1 Chemical Foundations

Chapter Objectives:

- Learn the development of the atomic theory.
- Understand the basic structure of the atom.

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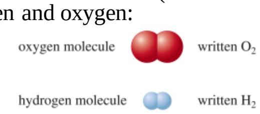
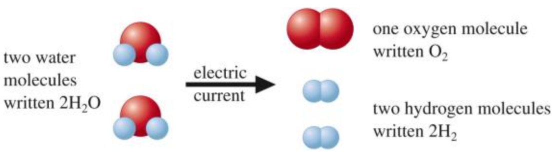
Chemistry: An Overview

- The tremendous variety of substances that we find around us result from combinations of about 100 different kinds of atoms.
- Water is composed of two types of atoms: hydrogen and oxygen, bound together (more later) to form a water molecule:

- When decomposed by an electric current, water is transformed into elemental (and diatomic) molecules of hydrogen and oxygen:



- This can be represented by a chemical equation:



- This illustrates two of the fundamental concepts of chemistry:
  – matter is composed of various types of atoms.
  – one substance changes to another by reorganizing the way the atoms are attached to each other.

> You only arrive at the right answer after making all possible mistakes. The mistakes began with the Greeks.
> Tony Rothman, *Instant Physics* (1995)

The Road to the Atomic Theory

> Nothing exists except atoms and empty space; everything else is opinion.
> Democritus

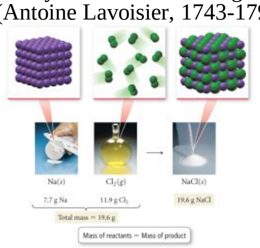
Atomos — *It's Greek to Me!*

- The ancient Greek philosopher Democritus (c. 460 - 370 BC) reasoned that if you cut a lump of matter into smaller and smaller pieces, you would eventually cut it down to a particle which could not be subdivided any further. He called these particles **atoms** (from the Greek *atomos*, "uncuttable")
- Aristotle (384-322 BC) believed that matter was continuous, and elaborated the idea that everything was composed of four elementary substances, assembled in varying proportions — earth, air, fire, and water, which possessed four properties — hot, dry, wet, and cold.
- The idea of atoms did not surface again until the 17^{th} and 18^{th} centuries.

PPT: **The Greek Periodic Table**

Law of Conservation of Mass

- In 1661, Robert Boyle redefined an **element** as a substance that cannot be chemically broken down further.
- **Law of Conservation of Mass** — Mass is neither created nor destroyed in chemical reactions (i.e., the total mass of a system does not change during a reaction). (Antoine Lavoisier, 1743-1794)



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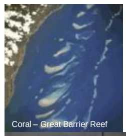
Law of Definite Proportions

- **Law of Definite Proportions** — All samples of a pure chemical substance, regardless of their source or how they were prepared, have the same proportions by mass of their constituent elements. (Joseph Proust, 1754-1826)



  – Calcium carbonate, which is found in coral, seashells, marble, limestone, chalk, and San Angelo tap water, is always 40.04% by mass calcium, 12.00% carbon, and 47.96% oxygen. (We now know that this results from the fact that calcium carbonate is $CaCO_3$.)

The Law of Multiple Proportions

- **Law of Multiple Proportions** — Elements can combine in different ways to form different substances, whose mass ratios are small whole-number multiples of each other. (John Dalton, 1804)

Compound	Sample Size	Mass of Sulfur	Mass of Oxygen
Sulfur oxide I	2.00 g	1.00 g	1.00 g
Sulfur oxide II	2.50 g	1.00 g	1.50 g

$$\frac{\text{mass of oxygen in sulfur oxide II per gram of sulfur}}{\text{mass of oxygen in sulfur oxide I per gram of sulfur}} = \frac{1.50\text{ g}}{1.00\text{ g}} = \frac{3}{2}$$

9

Dalton's Atomic Theory

- John Dalton (1766-1844) explained these observations in 1808 by proposing the *atomic theory*:
  - Each element consists of tiny indivisible (*not quite*) particles called **atoms**.
  - All atoms of the same element have the same mass (*not quite*), but atoms of different elements have different masses.
  - Atoms combine in simple, whole-number ratios to form compounds. A given compound always has the same relative numbers and types of atoms.
  - Atoms of one element cannot change into atoms of another element (*not quite*). In a chemical reaction, atoms change the way they are bound to other atoms, but the atoms themselves are unchanged.

10

Atomic Weights; Combining Gas Volumes

- Dalton prepared one of the earliest tables of atomic weights, later extended and corrected by Jons Jakob Berzelius (1779-1848).
- The key to determining absolute formulas for compounds came from the work on gases done by Joseph Gay-Lussac (1778-1850):

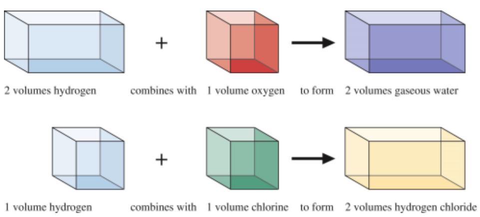


Figure 1.8

11

Avogadro's Hypothesis

- Amedeo Avogadro (1776-1856) explained these results by proposing that *at the same temperature and pressure, equal volumes of different gases contain the same number of particles* (Avogadro's hypothesis). This meant that water was formed by the reaction of diatomic molecules of hydrogen and oxygen, to form H_2O:

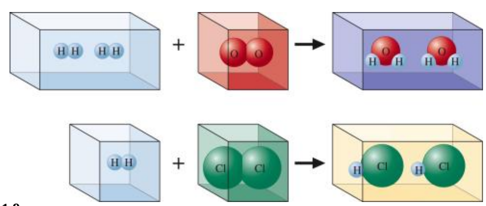


Figure 1.9

12

Dalton's Atomic Theory

- Dalton's atomic hypothesis had an uphill struggle — many scientists didn't like the idea of using small, invisible entities to explain phenomena.
- Most (but not all) chemists had accepted the existence of atoms by the early 20th century; however, many influential physicists did not accept the atomic theory until Einstein's landmark paper on Brownian motion (1905).

["On the Motion of Small Particles Suspended in a Stationary Liquid, as Required by the Molecular Kinetic Theory of Heat," *Annalen der Physik* in May 1905

- Dalton's original formulation of atoms as miniature billiard balls was incomplete: it did not explain *how* atoms combined to form compounds, or anything about their interior structure. The theory was modified greatly once charged particles coming from inside the atom (*radioactivity*) were discovered in the late 19th century.

The Electron

- In 1897, J. J. Thomson (1856-1940) investigated *cathode rays*, produced by passing an electric current through two electrodes in a vacuum tube (a *cathode ray tube*, CRT).
- The beam was produced at the negative electrode (cathode), and was deflected by the negative pole of an applied electrical field, implying that the rays were composed of negatively charged particles, with a very low mass. These particles were named **electrons**.

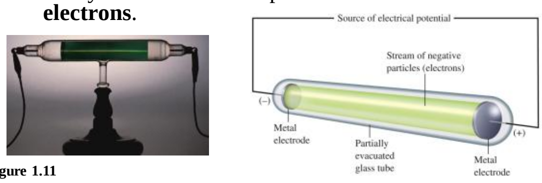
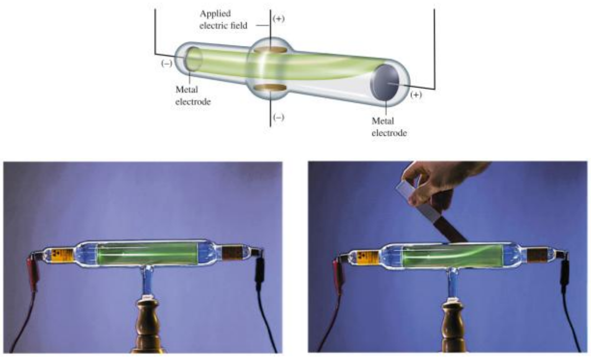


Figure 1.11

The Electron



sim. to Figure 1.12

The Electron

- Thomson's experiments showed that electrons were emitted by many different types of metals, so electrons must be present in all types of atoms.
- Although Thomson was unable to measure the mass of the electron directly, he was able to determine that the *charge-to-mass ratio*, *e*/*m*, was $-1.758820\ 10^8$ C/g.
  - This meant that the electron was about 2000 times lighter than hydrogen, the lightest element, and that atoms were *not* the smallest unit of matter.

The Mass of the Electron

- In 1909, Robert Millikan (1868-1953) measured the charge on the electron by observing the movement of tiny ionized droplets of oil passing between two electrically charged plates. Knowing the *e*/*m* ratio from Thomson's work, the mass of the electron could then be determined:

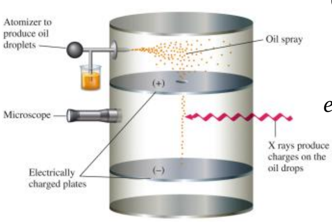


Figure 1.14

Charge of an electron:

$e = -1.60218\times10^{-19}$ C

Charge to mass ratio:

$e/m = -1.758820\times10^{8}$ C/g

Mass of an electron:

$m_e = 9.1093897\times10^{-28}$ g

Okay, Where's the Positive Charge?

- If there is negatively particle inside an electrically neutral atom, there must also be a positive charge.
- The model for the atom that Thomson proposed was of a diffuse, positively charged lump of matter with electrons embedded in it like "raisins in a plum pudding" (a watermelon or a blueberry muffin might be a more familiar analogy).

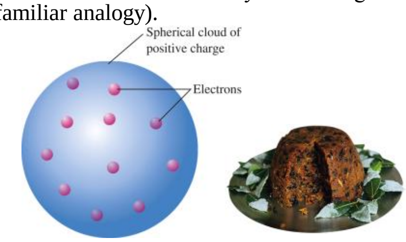


Figure 1.13

Radioactivity

- In the late 19th century, it was discovered that certain elements produce high-energy radiation.
  - In 1896, Henri Becquerel [Nobel Prize, 1903] found that uranium produces an image on a photographic plate in the absence of light.
  - Marie Curie [Nobel Prize, 1903 (Phys.) and 1911 (Chem.)] and Pierre Curie [Nobel Prize, 1903 (Phys.)] discovered radioactivity in thorium, and isolated previously unknown elements (radium, polonium) that were even more radioactive.
- There are three major types of radiation:
  - **alpha (α) particles** — consists of two protons and two neutrons (a helium nucleus), having a +2 charge and a mass 7300 times that of an electron.
  - **beta (β) particles** — a high-speed electron emitted from the nucleus of an atom (when a neutron turns into a proton).
  - **gamma (γ) rays** — high-energy electromagnetic radiation.

The Discovery of the Nucleus

- In 1910, Ernest Rutherford [Nobel Prize, 1908, Chem.] tested the "plum-pudding" model of the atom by firing a stream of alpha particles at a thin sheet of gold foil (about 2000 atoms thick).
- In the "plum-pudding" model, the mass of the atom is spread evenly through the volume of the atom. All of the alpha particles should sail right through the foil — but that's not what happened . . .

VIDEO

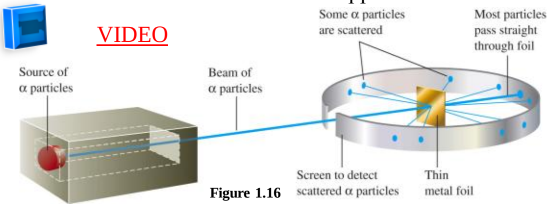


Figure 1.16

The Discovery of the Nucleus

- . . . instead, while most of the alpha-particles sailed through the gold foil, some were deflected at large angles, as if they had hit something massive, and some even bounced back toward the emitter.

> 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.
> — Ernest Rutherford, in E. N. da C. Andrade, *Rutherford and the Nature of the Atom* (1964)

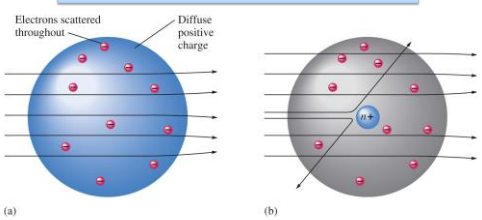


Figure 1.17

The Nuclear Atom Model

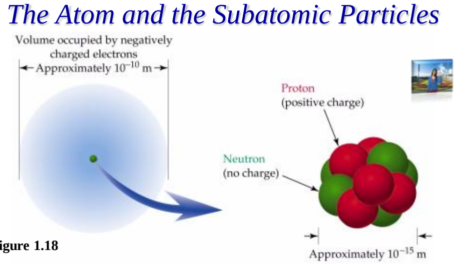
- Rutherford concluded that *all of the positive charge* and *most of the mass* (~99.9%) of the atom was concentrated in the center, called the **nucleus**. Most of the volume of the atom was empty space, through which the electrons were dispersed in some fashion.
- The positively charged particles within the nucleus are called **protons**; there must be one electron for each proton for an atom to be electrically neutral.
- This did not account for all of the mass of the atom, or the existence of *isotopes* (more later); the inventory of subatomic particles was "completed" (for the moment) by James Chadwick in 1932 [Nobel Prize, 1935], who discovered the **neutron**, an uncharged particle with about the same mass as the proton, which also resides in the nucleus.

The Modern View of Atomic Structure

The Atomic Theory Today

- An **atom** is an electrically neutral, spherical entity composed of a positively charged central **nucleus** surrounded by one or more negatively charged **electrons**.
- The nucleus contains the **protons**, which have positive charges, and **neutrons**, which are neutral. Neutrons are very slightly heavier than protons; protons are 1836 times heavier than electrons.
- The nucleus contains about 99.97% of the atom's mass, but occupies 1 ten-trillionth of the its volume.
- The **electrons** (e^-), which have negative charges, surround the nucleus, and account for most of the atomic volume. *The number of electrons equals the number of protons in the nucleus of a neutral atom.*

The Atom and the Subatomic Particles



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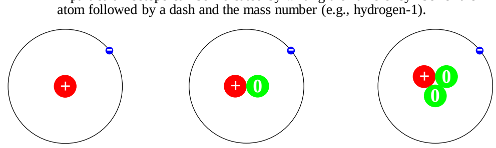
Particle	Mass in kilograms (kg)	Mass in atomic mass units (amu)	Charge in Coulombs (C)	Relative charge
Electron	9.109382×10^{-31} kg	5.485799×10^{-4} amu	-1.602176×10^{-19} C	-1
Proton	1.672622×10^{-27} kg	1.007276 amu	$+1.602176\times10^{-19}$ C	+1
Neutron	1.674927×10^{-27} kg	1.008665 amu	0	0

Atomic Number, Electrons

- What makes elements different from each another is the *number of protons* in their atoms, called the **atomic number** (**Z**). *All atoms of the same element contain the same number of protons.*
  - The number of protons determines the number of electrons in a neutral atom.
  - Since most of the volume of the atom is taken up by the electrons, when two atoms interact with each other, it is the outermost (valence) electrons that are making contact with each other.
  - The number and arrangement of the electrons in an atom determines its chemical properties. *Thus, the chemistry of an atom arises from its electrons.*

Mass Number, and Isotopes

- The **mass number** (A) is the sum of the number of protons (Z) and neutrons (N) in the nucleus of an atom: A = Z + N.
- **Isotopes** of an element have the same atomic number, but different #'s of neutrons. Isotopes of an element have nearly identical chemical behavior.
  - A particular isotope can be indicated by writing the name or symbol of the atom followed by a dash and the mass number (e.g., hydrogen-1).



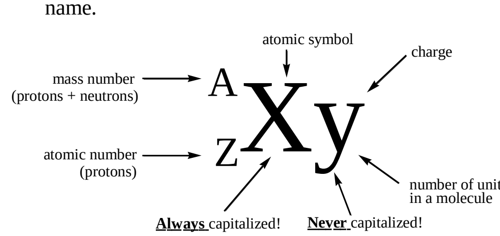
Hydrogen-1 — 1 proton, 0 neutrons — A = 1

Hydrogen-2 (deuterium) — 1 proton, 1 neutron — A = 2

Hydrogen-3 (tritium) — 1 proton, 2 neutrons — A = 3

Atomic Symbols

- The **atomic symbol** specifies information about the nuclear mass, atomic number, and charge on a particular element. Every element has a one- or two-letter symbol based on its English or Latin name.



CO ≠ Co !!!!

Ions

- Neutral atoms have the same number of electrons as protons. In many chemical reactions, atoms gain or lose electrons to form charged particles called **ions**.
- For example, sodium loses one electron, resulting in a particle with 11 protons and 10 electrons, having a +1 charge:

$$Na \rightarrow Na^{+} + e^{-}$$

- Positively charged ions are called **cations**.
- Fluorine gains one electron, resulting in a particle with 9 protons and 10 electrons, having a -1 charge:

$$F + e^{-} \rightarrow F^{-}$$

- Negatively charged ions are called **anions**.

Examples: Writing Element Symbols

1. Carbon-12 has how many protons? How many neutrons? How many electrons?

2. What would be the symbol for an element which has 14 protons and 15 neutrons?

3. What would be the symbol for an element which has 24 protons and 28 neutrons?

4. What would be the symbol for an element with 7 protons, 7 neutrons, and 10 electrons?

5. What would be the symbol for an element with 12 protons, 12 neutrons, and 10 electrons?

6. How many protons, neutrons, and electrons are there in ${}^{238}_{92}U$?

3 Chemical Foundations

Chapter Objectives (Sections 3.1-3.3):

- Learn how to find the atomic mass of an element.
- Learn how we can use the mole concept to count atoms and molecules by using their mass.
- Learn how to use the atomic mass of an element and the molecular weight of a compound to relate grams, moles, and the number of formula units.

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Atomic Mass and the Mole Concept

Counting By Weighing

- If the total mass of a sample of small objects is known, and the average mass of each small object is known, the number of objects in the sample can be determined.
- The same logic works for counting the number of atoms or molecules in a sample, but first we have to figure out how to weigh an atom.

Atomic Mass Units

- The mass of an atom is measured relative to the carbon-12 isotope, which is defined as weighing *exactly* 12 **atomic mass units** (**amu**, or **dalton**, **Da**).
  - 1 amu = 1 dalton = 1.660539×10^{-24} g.
  - Protons and neutrons each weigh about 1 amu.
  - Using carbon-12 as a reference allows the masses of other elements to be fairly close to whole numbers.
- The **isotopic mass** of a particular isotope is mass of one atom of that isotope measured in amu's. (Hydrogen-1 = 1.007825035 amu, hydrogen-2 = 2.014101779 amu.)

Atomic Masses

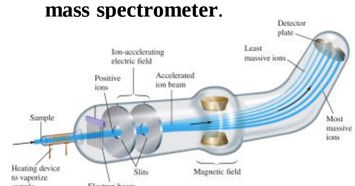
- When considering a sample of an element found in nature, we must take into account that the sample probably contains a number of different isotopes of the element. For instance, Hydrogen (H) is mostly ^{1}H (99.985%), but there is also a small percentage of ^{2}H present (deuterium, 0.015%).
- The **atomic mass** (or **atomic weight**) of an element is the average of the masses of each naturally-occurring isotopes of that element, weighted according to the isotopes' abundance.
  - This number is obtained by adding up the weights of all the naturally occurring isotopes multiplied by their relative abundances.

Atomic Masses

- For hydrogen,

0.99985	1.007825035 amu	=	1.0077 amu
0.00015	2.014101779 amu	=	0.00030 amu
			1.0080 amu

- This data can be obtained from a device called a **mass spectrometer**.



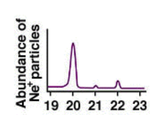


Figure 3.1

Examples: Calculating Atomic Masses

1. Use the following data to calculate the atomic mass of neon.

Isotope	Mass	Abundance
neon-20	19.992 amu	90.48%
neon-21	20.994 amu	0.27%
neon-22	21.991 amu	9.25%

```
Solution:
  0.9048   19.992 amu = 18.09 amu
  0.0027   20.994 amu =  0.057 amu
  0.0925   21.991 amu =  2.03 amu
                         20.177  amu
                         20.18   amu
```

7

The Mole Concept

- It is not possible to count the number of atoms or molecules involved in chemical reactions, since the molecules are so small, and so many are involved, even in a very small-scale reaction.
- Instead, it is necessary to measure amounts of molecules by using their mass.
- The relationship between sub-microscopic quantities like atoms and molecules, and macroscopic quantities like grams, is made using the **mole concept**.
- Using moles allows us to *count* particles by *weighing* them.

8

The Mole

- The **mole** (abbreviated **mol**) is the SI unit for *amount of substance*.
- A **mole** is defined as *the amount of a substance that contains the same number of entities as there are atoms in exactly 12 g of carbon-12*.
- 12 g of carbon-12 contains 6.022×10^{23} atoms. This number is known as **Avogadro's number, N_A**, in honor of Amedeo Avogadro (1776-1856, who first proposed the concept, and who also coined the word "molecule").

1 mole = 6.022×10^{23} units (*Avogadro's number, N_A*)

1 mol carbon-12	contains	6.022×10^{23} atoms
1 mol H_2O	contains	6.022×10^{23} molecules
1 mol NaCl	contains	6.022×10^{23} formula units

9

The Molar Mass of an Element

- The **molar mass of an element** is the mass of one mole of atoms of the element.

The value of an element's molar mass in grams per mole is numerically equal to the element's atomic mass in atomic mass units.

- 1 Fe atom has a mass of 55.847 amu.
  1 mole of Fe atoms has a mass of 55.847 grams.
- 1 O atom has a mass of 15.9994 amu.
  1 mole of O atoms has a mass of 15.9994 grams.
- 26.98 g Al = 1 mole Al = 6.022×10^{23} atoms Al
- 4.003 g He = 1 mole He = 6.022×10^{23} atoms He

10

The Molar Mass of a Compound

- The **molar mass of a compound** is obtained by adding together the atomic masses of all of the atoms in the molecule or formula unit. This number is either the mass of the compound in units of amu, or the mass of one mole of the compound in grams.
  - For molecular compounds, this is often referred to as the *molecular mass*, and for ionic compounds, it is sometimes referred to as the *formula mass*.
- To find the molar mass of water, H_2O:

Molar mass H_2O = (2×atomic mass H) + (1×atomic mass O)
= (2×1.00794) + (1×15.9994)
= 18.02 g/mol

- 1 H_2O molecule has a mass of 18.02 amu.
  1 mole of H_2O molecules has a mass of 18.02 grams.

11

Relating Moles, amu's and Grams

Molar mass of $Ca(NO_3)_2$ = (1 × Ca) + (2 × N) + (6 × O)
= (1 × 40.08) + (2 × 14.0067) + (6 × 15.9994)
= 164.09 g/mol

- 1 O_2 molecule has a mass of 32.00 amu
  1 mole of O_2 has a mass of 32.00g
- 1 NaCl formula unit has a mass of 58.44 amu
  1 mole of NaCl has a mass of 58.44 g
- 1 mole of $C_6H_{12}O_6$ = 180.16 g
- 1 mole of $Mg(C_2H_3O_2)_2$ = 83.35 g

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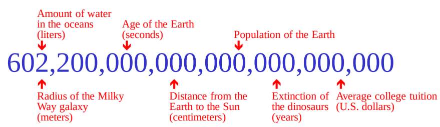
Just How Large is Avogadro's Number?

- A mole of marbles would cover the United States 70 miles deep.
- There are more atoms in a glass of water than there are glasses of water in the ocean.



- How much is a mole of water molecules?

Using Gram-Mole Conversions

- Thus, the molar mass (g/mol) is a conversion factor between numbers of moles and mass:
  - moles molar mass = mass in grams
  - grams molar mass = amount in moles
- and Avogadro's number (units/mol) is a conversion factor between numbers of units (molecules, atoms, or formula units) and moles:
  - moles N_A = number of units
  - number of units N_A = amount in moles

Examples: Gram-Mole Conversions

2. How many moles are present in 4.60 g of silicon?

Answer: 0.164 mol Si

Examples: Gram-Mole Conversions

3. How many g of Si are present in 9.0 mol of Si?

Answer: 250 g Si

Examples: Gram-Mole Conversions

4. How many atoms are in a sample of uranium with a mass of 1.000 μg?

Answer: 2.530×10^{15} atoms U

Examples: Gram-Mole Conversions

5. A pure silver ring contains 2.80×10^{22} silver atoms. How many grams of silver atoms does it contain?

Answer: 5.02 g Ag

Examples: Gram-Mole Conversions

6. How many moles of sucrose, $C_{12}H_{22}O_{11}$, are in a tablespoon of sugar that contains 2.85 g?

Answer: 0.00833 mol $C_{12}H_{22}O_{11}$

Examples: Gram-Mole Conversions

7. How many grams are in 0.0626 mol of $NaHCO_3$, the main ingredient in Alka-Seltzer tablets?

Answer: 5.26 g $NaHCO_3$

Examples: Gram-Mole Conversions

8. Isopentyl acetate, $C_7H_{14}O_2$, is the compound responsible for the scent of bananas. It is also released by bees when they sting. If a typical bee sting contains 1μg (1×10^{-6} g) of isopentyl acetate, how many molecules does this represent? How many atoms of carbon are present?

Answer: 5×10^{15} molecules $C_7H_{14}O_2$; 4×10^{16} atoms C

Examples: Gram-Mole Conversions

9. A sample of glucose, $C_6H_{12}O_6$, contains 1.52×10^{25} molecules. How many kilograms of glucose is this?

Answer: 4550 kg

The End