CHAPTER 5: INTRODUCTORY ATOMIC THEORY AND STRUCTURE

Enhanced Introductory College Chemistry

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Chapter Contents

Please visit the web version of Enhanced Introductory College Chemistry (https://ecampusontario.pressbooks.pub/enhancedchemistry/) to access the complete book, interactive activities and ancillary resources.

In this chapter, you will learn about

- Dalton's atomic theory
- Basic concepts in electric charge
- The discoveries of an atom's subatomic particles atom including electrons, protons, and neutrons
- The structure of the nuclear atom and how neutral atoms form ions
- Isotopes of elements and how to read and write isotope symbols
- Atomic masses of elements and their isotopes and how to calculate percent abundance.

To better support your learning, you should be familiar with the following concepts before starting this chapter:

- How to reference element names, symbols, and masses from the periodic table
- How to use your scientific calculator to perform percent calculations

Your overall health and susceptibility to disease depends upon the complex interaction between your genetic makeup and environmental exposure, with the outcome difficult to predict. Early detection of biomarkers, substances that indicate an organism's disease or physiological state, could allow diagnosis and treatment before a condition becomes serious or irreversible. Recent studies have shown that your exhaled breath can contain molecules that may be biomarkers for recent exposure to environmental contaminants or for pathological conditions ranging from asthma to lung cancer (Figure 5a). Scientists are working to develop biomarker "fingerprints" that could be used to diagnose a specific disease based on the amounts and identities of certain molecules in a patient's exhaled breath. An essential concept underlying this goal is that of a molecule's identity, which is determined by the numbers and types of atoms it contains, and how they are bonded together. This chapter will describe some of the fundamental chemical principles related to the composition of matter, including those central to the concept of molecular identity.

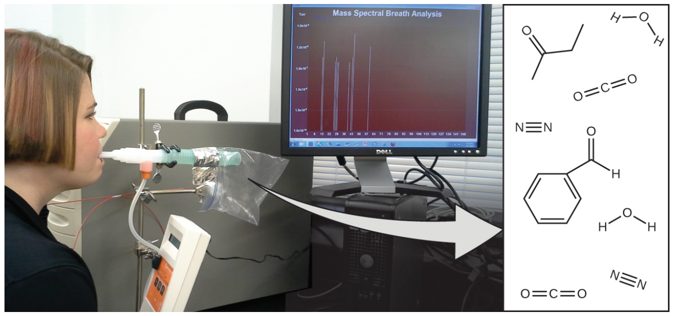

Figure 5a: Mass spectral breath analysis: analysis of molecules in an exhaled breath can provide valuable information, leading to early diagnosis of diseases or detection of environmental exposure to harmful substances. (credit: modification of work by Paul Flowers in *Chemistry (OpenStax)*, CC BY 4.0).

Attribution & References

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5.1 EARLY ATOMIC THEORY: DALTON'S MODEL OF THE ATOM

Learning Objectives

By the end of this section, you will be able to:

- Summarize the postulates of Dalton's atomic theory
- Apply Dalton's atomic theory to explain the laws of definite and multiple proportions

The language used in chemistry is seen and heard in many disciplines, ranging from medicine to engineering to forensics to art. The language of chemistry includes its own vocabulary as well as its own form of shorthand. Chemical symbols are used to represent atoms and elements. Chemical formulas depict molecules as well as the composition of compounds. Chemical equations provide information about the quality and quantity of the changes associated with chemical reactions.

This chapter will lay the foundation for our study of the language of chemistry. The concepts of this foundation include the atomic theory, the composition and mass of an atom, and the variability of the composition of isotopes.

Atomic Theory through the Nineteenth Century

Watch The 2,400-year search for the atom – Theresa Doud (6 mins) (https://www.youtube.com/watch?v=xazQRcSCRaY)

The earliest recorded discussion of the basic structure of matter comes from ancient Greek philosophers, the scientists of their day. In the fifth century BC, Leucippus and Democritus argued that all matter was composed of small, finite particles that they called *atomos*, a term derived from the Greek word for "indivisible." They thought of atoms as moving particles that differed in shape and size, and which could join together. Later, Aristotle and others came to the conclusion that matter consisted of various combinations of the four "elements"—fire, earth, air, and water—and could be infinitely divided. Interestingly, these philosophers thought about atoms and "elements" as philosophical concepts, but apparently never considered performing experiments to test their ideas.

The Aristotelian view of the composition of matter held sway for over two thousand years, until English schoolteacher John Dalton helped to revolutionize chemistry with his hypothesis that the behaviour of matter could be explained using an atomic theory. First published in 1807, many of Dalton's hypotheses about the microscopic features of matter are still valid in modern atomic theory. Here are the postulates of **Dalton's atomic theory**.

1. Matter is composed of exceedingly small, indivisible particles called atoms. An **atom** is the smallest unit of an element that can participate in a chemical change.
2. An **element** consists of only one type of atom, which has a mass that is characteristic of the element and is the same for all atoms of that element (Figure 5.1a). A macroscopic sample of an element contains an incredibly large number of atoms, all of which have identical chemical properties

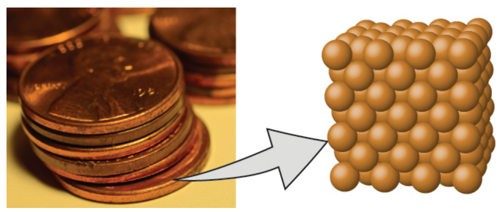


Figure 5.1a: Macroscopic vs microscopic visual of the element copper: a pre-1982 copper penny (left) contains approximately 3×10^{22} copper atoms (several dozen are represented as brown spheres at the right), each of which has the same chemical properties. (credit: modification of work by slgc, CC BY 2.0; in *Chemistry (OpenStax)*, CC BY 4.0).

3. Atoms of one element differ in properties from atoms of all other elements.
4. A compound consists of atoms of two or more elements combined in a small, whole-number ratio. In a given compound, the numbers of atoms of each of its elements are always present in the same ratio (Figure 5.1b).

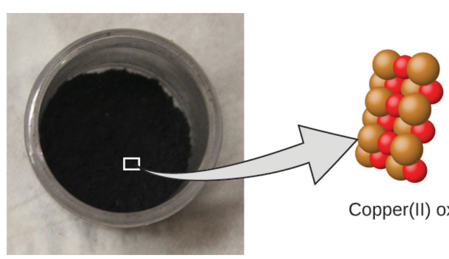


Figure 5.1b: Macroscopic vs microscopic visual of the compound copper(II) oxide: copper(II) oxide, a powdery, black compound, results from the combination of two types of atoms—copper (brown, larger spheres) and oxygen (red, smaller spheres)—in a 1:1 ratio. (credit: modification of work by Chemicalinterest, PD; in *Chemistry (OpenStax)*, CC BY 4.0).

5. Atoms are neither created nor destroyed during a chemical change, but are instead rearranged and regrouped to yield substances that are different from those present before the change (Figure 5.1c).

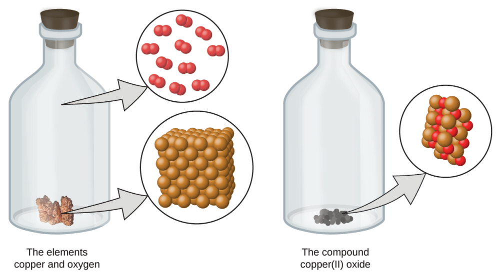


Figure 5.1c: Visual of Dalton's theory: atoms are neither created nor destroyed during a chemical change but are rearranged to yield new substances. When the elements copper (a shiny, red-brown solid, shown here as brown spheres shaped as a cube) and oxygen (a clear and colourless gas, shown here as red spheres that are present in pairs) react, their atoms rearrange to form a compound containing copper and oxygen (a powdery, black solid). (credit copper: modification of work by Anonymous, CC BY 3.0; in *Chemistry (OpenStax)*, CC BY 4.0).

Dalton's atomic theory laid the foundation in the development of chemistry. Most of his postulates remain valid; however, some of his conclusions have been revolutionized because further investigations have shown that

1. Atoms are composed of subatomic particles; they are not indivisible
2. Not all the atoms of a specific element have the exact same mass; an element can exist in different forms, called isotopes
3. Atoms, under special conditions, can be decomposed.

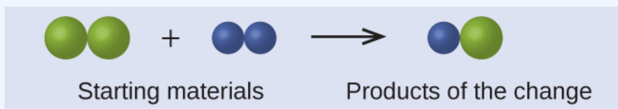
Source: (Hein & Arena, 2014, p. 83)

Despite these flaws, Dalton's atomic theory provides a microscopic explanation of the many macroscopic properties of matter that you've learned about. For example, if an element such as copper consists of only one kind of atom, then it cannot be broken down into simpler substances, that is, into substances composed of fewer types of atoms. And if atoms are neither created nor destroyed during a chemical change, then the total mass of matter present when matter changes from one type to another will remain constant (the law of conservation of matter).

Example 5.1a

Testing Dalton's Atomic Theory

In the following drawing, the green, larger spheres represent atoms of a certain element. The blue, smaller spheres represent atoms of another element. If the spheres touch, they are part of a single unit of a compound. Does the following chemical change represented by these symbols violate any of the ideas of Dalton's atomic theory? If so, which one?

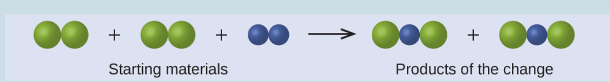


Solution

The starting materials consist of two green, larger spheres and two blue, smaller spheres. The products consist of only one green sphere and one blue sphere. This violates Dalton's postulate that atoms are neither created nor destroyed during a chemical change, but are merely redistributed. (In this case, atoms appear to have been destroyed.)

Exercise 5.1a

In the following drawing, the green, larger spheres represent atoms of a certain element. The blue, smaller spheres represent atoms of another element. If the spheres touch, they are part of a single unit of a compound. Does the following chemical change represented by these symbols violate any of the ideas of Dalton's atomic theory? If so, which one?



Check Your Answer

1

Exercise 5.1b

Check Your Learning Exercise (Text Version)
Choose the answer that best answers the questions for each of the multiple choice questions.

1. According to Dalton's Atomic Theory, matter consists of indivisible
   a. Molecules
   b. Cells
   c. Atoms
   d. Subatomic particles
2. Dalton's Atomic Theory postulates mass is neither created nor destroyed in a chemical reaction. This supports the
   a. Law of constant proportions
   b. Law of conservation of mass
   c. Law of multiple proportions
   d. Law of electric force
3. Fill in the blanks with the correct pair of terms from the choices below according to Dalton's Atomic Theory: All atoms of a given element have identical ________ including identical _______.
   a. Temperature, volume
   b. Force, pressure
   c. Weight, volume
   d. Physical and chemical properties, mass
4. Which of the following is NOT a postulate of Dalton's Atomic Theory?
   a. Atoms that combine to form new molecules do so in simple, whole number ratios
   b. A chemical reaction is a rearrangement of atoms. No atoms are created or destroyed.
   c. All elements are composed of small particles called atoms.
   d. Atoms of a given element are always identical.
   e. Atoms are always on motion
5. 5. Most of Dalton's postulates remain valid; however, some of Dalton's atomic theory postulates have been proven to be false. Which of the following postulates were determined

to be incorrect?

a. Matter is composed of exceedingly small, indivisible particles.
b. Elements consist of only one type of identical atom, which has the same mass for all atoms.
c. Theories stated in a) and b) were both proven to be false
d. Neither theory stated in a) nor b) were proven to be false

Check Your Answer

2

Source: "Exercise 5.1b" by Jackie MacDonald, licensed under CC BY-NC-SA 4.0

Dalton knew of the experiments of French chemist Joseph Proust, who demonstrated that *all samples of a pure compound contain the same elements in the same proportion by mass*. This statement is known as the **law of definite proportions** or the **law of constant composition**. The suggestion that the numbers of atoms of the elements in a given compound always exist in the same ratio is consistent with these observations. For example, when different samples of isooctane (a component of gasoline and one of the standards used in the octane rating system) are analyzed, they are found to have a carbon-to-hydrogen mass ratio of 5.33:1, as shown in Table 5.1a.

Table 5.1a: Constant Composition of Isooctane

Sample	Carbon	Hydrogen	Mass Ratio
A	14.82 g	2.78 g	$\frac{14.82\text{g carbon}}{2.78\text{g hydrogen}} = \frac{5.33\text{g carbon}}{1.00\text{g hydrogen}}$
B	22.33 g	4.19 g	$\frac{22.33\text{g carbon}}{4.19\text{g hydrogen}} = \frac{5.33\text{g carbon}}{1.00\text{g hydrogen}}$
C	19.40 g	3.64 g	$\frac{19.40\text{g carbon}}{3.64\text{g hydrogen}} = \frac{5.33\text{g carbon}}{1.00\text{g hydrogen}}$

It is worth noting that although all samples of a particular compound have the same mass ratio, the converse is not true in general. That is, samples that have the same mass ratio are not necessarily the same substance. For example, there are many compounds other than isooctane that also have a carbon-to-hydrogen mass ratio of 5.33 : 1.00.

Watch Thinking Reeds Chemistry – The Law of Definite Proportions (4 mins 13 sec) (https://www.youtube.com/watch?v=4-SjNzqFb5U)

Dalton also used data from Proust, as well as results from his own experiments, to formulate another interesting law. The **law of multiple proportions** states that *when two elements react to form more than one compound, a fixed mass of one element will react with masses of the other element in a ratio of small, whole numbers*. For example, copper and chlorine can form a green, crystalline solid with a mass ratio of 0.558 g chlorine to 1 g copper, as well as a brown crystalline solid with a mass ratio of 1.116 g chlorine to 1 g copper. These ratios by themselves may not seem particularly interesting or informative; however, if we take a ratio of these ratios, we obtain a useful and possibly surprising result: a small, whole-number ratio.

$$\frac{\frac{1.116\text{g Cl}}{1\text{g Cu}}}{\frac{0.558\text{g Cl}}{1\text{g Cu}}} = \frac{2}{1}$$

This 2-to-1 ratio means that the brown compound has twice the amount of chlorine per amount of copper as the green compound. When referencing Figure 5.1d, the above can be explained by atomic theory if the copper-to-chlorine ratio in the brown compound (Figure 5.1d (b)) is 1 copper atom to 2 chlorine atoms, and the ratio in the green compound (Figure 5.1d (a)) is 1 copper atom to 1 chlorine atom. The ratio of chlorine atoms in compound B compared to compound A (and thus the ratio of their masses) is therefore 2 to 1 (Figure 5.1d).

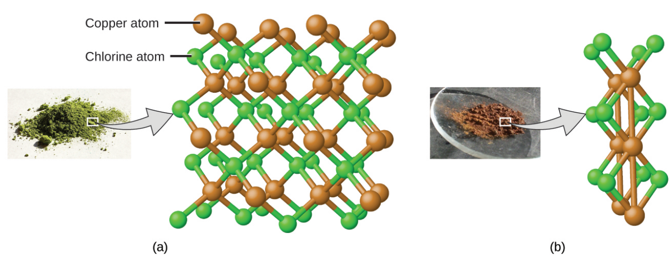


Figure 5.1d: Law of multiple proportions. Comparing two different compounds containing variable amounts of copper and chlorine atoms. Compared to the copper chlorine compound in (a), where copper is represented by brown, larger spheres and chlorine by green, smaller spheres, the copper chlorine compound in (b) has twice as many chlorine atoms per copper atom. (credit a: modification of work by Benjah-bmm27, PD; credit b: modification of work by Walkerma, PD; in *Chemistry (OpenStax),* CC BY 4.0).

Watch Multiple Proportions (1 min 11 sec) (https://www.youtube.com/watch?v=3dWdMqZ2UOU)

Example 5.1b

Laws of Definite and Multiple Proportions

A sample of compound A (a clear, colourless gas) is analyzed and found to contain 4.27 g carbon and 5.69 g oxygen. A sample of compound B (also a clear, colourless gas) is analyzed and found to contain 5.19 g carbon and 13.84 g oxygen. Are these data an example of the law of definite proportions, the law of multiple proportions, or neither? What do these data tell you about substances A and B?

Solution

In compound A, the mass ratio of carbon to oxygen is:

$$\frac{1.33\text{g O}}{1\text{g C}}$$

In compound B, the mass ratio of carbon to oxygen is:

$$\frac{2.67\text{g O}}{1\text{g C}}$$

The ratio of these ratios is:

$$\frac{\frac{1.33\text{g O}}{1\text{g C}}}{\frac{2.67\text{g O}}{1\text{g C}}} = \frac{1}{2}$$

The ratio of oxygen atoms of compound A to compound B (and thus the ratio of their masses) is 1 to 2. This supports the law of multiple proportions. This means that A and B are different compounds, with A having one-half as much oxygen per amount of carbon as compound B. A possible pair of compounds that would fit this relationship would be A = CO and B = CO_2.

Exercise 5.1c

A sample of compound X (a clear, colourless, combustible liquid with a noticeable odour) is analyzed and found to contain 14.13 g carbon and 2.96 g hydrogen. A sample of compound Y (a clear, colourless, combustible liquid with a noticeable odour that is slightly different from X's odour) is analyzed and found to contain 19.91 g carbon and 3.34 g hydrogen. Are these data an example of the law of definite proportions, the law of multiple proportions, or neither? What do these data tell you about substances X and Y?

Check Your Answer

3

Links to Interactive Learning Tools

Explore the Timeline of Atomic Discovery (https://h5pstudio.ecampusontario.ca/content/8507)from eCampusOntario H5P Studio (https://h5pstudio.ecampusontario.ca/).

Attribution & References

References

Hein, M., & Arena, S. (2014). *Foundations of College Chemistry (14th edition)*. Wiley & Sons.

Notes

1. The starting materials consist of four green, larger spheres and two blue, smaller spheres. The products consist of four green, larger spheres and two blue, smaller spheres. This does not violate any of Dalton's postulates: Atoms are neither created nor destroyed, but are redistributed in small, whole-number ratios.
2. 1c; 2b; 3d; 4e; 5c
3. In compound X, the mass ratio of carbon to hydrogen is $\frac{14.13\text{g C}}{2.96\text{g H}}$. In compound Y, the mass ratio of carbon to

oxygen is $\frac{19.91\text{g C}}{3.34\text{g H}}$. The ratio of these ratios is $\frac{\frac{14.13\text{g C}}{2.96\text{g H}}}{\frac{19.91\text{g C}}{3.34\text{g H}}} = \frac{4.77\text{g C/g H}}{5.96\text{g C/g H}} = 0.800 = \frac{4}{5}$. This small, whole-number ratio supports the law of multiple proportions. This means that X and Y are different compounds.

5.2 ELECTRIC CHARGE

Learning Objectives

By the end of this section, you will be able to:

- Describe the concept of electric charge and its properties

In the two centuries since Dalton developed his ideas, scientists have made significant progress in advancing our understanding of atomic theory. Some of this development came from the results of several pioneering experiments that revealed details of electric charge and discovery of ions. Before you learn about the internal structure of an atom and the experiments that led to their discovery, it is important to outline key concepts in electric charge and about the discovery of ions.

Electric Charge

You are certainly familiar with electronic devices that you activate with the click of a switch, from computers to cell phones to television. And you have certainly seen electricity in a flash of lightning during a heavy thunderstorm. But you have also most likely experienced electrical effects in other ways, maybe without realizing that an electric force was involved. Let's take a look at some of these activities and see what we can learn from them about electric charges and forces.

Discoveries

You have probably experienced the phenomenon of **static electricity**: When you first take clothes out of a dryer, many (not all) of them tend to stick together; for some fabrics, they can be very difficult to separate. Another example occurs if you take a woolen sweater off quickly—you can feel (and hear) the static electricity pulling on your clothes, and perhaps even your hair. If you comb your hair on a dry day and then put the comb close to a thin stream of water coming out of a faucet, you will find that the water stream bends toward (is attracted to) the comb (Figure 5.2a).

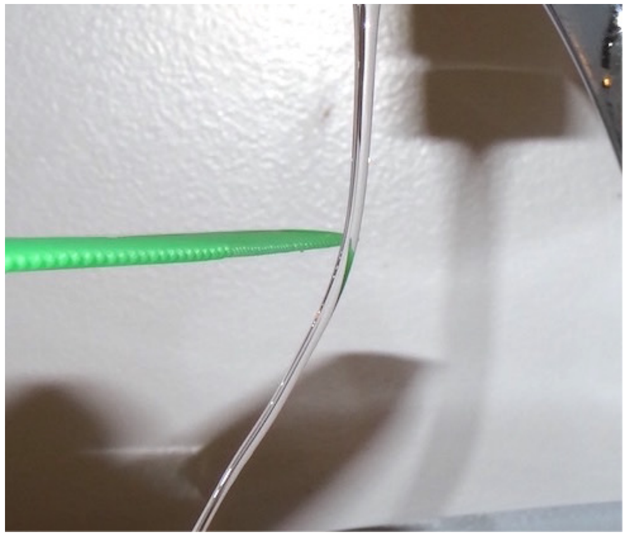


Figure 5.2a: Electric charge real life example 1: an electrically charged comb attracts a stream of water from a distance. Note that the water is not touching the comb. (credit: Jane Whitney in *University Physics Volume 2 (Open Stax)*, CC BY 4.0).

Suppose you bring the comb close to some small strips of paper; the strips of paper are attracted to the comb and even cling to it (Figure 5.2b). In the kitchen, quickly pull a length of plastic cling wrap off the roll; it will tend to cling to most any nonmetallic material (such as plastic, glass, or food). If you rub a balloon on a wall for a few seconds, it will stick to the wall. Probably the most annoying effect of static electricity is getting shocked by a doorknob (or a friend) after shuffling your feet on some types of carpeting.

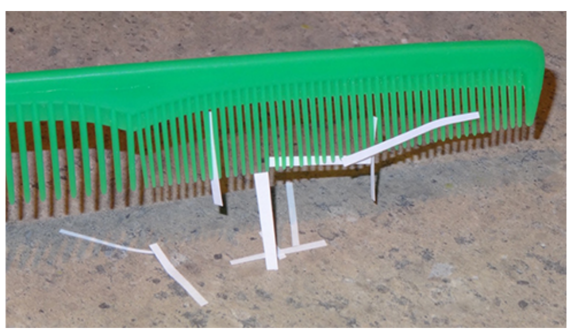


Figure 5.2b: Electric charge real life example 2: after being used to comb hair, this comb attracts small strips of paper from a distance, without physical contact. Investigation of this behaviour helped lead to the concept of the electric force (credit: Jane Whitney in *University Physics Volume 2 (Open Stax)*, CC BY 4.0).

Many of these phenomena have been known for centuries. The ancient Greek philosopher Thales of Miletus (624–546 BCE) recorded that when amber (a hard, translucent, fossilized resin from extinct trees) was vigorously rubbed with a piece of fur, a force was created that caused the fur and the amber to be attracted to each other. Additionally, he found that the rubbed amber (Figure 5.2c) would not only attract the fur, and the fur attract the amber, but they both could affect other (nonmetallic) objects, even if not in contact with those objects. (Figure 5.2d)

Figure 5.2c: Image of Borneo amber: Borneo amber is mined in Sabah, Malaysia, from shale-sandstone-mudstone veins. When a piece of amber is rubbed with a piece of fur, the amber gains more negative charge, giving it a net negative charge. At the same time, the fur, having lost what we now know to be called electrons, becomes positively charged. (credit: work by Sebakoamber, PD)

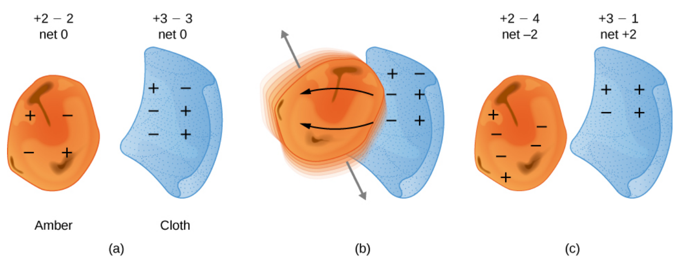


Figure 5.2d: Visual of attractive forces: when materials are rubbed together, charges can be separated, particularly if one material has a greater affinity for electrons than another. (a) Both the amber and cloth are originally neutral, with equal positive and negative charges. Only a tiny fraction of the charges are involved, and only a few of them are shown here. (b) When rubbed together, some negative charge is transferred to the amber, leaving the cloth with a net positive charge. (c) When separated, the amber and cloth now have net charges, but the absolute value of the net positive and negative charges will be equal (credit: *University Physics Volume 2 (Open Stax)*, CC BY 4.0).

The English physicist William Gilbert (1544–1603) also studied this attractive force, using various substances. He worked with amber, and, in addition, he experimented with rock crystal and various precious and semi-precious gemstones. He also experimented with several metals. He found that the metals never exhibited this force, whereas the minerals did. Moreover, although an electrified amber rod would attract a piece of fur, it would repel another electrified amber rod; similarly, two electrified pieces of fur would repel each other.

This suggested there were two types of an electric property, which eventually came to be called **electric charge**. It was concluded there were two types of electric charge – positive and negative. The difference between the two types of electric charge is in the directions of the **electric forces** that each type of charge causes:

- These forces are repulsive when the same type of charge exists on two interacting objects
- These forces are attractive when the charges are of opposite types

The most peculiar aspect of this new force is that it does not require physical contact between the two objects in order to cause an acceleration. This is an example of a so-called "long-range" force, (or, as James Clerk Maxwell later phrased it, "action at a distance"), which later became known as a form of induction.

The properties of electric charge are as follows:

- Charges can be positive and negative
- Electric force can be either attractive or repulsive
  - If two interacting objects carry the same sign of charge, the force is repulsive.
    - This interaction is referred to as **electrostatic repulsion**
  - If the charges are of opposite sign, the force is attractive.
    - This interaction is referred to as **electrostatic attraction**
- The magnitude of the force decreases (rapidly) with increasing separation distance between objects. The magnitude of the force increases (rapidly) with decreasing separation distance between the objects.
- The force acts by contact or induction (without physical contact between the two objects)
- Not all objects are affected by this force

Exercise 5.2a

Check Your Learning Exercise (Text Version)
Read the following statement about electric charge and determine whether the statement is True OR False.

1. Charges can be positive and neutral
2. Electric force can be either attractive or equal
3. If two interacting objects carry the same sign of charge, it results in electrostatic repulsion
4. If the charges are of opposite sign, the force is attractive.
5. If a balloon is rubbed on hair to gain charge and then is placed against a wall and sticks to the wall, the two objects have opposite charges
6. The magnitude of the force decreases (rapidly) with decreasing separation distance between objects
7. When two objects of similar charge repel each other without contact it is called induction.
8. All objects are affected by electric force

Check Your Answer[1]

Source: "Exercise 5.2a" by Jackie MacDonald, licensed under CC BY-NC-SA 4.0

The discovery that matter (and its atoms) has properties of electric charge and contain both positive and negative charges led to the theory that a given neutral atom may be able to lose or gain such charges and become positively or negatively charged atoms, respectively. These charged atoms were later defined as positive ions – cations and negative ions – anions. This concept will be discussed in more detail in upcoming sections and chapters.

Attribution & References

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Notes

1. For the following answers, any false answers, have been rewritten to show the correct statement. The bolded words (also noted with an *) were changed from the original false statement to make the statement true.
   1. False - Charges can be positive and ***negative**;
   2. False - Electric force can be either attractive or ***repulsive**;
   3. True;
   4. True;
   5. True;
   6. False - The magnitude of the force decreases (rapidly) with ***increasing** separation distance between objects OR The magnitude of the force ***increases** (rapidly) with decreasing separation distance between objects;
   7. True;
   8. False – ***Not** all objects are affected by electric force

5.3 SUBATOMIC PARTICLES OF THE ATOM

Learning Objectives

By the end of this section, you will be able to:

- Outline milestones in the development of modern atomic theory
- Summarize and interpret the results of the experiments of Thomson, Millikan, and Rutherford
- Describe the three subatomic particles that compose atoms

Scientists have made significant progress in furthering our understanding of atomic theory. Much of this came from the results of several seminal experiments that revealed the details of the internal structure of atoms. Here, we will discuss some of those key developments, with an emphasis on the application of the scientific method, as well as understanding how the experimental evidence was analyzed. While the historical persons and dates behind these experiments can be quite interesting, it is most important to understand the concepts resulting from their work.

Atomic Theory after the Nineteenth Century

If the matter were composed of atoms, what were atoms composed of? Was it just predominately European scientists questioning the properties of matter, its constituents, its behaviour, and why and how it exists? Or at that time, were other groups of people around the world, such as Indigenous communities, also asking similar questions? From their experiences, traditions, and cultural practices did they test their own hypotheses about the different matter around them – what makes mud, mud, and why is it different from sweetgrass before and after it is burned? Did they, too, ask the question what makes the smallest particles, or is there something smaller? Matter is made up of inconceivably small atoms, and yet scientists found atoms contain even smaller subatomic particles, including electrons, protons, and neutrons. The discovery of these subatomic particles is discussed next.

Revelation of the Electron

In the late 1800s, a number of scientists interested in questions like these investigated the electrical discharges that could be produced in low-pressure gases, with the most significant discovery made in 1897 by English physicist J. J. Thomson using a cathode ray tube. This apparatus consisted of a sealed glass tube from which almost all the air had been removed; the tube contained two metal electrodes. When high voltage was applied across the electrodes, a visible beam called a cathode ray appeared between them. This beam was deflected toward the positive charge and away from the negative charge, and was produced in the same way with identical properties when different metals were used for the electrodes. In similar experiments, the ray was simultaneously deflected by an applied magnetic field, and measurements of the extent of deflection and the magnetic field strength allowed Thomson to calculate the charge-to-mass ratio of the cathode ray particles. The results of these measurements indicated that these particles were much lighter than atoms (Figure 5.3a).

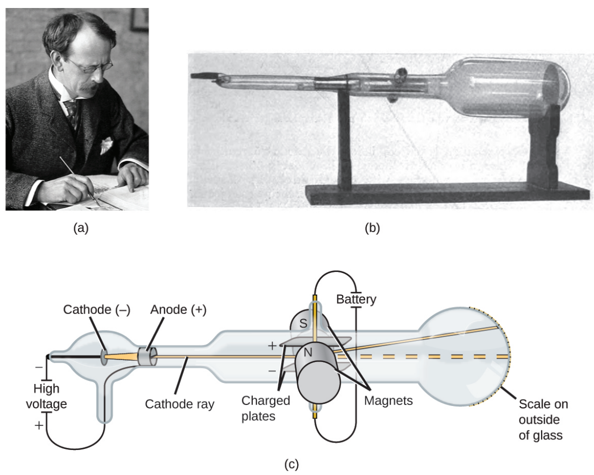


Figure 5.3a: Thomson cathode ray experiments: (a) J. J. Thomson produced a visible beam in a cathode ray tube. (b) This is an early cathode ray tube, invented in 1897 by Ferdinand Braun. (c) In the cathode ray, the beam (shown in yellow) comes from the cathode and is accelerated past the anode toward a fluorescent scale at the end of the tube. Simultaneous deflections by applied electric and magnetic fields permitted Thomson to calculate the mass-to-charge ratio of the particles composing the cathode ray. (credit a: modification of work by Nobel Foundation, PD; credit b: modification of work by Eugen Nesper, PD; credit c: modification of work by Kurzon, PD).

Based on his observations, here is what Thomson proposed and why:

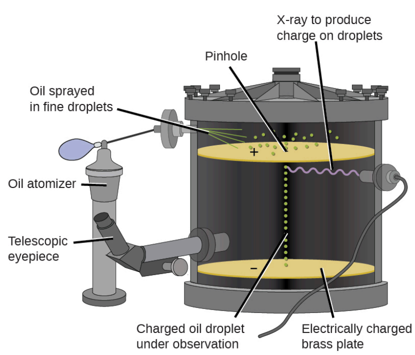
- The particles are attracted by positive (+) charges and repelled by negative (−) charges, so the particles of the cathode ray must be negatively charged since like charges repel and unlike charges attract.
- they are less massive than atoms and indistinguishable, regardless of the source material, so they must be fundamental, subatomic constituents of all atoms.

Although controversial at the time, Thomson's idea was gradually accepted, and his cathode ray particle is what we now call an **electron**, a negatively charged, subatomic particle with a mass more than one thousand-times less that of an atom. The term "electron" was initially coined in 1891 by Irish physicist George Stoney, meaning "electric ion." Stoney originally recognized the atom must have a unit of electricity and charge associated with the atom, but he had no experimental proof. It was J.J Thomson who took this theory and orchestrated scientific experiments using the cathode rays to prove atoms contained charged particles – the fundamental unit of charge – the electron.

J.J. Thomson Talks About the Size of the Electron

Listen (https://history.aip.org/exhibits/electron/jjsound.htm) to Thomson describe his discovery in his own voice.

In 1909, more information about the electron was uncovered by American physicist Robert A. Millikan via his "oil drop" experiments. Millikan created microscopic oil droplets, which could be electrically charged by friction as they formed or by using X-rays. These droplets initially fell due to gravity, but their downward progress could be slowed or even reversed by an electric field lower in the apparatus. By adjusting the electric field strength and making careful measurements and appropriate calculations, Millikan was able to determine the charge on individual drops (Figure 5.3b).



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

Figure 5.3b: Millikan's oil drop experiment: Millikan's experiment measured the charge of individual oil drops. The tabulated data are examples of a few possible values (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Looking at the charge data that Millikan gathered, you may have recognized that the charge of an oil droplet is always a multiple of a specific charge, 1.6×10^{-19} C. The symbol C – the coulomb – is the unit of electric charge in the International System of Units (SI). Millikan concluded that this electric charge value must therefore be a fundamental charge – the charge of a single electron – with his measured charges due to an excess of one electron (1 times 1.6×10^{-19} C), two electrons (2 times 1.6×10^{-19} C), three electrons (3 times 1.6×10^{-19} C), and so on, on a given oil droplet. Since the charge of an electron was now known due to

Millikan's research, and the charge-to-mass ratio was already known due to Thomson's research (1.759×10^{11} C/kg), it only required a simple calculation to determine the mass of the electron as well.

$$\text{Mass of electron} = 1.602 \times 10^{-19}\text{C} \times \frac{1\text{kg}}{1.759 \times 10^{11}\text{C}} = 9.107 \times 10^{-31}\text{kg}$$

A summary of the important concepts discovered about electrons include:

- They are negatively charged subatomic particles
- Mass of the electron is found to be 9.110 x 10^{-28} g or 9.110 x 10^{-31} kg
- The charge on the electron is -1.602 x 10^{-19} coulombs
  - when considering the chemical behaviour of subatomic particles in an atom and its ability to form ions, it is customary to consider these particles as having a relative charge: An electron carries a negative charge (-1) and is represented by the symbol, e^{-}.

Scientists had now established that the atom was not indivisible as Dalton had believed, but consisted of smaller subatomic charged particles. Due to the work of Thomson, Millikan, and others, the charge and mass of the negative, subatomic particles—the electrons—were known. However, the positively charged part of an atom was not yet well understood. It is suggested that the proton was observed by Eugen Goldstein in 1886 when he used anode rays of a hydrogen ion. However, it was the later experiments of J.J. Thomson and Ernest Rutherford that uncovered the nature of the positively charged proton. In 1904, Thomson proposed the "plum pudding" model of atoms, which described a positively charged mass with an equal amount of negative charge in the form of electrons embedded in it, since all atoms are electrically neutral. A competing model had been proposed in 1903 by Hantaro Nagaoka, who postulated a Saturn-like atom, consisting of a positively charged sphere surrounded by a halo of electrons (Figure 5.3c)

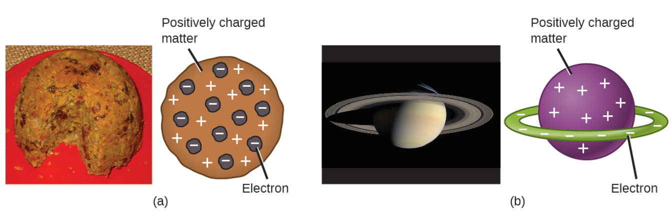


Figure 5.3c: Early models of the atom and its subatomic parts. (a) Thomson suggested that atoms resembled plum pudding, an English dessert consisting of moist cake with embedded raisins ("plums"). (b) Nagaoka proposed that atoms resembled the planet Saturn, with a ring of electrons surrounding a positive "planet." (credit a: modification of work by Man vyi, PD; credit b: modification of work courtesy of NASA/ JPL-Caltech, JPL Image Policy; in *Chemistry (OpenStax)*, CC BY 4.0).

Revelation of the Proton

The next major development in understanding the atom came from Ernest Rutherford, a physicist from New Zealand who largely spent his scientific career in Canada and England. He performed a series of experiments using a beam of high-speed, positively charged **alpha particles (α particles)** that were produced by the radioactive decay of radium; α particles consist of two protons and two neutrons (you will learn more about radioactive decay if you study nuclear chemistry). Rutherford and his colleagues Hans Geiger (later famous for the Geiger counter) and Ernest Marsden aimed a beam of α particles, the source of which was embedded in a lead block to absorb most of the radiation, at a very thin piece of gold foil and examined the resultant scattering of the α particles using a luminescent screen that glowed briefly where hit by an α particle.

What did they discover? Most particles passed right through the foil without being deflected at all. However, some were diverted slightly, and a very small number were deflected almost straight back toward the source (Figure 5.3d). Rutherford described finding these results: "It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you" [1]

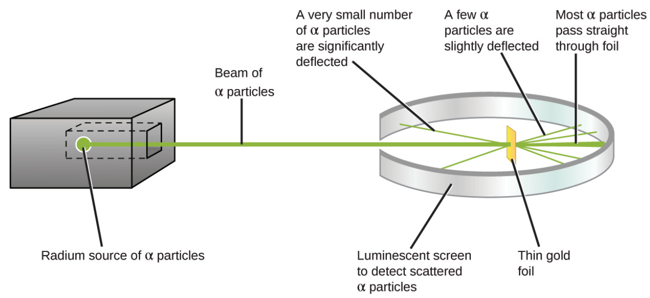


Figure 5.3d: Rutherford gold foil experiment schematic: Geiger and Rutherford fired α particles at a piece of gold foil and detected where those particles went, as shown in this schematic diagram of their experiment. Most of the particles passed straight through the foil, but a few were deflected slightly and a very small number were significantly deflected (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Here is what Rutherford deduced: Because most of the fast-moving α particles passed through the gold atoms undeflected, they must have traveled through essentially empty space inside the atom. Alpha particles are positively charged, so deflections arose when they encountered another positive charge. Since like charges repel one another, the few positively charged α particles that changed paths abruptly must have hit, or closely approached, another body that also had a highly concentrated, positive charge. Since the deflections occurred a small fraction of the time, this charge only occupied a small amount of the space in the gold foil. Analyzing a series of such experiments in detail, Rutherford drew two conclusions:

1. The volume occupied by an atom must consist of a large amount of empty space.
2. A small, relatively heavy, positively charged body, termed the **nucleus**, must be at the centre of each atom.

The Rutherford Experiment

This simulation (https://micro.magnet.fsu.edu/electromag/java/rutherford/) of the Rutherford gold foil experiment allows you to adjust the slit width to produce a narrower or broader beam of α particles to see how that affects the scattering pattern.

This analysis led Rutherford to propose a model in which an atom consists of a very small, positively charged nucleus, in which most of the mass of the atom is concentrated, surrounded by the negatively charged electrons, so that the atom is electrically neutral (Figure 5.3e). After many more experiments, Rutherford also discovered that the nuclei of other elements contain the hydrogen nucleus as a "building block," and he named this more fundamental particle the **proton**, the positively charged, subatomic particle found in the nucleus. With one addition, which you will learn next, this nuclear model of the atom, proposed over a century ago, is still used today.

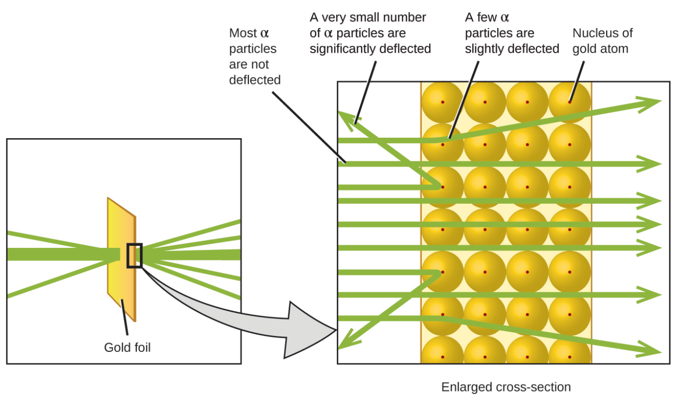


Figure 5.3e: Rutherford gold foil experiment – microscopic visual. The α particles are deflected only when they collide with or pass close to the much heavier, positively charged gold nucleus. Because the nucleus is very small compared to the size of an atom, very few α particles are deflected. Most pass through the relatively large region occupied by electrons, which are too light to deflect the rapidly moving particles (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Exercise 5.3a

Investigate the differences between a "plum pudding" atom and a Rutherford atom by firing α particles at each type of atom using the following PhET simulation: Rutherford

Scattering (https://phet.colorado.edu/sims/html/rutherford-scattering/latest/rutherford-scattering_en.html)

A summary of the important concepts discovered about protons include:

- Protons are positively charged subatomic particles
- Mass of the proton is found to be 1.673 x 10^{-24} g, which is 1,836 times the mass of an electron.
- Protons are found in the central part of an atom called the nucleus.
- The charge on the proton is +1.602 x 10^{-19} coulombs
    - When considering the chemical behaviour of subatomic particles in an atom and its ability to form ions, it is customary to consider these particles as having a relative charge: A proton carries a positive charge (+1) and is represented by the symbol p^+. As noted earlier, an electron carries a negative charge (-1) and is represented by the symbol, e^-. In a neutral atom, the number of protons equals the number of electrons, which means atoms of a given element are neutral.

The Neutron

One puzzle remained with regard to the subatomic particles of an atom. The nucleus of an atom was known to contain almost all of the mass of an atom, but the number of protons was only providing half, or less, of that atom's mass. Therefore, there must be some other type of subatomic matter present in the nucleus that had yet to be discovered. Different proposals were made to explain what constituted the remaining mass, while still maintaining the neutral charge of the atom – the existence of neutral particles in the nucleus. As you might expect, detecting uncharged particles is very challenging, and it was not until 1932 that James Chadwick found evidence of **neutrons**, uncharged, subatomic particles with a mass approximately the same as that of protons, 1.67493×10^{-24} g. Neutrons, having no relative electric charge and are represented by the symbol n^0. The existence of the neutron also explained isotopes, alternative forms of a given element. Isotopes of a given element differ in mass because they have different numbers of neutrons, but they are chemically identical because they have the same number of protons. The concept of isotopes will be explained in more detail later in this chapter.

Watch Subatomic Particles Explained in Under 4 Minutes (https://youtu.be/eD7hXLRqWWM?si=Pf9QrorY8Wl9bHVE) (3 mins 39 s)

One or more interactive elements has been excluded from this version of the text. You can view them online here: https://ecampusontario.pressbooks.pub/enhancedchemistry/?p=197#oembed-1

Exercise 5.3b

Check Your Learning Exercise (Text Version)
Review the scientist name list below. Match each of the seven scientists with their key discovery by filling in the [BLANK] with the correct scientist's name.

Scientist Name List (includes 7 names):

Democritus, Aristotle, John Dalton, J.J Thomson, Robert Millikan, Ernst Rutherford, James Chadwick

QUESTIONS:

1. The scientist who concluded that matter is composed of tiny, indivisible atoms that combine, separate, and rearrange in whole number ratios to form new matter is [BLANK].
2. [BLANK] is the scientist who postulated the nuclear model of the atom; the nuclear atom is mostly empty space with nearly all of its mass concentrated in the tiny central positively charged nucleus, which is surrounded by negatively charged electrons.
3. The scientist who hypothesized that all matter was composed of small, finite particles that they called atomos, meaning "indivisible" is [BLANK].
4. [BLANK] identified the mass of an electron to be 9.109×10^-31 kilograms
5. [BLANK] suggested a philosophical concept of matter such that it consisted of four elements – fire, earth, air, and water – and could be infinitely divided
6. This scientist, [BLANK], discovered the neutron, a subatomic particle with no charge and is in the tiny nucleus of an atom.
7. [BLANK] performed experiments with cathode ray tubes and discovered that all atoms contain tiny negatively charged subatomic particles and called them electrons. This scientist proposed the plum pudding model of the atom, which had a uniform sphere of positive charge with negatively charged electrons embedded within the sphere.

Check Your Answer[2]

Attribution & References

Notes

1. Ernest Rutherford, "The Development of the Theory of Atomic Structure," ed. J. A. Ratcliffe, in Background to Modern Science, eds. Joseph Needham and Walter Pagel, (Cambridge, UK: Cambridge University Press, 1938), 61–74. Accessed September 22, 2014, https://ia600508.us.archive.org/3/items/backgroundtomode032734mbp/backgroundtomode032734mbp.pdf. (p. 68).
2. (1) John Dalton; (2) Ernst Rutherford; (3) Democritus; (4) Robert Millikan; (5) Aristotle; (6) James Chadwick; (7) J.J Thomson

5.4 DEFINING THE NUCLEAR ATOM

Learning Objectives

By the end of this section, you will be able to:

- Summarize the structural characteristics of the nuclear atom
- Illustrate a simple model of the nuclear atom; locate its subatomic particles and their charges
- Explain why atoms have no overall charge
- Define ion and use the number of electrons lost or gained from an atom to calculate the overall charge of the ion and write the corresponding ion symbol.

The idea that matter is composed of tiny particles called atoms is at least 25 centuries old. It took until the twentieth century, however, for scientists to invent instruments that permitted them to probe inside an atom and find that it is not, as had been thought, hard and indivisible as Dalton theorized. Instead, experiments by Thomson, Rutherford, Chadwick, and other scientists revealed the atom is a complex structure composed of still smaller subatomic particles – electrons, protons, and neutrons.

Probing the Nuclear Atom

You have learned in the previous section that the first of these smaller particles were discovered by British physicist James (J. J.) Thomson in 1897. Named the *electron*, this particle is negatively charged. (It is the flow of these particles that produces currents of electricity, whether in lightning bolts or in the wires leading to your lamp.) Because an atom in its normal state is electrically neutral, each electron in an atom must be balanced by the same amount of positive charge.

The next step was to determine where in the atom the positive and negative charges were located. British physicist Ernest Rutherford devised the alpha-particle scattering, gold foil experiment that provided part of the answer to this question. The only way to account for the alpha particles that reversed direction when they hit the gold foil was to assume that nearly all of the mass, as well as all of the positive charge in each individual gold atom, is concentrated in a tiny centre or nucleus. When a positively charged alpha particle strikes a nucleus, it reverses direction, much as a cue ball reverses direction when it strikes another billiard ball. He termed this positive charge – a proton. Rutherford's model also placed the other type of charge—the negative electrons—in orbit around this nucleus.

Rutherford's model required that the electrons be in motion. Positive and negative charges attract each other, so stationary electrons would fall into the positive nucleus. Also, because both the electrons and the nucleus are extremely small, most of the atom is empty, which is why nearly all of Rutherford's particles were able to pass right through the gold foil without colliding with anything. Rutherford's model was a very successful explanation of the experiments he conducted, although eventually, scientists would discover that even the nucleus itself has structure.

Chadwick identified the neutron (n^0). It has neither a positive nor a negative charge, so it is considered neutral. It was determined that neutrons are also located in the nucleus (centre) of the atom with protons, and like protons, have a similar mass.

Collectively, these experimental observations, especially Rutherford's experiments, provided insight into the structure of the nuclear atom. The majority of the atom's structure is made up of empty space, with a centrally located, very concentrated nucleus. The nucleus contains positively charged protons and neutrally charged neutrons. Combined, these account for the majority of the mass in a given atom. The negative electrons, which contribute very little to the overall mass of the atom, are in orbit around the nucleus within the empty space.

The simplest possible atom (and the most common one in the sun and stars) is the element hydrogen (H). The nucleus of ordinary hydrogen contains a single proton. Moving around this proton is a single electron. Recall, the mass of an electron is nearly 2000 times smaller than the mass of a proton, and the electron carries an amount of charge exactly equal to that of the proton but opposite in sign (Figure 5.4a). Opposite charges attract each other, so it is an electromagnetic force that holds the proton and electron together. But what about the neutron(s)? The diagram below does not show hydrogen with any neutrons. In fact, hydrogen (the first element on the periodic table) has only one proton in the nucleus and one orbiting electron but does not contain any neutrons. Hydrogen actually has 3 naturally occurring forms that are all a little bit different from one another. The one shown below is the most abundant form of hydrogen in nature. These different forms of the same element are called **isotopes**. The concept of isotopes will be discussed in this section and in greater detail later in this chapter.

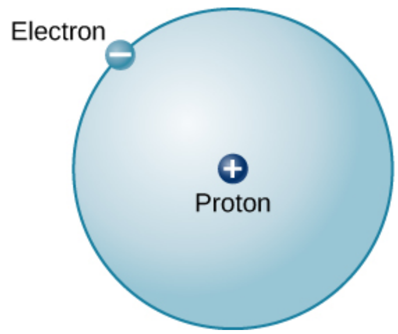


Figure 5.4a: Hydrogen-1 atom: this is a schematic diagram of a hydrogen atom. The proton and electron have equal but opposite charges, which exert an electromagnetic force that binds the hydrogen atom together. In the illustration, the size of the particles is exaggerated so that you can see them; they are not to scale. They are also shown much closer than they would actually be as it would take more than an entire page to show their actual distance to scale (credit: *Astronomy 2e (Open Stax)*, CC BY 4.0).

There are many other types of atoms in nature. Helium, for example, is the second-most abundant element in the Sun. Helium has two protons in its nucleus instead of the single proton that characterizes hydrogen. In addition, the helium nucleus contains two neutrons, particles with a mass comparable to that of the proton but with no electric charge. Moving around this nucleus are two electrons, so the total net charge of the helium atom is also zero (Figure 5.4b).

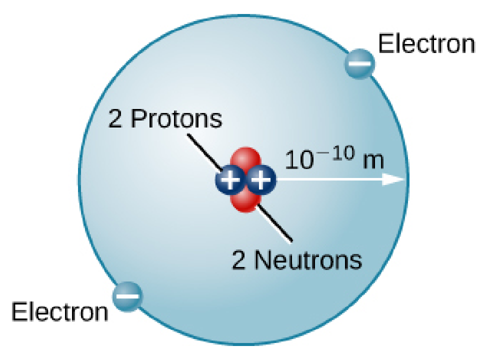


Figure 5.4b: Helium atom. Here we see a schematic diagram of a helium atom in its lowest energy state. Two protons are present in the nucleus of all helium atoms. In the most common variety of helium, the nucleus also contains two neutrons, which have nearly the same mass as the proton but carry no charge. Two electrons orbit the nucleus (credit: *Astronomy 2e (Open Stax)*, CC BY 4.0).

From this description of hydrogen and helium, perhaps you have guessed the pattern for building up all the elements (different types of atoms) that we find in the universe. The type of element is determined by the number of protons in the nucleus of the atom. Every element has a specific atomic number that equals the number of protons it contains; an element's atomic number can be sourced from the periodic table. For example, any atom with six protons is the element carbon, with eight protons is oxygen, with 26 is iron, and with 92 is uranium. On Earth, a typical atom has the same number of electrons as protons, and these electrons follow complex orbital patterns around the nucleus (which will be covered in more detail in chapter 10 – modern atomic theory).

Next, another model of the nuclear atom is shown in Figure 5.4c.

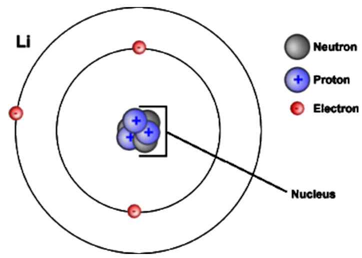


Figure 5.4c: Diagram of the atom: atoms contain protons (+) and neutrons, which are found in the nucleus (centre) of the atom. Atoms also contain electrons (-), which are found outside the nucleus. This is a model of a lithium atom. (credit: "Diagram of an Atom" from *Introduction to Biology* by Open Learning Initiative is licensed under CC BY-NC-SA 3.0).

Formation of Ions

The number of protons, neutrons, and electrons an atom contains differentiates one type of atom from the next. In any given atom of an element when the number of positively charged protons and the number of negatively charged electrons are the same, the atom is neutral. Interestingly, an atom can lose or gain electrons – this is what gives an element its chemical properties, including its ability to combine with other elements to form new matter. When an atom has more or less electrons than protons, it is electrically charged and is called an **ion**. Recall that the relative charge of one proton is +1 and one electron is -1. The amount of charge is dependent on the number of electrons lost or gained. Ions are written using the atomic symbol with its quantity of charge and charge sign (+ or -) in superscript. Note for any atoms having a +1 or -1 charge, the 1 is omitted when writing the ion symbol.

The charge of an atom is determined as follows:

$$\text{Atomic charge} = \text{number of protons} - \text{number of electrons}$$

As will be discussed in more detail later in this book, atoms (and molecules) typically acquire charge by gaining or losing electrons. An atom that gains one or more electrons will exhibit a negative charge and is called an **anion**. Positively charged atoms called **cations** are formed when an atom loses one or more electrons. For example, a neutral sodium atom (atomic number = 11) has 11 protons and 11 electrons. If this atom loses one electron, it will become a cation with a 1+ charge (11 – 10 = 1+); its ion symbol is Na^{+}. A neutral oxygen atom (atomic number = 8) has eight protons and 8 electrons, and if it gains two electrons it will become an anion with a 2– charge (8 – 10 = 2–); its ion symbol is O^{2-}.

In summary (also refer to Table 5.4a):

- Atoms that lose electrons will have less electrons than protons and form positive ions called cations.
- Atoms that gain electrons will have more electrons than protons and form negative ions called anions.

Table 5.4a: Summarizing Ion formation and Terminology

Starting Atom	Gain / Lose Electrons	More / Less Electrons than Protons	Type of Ion	Ion Classification	Determining Charge
Neutral Atom	loses electron(s)	LESS e^{-}s	forms positive ion	cation	= # electrons lost to form ion
Neutral Atom	gains electron(s)	MORE e^{-}s	forms negative ion	anion	= # electrons gained to form ion

Example 5.4a

Determining Ionic Symbols by Calculating Ionic Charge

A neutral lithium atom has 3 electrons and 3 protons. Determine the electric charge and write the ionic symbol for a lithium cation containing 2 electrons.

Solution

Lithium forms a CATION, a positive ion: The lithium ion has two electrons (2 e^{-}) and three protons (3 p^{+}), which means the neutral atom lost one electron (1 e^{-}).

CHARGE of Li ion = protons – electrons = 3 – 2 = +1; therefore, the ion has a net charge of +1.

ION SYMBOL for the lithium ion is Li^{+}

Example 5.4b

Determining Ionic Symbols by Calculating Ionic Charge

A neutral sulfur atom has 16 electrons and 16 protons. Determine the electric charge and write the ionic symbol for an sulfur anion containing 18 electrons.

Solution

Sulfur forms an ANION, a negative ion: The Sulfur ion has 18 electrons (18 e^-) and 16 protons (16 p^+), which means the neutral atom gained two electron (2 e^-).

CHARGE of S ion = protons – electrons = 16 – 18 = -2; therefore, the ion has an net charge of -2.

ION SYMBOL for the sulfur ion is S^{2-}

Example 5.4c

Determining Ionic Symbols by Calculating Ionic Charge

A neutral chlorine atom has 17 protons and 17 electrons. It gains one electron to form a chloride ion.

- Determine the electric charge on the ion and write the ionic symbol.
- Is the ion considered a cation or anion?
- How many total electrons does the chloride ion have?

Solution

A neutral chlorine atom gains one electron, so it will have a 1- charge; its symbol is Cl^-. Chlorine formed a negative ion, and is classified as an anion. Chloride ion has 18 electrons total (17 e^- + 1 e^- = 18 e^-)

Exercise 5.4a

Check Your Learning Exercise (Text Version)

From the options provided in brackets for each statement, decide which of the provided choices is correct.

1. Arsenic as a neutral atom contains 33 protons and (32 / 33 / 36) electrons. As an ion it can have 36 electrons. The charge on this ion would be (3^+ / $3^–$). This type of ion is a(n) (cation / anion). The ion symbol for the ion is (As^{3+} / As^{3-}).
2. Barium as a neutral atom contains 56 protons and (54 / 56 / 58) electrons. As an ion it has 54 electrons. The charge on this ion would be (2^+ / 1^+ / $1^–$ / $2^–$). This type of ion is a(n) (cation / anion). The ion symbol for the ion is (Ba^{2+} / Ba^{2-}).
3. Potassium as a neutral atom contains 19 protons and (18 / 19 / 20) electrons. Potassium loses one electron to form a(n) (cation / anion) that has a ($1^–$ / 1^+) charge. This ion will have total of (18 / 19 / 20) electrons. The ion symbol for the ion is (K^+ / P^+ / $K^–$ / $P^–$).

Check Your Answer[1]

Source: "Exercise 5.4a" by Jackie MacDonald is licensed under CC BY-NC-SA 4.0.

You will learn more about elements forming ions when reviewing concepts in chemical nomenclature, chemical reactions, and chemical bonding.

The Atomic Nucleus

The ratio of neutrons to protons increases as the number of protons increases, but each element is unique. The number of neutrons is not necessarily the same for all atoms of a given element. As aforementioned, most hydrogen atoms contain no neutrons at all. There are, however, hydrogen atoms that contain one proton and one neutron, and others that contain one proton and two neutrons. The various types of hydrogen nuclei with different numbers of neutrons are called isotopes of hydrogen (Figure 5.4d), and all other elements have isotopes as well. You can think of isotopes as siblings in the same element "family"—closely related but with different characteristics and behaviours. Turns out the element Hydrogen has three different isotopic forms, all containing one proton and one electron: Hydrogen – 1; Hydrogen – 2; and Hydrogen – 3, having 0, 1, and 2 neutrons in their atom nucleus, respectively. Hydrogen – 1 is the most abundant of the hydrogen isotopes and accounts for 99.9885% of all Hydrogen found in nature.

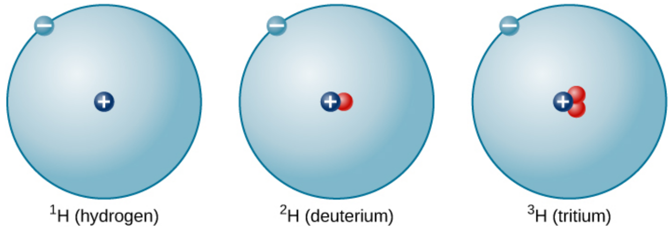


Figure 5.4d: Isotopes of Hydrogen: A single proton in the nucleus defines the atom to be hydrogen, but there may be zero, one, or two neutrons. The most common isotope of hydrogen is the one with only a single proton and no neutrons (credit: *Astronomy 2e (Open Stax),* CC BY 4.0).

Exercise 5.4b

Check Your Learning Exercise (Text Version)
Review the word list provided and choose a word or phase in the (BLANK) to make the statements correct. Each word will be used once.

WORD LIST: (11 words/phrases)

atomic number, protons, positively, electrons, neutron, isotopes, mass, empty space, ions, cations, anions

1. The nuclear model of the atom is made up of mostly (BLANK).
2. The atomic number tells you how many (BLANK) (and electrons) there are in a neutral atom.
3. The whole number found on the periodic table near an element symbol is called the (BLANK). It is unique for each element.
4. An atom with the same number of protons but different number of neutrons are called (BLANK).
5. The subatomic particle that has no charge is called a (BLANK).
6. The nucleus of the atom is (BLANK) charged. It contains protons and neutrons, which are the main subatomic particles that contribute to the overall (BLANK) of the atom.
7. Atoms can gain or lose (BLANK) to form positive or negative (BLANK).
8. A positive ion is called a(n) (BLANK); a negative ion is called a(n) (BLANK).

Check Your Answer[2]

Exercise 5.4c

Practice using the following PhET simulation: Build an Atom (https://phet.colorado.edu/sims/html/build-an-atom/latest/build-an-atom_en.html)

One or more interactive elements has been excluded from this version of the text. You can view them online here: https://ecampusontario.pressbooks.pub/enhancedchemistry/?p=203#iframe-phet-1

Links to Interactive Learning Tools

Explore Subatomic Particles (https://www.physicsclassroom.com/Concept-Builders/Chemistry/Subatomic-Particles) from the Physics Classroom. (https://www.physicsclassroom.com/)

Explore Subatomic Particle Terminology (https://h5pstudio.ecampusontario.ca/content/8510) from eCampusOntario H5P Studio (https://h5pstudio.ecampusontario.ca/).

Attribution & References

Except where otherwise noted, this page is adapted by Jackie MacDonald from "The structure of the atom " In *Astronomy 2e (Open Stax)*by Andrew Fraknoi, David Morrison, Sidney Wolff is licensed under CC BY 4.0. Access for free at *Astronomy 2e (Open Stax) (https://openstax.org/books/astronomy-2e/pages/1-introduction)* / Addition of paragraphs on atomic structure and ions created by Jackie MacDonald.

Notes

1. (1) Arsenic as a neutral atom contains 33 protons and 33 electrons. As an ion it can have 36 electrons. The charge on this ion would be 3^-. This type of ion is a(n) anion. The ion symbol for the ion is As^{3-}. (2) Barium as a neutral atom contains 56 protons and 56 electrons. As an ion it has 50 electrons. The charge on this ion would be 2^+. This type of ion is cation. The ion symbol for the ion is Ba^{2+}. (3) Potassium as a neutral atom contains 19 protons and 19 electrons.

Potassium loses one electron to form cation that has a 1^+ charge. This ion will have total of 18 electrons. The ion symbol for the ion is K^+.

2. (1) empty space; (2) protons; (3) atomic number; (4) isotopes; (5) neutron; (6) positively, mass; (7) electrons, ions; (8) cation, anion

5.5 ISOTOPES OF THE ELEMENTS

Learning Objectives

By the end of this section, you will be able to:

- Define isotopes and identify examples of isotopes for several elements
- Write and interpret symbols that depict the atomic number, mass number of isotopes
- Write isotope names using common naming methods

What is an Isotope?

Isotopes are various forms of the same element that have the same number of protons but a different number of neutrons. As the number of neutrons of an atom changes, so does its relative isotopic mass. The relative isotopic mass (also called mass number) is the sum of the protons and neutrons present in that isotope.

Mass Number (A) = Number of Protons + Number of Neutrons

Isotope symbols for elements are used to represent specific isotopes of atoms and include **mass number (A)** in superscript, **atomic number (Z)** in subscript, followed by the element symbol (X) in normal case (Figure 5.5a).

The number of protons in the nucleus of an atom is its atomic number (Z). This is the defining trait of an element. Its atomic number (Z) determines the identity of the atom. For example, any atom that contains six protons is the element carbon and has the atomic number 6, regardless of how many neutrons or electrons it may have. If you change the atomic number to 7, you are no longer dealing with carbon atoms, but nitrogen atoms. A neutral atom must contain the same number of positive and negative charges, so the number of protons equals the number of electrons. Therefore, the atomic number also indicates the number of electrons in an atom. The total number of protons and neutrons in an atom is called its mass number (A). The number of neutrons in that atom is therefore the difference between the mass number (A) and the atomic number (Z).

In summary:

$$\text{atomic number (Z)} = \text{number of protons}$$

$$\text{mass number (A)} = \text{number of protons} + \text{number of neutrons}$$

$$\text{A - Z} = \text{number of neutrons}$$

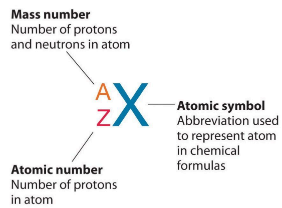


Figure 5.5a: Isotope Symbols of Elements: A symbol template is used to differentiate one isotope from another. The element symbol is written to identify the element. A represents the isotope's mass number and symbol Z represents the isotope's atomic number. All isotopes of an element have the same number of protons and electrons, which means they exhibit similar chemical properties. (credit**:** "1.6 Isotopes and Atomic Masses" In *Principles of General Chemistry (v. 1.0)* by Anonymous, CC BY-NC-SA 3.0./ Adapted by Jackie MacDonald.

Examples of Isotopes and their Properties:

As mentioned above, the symbol for a specific isotope of any element is written by placing the mass number as a superscript to the left of the element symbol. The atomic number is sometimes written as a subscript preceding the symbol, but since this number defines the element's identity (atomic number), as does its symbol, it is sometimes omitted, as shown in Figure 5.5b. The various isotopes for the element carbon and the number of each subatomic particle in that isotope are shown below:

- Carbon-12 (or ^{12}C) has the atomic number 6 and mass number 12 (six protons and six neutrons). It contains six protons, six neutrons, and six electrons
- Carbon-13 (or ^{13}C) has the atomic number 6 and mass number 13 (six protons and seven neutrons). It contains six protons, seven neutrons, and six electrons.
- Carbon-14 (or ^{14}C) has the atomic number 6 and mass number 14 (six protons and eight neutrons). It contains six protons, eight neutrons, and six electrons

When reading a specific isotope symbol, it is read as "element, mass number." For instance, in the case of magnesium, ^{24}Mg is read as "magnesium 24," and can be written as "magnesium-24" or "Mg-24." All magnesium atoms have the atomic number 12, which means they have 12 protons in their nucleus. These isotopes differ only because a ^{24}Mg atom has 12 neutrons in its nucleus, a ^{25}Mg atom has 13 neutrons, and a ^{26}Mg atom has 14 neutrons.

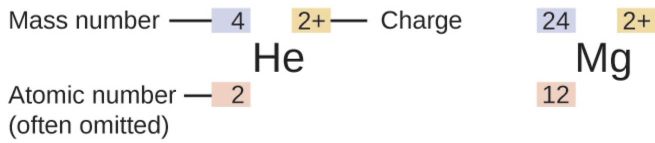


Figure 5.5b: Atomic Symbols used to Represent Isotopes of Elements: The symbol for an atom indicates the element via its usual two-letter symbol, the mass number as a left superscript, the atomic number as a left subscript (sometimes omitted), and the charge of its ion as a right superscript. Isotopes for Helium-4 and Magnesium-24 are shown in this figure (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Example 5.5a

Provide the isotope short form and isotope symbol (include the mass number only, omit atomic number) for the following isotopes: Arsenic-74, Calcium-44. Refer to the periodic table if needed.

Solution

Arsenic-74: Isotope short form = As-74; isotope symbol = ^{74}As

Calcium-44: Isotope short form = Ca-44; isotope symbol = ^{44}Ca

Example 5.5b

Provide the correct name for the isotope having the symbol ^{202}Hg; ^{56}Fe. Refer to the periodic table if needed.

Solution

^{202}Hg; Isotope name = Mercury-202

^{56}Fe; Isotope name = Iron – 56

Example 5.5c

Write the isotope name and symbol (mass number only) for the following two elements. Refer to the periodic table if needed.

- Element 1 has 12 protons and 13 neutrons
- Element 2 has 17 protons and 20 neutrons

Solution

Element 1 has 12 protons (therefore its atomic number is 12). Using a periodic table for reference, the element with atomic number 12 is Magnesium (Mg). The mass number will be 12 p^+ + 13 n^0 = 25.

Isotope Name = Magnesium-25 or Mg-25

Isotope Symbol = ^{25}Mg

Element 2 has 17 protons (therefore its atomic number is 17). Using a periodic table for reference, the element with atomic number 17 is chlorine (Cl). The mass number will be 17 p^+ + 20 n^0 = 37.

Isotope Name = Chlorine-37 or Cl-37

Isotope Symbol = ^{37}Cl

Exercise 5.5a

Check Your Learning Exercise (Text Version)
Choose the option that best answers the statements for each of the provided multiple choice questions.

1. Isotopes differ in the number of
   a. orbitals
   b. protons
   c. neutrons
   d. electrons
   e. charges

2. Hydrogen has three naturally occurring isotopes. Which of the following is not an isotope of hydrogen?
    a. Quatrium, ^{4}H, contains one electron, one proton, and three neutrons
    b. Tritium, ^{3}H, contains one electron, one proton, and two neutrons
    c. Deuterium, has symbol ^{2}H, contains one electron, one proton, and one neutron
    d. Protium, has symbol ^{1}H, contains one electron, one proton, and no neutrons
3. The isotope short form and isotope symbol (include the mass number only, omit atomic number) for Iodine-127 is
    a. I-127; ^{127}I
    b. I-127; $_{127}I$
    c. Io-127; ^{127}Io
    d. I-53; ^{53}I
    e. I-74; ^{74}I
4. Which of the following is the isotope symbol (include the mass number only, omit atomic number) for the atom that has 15 protons and 16 neutrons.
    a. ^{16}O
    b. ^{31}Ga
    c. ^{15}P
    d. ^{16}P
    e. ^{31}P
5. The isotope name and short form for a neutral atom that has with 7 electrons and 8 neutrons is
    a. nitrogen – 8; N-8
    b. nitrogen – 14; N-14
    c. nitrogen – 15; N-15
    d. oxygen – 14; O-14
    e. oxygen – 15; O-15
6. Tin is a silvery malleable metallic element belonging to group 14 of the periodic table. Tin has many stable naturally occurring isotopes, including tin–120, which has a natural abundance of 32.58%. Which of the following is true of the isotope tin-120?

a. As a neutral isotope, it contains 50 protons, 50 electrons, and 70 neutrons
b. Its mass number is 120
c. Its atomic number is 50
d. Its isotope symbol is 120Sn
e. all of these options contain true information about tin-120

7. The element X in an atom with a mass number (A) of 33 and atomic number (Z) 16 is

a. Arsenic
b. Indium
c. Sulfur
d. Chlorine
e. none of these elements represent this isotope

8. Which of the following is element X in an atom with a mass number of 58 and contains 30 neutrons?

a. Cobalt
b. Copper
c. Cerium
d. Nickel
e. Zinc

Check Your Answer[1]

While the atomic mass (mass number) of individual isotopes of a given element is different, their physical and chemical properties remain mostly the same. However, isotopes do differ in their stability. Carbon-12 (^{12}C) is the most abundant of the carbon isotopes, accounting for 98.93% of carbon on Earth. Carbon-13 (^{13}C) is stable but only accounts for about 1.07% abundance in nature. Carbon-14 (^{14}C) isotopes are found in trace amounts in nature and are unstable or radioactive. Radioactive isotopes may decay over time by emitting neutrons, protons, and/or electrons and energy to obtain a more stable form. Further information relating to radioactivity and radioactive decay would be covered in more detail in nuclear chemistry or nuclear physics.

Information about the naturally occurring isotopes of elements with atomic numbers 1 through 10 is given in Table 5.5a. Note that in addition to standard names and symbols, the isotopes of hydrogen are often referred to using common names and accompanying symbols. Hydrogen-2, symbolized ^{2}H, is also called deuterium and sometimes symbolized D. Hydrogen-3, symbolized ^{3}H, is also called tritium and sometimes symbolized T.

Table 5.5a: Nuclear Compositions of Atoms of the Very Light Elements (modified by Jackie MacDonald to include "mass number" column)

Isotope Name	Symbol	Atomic Number	Number of Protons	Number of Neutrons	Mass Number	Mass (amu)	% Natural Abundance
hydrogen-1 (protium)	${}^{1}_{1}H$	1	1	0	1	1.0078	99.989
hydrogen-2 (deuterium)	${}^{2}_{1}H$	1	1	1	2	2.0141	0.0115
hydrogen-3 (tritium)	${}^{3}_{1}H$	1	1	2	3	3.01605	— (trace)
helium-3	${}^{3}_{2}He$	2	2	1	3	3.01603	0.00013
helium-4	${}^{4}_{2}He$	2	2	2	4	4.0026	100
lithium-6	${}^{6}_{3}Li$	3	3	3	6	6.0151	7.59
lithium-7	${}^{7}_{3}Li$	3	3	4	7	7.0160	92.41
beryllium-9	${}^{9}_{4}Be$	4	4	5	9	9.0122	100
boron-10	${}^{10}_{5}B$	5	5	5	10	10.0129	19.9
boron-11	${}^{11}_{5}B$	5	5	6	11	11.0093	80.1
carbon-12	${}^{12}_{6}C$	6	6	6	12	12.0000	98.89
carbon-13	${}^{13}_{6}C$	6	6	7	13	13.0034	1.11
carbon-14	${}^{14}_{6}C$	6	6	8	14	14.0032	— (trace)
nitrogen-14	${}^{14}_{7}N$	7	7	7	14	14.0031	99.63
nitrogen-15	${}^{15}_{7}N$	7	7	8	15	15.0001	0.37
oxygen-16	${}^{16}_{8}O$	8	8	8	16	15.9949	99.757
oxygen-17	${}^{17}_{8}O$	8	8	9	17	16.9991	0.038
oxygen-18	${}^{18}_{8}O$	8	8	10	18	17.9992	0.205

Isotope Name	Symbol	Atomic Number	Number of Protons	Number of Neutrons	Mass Number	Mass (amu)	% Natural Abundance
fluorine-19	$^{19}_{9}F$	9	9	10	19	18.9984	100
neon-20	$^{20}_{10}Ne$	10	10	10	20	19.9924	90.48
neon-21	$^{21}_{10}Ne$	10	10	11	21	20.9938	0.27
neon-22	$^{22}_{10}Ne$	10	10	12	22	21.9914	9.25

In Exercise 5.5b, choose the ISOTOPES simulator option to explore different isotopes of the first 10 elements. Be sure to expand and review the atomic symbol and abundance in nature data when comparing isotopes. To create isotopes of a given element, choose your element, for example carbon, then add or remove neutrons. Observe what happens to the element symbol, the abundance in nature (%), and its stability.

Exercise 5.5b

Practice using the following PhET simulation: Isotopes and Atomic Mass (https://phet.colorado.edu/sims/html/isotopes-and-atomic-mass/latest/isotopes-and-atomic-mass_en.html)

Scientists in Action: Mildred Cohn, PhD.

As you will learn, isotopes are important in nature and especially in human understanding of science and medicine. Let's consider just one natural, stable isotope: Oxygen-18, which is noted in the Table 5.5a above and is referred to as one of the environmental isotopes. It is important in paleoclimatology, for example, because scientists can use the ratio between Oxygen-18 and Oxygen-16 in an ice core to determine the temperature of precipitation over time. Oxygen-18 (Figure 5.5c) was also critical to the discovery of metabolic pathways and the mechanisms of enzymes.

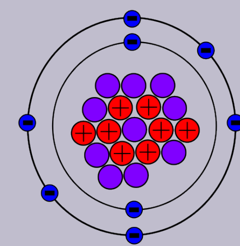


Figure 5.5c: Oxygen-18 Atom: The isotope, oxygen-18, includes 8 protons and 10 neutrons in the core nucleus, and 8 electrons around the nucleus. (credit: work by SM358, PD).

Mildred Cohn pioneered the usage of these isotopes to act as tracers, so that researchers could follow their path through reactions and gain a better understanding of what is happening. One of her first discoveries provided insight into the phosphorylation of glucose that takes place in mitochondria. And the methods of using isotopes for this research contributed to entire fields of study.

Learn more about Mildred Cohn's work and her motivating story of overcoming gender and religious biases in the article Mildren Cohn (1913-2009) [New Tab] (https://www.acs.org/education/whatischemistry/women-scientists/mildren-cohn.html).

WATCH Women in Chemistry: Mildred Cohn (18min 42sec) (https://youtu.be/8bV1Gu9Iy1Y).

Attribution & References

Notes

1. (1) c; (2) a; (3) a; (4) e; (5) c; (6) e; (7) c; (8)d

5.6 ATOMIC MASS

Learning Objectives

By the end of this section, you will be able to:

- Define the atomic mass unit and average atomic mass
- Calculate average atomic mass and isotopic abundance

The development of modern atomic theory revealed much about the inner structure of atoms. It was learned that an atom contains a very small nucleus composed of positively charged protons and uncharged neutrons, surrounded by a much larger volume of space containing negatively charged electrons. The nucleus contains the majority of an atom's mass because protons and neutrons are much heavier than electrons, whereas electrons occupy almost all of an atom's volume. The diameter of an atom is on the order of 10^{-10} m, whereas the diameter of the nucleus is roughly 10^{-15} m—about 100,000 times smaller. For a perspective about their relative sizes, consider this: If the nucleus were the size of a blueberry, the atom would be about the size of a football stadium (Figure 5.6a).

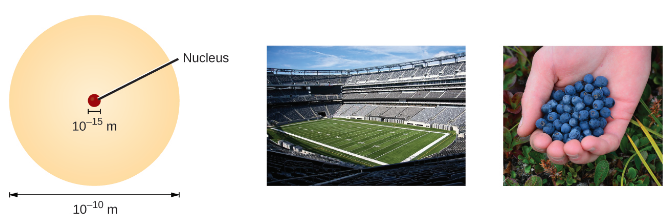


Figure 5.6a: Perspective about an atom's relative size: if an atom could be expanded to the size of a football stadium, the nucleus would be the size of a single blueberry. (credit middle: modification of work by babyknight, CC BY 2.0; credit right: modification of work by Paxson Woelber, CC BY 2.0)

Watch Just How Small is an Atom? (5:27 min) (https://www.youtube.com/watch?v=yQP4UJhNn0I)

Atoms—and the protons, neutrons, and electrons that compose them—are extremely small. For example, a carbon atom weighs less than 2×10^{-23} g, and an electron has a charge of less than 2×10^{-19} C (coulomb).

When describing the properties of tiny objects such as atoms, we use appropriately small units of measure, such as the **atomic mass unit (amu)** and the **fundamental unit of charge (e)**. The amu was originally defined based on hydrogen, the lightest element, then later in terms of oxygen. Since 1961, it has been defined with regard to the most abundant isotope of carbon, atoms of which are assigned masses of exactly 12 amu. (This isotope is known as "carbon-12"). Thus, one amu is exactly $\frac{1}{12}$ of the mass of one carbon-12 atom: 1 amu = 1.6605 × 10^{-24} g. (The **Dalton (Da)** and the **unified atomic mass unit (u)** are alternative units that are equivalent to the amu.) The fundamental unit of charge (also called the elementary charge) equals the magnitude of the charge of an electron (e) with e = 1.602 × 10^{-19} C.

A proton has a mass of 1.0073 amu and a charge of 1+. A neutron is a slightly heavier particle with a mass 1.0087 amu and a charge of zero; as its name suggests, it is neutral. The electron has a charge of 1− and is a much lighter particle with a mass of about 0.00055 amu (it would take about 1800 electrons to equal the mass of one proton). The properties of these fundamental particles are summarized in Table 5.6a. (An observant student might notice that the sum of an atom's subatomic particles does not equal the atom's actual mass. The total mass of six protons, six neutrons, and six electrons is 12.0993 amu, slightly larger than 12.00 amu. This "missing" mass is known as the mass defect, which you can learn about it if you study nuclear chemistry.)

Table 5.6a: 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}

Atomic Mass

Because each proton and each neutron contribute approximately one amu to the mass of an atom, and each electron contributes far less, the **atomic mass** of a single atom is approximately equal to its mass number (a whole number). However, the average masses of atoms of most elements are not whole numbers because most elements exist naturally as mixtures of two or more isotopes, each with their own slightly different masses due to the different number of neutrons they contain.

The mass of an element shown in a periodic table or listed in a table of atomic masses is a weighted, average mass of all the isotopes present in a naturally occurring sample of that element. This is equal to the sum of each individual isotope's mass multiplied by its fractional abundance.

$$\text{average mass} = \sum_i (\text{fractional abundance} \times \text{isotopic mass})_i$$

For example, the element boron is composed of two isotopes: About 19.9% of all boron atoms are ^{10}B with a mass of 10.0129 amu, and the remaining 80.1% are ^{11}B with a mass of 11.0093 amu. The average atomic mass for boron is calculated to be:

$$\begin{aligned} \text{boron average mass} &= (0.199 \times 10.0129 \text{ amu}) + (0.801 \times 11.0093 \text{ amu}) \\ &= 1.99 \text{ amu} + 8.82 \text{ amu} \\ &= 10.81 \text{ amu} \end{aligned}$$

It is important to understand that no single boron atom weighs exactly 10.8 amu; 10.8 amu is the average mass of all boron atoms, and individual boron atoms weigh either approximately 10 amu or 11 amu.

Exercise 5.6a

Calculating Average Atomic Mass

5.6a – Calculating Average Atomic Mass (Text version)

A meteorite found in central Indiana contains traces of the noble gas neon picked up from the solar wind during the meteorite's trip through the solar system. Analysis of a sample of the gas showed that it consisted of 91.84% ^{20}Ne (mass 19.9924 amu), 0.47% ^{21}Ne (mass 20.9940 amu), and 7.69% ^{22}Ne (mass 21.9914 amu). What is the average mass of the neon in the solar wind?

Solution

STEP 1 – List known quantities and identify what you are asked to find

Known information: Isotopes of neon – 91.84% ^{20}Ne (mass 19.9924 amu), 0.47% ^{21}Ne (mass 20.9940 amu), and 7.69% ^{22}Ne (mass 21.9914 amu).

What am I asked to Find: average mass of the neon in the solar wind

Step 2 – Determine how you will solve the problem

Use the following formula to solve problem:

$$\text{average mass} = \sum_i (\text{fractional abundance} \times \text{isotopic mass})_i$$

Step 3 – Solve the Problem

$$\begin{aligned} \text{average mass} &= (0.9184 \times 19.9924 \text{ amu}) + (0.0047 \times 20.9940 \text{ amu}) + (0.0769 \times 21.9914 \text{ amu}) \\ &= (18.36 + 0.099 + 1.69) \text{ amu} \\ &= 20.15 \text{ amu} \end{aligned}$$

The average mass of a neon atom in the solar wind is 20.15 amu. (The average mass of a terrestrial neon atom is 20.1796 amu. This result demonstrates that we may find slight differences in the natural abundance of isotopes, depending on their origin.)

Step 4 – Does the answer make sense?

- Yes it does. It is similar to the average atomic mass of Ne on the periodic table. The average atomic mass is in between the given amu values for Ne-20, Ne-21 and Ne-22, and the average amu value calculated is closest to the most abundant isotope, Ne-20.

Source: "Exercise 5.6a" by Jackie MacDonald, licensed under CC BY-SA 4.0.

If you want to review additional examples of how to calculate average atomic mass, **Watch How to Calculate Atomic Mass Practice Problems (https://www.youtube.com/watch?v=ULRsJYhQmlo) (6 mins 10s)**

Exercise 5.6b

Calculating Average Atomic Mass

A sample of magnesium is found to contain 78.70% of ^{24}Mg atoms (mass 23.98 amu), 10.13% of ^{25}Mg atoms (mass 24.99 amu), and 11.17% of ^{26}Mg atoms (mass 25.98 amu). Calculate the average mass of a Mg atom.

Check Your Answer[1]

We can also do variations of this type of calculation where you calculate the percent abundance of each isotope, as shown in the next example:

Example 5.6b

Calculating Percent Abundance

Naturally occurring chlorine consists of ^{35}Cl (mass 34.96885 amu) and ^{37}Cl (mass 36.96590 amu), with an average mass of 35.453 amu. What is the percent composition of Cl in terms of these two isotopes?

Solution

Step 1 – List known quantities and identify what you are asked to find

Known information: ^{35}Cl (mass 34.96885 amu) and ^{37}Cl (mass 36.96590 amu), with chlorine having an average atomic mass of 35.453 amu

What am I asked to Find: What is the percent composition of Cl in terms of these two isotopes

Step 2 – Determine how you will solve the problem

Use the following formula to solve problem:

The average mass of chlorine is the fraction that is ^{35}Cl times the mass of ^{35}Cl plus the fraction that is ^{37}Cl times the mass of ^{37}Cl.

$$\text{average mass} = (\text{fraction of } ^{35}\text{Cl} \times \text{ mass of } ^{35}\text{Cl}) + (\text{fraction of } ^{37}\text{Cl} \times \text{ mass of } ^{37}\text{Cl})$$

If we let *x* represent the fraction that is ^{35}Cl, then the fraction that is ^{37}Cl is represented by 1.00 − *x*.

(The fraction that is ^{35}Cl + the fraction that is ^{37}Cl must add up to 1, so the fraction of ^{37}Cl must equal 1.00 − the fraction of ^{35}Cl.)

Step 3 – Solve the Problem

Substituting this into the average mass equation, we have:

$$\begin{aligned} 35.453 \text{ amu} &= (x \times 34.96885 \text{ amu}) + [(1.00 - x) \times 36.96590 \text{ amu}] \\ 35.453 &= 34.96885x + 36.96590 - 36.96590x \\ 1.99705x &= 1.513 \\ x &= \frac{1.513}{1.99705} = 0.7576 \end{aligned}$$

So solving yields: *x* = 0.7576, which means that 1.00 − 0.7576 = 0.2424. Therefore, chlorine consists of 75.76% ^{35}Cl and 24.24% ^{37}Cl.

Step 4 – Does the answer make sense?

Yes it does. The highest % abundance found is ^{35}Cl ; which has a mass of 34.96885 amu is closest to the average mass of Cu, which is 35.453 amu.

If you want to watch the full work through of Example 5.6b – Calculation of Percent Abundance, **Watch**

How to Calculate Isotope Abundance (https://www.youtube.com/watch?v=6CfSyGd5Ry4) (11 mins 48s)

Exercise 5.6c

Calculation of Percent Abundance

Naturally occurring copper consists of ^{63}Cu (mass 62.9296 amu) and ^{65}Cu (mass 64.9278 amu), with an average mass of 63.546 amu. What is the percent composition of Cu in terms of these two isotopes?

Check Your Answer[2]

Try the interactive learning activity suggested below to reinforce your learning about isotope ratios and atomic mass. Try make mixtures of the main isotopes of the first 18 elements, gain experience with average atomic mass, and check naturally occurring isotope ratios using the Isotopes and Atomic Mass simulation.

Exercise 5.6d

Practice using the following PhET simulation: Isotopes and Atomic Mass

Determining Natural Abundances of Isotopes using Mass Spectrometry.

The occurrence and natural abundances of isotopes can be experimentally determined using an instrument called a mass spectrometer. Mass spectrometry (MS) is widely used in chemistry, forensics, medicine, environmental science, and many other fields to analyze and help identify the substances in a sample of material. In a typical mass spectrometer (Figure 5.6b), the sample is vaporized and exposed to a high-energy electron beam that causes the sample's atoms (or molecules) to become electrically charged, typically by losing one or more electrons. These cations then pass through a (variable) electric or magnetic field that deflects each cation's path to an extent that depends on both its mass and charge (similar to how the path of a large steel ball bearing rolling past a magnet is deflected to a lesser extent that that of a small steel BB). The ions are detected, and a plot of the relative number of ions generated versus their mass-to-charge ratios (a *mass spectrum*) is made. The height of each vertical feature or peak in a mass spectrum is proportional to the fraction of cations with the specified mass-to-charge ratio. Since its initial use during the development of modern atomic theory, MS has evolved to become a powerful tool for chemical analysis in a wide range of applications.

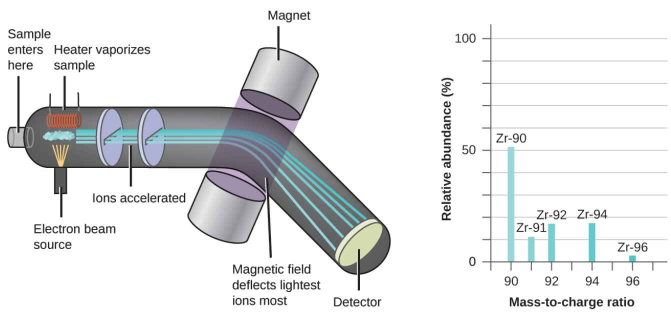


Figure 5.6b: A typical mass spectrometer: analysis of zirconium in a mass spectrometer produces a mass spectrum with peaks showing the different isotopes of Zr (credit: *Chemistry (OpenStax)*, CC BY 4.0).

Watch Mass Spectometry (8:19 min) (https://www.youtube.com/watch?v=mBT73Pesiog)

Watch Mass Spectrometry MS (7:58 min) (https://www.youtube.com/watch?v=J-wao0O0_qM)

Key Equations

- $\text{average mass} = \sum_i (\text{fractional abundance} \times \text{isotopic mass})_i$

Attribution & References

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Notes

1. Step 1 - List known quantities and identify what you are asked to find

Known information: 78.70% of ^{24}Mg atoms (mass 23.98 amu), 10.13% of ^{25}Mg atoms (mass 24.99 amu), and 11.17% of ^{26}Mg atoms (mass 25.98 amu)
What am I asked to Find: average mass of a Mg atom
Step 2 - Determine how you will solve the problem
Use the following formula to solve problem:

$$\text{average mass} = \sum_i (\text{fractional abundance} \times \text{isotopic mass})_i$$

Step 3 - Solve the Problem

$$\begin{aligned} \text{average mass} &= (0.7870 \times 23.98 \text{ amu}) + (0.1013 \times 24.99 \text{ amu}) + (0.1117 \times 25.98 \text{ amu}) \\ &= (18.872 + 2.531 + 2.902) \text{ amu} \\ &= 24.305 \text{ amu} \end{aligned}$$

Step 4 - Does the answer make sense?

- Yes it does. It is similar to the average atomic mass of Mg on the periodic table. The average atomic mass is in between the given amu values for Ne-20, Ne-21 and Ne-22, and the average amu value calculated is a bit more than the most abundant isotope, Mg-24.

2. Step 1 - List known quantities and identify what you are asked to find
Known information: ^{63}Cu (mass 62.9296 amu) and ^{65}Cu (mass 64.9278 amu), with an average mass of 63.546 amu.
What am I asked to Find:What is the percent composition of Cu in terms of these two isotopes?
Step 2 - Determine how you will solve the problem
Use the following formula to solve problem:
The average mass of Copper is the fraction that is ^{63}Cu times the mass of ^{63}Cu plus the fraction that is ^{65}Cu times the mass of ^{65}Cu

$$\text{average mass} = (\text{fraction of } ^{63}\text{Cu } \times \text{ mass of } ^{63}\text{Cu}) + (\text{fraction of } ^{65}\text{Cu } \times \text{ mass of } ^{65}\text{Cu})$$

If we let *x* represent the fraction that is ^{63}Cu, then the fraction that is ^{65}Cu is represented by 1.00 – *x*.
(The fraction that is ^{63}Cu + the fraction that is ^{65}Cu must add up to 1, so the fraction of ^{63}Cu must equal 1.00 – the fraction of ^{65}Cu.) Step 3 - Solve the Problem

$$\begin{aligned} 63.546 \text{ amu} &= (x \times 62.9296 \text{ amu}) + [(1.00 - x) \times 64.9278 \text{ amu}] \\ 63.546 &= 62.9296x + 64.9278 - 64.9278x \\ 1.9982x &= 1.3818 \\ x &= \frac{1.3818}{1.9982} = 0.6915 \end{aligned}$$

So solving yields: *x* = 0.6915, which means that 1.00 – 0.6915= 0.3085. Therefore, chlorine consists of 69.15% of Cu-63 and and 30.85% Cu-65
Step 4 - Does the answer make sense?
Yes it does. The highest % abundance found is Cu-63, which has a mass of 62.9296 amu is closest to the average mass of Cu, which is 63.546 amu.

CHAPTER 5 - SUMMARY

5.1 Early Atomic Theory: Dalton's Model of the Atom

The ancient Greeks proposed that matter consists of extremely small particles called atoms. Dalton postulated that matter is composed of exceedingly small, indivisible particles called atoms. He also stated an atom is the smallest unit of an element that can participate in a chemical change and that each element has a characteristic type of atom that differs in properties from atoms of all other elements, and that atoms of different elements can combine in fixed, small, whole-number ratios to form compounds. Samples of a particular compound all have the same elemental proportions by mass. When two elements form different compounds, a given mass of one element will combine with masses of the other element in a small, whole-number ratio. During any chemical change, atoms are neither created nor destroyed.

5.2 Electric Charge

Ancient Greek philosopher, Thales of Miletus (624–546 BCE), recorded that when amber (a hard, translucent, fossilized resin from extinct trees) was vigorously rubbed with a piece of fur, a force was created that caused the fur and the amber to be attracted to each other. The rubbed amber would not only attract the fur, and the fur attract the amber, but they both could affect other (nonmetallic) objects, even if not in contact with those objects.

The English physicist William Gilbert (1544–1603) completed similar experiments with similar findings. It was concluded there were two types of electric charge – positive and negative. These forces are repulsive when the same type of charge exists on two interacting objects. These forces are attractive when the charges are of opposite types. The force acts by contact or induction. The magnitude of the force decreases (rapidly) as the objects move further away from each other. Whereas, the magnitude of the force increases (rapidly) as the objects move closer to each other. Finally, not all matter is affected by this electric force.

5.3 Subatomic Particles of the Atom

Although no one has actually seen the inside of an atom, experiments have demonstrated much about atomic structure. Thomson's cathode ray tube showed that atoms contain small, negatively charged particles called electrons embedded in a positive atomic space. Millikan discovered that there is a fundamental electric charge—the charge of an electron. Rutherford's gold foil experiment showed that atoms have a small, dense, positively charged nucleus; the positively charged particles within the nucleus are called protons. Chadwick discovered that the nucleus also contains neutral particles called neutrons.

5.4 Defining the Nuclear Atom

Experimental observations provided insight into the structure of the nuclear atom. The majority of the atom's structure is made up of empty space, with a centrally located, very concentrated nucleus. The nucleus contains positively charged protons and neutrally charged neutrons. Combined, these two subatomic particles account for the majority of the mass in a given atom. The negative electrons, which contribute very little to the overall mass of the atom, are in orbit around the nucleus within the empty space. The type of element is determined by the number of protons in the nucleus of the atom. Every element has a specific atomic number that equals the number of protons it contains; an element's atomic number can sourced from the periodic table.

The number of protons, neutrons, and electrons an atom contains differentiates one type of atom from the next. In any given atom of an element when the number of positively charged protons and the number of negatively charged electrons are the same, the atom is neutral. An atom that loses or gains electrons is called an ion. An atom that gains one or more electrons will exhibit a negative charge and is called an anion. Positively charged atoms called cations are formed when an atom loses one or more electrons.

5.5 Isotopes of the Elements

Isotopes are various forms of the same element that have the same number of protons but a different number of neutrons. A neutral atom must contain the same number of positive and negative charges, so the number of protons equals the number of electrons. Therefore, the atomic number not only indicates the number of protons but also the number of electrons in an atom. The total number of protons and neutrons in an atom is called its mass number (A). Atomic symbols for elements are used to represent specific isotopes of atoms and include their mass number (A) in superscript, atomic number (Z) in subscript, followed by the element symbol.

In summary:

$$\begin{aligned} \text{atomic number (Z)} &= \text{number of protons} \\ \text{mass number (A)} &= \text{number of protons} + \text{number of neutrons} \\ \text{A - Z} &= \text{number of neutrons} \end{aligned}$$

When reading a specific isotope symbol, it is read as "element, mass number." For instance, in the case of magnesium, ^{24}Mg is read as "magnesium 24," and can be written as "magnesium-24" or "Mg-24."

5.6 Atomic Mass

An atom consists of a small, positively charged nucleus surrounded by electrons. The nucleus contains protons and neutrons; its diameter is about 100,000 times smaller than that of the atom. The mass of one atom is usually expressed in atomic mass units (amu), which is referred to as the atomic mass. An amu is defined as exactly $\frac{1}{12}$ of the mass of a carbon-12 atom and is equal to 1.6605×10^{-24} g.

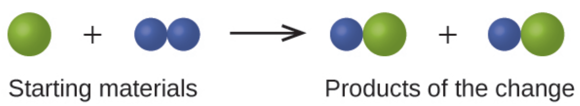
Protons are relatively heavy particles with a charge of 1+ and a mass of 1.0073 amu. Neutrons are relatively heavy particles with no charge and a mass of 1.0087 amu. Electrons are light particles with a charge of 1– and a mass of 0.00055 amu. The number of protons in the nucleus is called the atomic number (Z) and is the property that defines an atom's elemental identity. The sum of the numbers of protons and neutrons in the nucleus is called the mass number and, expressed in amu, is approximately equal to the mass of the atom.

Isotopes of an element are atoms with the same atomic number but different mass numbers; isotopes of an element, therefore, differ from each other only in the number of neutrons within the nucleus, so each isotope will have a slightly different atomic mass. When a naturally occurring element is composed of several isotopes, the atomic mass of the element represents the average of the masses of the isotopes involved, and this number is represented on a periodic table for each element. Average atomic masses and percent abundance for each element's isotopes can be calculated using formulas.

Attribution & References

CHAPTER 5 - REVIEW

5.1 Early Atomic Theory: Dalton's Model of the Atom

1. In the following drawing, the green, larger spheres represent atoms of a certain element. The blue, smaller spheres represent atoms of another element. If the spheres of different elements touch, they are part of a single unit of a compound. The following chemical change represented by these spheres may violate one of the ideas of Dalton's atomic theory. Which one?


Check Answer: [1]
2. Which postulate of Dalton's theory is consistent with the following observation concerning the weights of reactants and products? When 100 grams of solid calcium carbonate is heated, 44 grams of carbon dioxide and 56 grams of calcium oxide are produced.
Check Answer: [2]
3. Identify the postulate of Dalton's theory that is violated by the following observations: 59.95% of one sample of titanium dioxide is titanium; 60.10% of a different sample of titanium dioxide is titanium.
Check Answer: [3]
4. Samples of compound X, Y, and Z are analyzed, with results shown in the data table below.

Qualitative and Quantitative Analysis of Compounds X, Y, Z

Compound	Description	Mass of Carbon	Mass of Hydrogen
X	clear, colourless, liquid with strong odour	1.776 g	0.148 g
Y	clear, colourless, liquid with strong odour	1.974 g	0.329 g
Z	clear, colourless, liquid with strong odour	7.812 g	0.651 g

Do these data provide example(s) of the law of definite proportions, the law of multiple proportions, neither, or both? What do these data tell you about compounds X, Y, and Z? **Check Answer:** [4]
5. For the following questions, state whether the statement regarding Dalton's Atomic Theory is True or False. If it is false, provide the correct word or phrase for the **bolded** term to make that statement true.
   a. According to Dalton's Atomic Theory, matter consists of indivisible particles called **atoms**.
   b. Dalton's Atomic Theory postulates mass is neither created nor destroyed in a chemical reaction. This supports the **law of constant proportions**.
   c. According to Dalton's atomic theory, all atoms of a given element have identical **properties**

including identical **mass**.

d. Two of Dalton's postulates include (1) matter is composed of exceedingly small, indivisible particles and (2) elements consist of only one type of identical atom, which has the same mass for all atoms. These two theories have been proven to be **valid** theories by later scientists.
Check Answer: [5]

5.2 Electric Charge

1. For the following questions, state whether the statement regarding electric charge is True or False. If it is false, provide the correct word or phrase for the **bolded** term to make that statement true.
   a. Electric charges can be **positive** and negative.
   b. Electric force can be either be **equal** or repulsive.
   c. If two interacting objects carry opposite charges, it results in electrostatic **repulsion**.
   d. A balloon is rubbed on hair to gain charge. Then the comb is placed near running water from a tap. You observe that the stream of water is bending towards the comb. You conclude that the comb and water must have **similar** charges.
   Check Answer: [6]
2. Does the force of attraction increase, decrease, or stay the same as the distance decreases between two oppositely charged objects (as the objects near one another)?
   Check Answer: [7]
3. Does the force of attraction increase, decrease, or stay the same as the distance increases between two oppositely charged objects (as the objects get further from one another)?
   Check Answer: [8]
4. Explain what is meant by induction in reference to electric forces.
   Check Answer: [9]

5.3 Subatomic Particles of the Atom

1. How are electrons and protons similar? How are they different?
   Check Answer: [10]
2. How are protons and neutrons similar? How are they different?
3. **Check Answer:** [11]
4. Predict and test the behaviour of α particles fired at a "plum pudding" model atom.
   a. Predict the paths taken by α particles that are fired at atoms with a Thomson's plum pudding model structure. Explain why you expect the α particles to take these paths.
   b. If α particles of higher energy than those in (a) are fired at plum pudding atoms, predict how their

paths will differ from the lower-energy α particle paths. Explain your reasoning.

c. Now test your predictions from (a) and (b). Open the Rutherford Scattering simulation (http://phet.colorado.edu/sims/html/rutherford-scattering/latest/rutherford-scattering_en.html) and select the "Plum Pudding Atom" tab. Set "Alpha Particles Energy" to "min," and select "show traces." Click on the gun to start firing α particles. Does this match your prediction from (a)? If not, explain why the actual path would be that shown in the simulation. Hit the pause button, or "Reset All." Set "Alpha Particles Energy" to "max," and start firing α particles. Does this match your prediction from (b)? If not, explain the effect of increased energy on the actual paths as shown in the simulation.

Check Answer: [12]

5. Predict and test the behaviour of α particles fired at a Rutherford atom model.
    a. Predict the paths taken by α particles that are fired at atoms with a Rutherford atom model structure. Explain why you expect the α particles to take these paths.
    b. If α particles of higher energy than those in (a) are fired at Rutherford atoms, predict how their paths will differ from the lower-energy α particle paths. Explain your reasoning.
    c. Predict how the paths taken by the α particles will differ if they are fired at Rutherford atoms of elements other than gold. What factor do you expect to cause this difference in paths, and why?
    d. Now test your predictions from (a), (b), and (c). Open the Rutherford Scattering simulation (http://phet.colorado.edu/sims/html/rutherford-scattering/latest/rutherford-scattering_en.html) and select the "Rutherford Atom" tab. Due to the scale of the simulation, it is best to start with a small nucleus, so select "20" for both protons and neutrons, "min" for energy, show traces, and then start firing α particles. Does this match your prediction from (a)? If not, explain why the actual path would be that shown in the simulation. Pause or reset, set energy to "max," and start firing α particles. Does this match your prediction from (b)? If not, explain the effect of increased energy on the actual path as shown in the simulation. Pause or reset, select "40" for both protons and neutrons, "min" for energy, show traces, and fire away. Does this match your prediction from (c)? If not, explain why the actual path would be that shown in the simulation. Repeat this with larger numbers of protons and neutrons. What generalization can you make regarding the type of atom and effect on the path of α particles? Be clear and specific.

    Check Answer: [13]

5.4 Defining the Nuclear Atom

1. Describe the structure of the nuclear atom as proposed by Rutherford.

   Check Answer: [14]
2. Write the ion symbol for each of the following ions and state whether it is an anion or a cation.
    a. the ion with atomic number 55, 54 electrons, and mass number 133

   b. the ion with 54 electrons, 53 protons
   c. the ion with atomic number 15 and a 3– charge
   d. the ion with 24 electrons, and a 3+ charge (Hint – the neutral atom would have lost electrons to form this positive ion).
   Check Answer: [15]
3. Write the ion symbol for each of the following ions:
   a. the ion with a 3+ charge, 31 protons
   b. the ion with 36 electrons, 35 protons
   c. the ion with 86 electrons, and a 4+ charge
   d. the ion with a 2+ charge, atomic number 38
   Check Answer: [16]
4. Open the Build an Atom simulation (https://phet.colorado.edu/sims/html/build-an-atom/latest/build-an-atom_en.html) – Choose the ATOM simulation. Expand the net charge and mass number options by clicking on the + icon.
   a. Drag protons, neutrons, and electrons onto the atom template to make a neutral atom of Lithium that gives a mass number of 6. Write the element symbol for a neutral lithium atom.
   b. Now remove one electron to make an ion. What is the ion's net charge? Give the symbol for the ion you have created.
   Check Answer: [17]
5. Open the Build an Atom simulation (https://phet.colorado.edu/sims/html/build-an-atom/latest/build-an-atom_en.html) – Choose the ATOM simulation. Expand the net charge and mass number options by clicking on the + icon.
   a. Drag protons, neutrons, and electrons onto the atom template to make a neutral atom of Nitrogen that gives a mass number of 14. How many protons, neutrons, and electrons are there in this nitrogen atom?
   b. Write the element symbol for a neutral Nitrogen atom.
   c. Now add three electrons to make an ion. How many TOTAL electrons are now in this nitrogen atom? What is the ion's net charge? Give the symbol for the ion you have created.
   Check Answer: [18]

5.5 Isotopes of Elements

1. In what way are isotopes of a given element always different? In what way(s) are they always the same?
   Check Answer: [19]
2. Write the isotope symbol for each of the following neutral elements:
   a. atomic number 55, mass number 133
   b. 53 protons, 74 neutrons

c. 15 electrons, mass number 31
d. atomic number 27, 30 neutrons

Check Answer: [20]

3. Write the isotope symbol for each of the following neutral elements:
   a. 31 electrons, and a mass number of 71
   b. 35 protons, and 45 neutrons
   c. atomic number 90, 142 neutrons
   d. 38 protons, and mass number 87

   Check Answer: [21]

4. Open the Build an Atom simulation (https://phet.colorado.edu/sims/html/build-an-atom/latest/build-an-atom_en.html) and click on the Atom icon.
   a. Pick any one of the first 10 elements that you would like to build and state its symbol.
   b. Drag protons, neutrons, and electrons onto the atom template to make an atom of your element. State the numbers of protons, neutrons, and electrons in your atom, as well as the net charge and mass number.
   c. Click on "Net Charge" and "Mass Number," check your answers to (b), and correct, if needed.
   d. Predict whether your atom will be stable or unstable. State your reasoning.
   e. Check the "Stable/Unstable" box. Was your answer to (d) correct? If not, first predict what you can do to make a stable atom of your element, and then do it and see if it works. Explain your reasoning.

   Check Answer: [22]

5. Open the Build an Atom simulation (https://phet.colorado.edu/sims/html/build-an-atom/latest/build-an-atom_en.html)
   a. Drag protons, neutrons, and electrons onto the atom template to make a neutral atom of Oxygen-16 and give the isotope symbol for this atom.
   b. Now add two more electrons to make an ion and give the symbol for the ion you have created.

   Check Answer: [23]

6. Open the Build an Atom simulation (https://phet.colorado.edu/sims/html/build-an-atom/latest/build-an-atom_en.html)
   a. Drag protons, neutrons, and electrons onto the atom template to make a neutral atom, stable atom of Fluorine. How many protons, neutrons, and electrons does this stable isotope have? What is the mass number of this stable isotope? Give the isotope symbol for this atom.

   Check Answer: [24]

7. Determine the number of protons, neutrons, and electrons in the following isotopes that are used in medical diagnoses and name each element. Then, give the isotope symbol including atomic number, mass number and ion charge:
   a. atomic number 9, mass number 18, charge of 1–

b. atomic number 43, mass number 99, charge of 7+
c. atomic number 53, atomic mass number 131, charge of 1–
d. atomic number 81, atomic mass number 201, charge of 1+
e. Name the elements in parts (a), (b), (c), and (d).
Check Answer: [25]

8. The following are properties of isotopes of two elements that are essential in our diet. Determine the number of protons, neutrons and electrons in each and name the isotope.
   a. atomic number 26, mass number 58, charge of 2+
   b. atomic number 53, mass number 127, charge of 1–
   Check Answer: [26]
9. Give the number of protons, electrons, and neutrons in neutral atoms of each of the following isotopes. Name the element:
   a. $^{10}_{5}B$
   b. $^{199}_{80}Hg$
   c. $^{63}_{29}Cu$
   d. $^{13}_{6}C$
   e. $^{77}_{34}Se$
   Check Answer: [27]
10. Give the number of protons, electrons, and neutrons in neutral atoms of each of the following isotopes:
   a. $^{7}_{3}Li$
   b. $^{125}_{52}Te$
   c. $^{109}_{47}Ag$
   d. $^{15}_{7}N$
   e. $^{31}_{15}P$
   Check Answer: [28]
11. Click on the Isotopes and Atomic Mass Simulation (https://phet.colorado.edu/sims/html/isotopes-and-atomic-mass/latest/isotopes-and-atomic-mass_en.html) and select the "Mixtures" simulation option. Then under the isotope mixture option choose "My Mix". Hide the "Percent Composition" and "Average Atomic Mass" boxes, and then select the element boron (B).
   a. Write the symbols of the isotopes of boron that are provided (naturally occurring) in significant amounts.
   b. Predict the relative amounts (percentages) of these boron isotopes found in nature. % of isotope 1 vs % of isotope 2 of boron. Explain the reasoning behind your choice. (hint – reference your periodic table for an elements atomic mass).
   c. Add isotopes to the black box to make a mixture that matches your prediction in (b). You may drag isotopes from their bins or click on "More" and then move the sliders to the appropriate amounts.

d. Reveal the “Percent Composition” and “Average Atomic Mass” boxes. How well does your mixture match with your prediction? If necessary, adjust the isotope amounts to match your prediction.
e. Select “Nature’s” mix of isotopes and compare it to your prediction. How well does your prediction compare with the naturally occurring mixture? Explain. If necessary, adjust your amounts to make them match “Nature’s” amounts as closely as possible.
Check Answer: [29]

12. Repeat Chemistry End of Chapter Exercise 11 (a-e) using the PhET isotopes and atomic mass simulation (https://phet.colorado.edu/sims/html/isotopes-and-atomic-mass/latest/isotopes-and-atomic-mass_en.html) and choose the element Silicon, which has three naturally occurring isotopes.
Check Answer: [30]

5.6 Atomic Mass

1. An element has the following natural abundances and isotopic masses: 90.92% abundance with 19.99 amu, 0.26% abundance with 20.99 amu, and 8.82% abundance with 21.99 amu. Calculate the average atomic mass of this element. Write the name and element symbol for this element.
Check Answer: [31]
2. Average atomic masses listed by IUPAC are based on a study of experimental results. Bromine has two isotopes ^{79}Br and ^{81}Br, whose masses (78.9183 and 80.9163 amu) and abundances (50.69% and 49.31%) were determined in earlier experiments. Calculate the average atomic mass of bromine based on these experiments.
Check Answer: [32]
3. Variations in average atomic mass may be observed for elements obtained from different sources. Lithium provides an example of this. The isotopic composition of lithium from naturally occurring minerals is 7.5% ^{6}Li and 92.5% ^{7}Li, which have masses of 6.01512 amu and 7.01600 amu, respectively. A commercial source of lithium, recycled from a military source, was 3.75% ^{6}Li (and the rest ^{7}Li). Calculate the average atomic mass values for each of these two sources.
Check Answer: [33]
4. The average atomic masses of some elements may vary, depending upon the sources of their ores. Naturally occurring boron consists of two isotopes with accurately known masses (^{10}B, 10.0129 amu and ^{11}B, 11.0931 amu). The actual atomic mass of boron can vary from 10.807 to 10.819, depending on whether the mineral source is from Turkey or the United States. Calculate the percent abundances leading to the two values of the average atomic masses of boron from these two countries.
Check Answer: [34]
5. The ^{18}O:^{16}O abundance ratio in some meteorites is greater than that used to calculate the average atomic mass of oxygen on earth. Is the average mass of an oxygen atom in these meteorites greater than, less than, or equal to that of a terrestrial oxygen atom?

Check Answer: [35]

Attribution & References

Except where otherwise noted, this page is adapted by Jackie MacDonald from "2. 1 Early ideas in atomic theory (https://openstax.org/books/chemistry-2e/pages/2-1-early-ideas-in-atomic-theory)", "2.2 Evolution of Atomic Theory" and "2.3 Atomic Structure and Symbolism" In *Chemistry 2e (Open Stax)* by Paul Flowers, Klaus Theopold, Richard Langley & William R. Robinson is licensed under CC BY 4.0. Access for free at *Chemistry 2e (Open Stax) (https://openstax.org/details/books/chemistry-2e)*. / Additional questions created and added by Jackie MacDonald.

Notes

1. The starting materials consist of one green, larger sphere and two blue, smaller spheres. The products consist of two green spheres and two blue spheres. This violates Dalton's theory that that atoms are not created during a chemical change, but are merely redistributed.
2. Atoms are neither created nor destroyed during a chemical change
3. This statement violates Dalton's fourth postulate: In a given compound, the numbers of atoms of each type (and thus also the percentage) always have the same ratio.
4. X+Z have the same ratios of C and H and are therefore similar compounds, and is an example of the Law of Definite Proportions; however, they may or may not be the same compound. Samples from two unknown compounds that have the same mass ratio are not necessarily the same substance. Example: formaldehyde (CH_2O) and glyceraldehyde ($C_3H_6O_3$) have the same mass ratios but are not the same compound. X+Y and Y+Z have differing ratios of C and H and are therefore different compounds, aligning with the Law of Multiple Proportions. View an explanation of the solution in the video "Problem 2.1.3 (https://www.youtube.com/watch?v=FQU4DdD6Ldo&t=89s)"
5. (a) TRUE; (b) FALSE - Law of conservation of mass; (c) TRUE (d) FALSE - Invalid
6. (a) TRUE; (b) FALSE – attractive; (c)FALSE – attraction; (d) FALSE - opposite
7. The force of attraction **increases** as the distance decreases between two oppositely charged objects (as the objects near one another the force of attraction is greater / gets stronger).
8. The force of attraction **decreases** as the distance increases between two oppositely charged objects (as the objects get further from one another the attractive force between the two objects lessens / gets weaker)
9. Induction is a method used to charge an object without actually touching the object to any other charged object.
10. Protons reside in an atom's nucleus; whereas electrons reside outside the nucleus in the space around the nucleus. Electrons can freely move around nucleus of the atom. Electrons are a type of subatomic particle with a negative charge; whereas protons are a type of subatomic particle with a positive charge. Mass of the proton is found to be 1.673×10^{-24} g, which is 1,836 times the mass of an electron.
11. Both protons and neutrons are subatomic particles that reside in an atom's nucleus. Both have approximately the same mass. Protons are positively charged, whereas neutrons are uncharged.
12. See written answers below. For a video demonstration of this answer view "2.8 | Predict and test the behaviour of α particles fired at a "plum pudding" model atom. (a) Predict (https://youtu.be/e2A6I48ptnQ)" 3(a) This is a

prediction; your prediction may be different than shown here. Most of the α alpha particles should repel the atoms in the structure and would not directly go through since the positive charge in the plum pudding model was assumed to be spread out throughout the entire volume of the atom. They would be redirected and deflected in many areas.3(b) This is a prediction; your prediction may be different than shown here. α alpha particles of higher energy would be moving faster and reacting faster. They would still come in contact with the positive charges scattered throughout the atom and will be deflected faster in many directions due to higher energy state.3(c) Your answer may be YES or NO depending on the predictions you made. The predictions listed above did not match what happened in the simulation for question (a) and (b). In the simulation, the majority of the α alpha particles passed straight through the plum pudding model. This is because the positive charge in the plum pudding model was assumed to be spread out throughout the entire volume of the atom and so were the oppositely charged electrons. Therefore, the electric field from the positively charged "pudding" would be too weak to significantly affect the path of the relatively massive particles moving at slower (lower energy) or faster speeds (higher energy).

13. See written answers below. For a video demonstration of this answer view the video "2.9 | Predict and test the behaviour of α particles fired at a Rutherford atom model. (https://youtu.be/DJt819YdeU44)" (a) The Rutherford atom has a small, positively charged nucleus, so most α particles will pass through empty space far from the nucleus and be undeflected. Those α particles that pass near the nucleus will be deflected from their paths due to positive-positive repulsion. The more directly toward the nucleus the α particles are headed, the larger the deflection angle will be.4(b) Higher-energy α particles that pass near the nucleus will still undergo deflection, but the faster they travel, the less the expected angle of deflection.4(c) If the nucleus is smaller, the positive charge is smaller and the expected deflections are smaller—both in terms of how closely the α particles pass by the nucleus undeflected and the angle of deflection. If the nucleus is larger, the positive charge is larger and the expected deflections are larger—more α particles will be deflected, and the deflection angles will be larger.4(d) The paths followed by the α particles match the predictions from (a), (b), and (c).
14. The majority of the atom's structure is made up of empty space, with a centrally located, very concentrated nucleus. The three major subatomic particles are protons, neutrons, and electrons. The nucleus contains positively charged protons and neutrally charged neutrons. Combined, these account for most of the mass in a given atom. The negative electrons, which contribute very little to the overall mass of the atom, are in orbit around the nucleus within the empty space.
15. (a) Cs^{+} - cation; (b) I^{-} - anion; (c) P^{3-} - anion; (d) Co^{3+} - cation
16. (a) Ga^{3+} (b) Br^{-} (c) Th^{4+} (d) Sr^{2+}
17. (a) Symbol for neutral lithium atom (3 protons, 3 neutrons and 3 electrons) is Li; (b) By removing one electron lithium atom now has only 2 electrons but still 3 protons. The ion's net charge is +1. The symbol for the lithium ion is Li^{+}.
18. (a) There are 7 protons, 7 neutrons, and 7 electrons in the nitrogen atom; (b) N symbol for neutral atom is N; (c) TOTAL electrons = 7+3 = 10 electrons in nitrogen ion. Its net ionic charge is 3- or -3. Ion symbol is N^{3-}.
19. Isotopes of a given element always have different masses due to different numbers of neutrons. They always have the same number of protons, which determines the identity of the element.
20. (a) $^{133}_{55}\mathrm{Cs}$ (b) $^{127}_{53}\mathrm{I}$ (c) $^{31}_{15}\mathrm{P}$ (d) $^{57}_{27}\mathrm{Co}$
21. (a) $^{71}_{31}\mathrm{Ga}$ (b) $^{80}_{35}\mathrm{Br}$ (c) $^{232}_{90}\mathrm{Th}$ (d) $^{87}_{38}\mathrm{Sr}$
22. This is just one example of an answer the first ten elements - the element Nitrogen was chosen. You may have chosen a different atom and therefore would have different answers. (a) Nitrogen symbol is N; (b) We built an atom with 7 protons, 7 electrons and 7 neutrons; the net charge is 0; mass number is 14 (number of protons + number of neutrons); (c) our answers check correctly with the simulation. (c) We predict the nitrogen atom we build with be stable. The number of protons and neutrons are a similar number and are likely to be a stable isotope that occurs naturally. (d) when stable/stable is checked off, our element is STABLE; if it was unstable, you would add or take away neutrons until the simulation shows a stable isotope. Nitrogen-15 (has 8 neutrons) is also a stable isotope of nitrogen. However nitrogen atoms with 6 or less or 9 or more neutrons are unstable. As a result, they are not abundant in nature.

23. (a) Oxygen-16 isotope symbol is $^{16}_{8}\mathrm{O}$ (b) O^{2-}
24. A neutral, stable atom of fluorine will have 9 protons, 9 electrons, and 10 neutrons. Mass number is 19 (number of protons + number of neutrons); Isotope symbol for Fluorine-19 is $^{19}_{9}\mathrm{F}$
25. (a) 9 protons, 9 neutrons, 10 electrons in fluorine; $^{18}_{9}\mathrm{F}^{-}$ (b) 43 protons, 56 neutrons, 36 electrons in Technetium; $^{99}_{43}\mathrm{Tc}^{7+}$ (c) 53 protons, 78 neutrons, 54 electrons in Iodine $^{131}_{53}\mathrm{I}^{-}$ (d) 81 protons, 120 neutrons, 80 electrons in Thallium; $^{201}_{81}\mathrm{Tl}^{+}$
26. (a) 26 protons, 32 neutrons, 24 electrons; Iron-58 (b) 53 protons, 74 neutrons, 54 electrons; Iodine-127
27. (a) 5 protons, 5 electrons, 5 neutrons; Boron (b) 80 protons, 80 electrons, 119 neutrons; Mercury (c) 29 protons, 29 electrons, 34 neutrons; Copper (d) 6 protons, 6 electrons, 7 neutrons; Carbon (e) 34 protons, 34 electrons, 43 neutrons; Selenium
28. (a) 3 protons, 3 electrons, 4 neutrons; Lithium (b) 52 protons, 52 electrons, 73 neutrons; Tellurium (c) 47 protons, 47 electrons, 62 neutrons; Silver (d) 7 protons, 7 electrons, 8 neutrons; Nitrogen (e) 15 protons, 15 electrons, 16 neutrons; Phosphorus
29. (a) There are two isotopes, boron-10 and boron-11. (b) There is no way to be sure to accurately predict the abundances to give you an answer of 10.81 amu average atomic mass for boron, as shown on periodic table. However, we can see that the average of 10.81 amu is closest to 11 amu, so the natural abundances will contain more boron-11 isotopes than boron-10 isotopes. Let us guess that the abundances are 15% B-10, 85% B-11. The average mass would be (.15x10 + .85x11) 10.85 amu (you may have predicted different percentages - it is just a prediction). (e) Checking the nature's mix of isotopes for boron shows that the abundances are 19.9% B-10 and 80.1% for B-11, so our guessed amounts only had to be slightly adjusted.
30. Let us use neon as an example. Since there are three isotopes (Ne-20, Ne21, Ne-23), there is no way to be sure to accurately predict the abundances to make the total of 20.18 amu average atomic mass. Let us guess that the abundances are 9% Ne-22, 91% Ne-20, and only a trace of Ne-21. The average mass would be 20.18 amu. Checking the nature's mix of isotopes shows that the abundances are 90.48% Ne-20, 9.25% Ne-22, and 0.27% Ne-21, so our guessed amounts have to be slightly adjusted.
31. Average atomic mass of element = (0.9092 x 19.99) + (0.0026x20.99) + (0.0882 x 21.99) = 20.169 amu. This average atomic mass is similar to Neon (Ne).
32. Average atomic mass of Bromine = (0.5069 x 78.9183) + (0.4931 x 80.9163) = 79.904 amu
33.
34. Average atomic mass of naturally occurring lithium = (0.075x 6.01512) + (0.925 x 7.01600) = 6.941 amu; Average atomic mass of commercial sourced lithium = (0.0375x 6.01512) + (0.9625 x 7.01600) = 6.978 amu
35. Turkey source: Equation used to solve is 10.807 = (x)10.0129x + (1-x)(11.0931) ANSWER: 26.49% of B-10 and 73.51% of B-11. US Source: 10.819 = (x)10.0129x + (1-x)(11.0931) ANSWER: 25.37 of B-10 and 74.63% of B-11.
36. The average mass of an oxygen atom in these meteorites will be great than an oxygen atom from earth since the meteorites contain more O-18 and less O-16 isotopes. Whereas on earth, the majority of the oxygen isotopes on earth are O-16. For a video demonstration of this answer view the video "2.26 | The O-18:O-16 abundance ratio (https://youtu.be/sTOjQiVbX0o)"