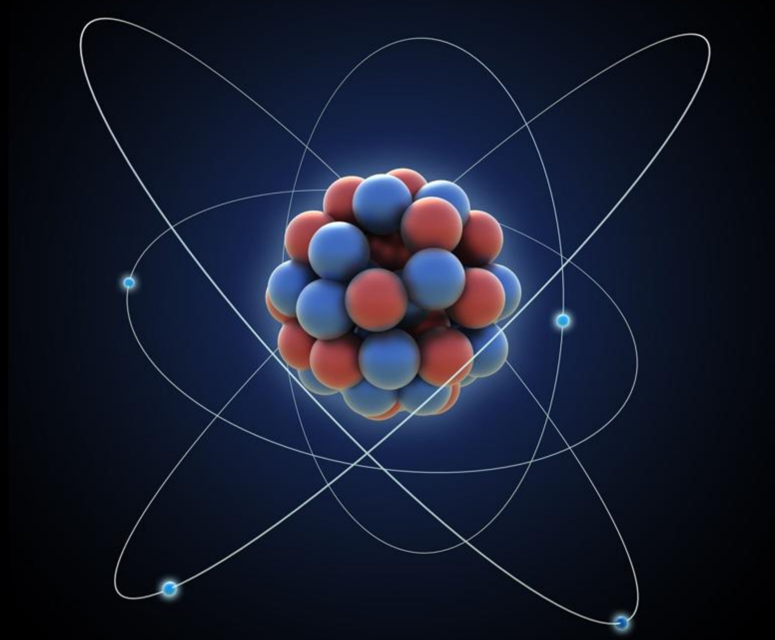


Chapter 2. Atoms, Molecules and Ions

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 conclusitons 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.

Reading from book pages 59-77.

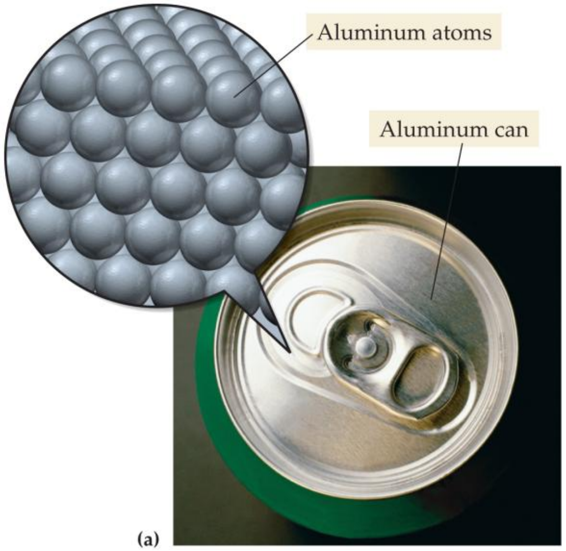
Matter

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.



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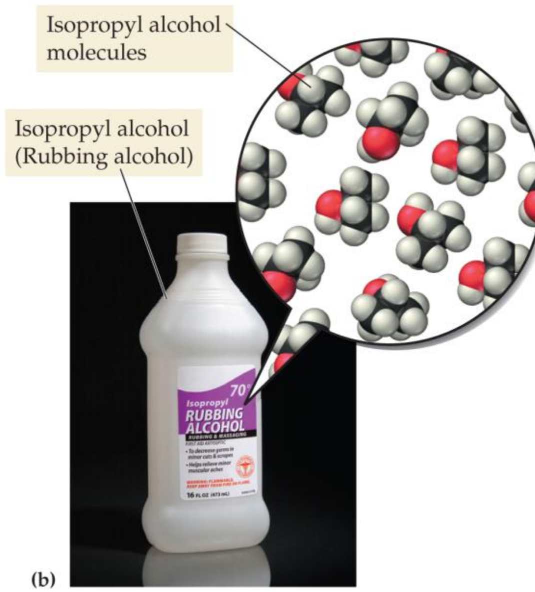
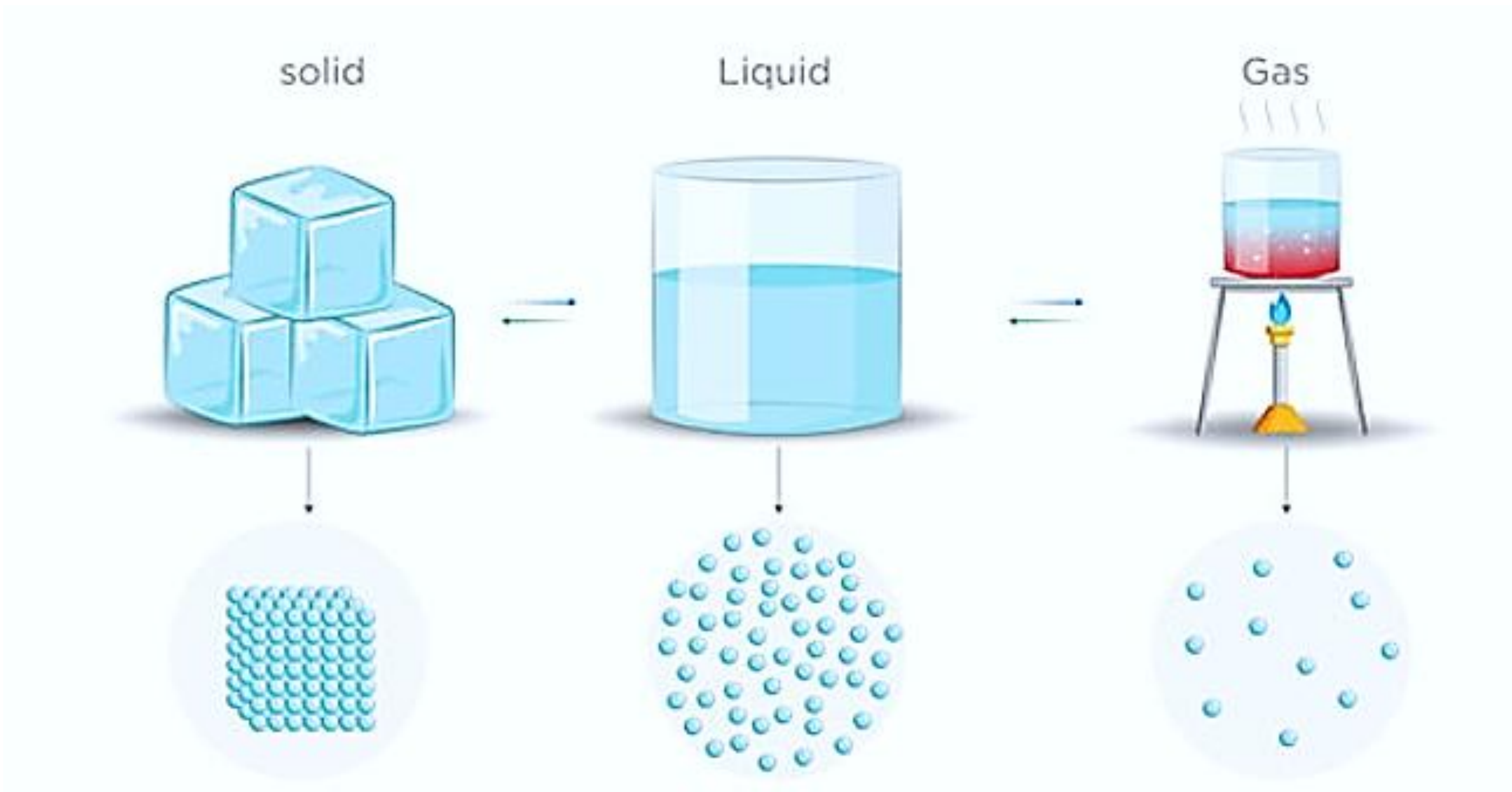


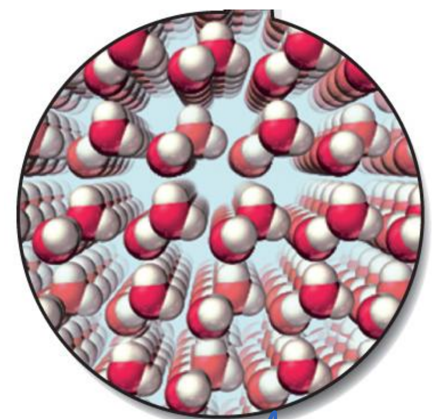
(b)

Three States of Matter



States of Matter: Solid

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



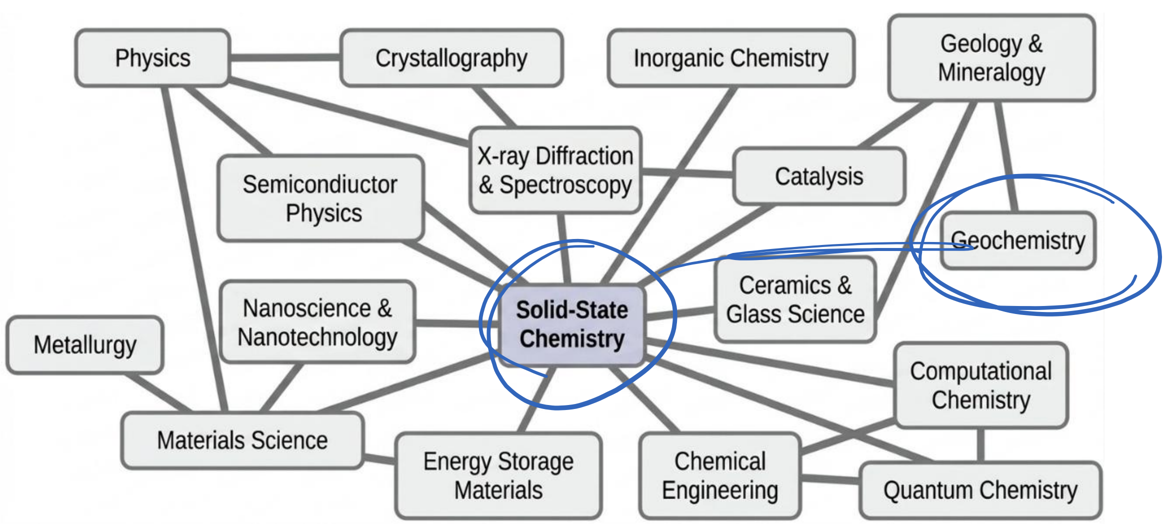
crystal

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

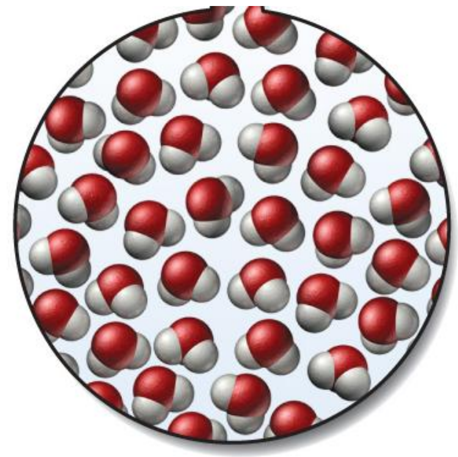
minerals



States of Matter: Liquid

standing in a crowded room

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

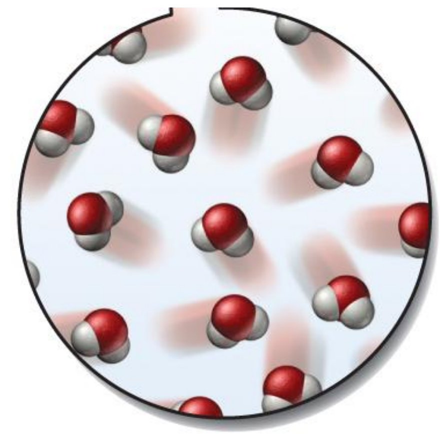
Atoms or molecules are separated by large distances and are free to move relative to one another.

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

non fixed

Gases always assume the shape and volume of their container.

Ex: Oxygen, helium, carbon dioxide



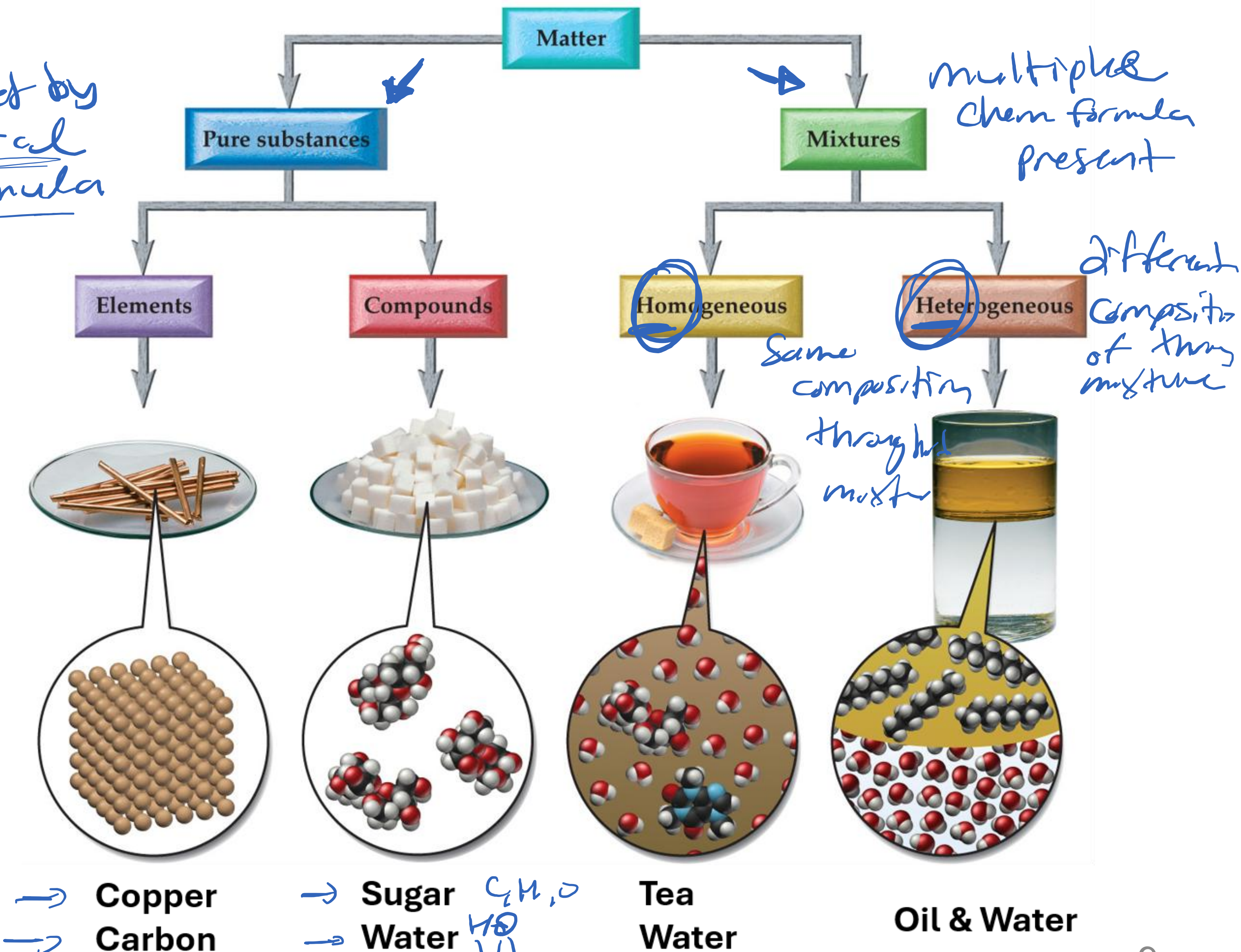
Properties of Solids, Liquids, and Gases

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

represented by
a chemical
formula

But what is everything fundamentally made of?



$C_6H_{12}O_6$
H_2O

Properties of Matter

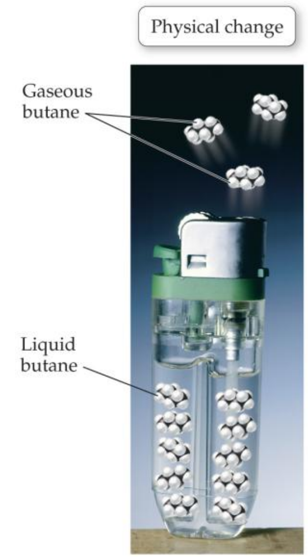
- Physical change

  chemical formula unchanged

  Matter changes appearance
  Composition stays the same

  state of matter change

  Can (often) be reversed
  Phase change



$\boxed{H_2O}$

solid / liquid

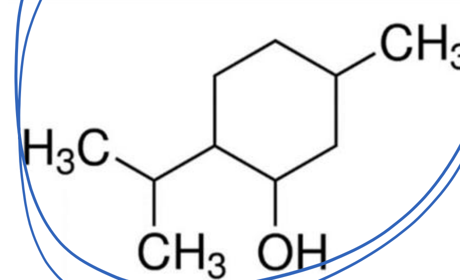
melting ice (solid water) to form liquid water

liquid / gas

boiling liquid water to form steam (gaseous water)

tearing a piece of paper

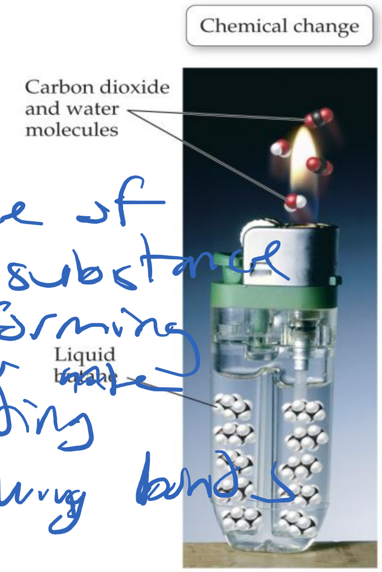
Menthol



Properties of Matter

chemical formula changes

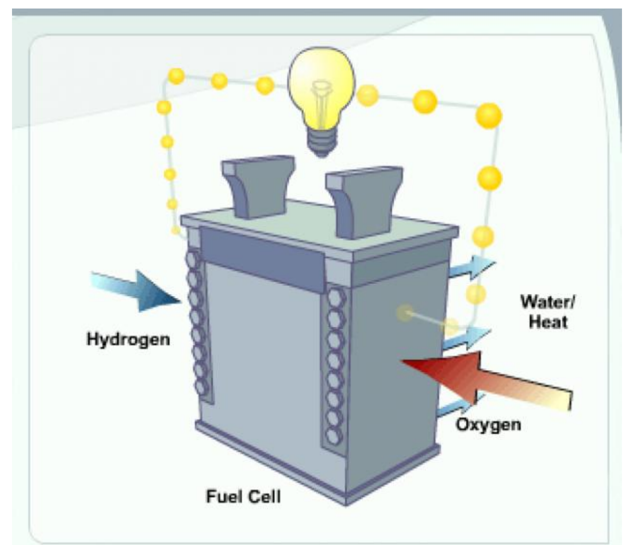
evidence of new substance forming new bonding & breaking bonds



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 P
2. Burning of coal C
3. Dissolving sugar into coffee P
4. Breaking a beaker made of glass P
5. Rusting of iron C

Click:
https://docs.google.com/forms/d/e/1FAIpQLSdPmLt8PjgS_c0go58ERDoGNFJmaqncpBdKpOw0qzXGvKMhzg/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.

Differences

CHEM 303

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 nucelar energy good or bad?

https://docs.google.com/forms/d/e/1FAIpQLSdbEW0bhEBhPfqYwx_0ghuEG_BcWdJ1G1t_EzQc7SMA_SOmFA/viewform?usp=header

What are the benefits of nuclear technology?

- Nuclear reactions are useful for generating vast amounts of clean, reliable electricity, powering industries, producing medical isotopes for cancer treatment and diagnostics (like PET scans), sterilizing equipment, advancing scientific research, and even enabling space exploration.

Example: Conservation of Mass

Butane + Oxygen → Carbon Dioxide + Water

$\underbrace{58\text{ g} + 208\text{ g}}_{266\text{ g}}$ $\underbrace{176\text{ g} \qquad\qquad 90\text{ g}}_{266\text{ g}}$

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?

Natural gas + oxygen → carbon dioxide + water

$\underbrace{12\text{ g} \qquad 48\text{ g}}_{60\text{ g}}$ $33\text{ g} + ?$ $x = 60\text{g} \Rightarrow \underline{x = 27\text{g}}$

Early Ideas About Atoms

<u>Around 400 BCE</u>: 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

<u>1808 CE</u>: 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).

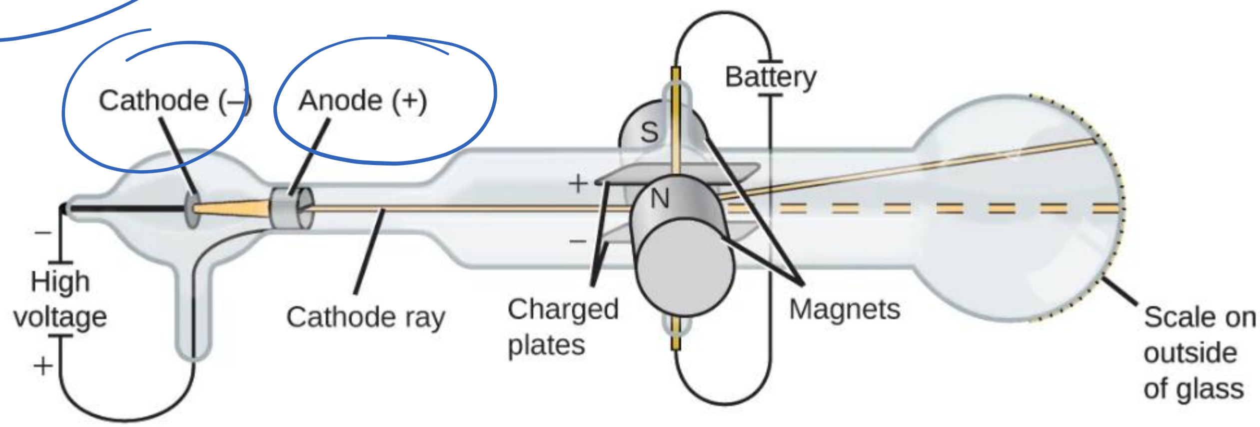
Revising Dalton's Atomic Theory

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

<u>1897 CE:</u> J.J. Thompson built on (and challenged) Dalton's atomic theory ⇒ electron

subatomic particles

electric ion



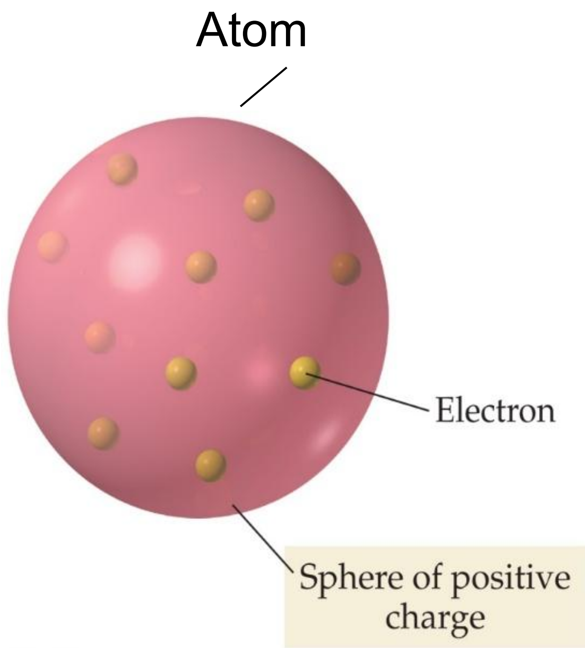
Electricity applied in a vacuum generates a beam of negatively charged particles.

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!

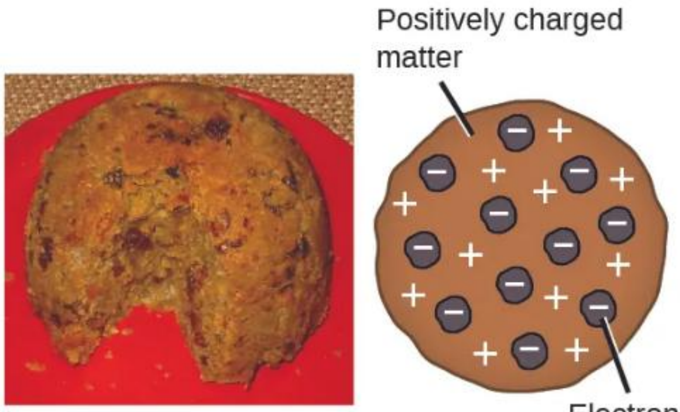
Atomic theory was revised to account for new observations.

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"

<u>1904 CE</u>: "Plum Pudding Model"
The atom is a sphere of positive charge with negative particles (electrons) mixed in



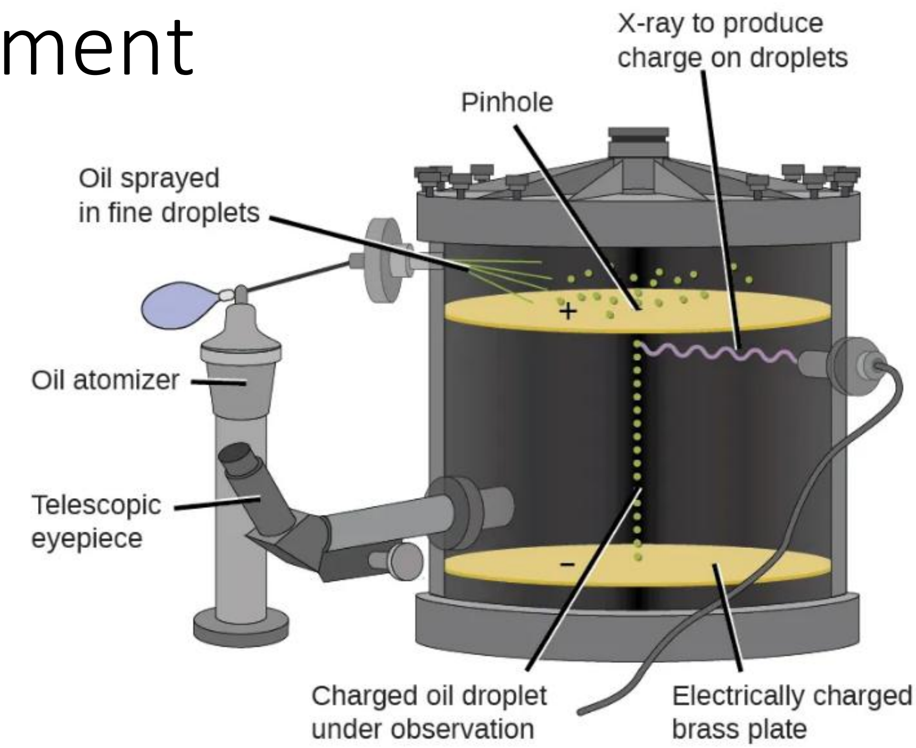
(a)

Millikan's Oil Drop Experiment

<u>1909 CE</u>: 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!



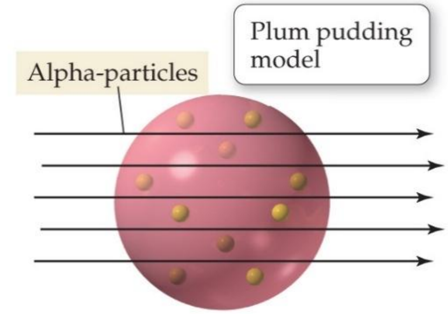
Oil drop	Charge in coulombs (C)
A	4.8×10^{-19} C
B	3.2×10^{-19} C
C	6.4×10^{-19} C
D	1.6×10^{-19} C
E	4.8×10^{-19} C

Determined the smallest unit of negative charge must be 1.6×10^{-19} C

Combined with Thompson's data, allowed calculation of mass of one electron atom ~~is~~ was a tiny dense positive nucleus

Rutherford's Gold Foil Experiment

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



(a) Rutherford's expected result

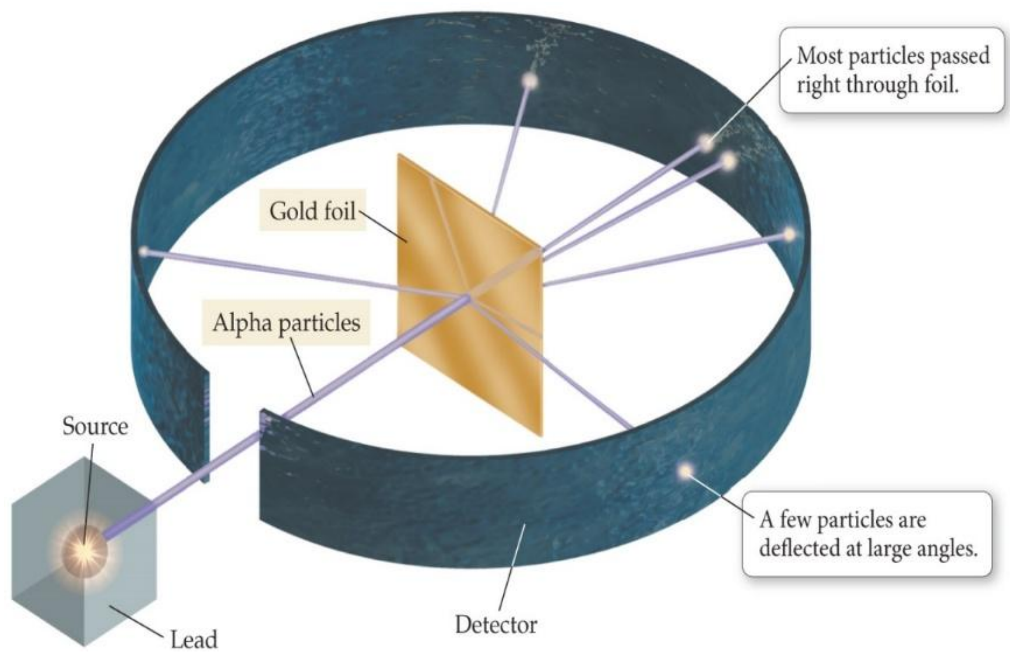
Tiny alpha-particles were directed at a thin sheet of gold foil.

high charge positively

Most of the particles passed directly through the foil.

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

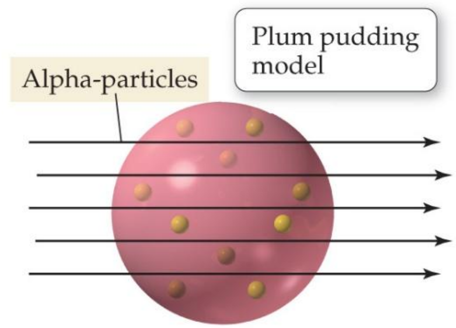
empty space



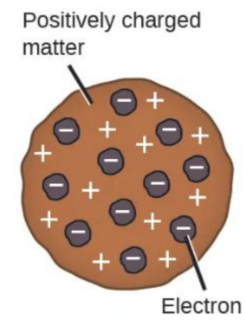
Nuclear Theory of the Atom (<u>1911 CE</u>)

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

Deflections should be roughly uniform.



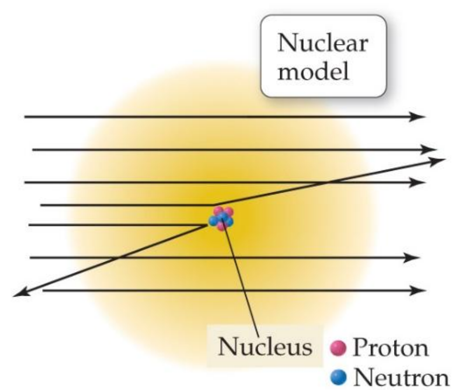
(a) Rutherford’s expected result



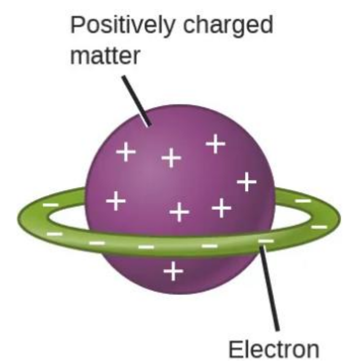
“Plum Pudding” model

(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



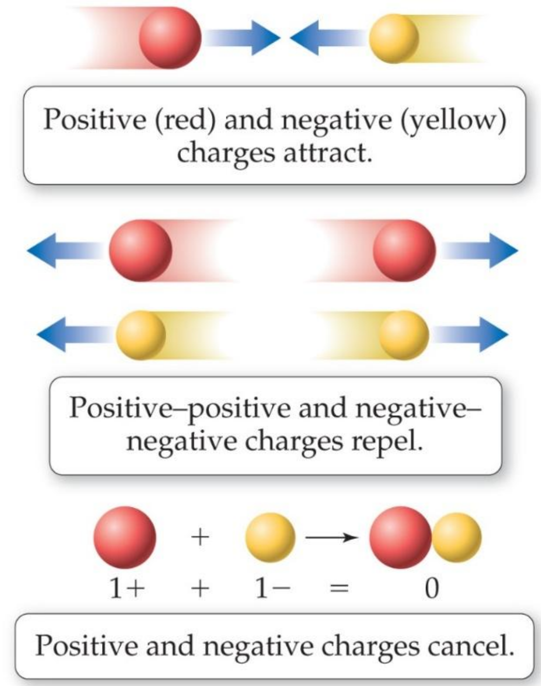
“planetary” model

Nuclear Theory of the Atom (<u>1911 CE</u>)

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 **<u>all elements</u>** contained Hydrogen nuclei as "building blocks"

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

James Chadwick Discovering the Neutron

uncharged subatomic particles

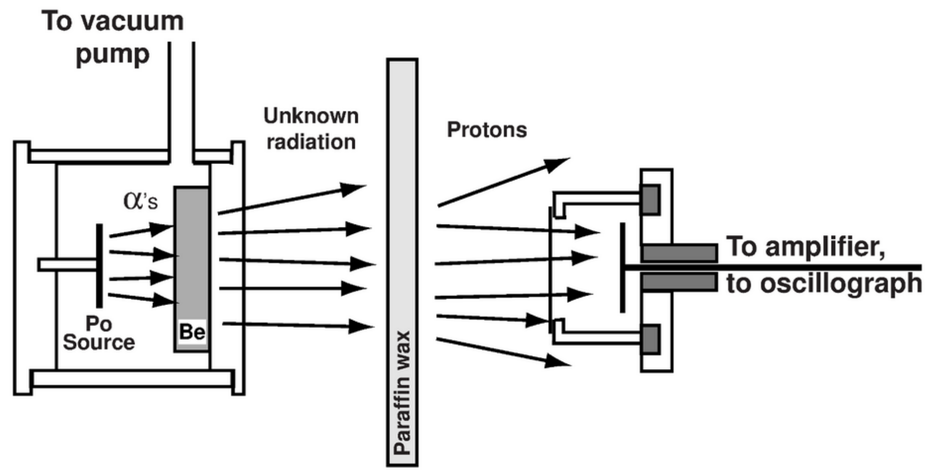
Experiments to this point had determined:

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.

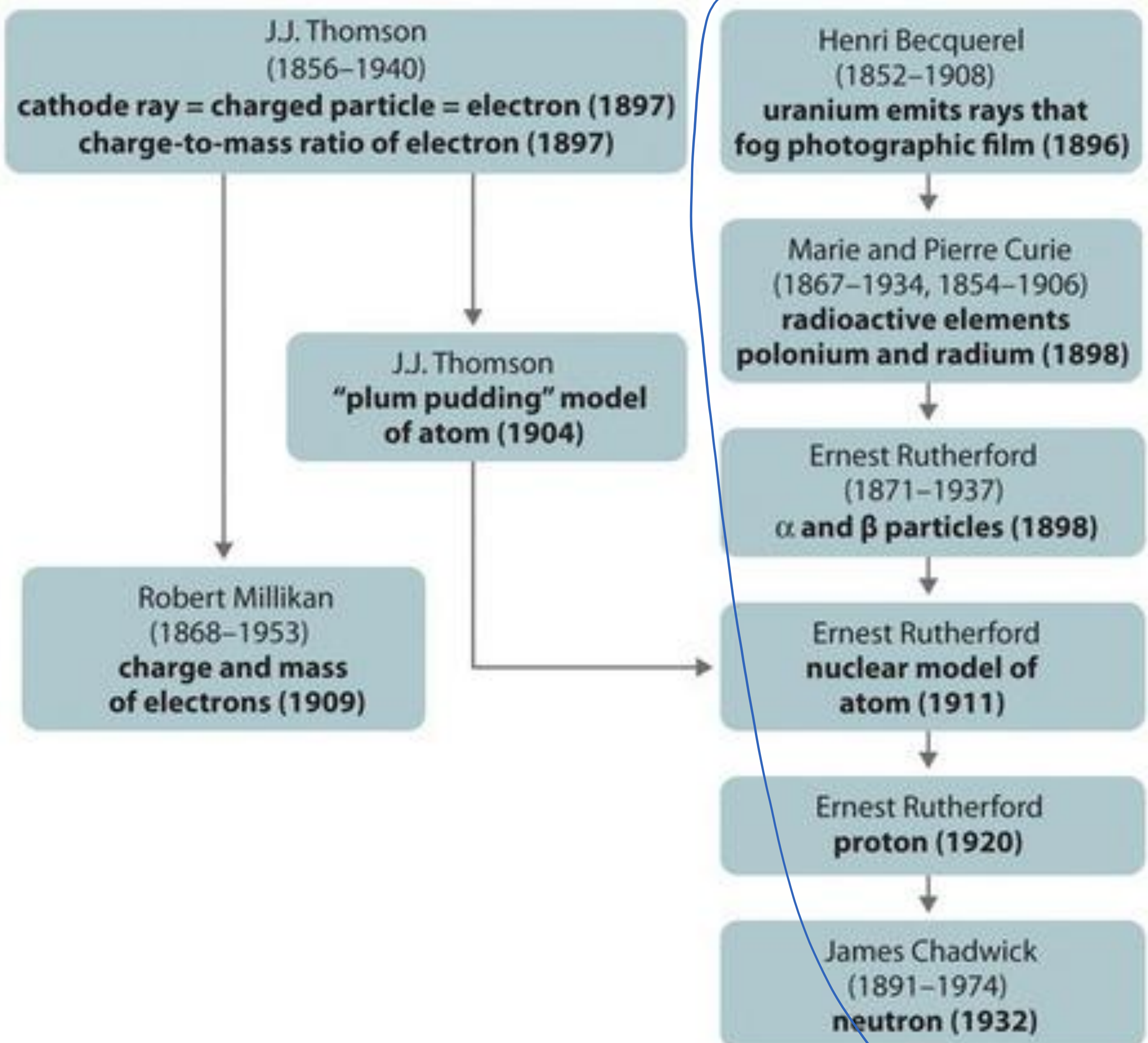


1932 CE: James Chadwick discovered a new type of radiation.

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

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

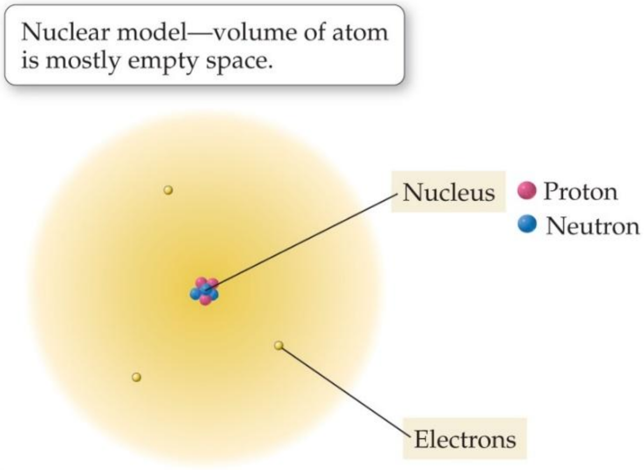
A Summary of the Historical Development of Models of the Components and Structure of the Atom



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

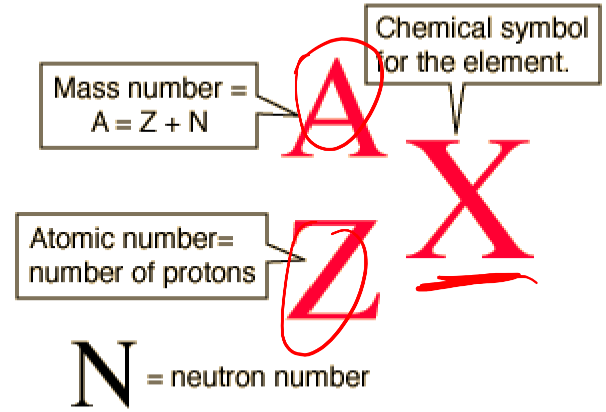
$$1 \text{ amu} = 1.66 \times 10^{-27} \text{ kg}$$

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

amu is also called _______________ , especially

by __________________.

Atomic Numbers & Symbols



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

Z

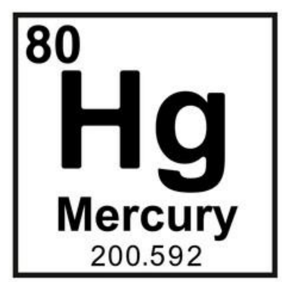
- All atoms of the same element have the same number of protons.

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: <u>ὑδράργυρος</u> *(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

2
He
4.00
Helium

Carbon Isotope:

6 prot, 6 neutrons → C-12
Z=6, A=12

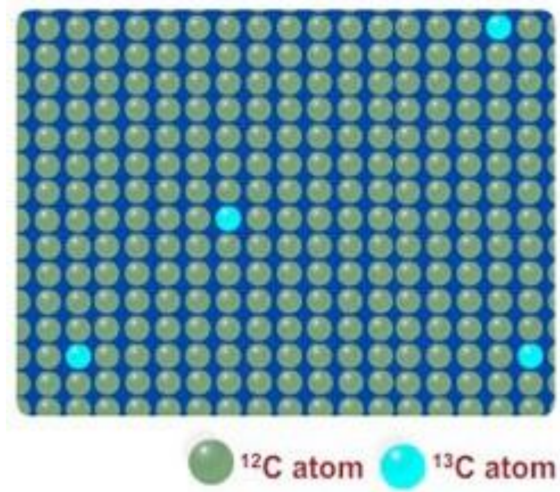
6 prot, 7 neutrons → C-13
Z=6, A=13

6 prot, 8 neutrons → C-14
Z=6, A=14

Isotopes →

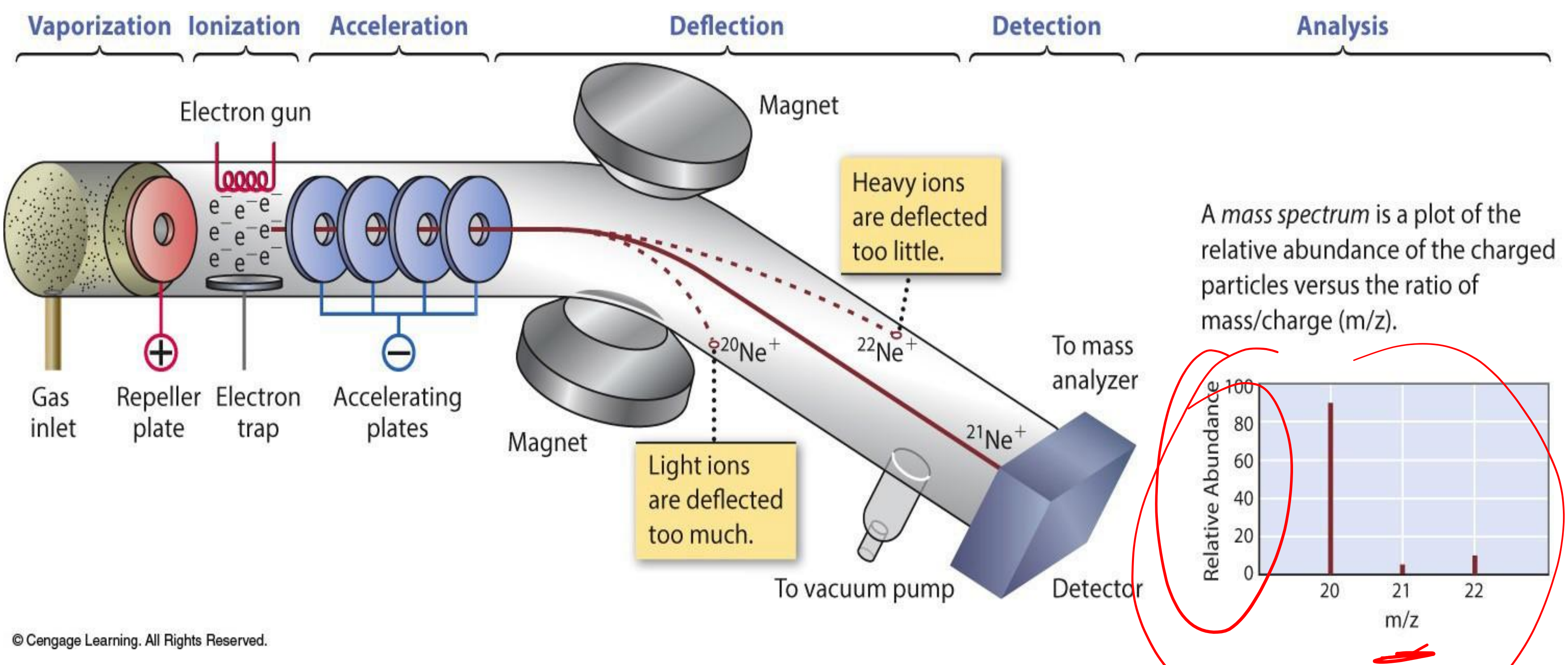
various # of neutrons, same # protons

- 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



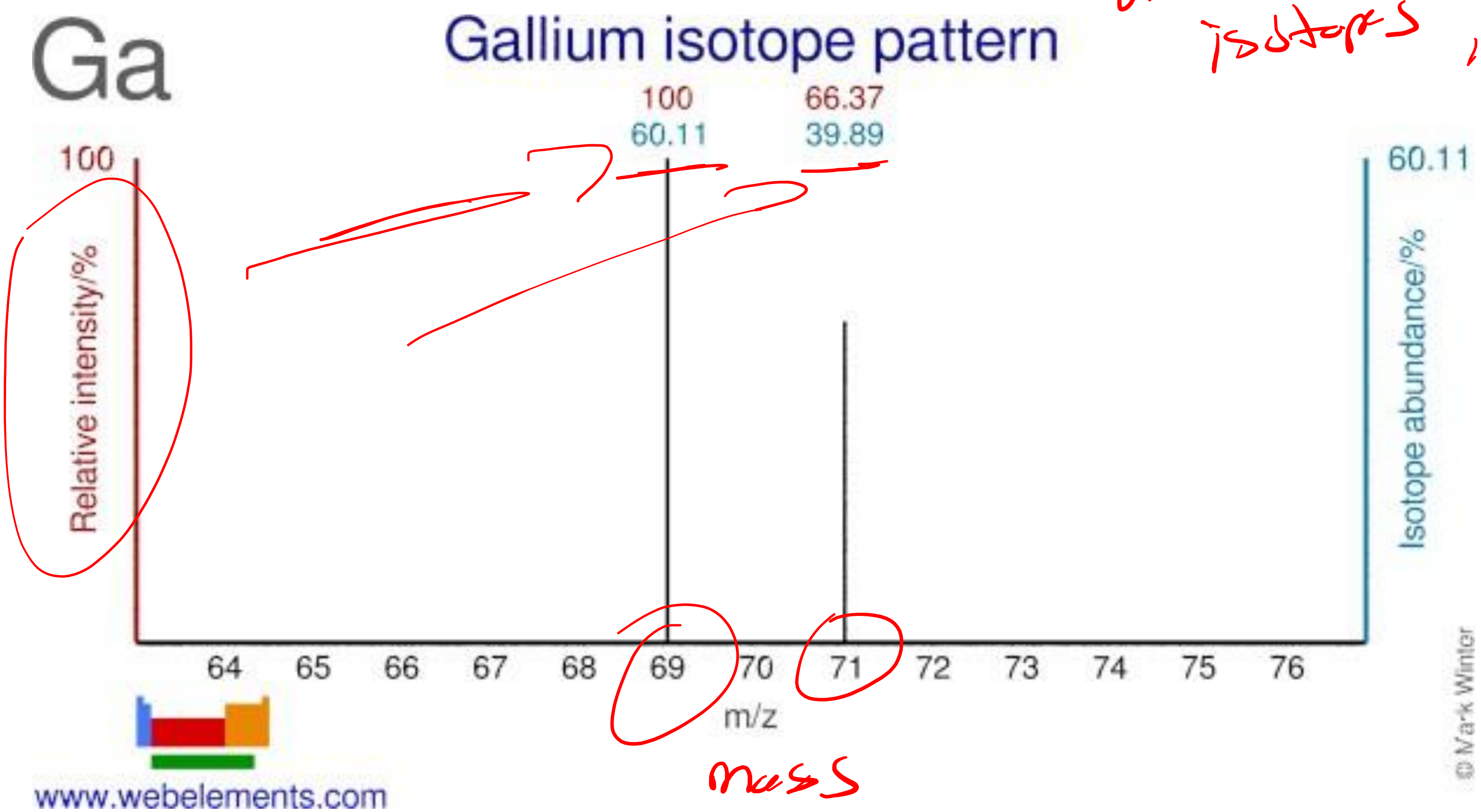
Isotopes and mass spectrometry

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



Example element mass spectrum: gallium (Ga)

different neutrons in isotopes, difference weight



mass

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?

= 0.0115/100 = 0.000115

average mass = H1 + H2 ⇒

= (0.999885 × 1.007825 amu) + (0.000115 × 2.014102 amu)

6SF 7SF 3SF 7SF

5DP 6DP

= 1.0077091 amu + 0.000231620 amu

= 1.00794072

1.00794 amu

Round answer to fewest decimal places found in original number

Atomic Numbers & Symbols

2 = Atomic Number (Z)

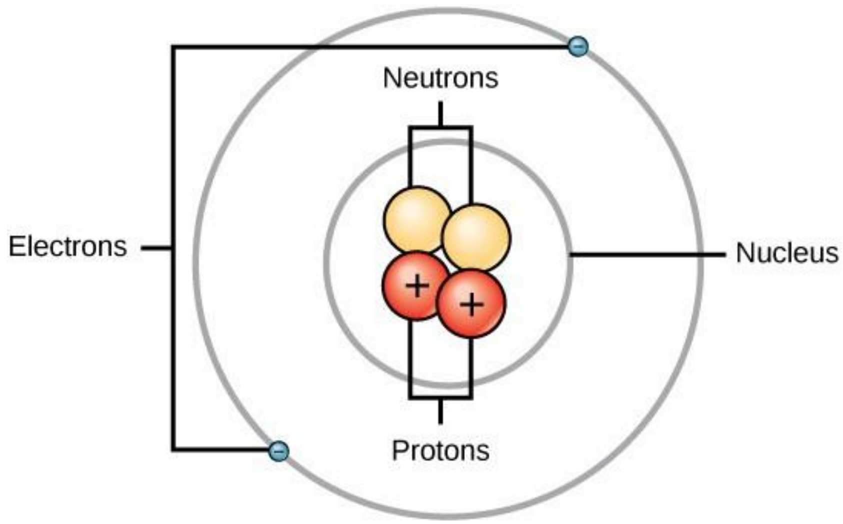
He = Chemical Symbol

4.00 = Atomic Mass (in amu)

Helium = Name

2
He
4.00
Helium

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.