Physical Science 2020

(This fourth quarter is an Intro to Chemistry)

LEWISBURG HIGH SCHOOL

Week 13 Chapter 15 (Spring Semister 2020)

April 6 – April 10, 2020

Please Only Take ONE Packet Per Student

Chapter 15 Atoms The Building Blocks of Matter

The particle theory of matter was supported as early as 400 B.C. by certain Greek thinkers, such Democritus. He called nature's basic particle an atom, based on the Greek word meaning "indivisible".

- Aristole on the other hand, did not believe in atoms. He thought that all matter was continuous, and his opinion was accepted for nearly 2000 years.
- Neither view was supported by experimental evidence, so each remained speculation until the eighteenth century. Then scientists began to gather evidence favoring the atomic theory of matter.

Foundations of Atomic Theory

- The transformation of a substance or substances into one or more new substances is known as a chemical reaction.
- In the 1790s, the study of matter was revolutionized by a new emphasis on the quantitative analysis of chemical reactions. Aided by improved balances, investigators began to accurately measure the masses of the elements and compounds they were studying. This lead to the discovery of several basic laws:
  - The law of conservation of mass states that mass is neither destroyed nor created during ordinary chemical reactions or physical changes.

  The law of definite proportions states that a chemical compound contains the same elements in exactly the same proportions by mass regardless of the size of the sample or the source of the compound. For example, sodium chloride also known as table salt, always consists of 39.34% by mass of the element Na and 60.66% by mass of the element Cl.
  - The law of multiple proportions states if two or more different compounds are composed of the same two elements, then the ratio of the masses of the second element combined with a certain mass of the first element is always a ratio of small whole numbers. For example, the elements carbon and oxygen form two compounds, carbon dioxide and carbon monoxide. Consider samples each containing 1.0 g of carbon. In carbon dioxide, 2.66 g of oxygen combine with 1.0 g of carbon. In carbon monoxide 1.33 g of oxygen combine with 1.0 g of carbon. The ratio of the masses of oxygen in these two compounds is exactly 2.66 to 1.33, or 2 to 1.

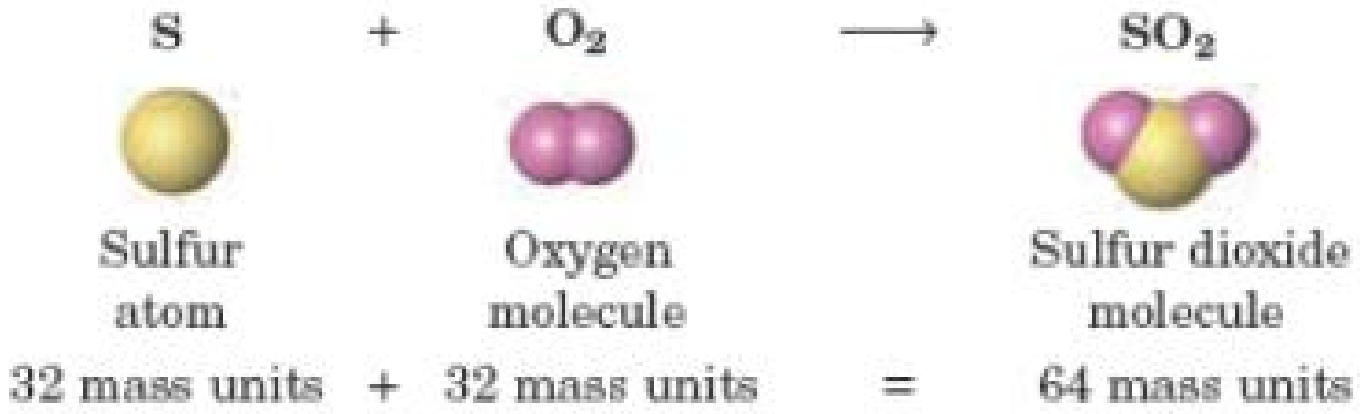
Dalton's Atomic Theory

- In 1808, an English schoolteacher named John Dalton proposed an explanation for these three laws. He reasoned that elements were composed of atoms and that only whole numbers of atoms can combine to form compounds. His theory can be summed up by the following statements.

1. All matter is composed of extremely small particles called atoms.
2. Atoms of a given element are identical in size, mass, and other properties. Atoms of different elemetns differ in size, mass, and other properties.
3. Atoms cannot be subdivided, created, or destroyed.
4. Atoms of different elements combine in simple whole number ratios to form chemical compounds.

5. In chemical reactions, atoms are combined, separated, or rearranged.

Which statement explains the law of conservation of mass?



Modern Atomic Theory

- By relating atoms to the measurable property of mass, Dalton turned Democritus's idea into a scientific theory that could be tested by experiment. But not all aspects of Dalton's atomic theory have proven to be correct.
  - For example, today we know that atoms are divisible into even smaller particles (although the law of conservation of mass still holds true for chemical reaction). Also a given element can have atoms with different masses, called isotopes.
  - Not all of the atomic theory has been discarded. Instead it has been modified to explain the new observations. The important concepts that (1) all matter is composed of atoms and that (2) atoms of any one element differ in properties atoms of another element, remain unchanged.

Circle the letter of the best answer.

1. The basic laws of chemistry were discovered when

a. improved equipment allowed investigators to accurately measure elements.

b. Aristotle determined that matter is continuous.

c. Democritus studied particle theory in 400 B.C.

d. carbon and oxygen combined to form carbon monoxide and carbon dioxide.

2. Which of the following proves the law of definite proportions?

a. Oxygen is heavier than hydrogen.

b. As the pressure on oxygen increases, its volume decreases.

c. Water is always composed of twice as many hydrogen atoms as oxygen atoms.

3. Whether you examine a small amount or a large amount, calcium fluoride always contains one atom of calcium for every two atoms of fluorine. This phenomenon is explained by

a. Aristotle's theory of atoms.

b. the law of conservation of mass.

c. the law of definite proportions.

d. the law of multiple proportions.

4. When elements react to form compounds,
a. they always create a flame or spark.
b. some of the elements' mass is destroyed.
c. new atoms are formed by the reaction.
d. they combine in fixed proportions by mass.

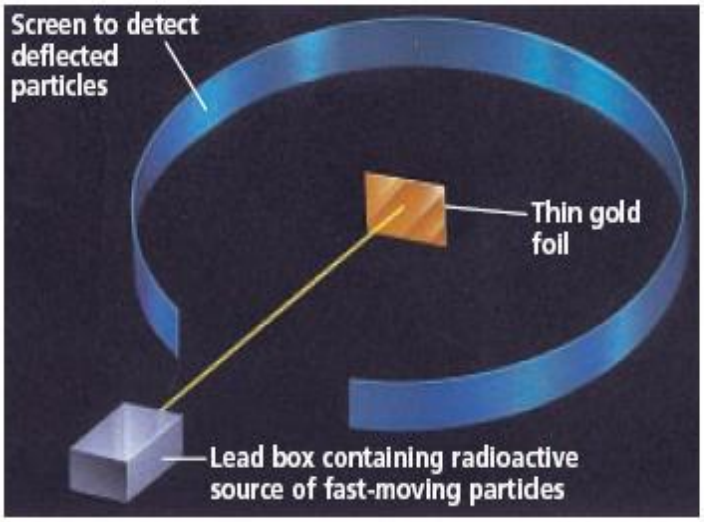
Use the following values to complete the table. Show your work. An example is done for you.

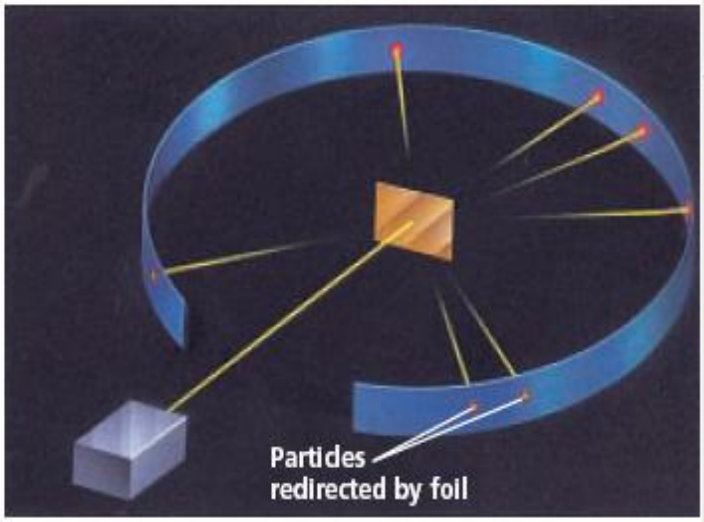
Weight of element A = 4 units Weight of element C = 6 units
Weight of element B = 3 units Weight of element D = 5 units

Compound	Weight of Compound
ABC_2	1(A) + 1(B) + 2(C) = 1(4) + 1(3) + 2(6) = 19
A_3BD_3	
AD_2	
BC_6D_4	
B_2CD_5	
A_2CD_3	
ABCD	

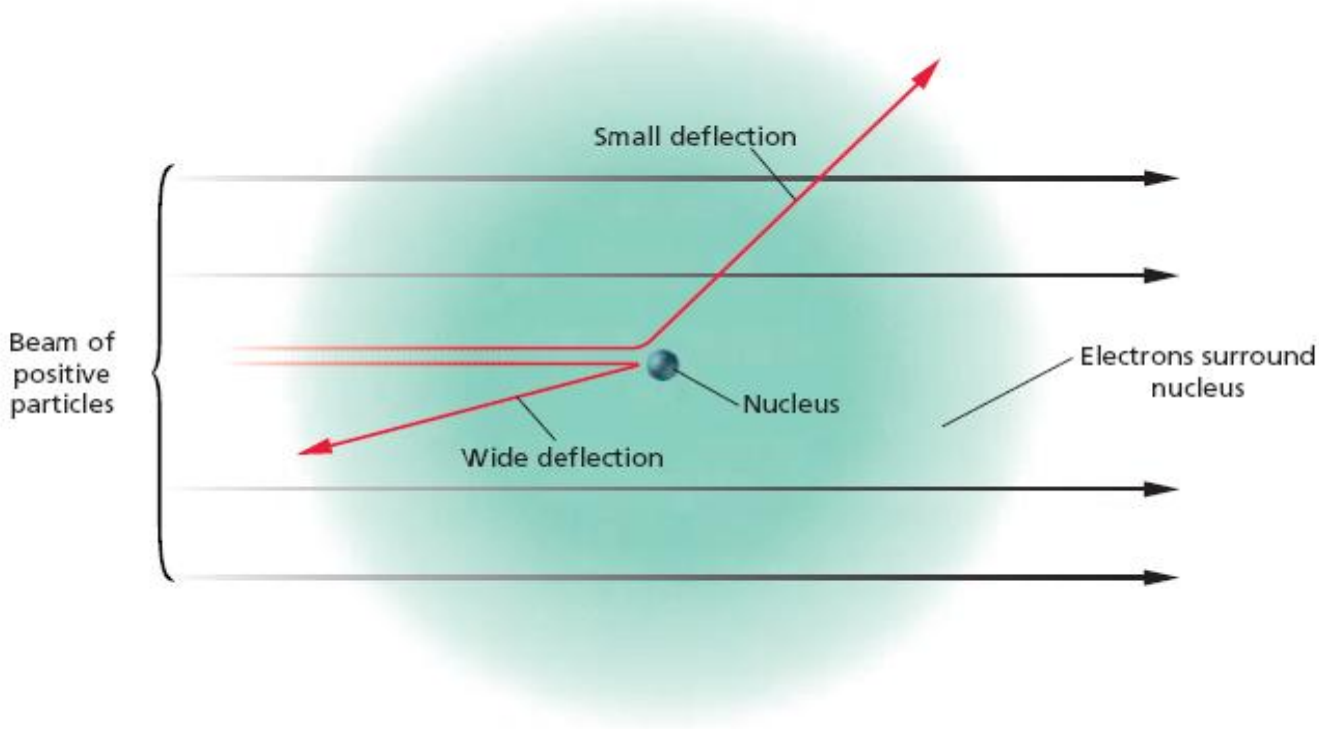
Atomic Structure:
An atom is the smallest particle of an element that retains the chemical properties of that element.
2 Regions of the Atom:
1) Nucleus: very small region located near the center of an atom.

- Contains:
  protons: positively charged particles
  neutrons: neutral (no charge) particles
- Ernest Rutherford performed the "Gold Foil Experiment"
  - discovered the nucleus by shooting a thin, gold foil with fast moving alpha particles (positively charged particles)
  - suggested that the electrons surrounded the positively charged nucleus like planets around the sun.





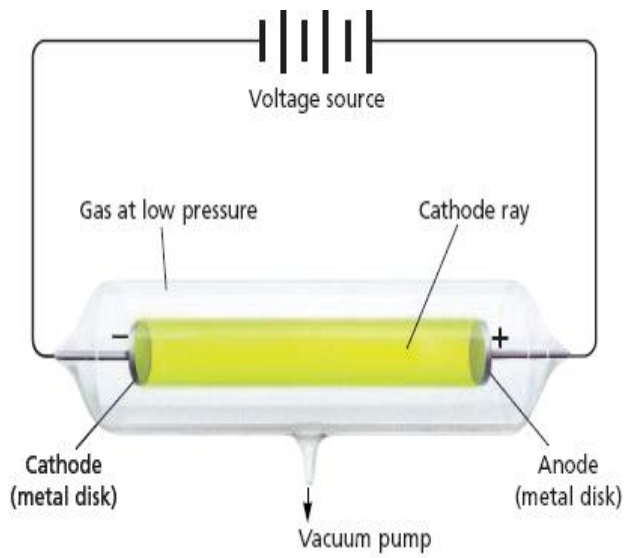
Rutherford and his associates assumed mass and charge were evenly distributed throughout the nucleus. If that were the case, most alpha particles would pass through with a few being slightly deflected.
However, when the scientists checked for the possibility of wide-angle deflections, they were shocked to find that
roughly 1 in 8000 of the alpha particles had actually been redirected back toward the source.
As Rutherford later exclaimed,
it was "as if you had fired a 15-inch [artillery] shell at a piece of tissue paper and it came back and hit you."

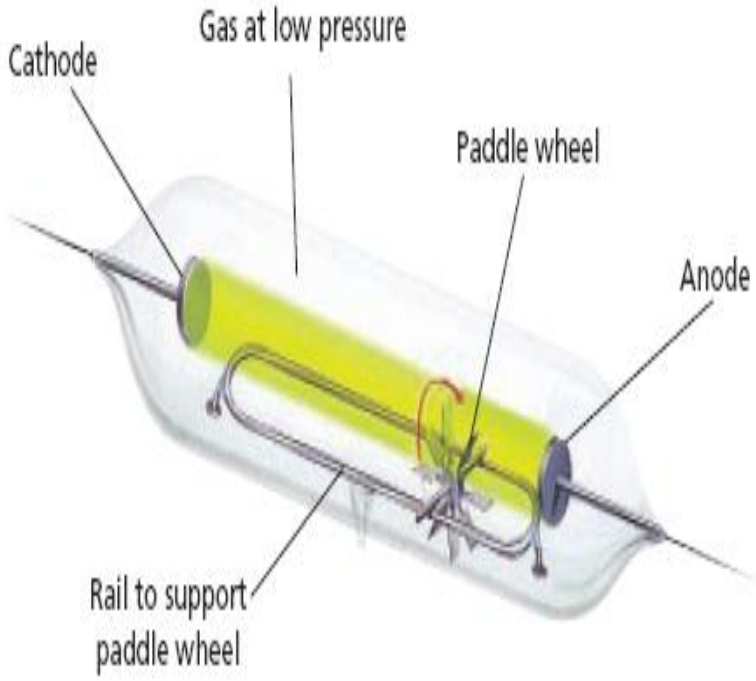


2 Regions of the Atom:

2) Surrounding the nucleus: negatively charged particles

- discovered by J.J. Thomson
- performed the cathode ray experiment (concluded that the electron had a negative charge)
- Robert A. Millikan also performed experiments to confirm Thomson's discovery and found the mass of the electron.





Chapter 15 Counting Atoms

1) Atomic number: the number of protons in the nucleus of each atom of that element.

- the number on the periodic table

2) Isotopes: atoms of the same element that have different masses.

- have the same number of protons and neutrons but different numbers of electrons.
- designating isotopes:
  - Hyphen notation: the mass number is written with a hyphen after the name of the element.
    - Example: uranium-235
  - Nuclear symbol: the element symbol with the subscript indicating the atomic number and the superscript indicating the mass number.
    - Example:

3) Mass number: the total number of protons and neutrons in the nucleus of an isotope.

- mass number - atomic number = number of neutrons

uranium-235:

235(protons + neutrons) - 92(protons) =143(neutrons)

Practice Problems:

1. How many protons, electrons, and neutrons are there in an atom of chlorine-37?
2. How many protons, electrons, and neutrons are there in an atom of bromine-80?
3. Write the nuclear symbol for carbon-13
4. Write the hyphen notation for the element that contains 15 electrons and 15 neutrons?

Atomic mass unit (1 amu): exactly 1/12 the mass of a carbon-12 atom. Each element is compared to this number.

Average atomic mass: the weighted average of the atomic masses of the naturally occurring isotopes of an element.

- A simple example of how to calculate a weighted average: Suppose you have a box of marbles. 25% of the marbles have masses of 2.00 g each and 75% have masses of 3.00 g each, calculate the weighted average.

.25 marbles x 2.00 g = .50 g

.75 marbles x 3.00 g = 2.25 g

Add the two masses to get the average mass = 2.75 g

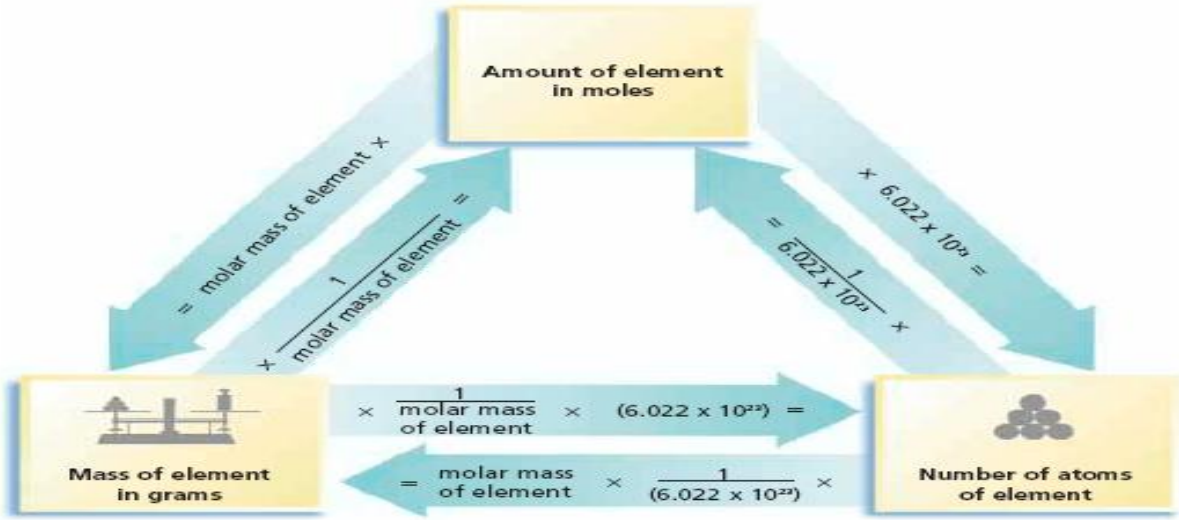
Copper has 69.17% copper-63 (atomic mass of 62.929598 amu) and 30.83% copper-65 (atomic mass of 64.927793 amu). Find the average atomic mass. Round to 2 decimal places.

Relating mass to numbers of atoms:

The mole is the SI unit for amount of substance. A mole is the amount of substance that contains as many particles as there are atoms in exactly 12 g of carbon-12. The mole is a counting unit, just like a dozen is.

Avogadro's number: 6.022×10^{23} -is the number of particles in exactly one mole of a pure substance.

Molar mass is the mass of one mole of a pure substance. The unit is g/mol. The molar mass is numerically equal to the atomic mass of the element in amu (which can be found on the periodic table. A molar mass of an element contains one mole of atoms.



Indicate the letter of the best answer. (Email the question number with the letter answer and 3 sentence essay to John Estes at the School Email Address)(This is a open note assignment.)

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5. What is referred to atoms of the same atomic number, but different masses?

    a. Ions

    b. Isomers

    c. Cations

    d. Anions

After you finish the questions leave a three sentence or more essay on what you know, what you wanted to know, or what you learned. Thank you