THE ATOMIC THEORY

Democritus

- The smallest particle of matter that still retains its properties is called an atom.
- In the fifth century B.C., the Greek philosopher ***Democritus*** proposed that matter is composed of a ***finite number of discrete particles***, named ***atomos*** (meaning uncuttable or indivisible)

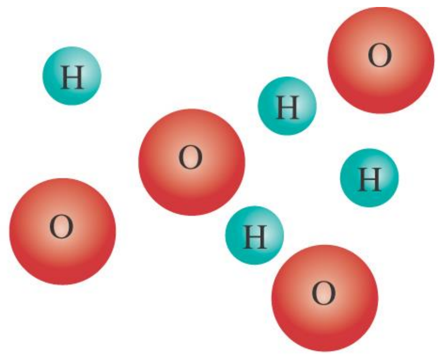
John Dalton

- In 1808, ***John Dalton***, built on ideas of Democritus, and formulated a precise definition of the building blocks of matter.
- Dalton's model represented the atom as a featureless ball of uniform density.

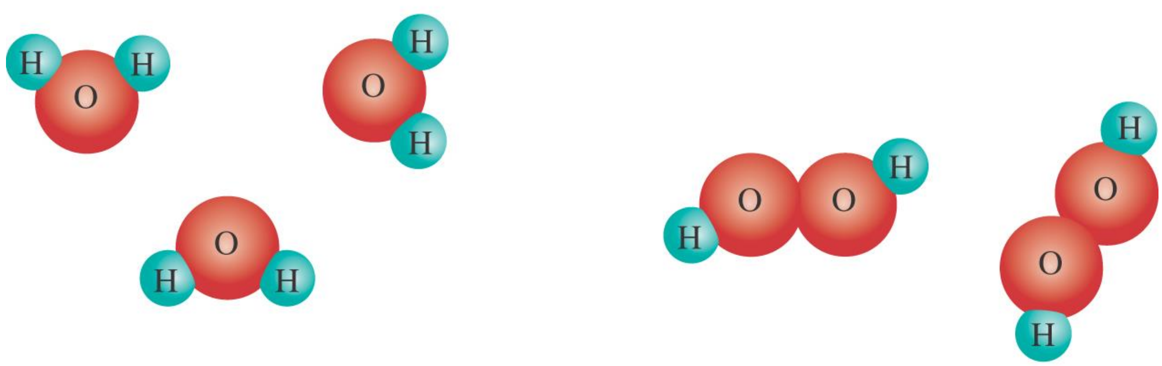
- This model is referred to as the "***soccer ball***" model.

DALTON'S ATOMIC THEORY

- Dalton's Atomic Theory consisted of three parts:
  - Each element is composed of tiny indestructible particles called atoms.
  - All atoms of a given element are similar to one another, but different from atoms of other elements.

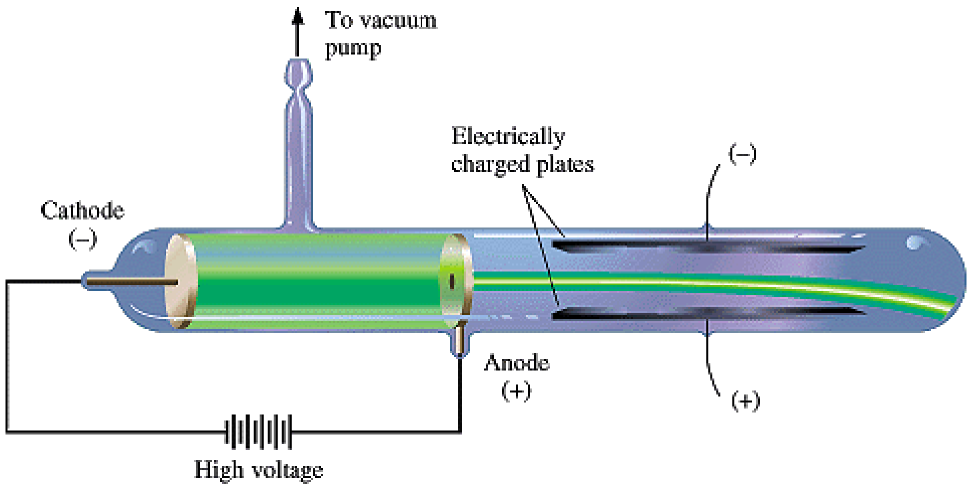


  - Atoms combine in simple, whole-number ratios to form compounds.



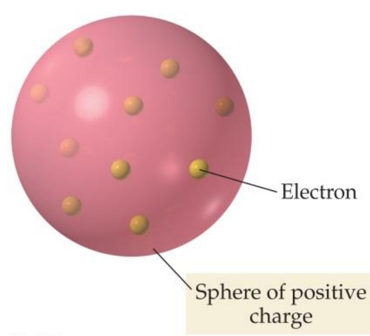
DISCOVERY OF THE ELECTRON

- In 1897, ***J.J. Thomson*** performed experiments with a ***cathode ray*** tube. Negatively charged particles from ***cathode*** were pulled towards positively charged plate, ***anode***, and allowed to pass through and be detected on a fluorescent screen.



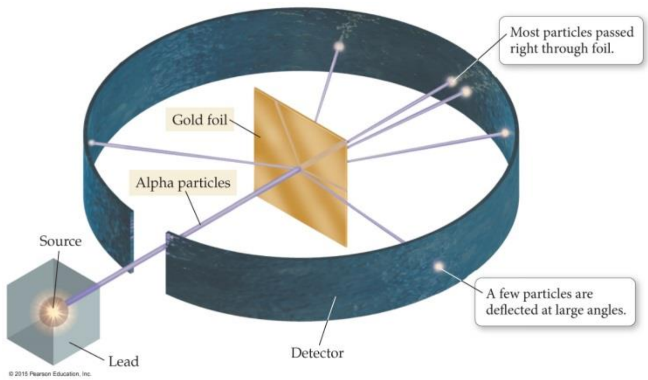
- In ***absence*** of a ***magnetic field***, the cathode rays were not deflected.
- In ***presence of a magnetic and electric fields***, the cathode rays were ***deflected*** towards the positive plate, indicating a ***negative charge.***
- The cathode rays were later named ***electrons***.

- Based on these findings, ***Thomson*** proposed an atomic model composed of ***negatively charged electrons*** embedded in a uniform ***positively charged sphere***.
- This model is called the "***plum pudding***" model.

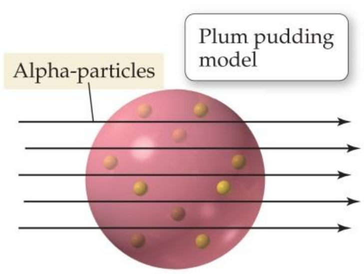


NUCLEAR MODEL OF THE ATOM

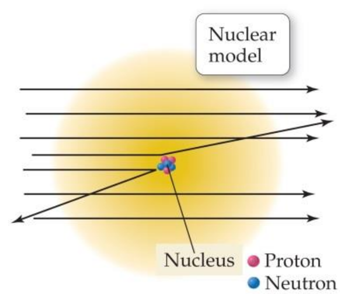
- In 1910, ***Ernest Rutherford*** carried out a number of experiments to further ***probe the nature of the atom***.
- In these experiments he bombarded a ***thin sheet of gold foil*** with ***α-particles*** (***large, positively charged***) emitted from a radioactive source.



- The ***majority*** of the particles were observed to ***pass through undeflected*** or ***slightly deflected***.
- ***Some*** of the particles were observed to be ***deflected at large angles***.
- ***Few*** of the particles were observed to be ***turned back*** towards the direction they came from.
- Thomson's model of the atom could not account for these results.



(a) Rutherford's expected result



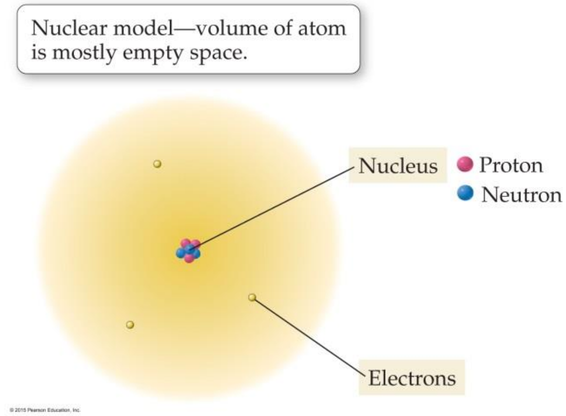
(b) Rutherford's actual result

- Based on these observations, ***Rutherford*** proposed a ***model of the atom*** consisting of a ***small, massive positive center*** (***nucleus***), ***surrounded*** by ***electrons*** in ***mostly empty space***.
- This model is called the ***nuclear model***.

THE STRUCTURE OF THE ATOM

- In 1932, ***James Chadwick*** discovered the existence of a second nuclear particle. This ***neutral particle*** was named ***neutron***.

Current Model of the Atom:



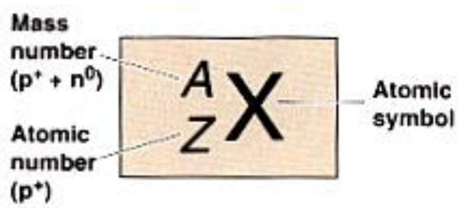
- The atom is an electrically ***neutral*** spherical entity.
- It is composed of a ***positively charged center*** surrounded by ***negatively charged electrons***.
- The ***electrons (e^-) move rapidly*** through the atomic volume, ***held*** by the ***attractive forces to the nucleus***.
- The ***atomic nucleus*** consists of positively charged ***protons*** (p^+) and neutrally charged ***neutrons*** (n^0).

- The modern atom consists of 3 subatomic particles:

TABLE 4.1 Subatomic Particles

	Mass (kg)	Mass (amu)	Charge
proton	1.67262×10^{-27}	1.0073	1+
neutron	1.67493×10^{-27}	1.0087	0
electron	0.00091×10^{-27}	0.00055	1–

Mass Relationships in the Atom:

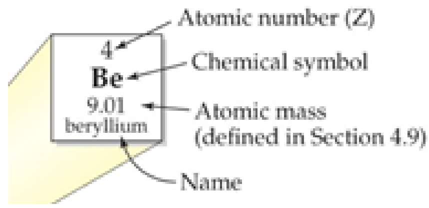


- The number of protons in an atom determines its identity, and is called ***atomic number (Z).***
- In a neutral atom, the number of protons (+) are equal to the number of electrons (–).
- Almost all the mass of the atom rests in the nucleus.
- The ***number of protons and neutrons*** in an atom is called the ***mass number (A).***
- The number of neutrons in an atom can be determined by the difference between mass number (A) and atomic number (Z):

$$\text{\# of neutrons} = \text{mass number} - \text{atomic number}$$
$$\text{\# of } n^0 = A - Z$$

ELEMENTS & THE PERIODIC TABLE

- The number of ***protons*** in an atom identify it as a particular element, and is called ***atomic number.***



- Elements are arranged on the periodic table according to their atomic number.

- Over time some elements have been named for planets, mythological figures, minerals, colors, geographic locations and famous people. Some examples are shown below:

Element	Source of Name
Uranium	The planet Uranus
Titanium	Titans (mythology)
Chlorine	*Chloros,* "greenish-yellow" (Greek)
Iodine	*Ioeides,* "violet" (Greek)
Magnesium	Magnesia, a mineral
Californium	California
Curium	Marie and Pierre Curie

- The ***symbol*** for most elements is the one- or two-letter abbreviation of the name of the element. Only the first letter of an elements symbol is capitalized. If the symbol has a second letter, it is written as lowercase.

Co (cobalt)
CO (carbon and oxygen)

- Although most of the symbols use letters from current names, some of the symbols of the elements are based on their Latin names.

K	potassium (kalium)
Na	sodium (natrium)
Fe	iron (ferrum)

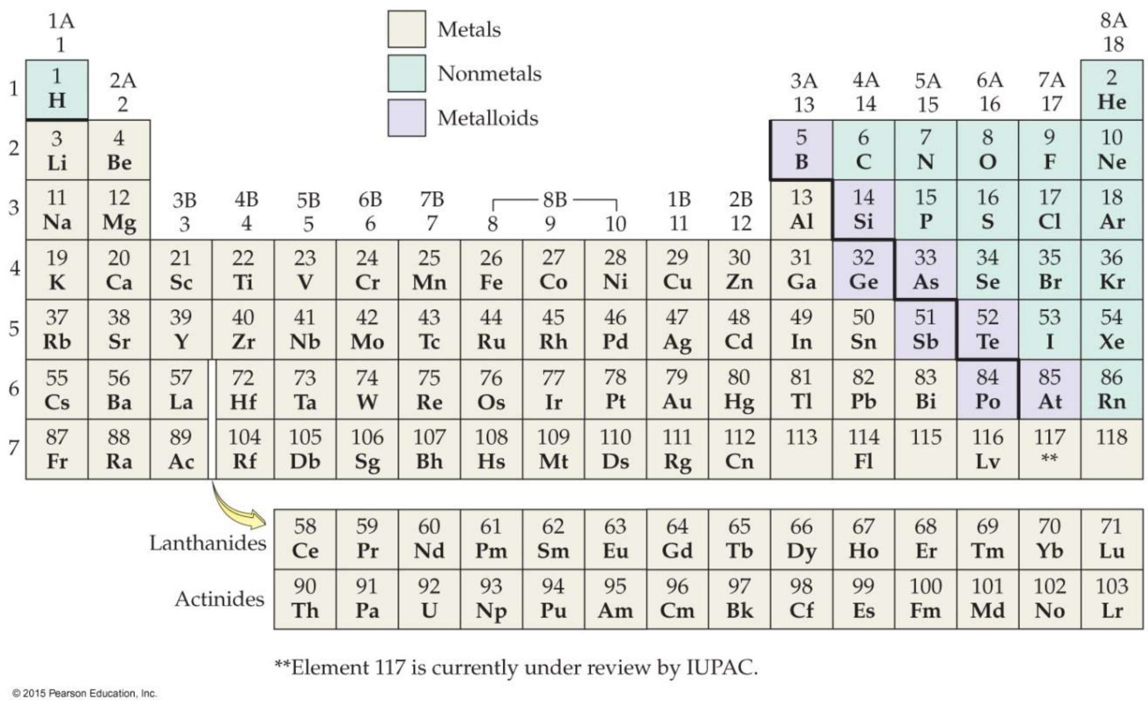
- Some elements have formulas that are ***not single*** atoms. ***Seven*** of these elements have ***diatomic*** (2-atoms) molecules.

Hydrogen	H_2	Chlorine	Cl_2
Oxygen	O_2	Fluorine	F_2
Nitrogen	N_2	Bromine	Br_2
		Iodine	I_2

ELEMENTS & THE PERIODIC TABLE

- Arrangement of elements based on their atomic masses was first proposed by the Russian chemist, ***Dmitri Mendeleev*** in 1869.

- The elements are generally classified as ***metals***, ***nonmetals*** and ***metalloids***.



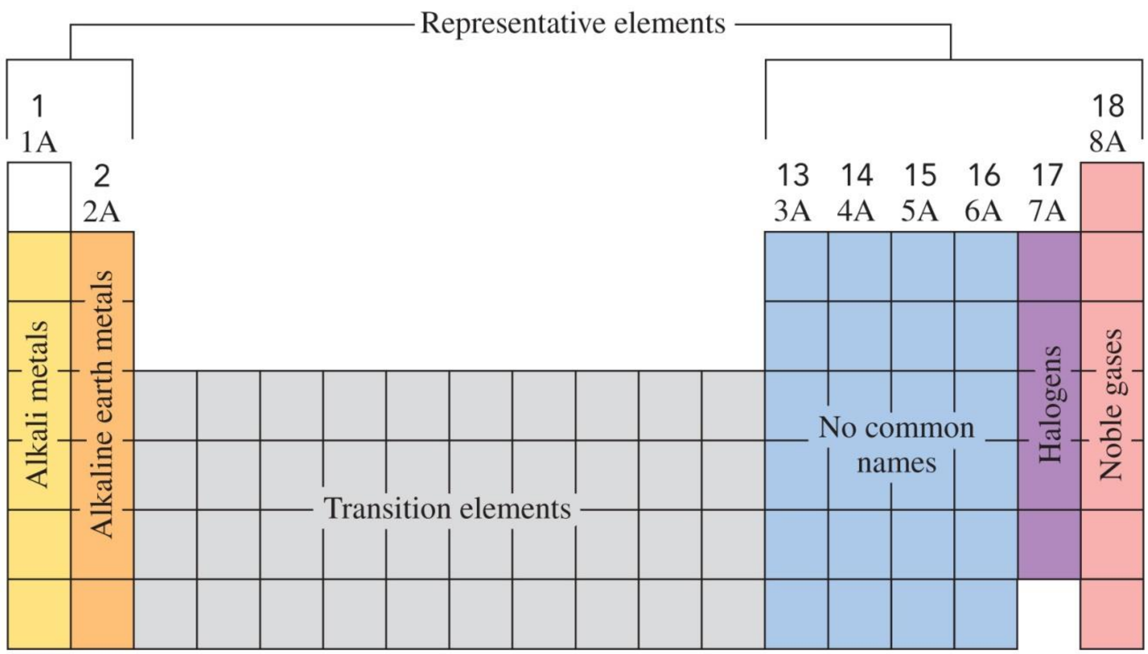
Metals	***Nonmetals***
1. Mostly solid	1. Can be solid, liquid or gas
2. Have shiny appearance	2. Have dull appearance
3. Good conductors of heat and electricity	3. Poor conductors of heat and electricity
4. Are malleable and ductile	4. Are brittle (if solid)
5. Lose electrons in a chemical reaction	5. Gain or share electrons in a chemical reaction

- **Metalloids** are elements that possess some properties of metals and some of non-metals. The most important metalloids are silicon (Si) and germanium (Ge) which are used extensively in computer chips.

ELEMENTS & THE PERIODIC TABLE

- The periodic table is composed of **periods** (rows) and **groups or families** (columns).

- Elements in the same family have similar properties, and are commonly referred to by their traditional names.



- Elements in groups ***1-2*** and ***13-18*** are referred to as ***main-group*** or ***representative elements***.

- Group 1 elements are called ***alkali metal.*** These are ***soft*** metals that are ***very reactive***, and often react explosively with other elements.

- Group 8 elements are called n***oble gases.*** These are ***un-reactive*** gases that are commonly used in light bulbs.

- Group 7 elements are called h***alogens.*** These are the most ***reactive nonmetals***, and occur in nature only as compounds.

- Group 2 elements are called ***alkaline-earth metals***. These metals are less reactive than alkali metals.

- The group of metals in-between the main group elements are called the ***transition metals***.

FORMATION OF IONS

- When reacting, atoms often ***lose or gain electrons*** to form charged particles, called ***ions***.

- For example, when reacting, lithium atom (3 protons and 3 electrons) loses one electron to form a Li^+ ion:

$$Li \longrightarrow Li^+ + e^-$$

- The charge of an ion is determined by the imbalance of the number of protons and electrons in the atom. Therefore, for lithium ion, the charge is determined as follows:

Ion charge = number of protons – number of electrons
Ion charge = 3 – 2 = 1+

- Similarly, when reacting, fluorine atom (9 protons and 9 electrons) gains an electron to form a F^- ion:

$$F + e^- \longrightarrow F^-$$

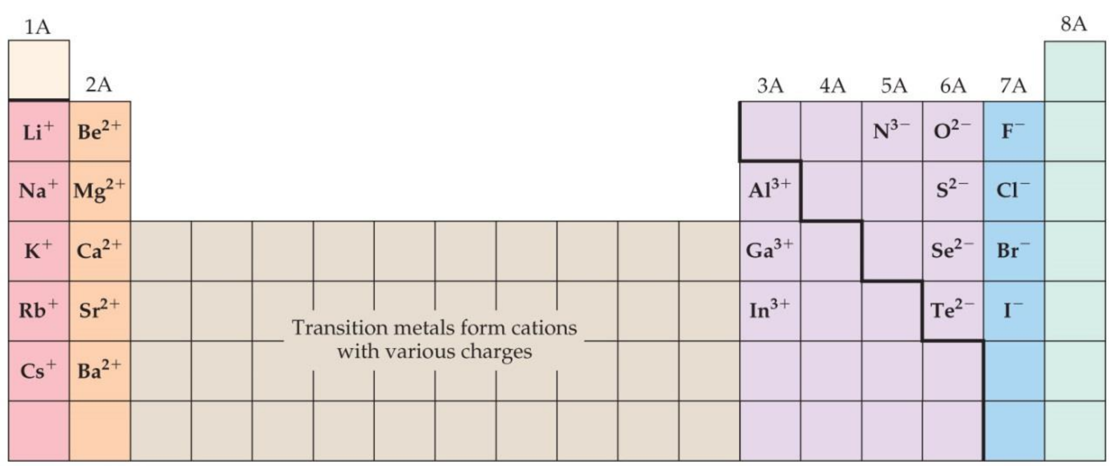
Ion charge = number of protons – number of electrons
Ion charge = 9 – 10 = 1–

- ***Positive ions*** are called ***cations***. Cations are formed when a neutral atom ***loses electrons***.

- ***Negative ions*** are called ***anions***. Anions are formed when a neutral atom ***gains electrons***.

- In nature, cations and anions occur together so that matter remains charge-neutral. For example, in table salt, sodium cations (Na^+) occur together with chloride anions (Cl^-).

- For most main-group elements, the periodic table can be used to predict the number of electrons lost or gained to form ions.



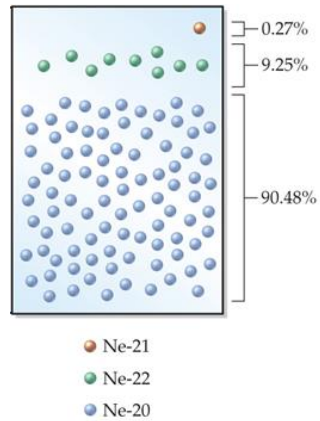
ISOTOPES & ATOMIC MASS

- Atoms of the ***same element*** that posseess ***different*** number of ***neutrons*** are called ***isotopes***.

- ***Isotopes*** of an element have the ***same atomic number (Z),*** but ***different mass number (A).***

$^{20}_{10}\mathbf{Ne}$ $^{21}_{10}\mathbf{Ne}$ $^{22}_{10}\mathbf{Ne}$

Ne-20 (90.48%) Ne-21 (0.27%) Ne-22 (9.25%)



- For a given element, the relative amount of each isotope in a naturally occurring sample of that element is always the same.

- The ***mass*** of an atom is measured ***relative*** to the mass of a chosen ***standard*** (***carbon-12 atom***), and is expressed in ***atomic mass units*** (***amu***).

- The ***average atomic mass*** of an element–called ***atomic mass***– is the ***mass*** of that element's natural occurring ***isotopes weighted*** according to their ***abundance***. Therefore the atomic mass of an element is closest to the mass of its most abundant isotope.

<u>Examples:</u>

1. Determine the number of protons, neutrons and electrons in $^{35}_{17}Cl$.

   number of p^+=
   number of e^- =
   number of n^0 =

2. Which two of the following are isotopes of each other?

   $^{410}_{186}X$ $^{410}_{185}Y$ $^{412}_{183}Z$ $^{412}_{185}R$

3. Based on the information below, which is the most abundant isotope of boron (atomic mass = 10.8 u)?

Isotope	^{10}**B**	^{11}**B**
Mass (amu)	10.0	11.0

CALCULATING ATOMIC MASSES FROM ISOTOPIC DATA

- The ***atomic mass*** of an element can be calculated from the mass and abundance of its isotopes as shown below:

$$\left\{\begin{matrix}\textbf{Atomic mass}\\ \textbf{of an element}\end{matrix}\right\} = \left\{\begin{pmatrix}\textbf{Abundance}\\ \textbf{of isotope 1}\end{pmatrix} \mathbf{x} \begin{pmatrix}\textbf{Mass of}\\ \textbf{isotope 1}\end{pmatrix}\right\} + \left\{\begin{pmatrix}\textbf{Abundance}\\ \textbf{of isotope 2}\end{pmatrix} \mathbf{x} \begin{pmatrix}\textbf{Mass of}\\ \textbf{isotope 2}\end{pmatrix}\right\}$$

<u>Examples:</u>

1. Calculate the average atomic mass of silver from the following isotopic data:

Isotope	Mass (amu)	Abundance (%)
^{107}Ag	106.90509	51.84
^{109}Ag	108.90476	48.16

^{107}Ag=106.90509 x (0.5184)= 55.42 amu
^{109}Ag=108.90476 x (0.4816)= 52.45 amu

Atomic mass of Ag=55.42 amu+52.45 amu=107.87 amu

2. Calculate the average atomic mass of gallium from the following isotopic data:

Isotope	Mass (amu)	Abundance (%)
Ga-69	68.9256	60.11
Ga-71	70.9247	39.89

3. A fictitious element is composed of isotopes A and B with masses of 61.9887 and 64.9486 amu, respectfully. The atomic mass of the element is 64.52 amu. What can you conclude about the natural abundances of the two isotopes? Briefly explain your answer.