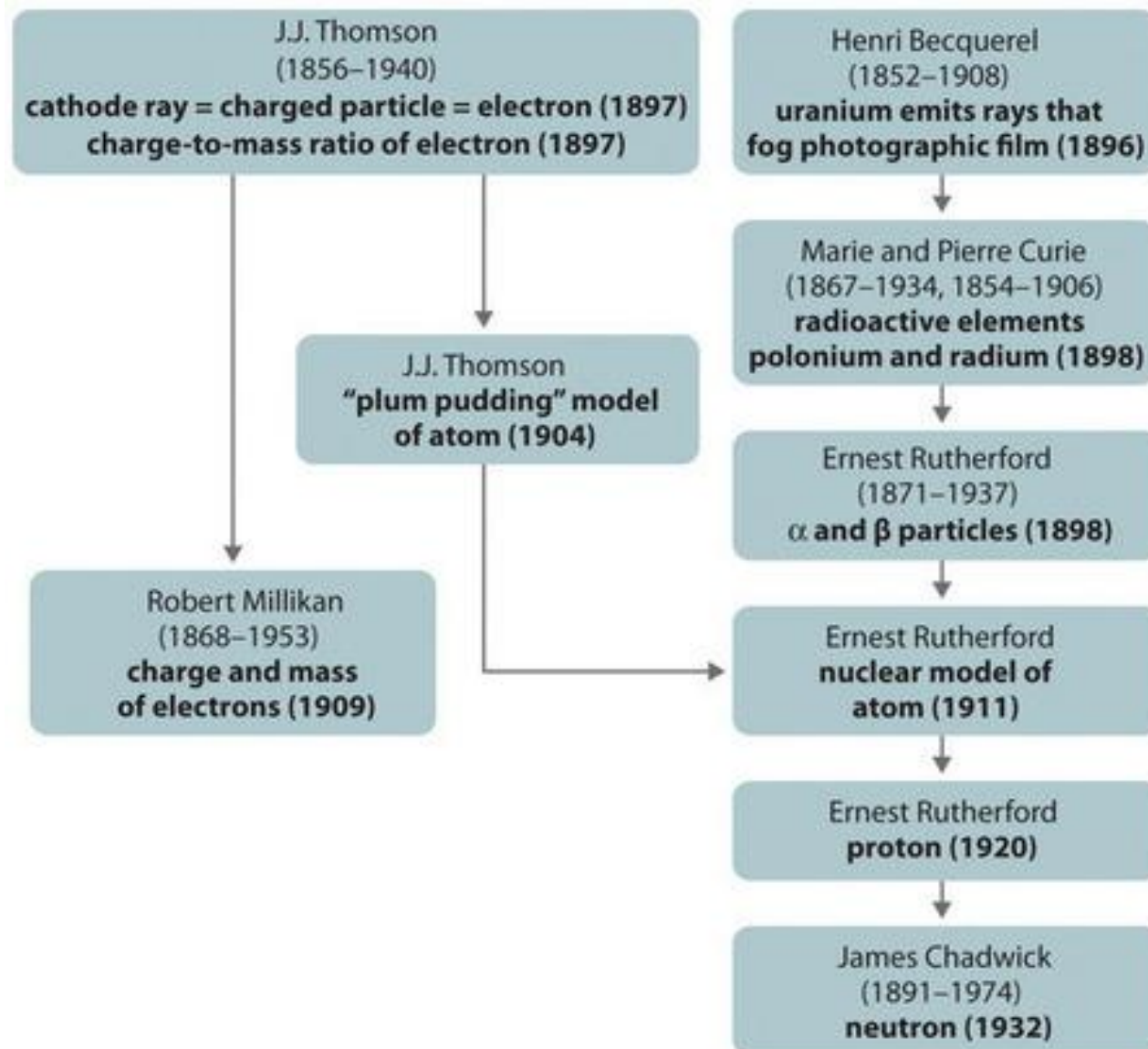
Data suggested there was a neutral particle contributing to the nucleus mass!



1932 CE: James Chadwick discovered a new type of radiation.

Particles had similar mass to a Hydrogen nucleus, but no charge.

A Summary of the Historical Development of Models of the Components and Structure of the Atom



Protons, Neutrons, & Electrons (Subatomic Particles)

Protons (+ charge) and **neutrons (neutral)** make up **nucleus** (99.9% of atom's mass)

Electrons (- charge): cloud of negatively charged particles around nucleus; very little mass

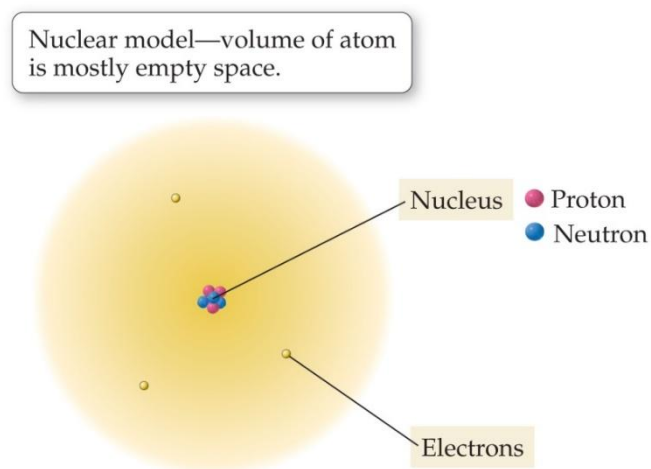
Atomic mass unit (amu): convenient unit for the mass of subatomic particles

$$1 \text{ amu} = 1.66 \times 10^{-27} \text{ kg}$$

Based on 1/12 mass of a ^{12}C atom.

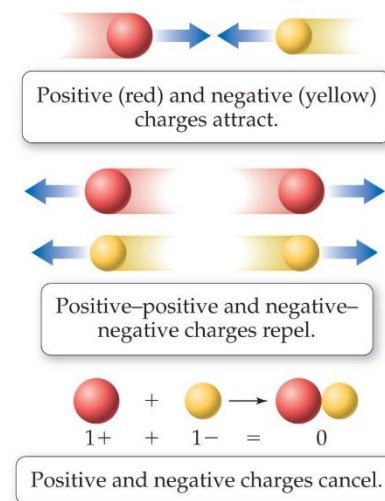
amu is also called _____, especially

by _____.



Subatomic Particles

	Mass (kg)	Mass (amu)	Charge
Proton	1.67273×10^{-27}	1.0073	+1
Neutron	1.67493×10^{-27}	1.0087	0
Electron	9.1×10^{-31}	0.00055	-1

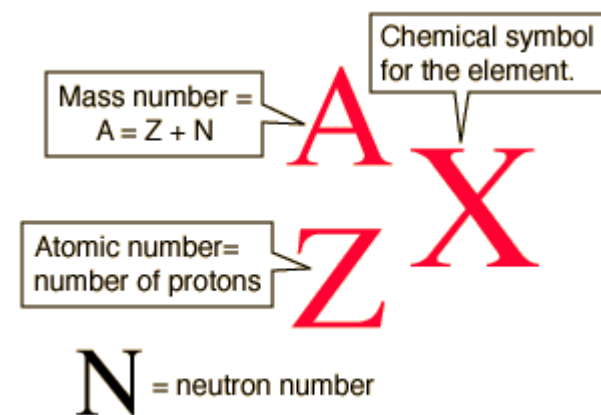


If a **proton** had the mass of a **baseball**, an **electron** would have the mass of a **rice grain**.

A **proton** is nearly **2000 times more massive** than an **electron**.



Atomic Numbers & Symbols



- 118 known elements each have a 1 or 2 letter symbol (first letter always capitalized).

- All atoms of the **same element** have **the same number of protons**.

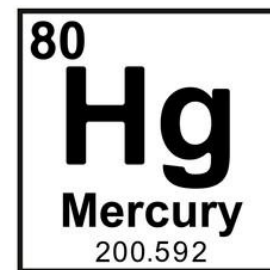
In other words, Elements are defined by their number of protons



- **Atomic number** (Z) = number of **protons**
 - Often omitted – element symbol implies # of protons
- **Mass number** (A) = number of **protons + neutrons**



- **Chemical symbol** = abbreviation for an element.
Example: Carbon (C), Silicon (Si), Lead (Pb), Mercury (Hg)



Ancient Greek word for mercury:
ὕδραργυρος
(*hydrargyros*)
meaning
water-silver

Some symbols are based on Greek and Latin names

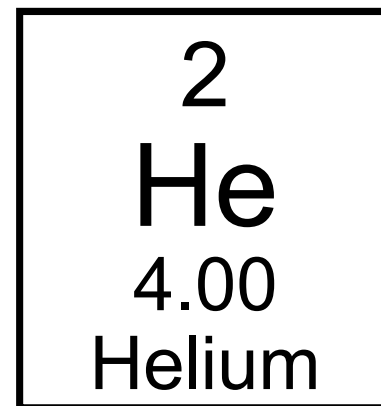
Atomic Numbers & Symbols

2 = Atomic Number (Z)

He = Chemical Symbol

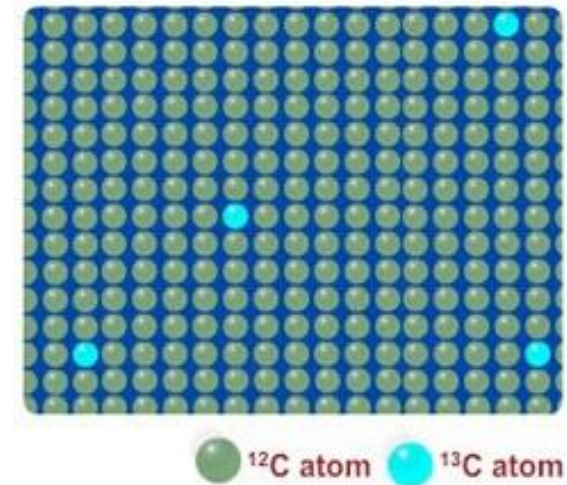
4.00 = Atomic Mass (in amu)

Helium = Name



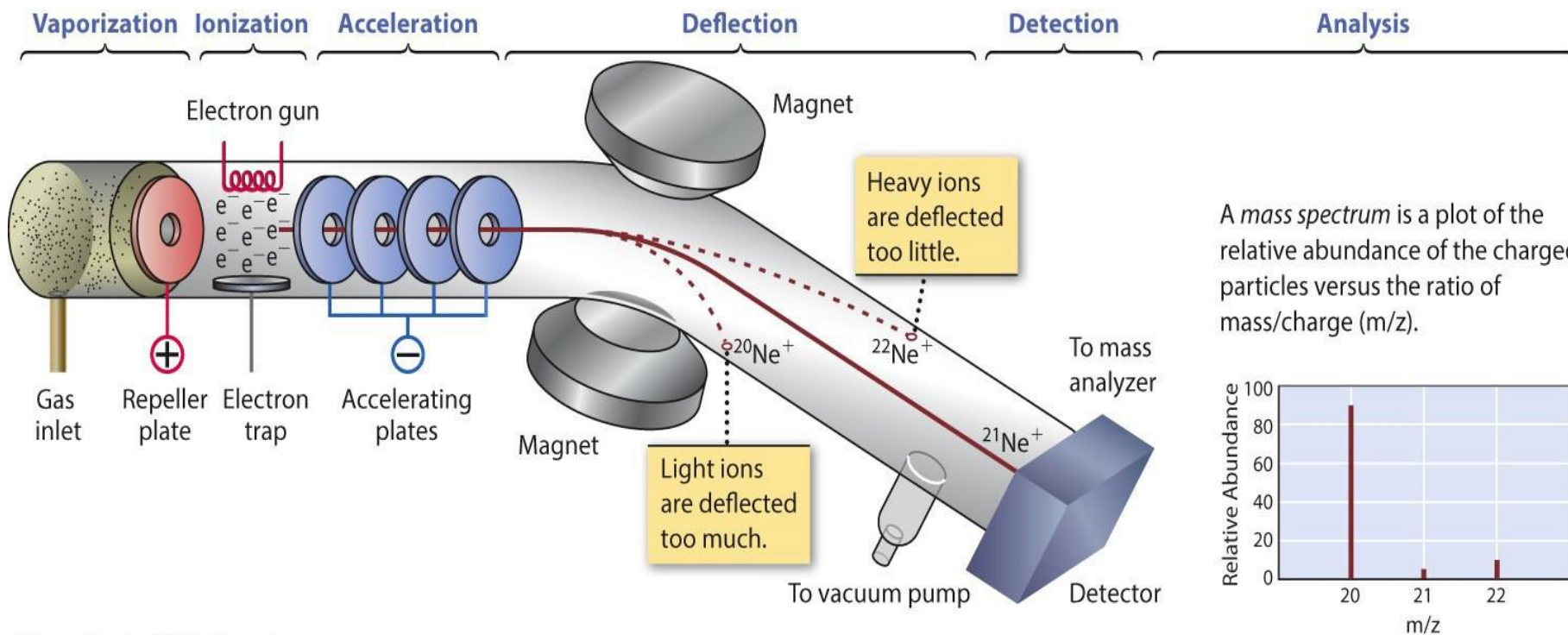
Isotopes

- Atoms of the same element that **differ by the number of neutrons** are called **isotopes**
- For each naturally occurring element, we can refer to its natural isotope distribution on Earth
 - Carbon (C): all C atoms have 6 protons
 - 98.9% of C atoms on Earth have 6 neutrons; 1.1% have 7 neutrons
- Two common ways to indicate an isotope – e.g., for a carbon atom with 6 neutrons:
 - Carbon-12
 - ^{12}C

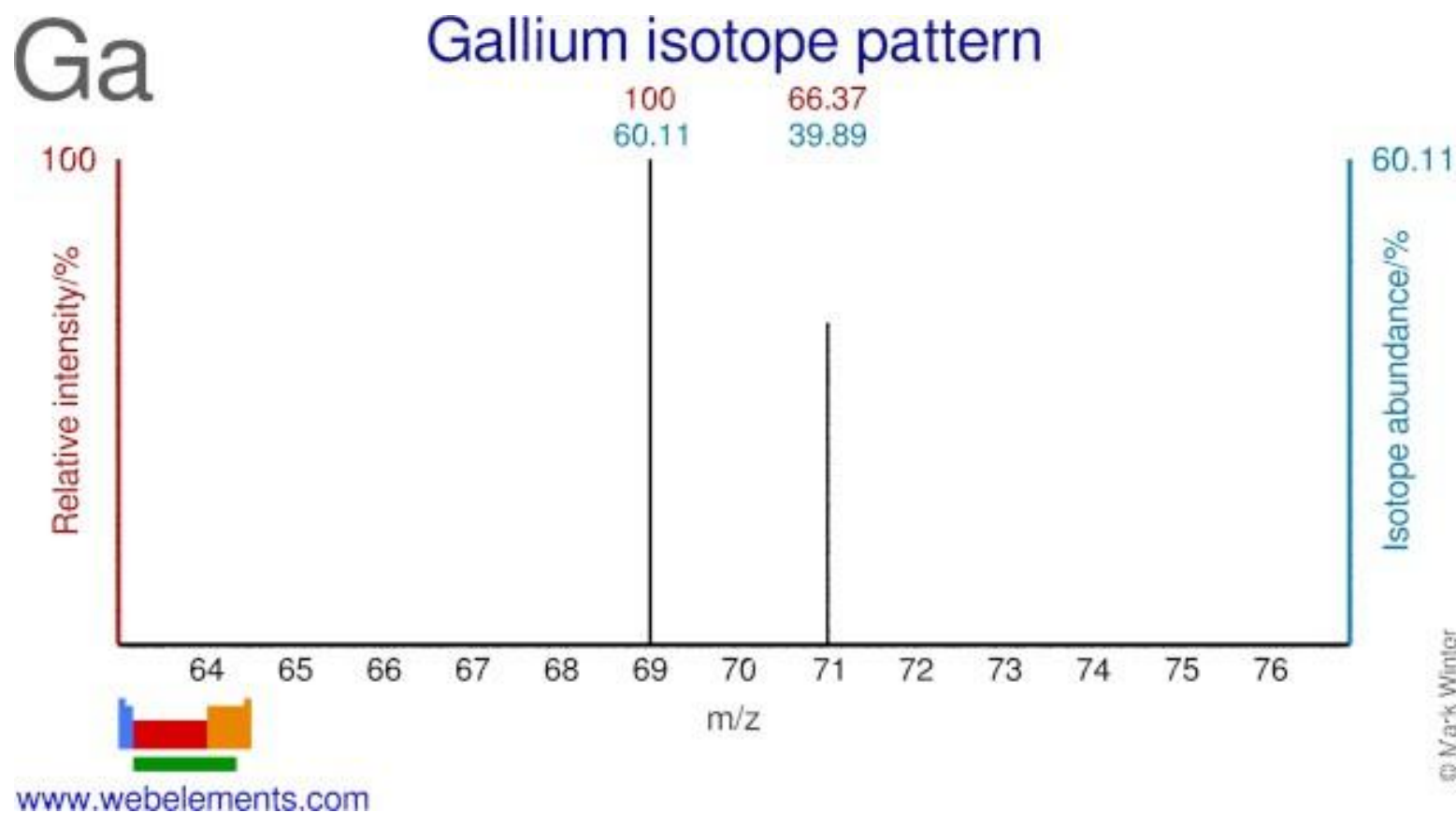


Isotopes and mass spectrometry

The masses of isotopes and their abundances are determined experimentally by mass spectrometry



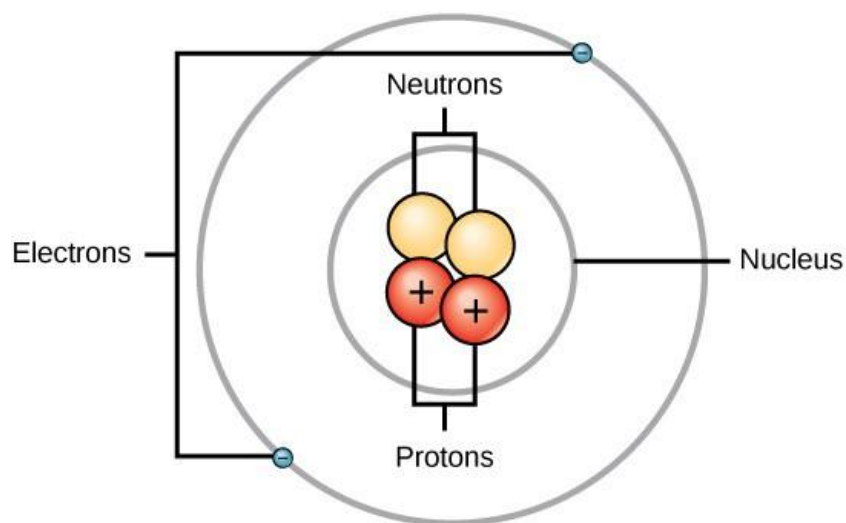
Example element mass spectrum: gallium (Ga)



Atomic mass unit and atomic weight

- Atomic mass unit (amu) = mass of 1/12 of a carbon-12 atom
 - 1 amu = 1.66054×10^{-24} g
 - Not an exact constant, but known to high precision
- Atomic weight = average mass of all naturally occurring isotopes of an element
 - Weighted average
$$\text{average mass} = \sum_i (\text{fractional abundance} \times \text{isotopic mass})_i$$
 - Example: hydrogen (H) is 99.9885% hydrogen-1 (1.007825 amu) and 0.0115% hydrogen-2 (2.014102 amu). What is the average mass?

Atomic structure: summary



Atoms of the same element have the same number of protons

e.g., all helium atoms have 2 protons

Neutral (uncharged) atoms:

of protons = # of electrons

Properties of Subatomic Particles

Name	Location	Charge (C)	Unit Charge	Mass (amu)	Mass (g)
electron	outside nucleus	-1.602×10^{-19}	1-	0.00055	0.00091×10^{-24}
proton	nucleus	1.602×10^{-19}	1+	1.00727	1.67262×10^{-24}
neutron	nucleus	0	0	1.00866	1.67493×10^{-24}

Table 2.2

Recitation Worksheet Instructions

- Must complete all questions to get full points.
- **Closed book, no AI, only your calculator allowed.**
- **Must complete on your own.**
- Pause, Pair and Partner Moment:
 - Last 15 minutes before end of recitation at 1:45 PM you can turn to a neighbor and you have 5 minutes to discuss the worksheet.
 - 1:50 PM regroup on your own.
- Turn in your worksheet at the end of recitation time at 2PM for full credit.