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Objectives

• **Explain** the *law of conservation of mass*, the *law of definite proportions*, and the *law of multiple proportions*.

• **Summarize** the five essential points of Dalton’s atomic theory.

• **Explain** the relationship between Dalton’s atomic theory and the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.
Foundations of Atomic Theory

• The transformation of a substance or substances into one or more new substances is known as a chemical reaction.

• Law of conservation of mass: mass is neither created nor destroyed during ordinary chemical reactions or physical changes.
Foundations of Atomic Theory, continued

- **Law of definite proportions**: a chemical compound contains the same elements in exactly the same proportions by mass regardless of the size of the sample or source of the compound.

- **Law of multiple proportions**: if two or more different compounds are composed of the same two elements, then the ratio of the masses of the second element combined with a certain mass of the first element is always a ratio of small whole numbers.
Law of Conservation of Mass

Section 1 The Atom: From Philosophical Idea to Scientific Theory

Hydrogen molecule: $3.348 \times 10^{-27}$ kg
Oxygen atom: $2.657 \times 10^{-26}$ kg
Water molecule: $2.992 \times 10^{-26}$ kg

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$

Sulfur atom: $5.325 \times 10^{-26}$ kg
Oxygen molecule: $5.314 \times 10^{-26}$ kg
Sulfur dioxide molecule: $1.064 \times 10^{-25}$ kg

$S + O_2 \rightarrow SO_2$

Phosphorus pentachloride molecule: $3.458 \times 10^{-25}$ kg
Phosphorus trichloride molecule: $2.280 \times 10^{-25}$ kg
Chlorine molecule: $1.177 \times 10^{-25}$ kg

$PCL_3 \rightarrow PCL_3 + Cl_2$
## Law of Multiple Proportions

<table>
<thead>
<tr>
<th>Name of compound</th>
<th>Description</th>
<th>As shown in figures</th>
<th>Formula</th>
<th>Mass O (g)</th>
<th>Mass N (g)</th>
<th>( \frac{\text{Mass O (g)}}{\text{Mass N (g)}} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nitrogen monoxide</td>
<td>colorless gas that reacts readily with oxygen</td>
<td></td>
<td>NO</td>
<td>16.00</td>
<td>14.01</td>
<td>( \frac{16.00 \text{ g O}}{14.01 \text{ g N}} = \frac{1.14 \text{ g O}}{1 \text{ g N}} )</td>
</tr>
<tr>
<td>Nitrogen dioxide</td>
<td>poisonous brown gas in smog</td>
<td></td>
<td>NO₂</td>
<td>32.00</td>
<td>14.01</td>
<td>( \frac{32.00 \text{ g O}}{14.01 \text{ g N}} = \frac{2.28 \text{ g O}}{1 \text{ g N}} )</td>
</tr>
</tbody>
</table>
Dalton’s Atomic Theory

• All matter is composed of extremely small particles called atoms.

• Atoms of a given element are identical in size, mass, and other properties; atoms of different elements differ in size, mass, and other properties.

• Atoms cannot be subdivided, created, or destroyed.
Dalton’s Atomic Theory, *continued*

- Atoms of different elements combine in simple whole-number ratios to form chemical compounds.

- In chemical reactions, atoms are combined, separated, or rearranged.
Modern Atomic Theory

• Not all aspects of Dalton’s atomic theory have proven to be correct. We now know that:
  • Atoms are divisible into even smaller particles.
  • A given element can have atoms with different masses.
• Some important concepts remain unchanged.
  • All matter is composed of atoms.
  • Atoms of any one element differ in properties from atoms of another element.
Objectives

• **Summarize** the observed properties of cathode rays that led to the discovery of the electron.

• **Summarize** the experiment carried out by Rutherford and his co-workers that led to the discovery of the nucleus.

• **List** the properties of protons, neutrons, and electrons.

• **Define** *atom*. 
The Structure of the Atom

• An atom is the smallest particle of an element that retains the chemical properties of that element.

• The nucleus is a very small region located at the center of an atom.

• The nucleus is made up of at least one positively charged particle called a proton and usually one or more neutral particles called neutrons.
The Structure of the Atom, continued

- Surrounding the nucleus is a region occupied by negatively charged particles called *electrons*.

- Protons, neutrons, and electrons are often referred to as *subatomic particles*.
# Properties of Subatomic Particles

<table>
<thead>
<tr>
<th>Particle</th>
<th>Symbols</th>
<th>Relative electric charge</th>
<th>Mass number</th>
<th>Relative mass (amu*)</th>
<th>Actual mass (kg)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Electron</td>
<td>$e^-, , _{0}^{1}e$</td>
<td>-1</td>
<td>0</td>
<td>0.000 5486</td>
<td>$9.109 \times 10^{-31}$</td>
</tr>
<tr>
<td>Proton</td>
<td>$p^{+}, , _{1}^{1}H$</td>
<td>+1</td>
<td>1</td>
<td>1.007 276</td>
<td>$1.673 \times 10^{-27}$</td>
</tr>
<tr>
<td>Neutron</td>
<td>$n^{0}, , _{0}^{1}n$</td>
<td>0</td>
<td>1</td>
<td>1.008 665</td>
<td>$1.675 \times 10^{-27}$</td>
</tr>
</tbody>
</table>

*1 amu (atomic mass unit) = $1.660 \, 540 \times 10^{-27}$ kg
Discovery of the Electron

Cathode Rays and Electrons

- Experiments in the late 1800s showed that cathode rays were composed of negatively charged particles.

- These particles were named *electrons*. 
Discovery of the Electron, continued

Charge and Mass of the Electron

- Joseph John Thomson’s cathode-ray tube experiments measured the charge-to-mass ratio of an electron.

- Robert A. Millikan’s oil drop experiment measured the charge of an electron.

- With this information, scientists were able to determine the mass of an electron.
Discovery of the Electron, *continued*
Discovery of the Atomic Nucleus

• More detail of the atom’s structure was provided in 1911 by Ernest Rutherford and his associates Hans Geiger and Ernest Marsden.

• The results of their gold foil experiment led to the discovery of a very densely packed bundle of matter with a positive electric charge.

• Rutherford called this positive bundle of matter the nucleus.
Gold Foil Experiment
Rutherford reasoned that each atom in the gold foil contained a small, dense, positively charged nucleus surrounded by electrons. A small number of the alpha particles directed toward the foil were deflected by the tiny nucleus (red arrows). Most of the particles passed through undisturbed (black arrows).
Composition of the Atomic Nucleus

• Except for the nucleus of the simplest type of hydrogen atom, all atomic nuclei are made of *protons* and *neutrons*.

• A proton has a positive charge equal in magnitude to the negative charge of an electron.

• Atoms are electrically neutral because they contain equal numbers of protons and electrons.

• A neutron is electrically neutral.
Composition of the Atomic Nucleus, *continued*

- The nuclei of atoms of different elements differ in their number of protons and therefore in the amount of positive charge they possess.

- Thus, the number of protons determines that atom’s identity.
Composition of the Atomic Nucleus, continued

**Forces in the Nucleus**

- When two protons are extremely close to each other, there is a strong attraction between them.
- A similar attraction exists when neutrons are very close to each other or when protons and neutrons are very close together.
- The short-range proton-neutron, proton-proton, and neutron-neutron forces that hold the nuclear particles together are referred to as **nuclear forces**.
Nuclear Forces

- Neutrons
- Protons

**Strong nuclear force**
(Acts on protons and neutrons alike)

**Electric repulsion**
(Acts on protons)
The Sizes of Atoms

• The radius of an atom is the distance from the center of the nucleus to the outer portion of its electron cloud.

• Because atomic radii are so small, they are expressed using a unit that is more convenient for the sizes of atoms.

• This unit is the *picometer, pm*. 
Objectives

- **Explain** what isotopes are.

- **Define** *atomic number* and *mass number*, and **describe** how they apply to isotopes.

- Given the identity of a nuclide, **determine** its number of protons, neutrons, and electrons.

- **Define** *mole*, *Avogadro’s number*, and *molar mass*, and state how all three are related.

- **Solve** problems involving mass in grams, amount in moles, and number of atoms of an element.
Atomic Number

- Atoms of different elements have different numbers of protons.

- Atoms of the same element all have the same number of protons.

- The **atomic number** \((Z)\) of an element is the number of protons of each atom of that element.
Atomic Number

Oxygen

15.9994

[He]2s²2p⁴

8 Protons

8 Neutrons

8 Electrons
Isotopes

- **Isotopes** are atoms of the same element that have different masses.

- The isotopes of a particular element all have the same number of protons and electrons but different numbers of neutrons.

- Most of the elements consist of mixtures of isotopes.
Mass Number

• The **mass number** is the total number of protons and neutrons that make up the nucleus of an isotope.
Mass Number

Oxygen

8 Protons
8 Neutrons
8 Electrons
Designating Isotopes

- **Hyphen notation**: The mass number is written with a hyphen after the name of the element.
  - uranium-235

- **Nuclear symbol**: The superscript indicates the mass number and the subscript indicates the atomic number.
  - $^{235}_{92}U$
Designating Isotopes, continued

- The number of neutrons is found by subtracting the atomic number from the mass number.

\[
\text{mass number} - \text{atomic number} = \text{number of neutrons}
\]

\[
235 \text{ (protons } + \text{ neutrons)} - 92 \text{ protons} = 143 \text{ neutrons}
\]

- **Nuclide** is a general term for a specific isotope of an element.
Isotopes and Nuclides

Carbon-12, Carbon-14

$^{12}_6\text{C}$  $^{14}_6\text{C}$

mass number  atomic number  Element Symbol
Designating Isotopes, continued

Sample Problem A

How many protons, electrons, and neutrons are there in an atom of chlorine-37?
Designating Isotopes, *continued*

**Sample Problem A Solution**

**Given:** name and mass number of chlorine-37

**Unknown:** numbers of protons, electrons, and neutrons

**Solution:**
atomic number = number of protons = number of electrons
mass number = number of neutrons + number of protons
Designating Isotopes, *continued*

Sample Problem A Solution, *continued*

mass number of chlorine-37 − atomic number of chlorine = number of neutrons in chlorine-37

mass number − atomic number = 37 (protons plus neutrons) − 17 protons = 20 neutrons

An atom of chlorine-37 is made up of 17 electrons, 17 protons, and 20 neutrons.
Relative Atomic Masses

- The standard used by scientists to compare units of atomic mass is the carbon-12 atom, which has been arbitrarily assigned a mass of exactly 12 atomic mass units, or 12 amu.

- One **atomic mass unit**, or 1 **amu**, is exactly 1/12 the mass of a carbon-12 atom.

- The atomic mass of any atom is determined by comparing it with the mass of the carbon-12 atom.
Average Atomic Masses of Elements

- **Average atomic mass** is the weighted average of the atomic masses of the naturally occurring isotopes of an element.

Calculating Average Atomic Mass

- The average atomic mass of an element depends on both the mass and the relative abundance of each of the element’s isotopes.
Average Atomic Masses of Elements, *continued*

**Calculating Average Atomic Mass, *continued***

- Copper consists of 69.15% copper-63, which has an atomic mass of 62.929 601 amu, and 30.85% copper-65, which has an atomic mass of 64.927 794 amu.

- The average atomic mass of copper can be calculated by multiplying the atomic mass of each isotope by its relative abundance (expressed in decimal form) and adding the results.
Average Atomic Masses of Elements, continued

Calculating Average Atomic Mass, continued

- \((0.6915 \times 62.929\,601\,\text{amu}) + (0.3085 \times 64.927\,794\,\text{amu}) = 63.55\,\text{amu}\)

- The calculated average atomic mass of naturally occurring copper is 63.55 amu.
Relating Mass to Numbers of Atoms

The Mole

- The mole is the SI unit for amount of substance.
- A **mole** (abbreviated mol) is the amount of a substance that contains as many particles as there are atoms in exactly 12 g of carbon-12.

Avogadro’s Number

- **Avogadro’s number**—$6.022 \ 1415 \times 10^{23}$—is the number of particles in exactly one mole of a pure substance.
Relating Mass to Numbers of Atoms, continued

Molar Mass

- The mass of one mole of a pure substance is called the molar mass of that substance.

- Molar mass is usually written in units of g/mol.

- The molar mass of an element is numerically equal to the atomic mass of the element in atomic mass units.
**Relating Mass to Numbers of Atoms, continued**

**Gram/Mole Conversions**

- Chemists use molar mass as a conversion factor in chemical calculations.

- For example, the molar mass of helium is 4.00 g He/mol He.

- To find how many grams of helium there are in two moles of helium, multiply by the molar mass.

\[
2.00 \text{ mol He} \times \frac{4.00 \text{ g He}}{1 \text{ mol He}} = 8.00 \text{ g He}
\]
Relating Mass to Numbers of Atoms, continued

Conversions with Avogadro’s Number

• Avogadro’s number can be used to find the number of atoms of an element from the amount in moles or to find the amount of an element in moles from the number of atoms.

• In these calculations, Avogadro’s number is expressed in units of atoms per mole.
Solving Mole Problems

\[\text{Amount of element in moles} = \frac{1\ \text{mol}}{\text{molar mass of element}} \times \frac{6.022 \times 10^{23}\ \text{atoms}}{1\ \text{mol}}\times \frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{atoms}} = \frac{\text{mass of element in grams}}{\text{molar mass of element}}\times \frac{6.022 \times 10^{23}\ \text{atoms}}{1\ \text{mol}}\times \frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{atoms}}\times \frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{atoms}} = \frac{\text{number of atoms of element}}{1\ \text{mol}}\times \frac{6.022 \times 10^{23}\ \text{atoms}}{1\ \text{mol}}\times \frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{atoms}}\]
Determining the Mass from the Amount in Moles

1. \( ? \text{ mol} \) to \( ? \text{ g} \)

2. \( ? \text{ g} \)

3. \( \frac{1 \text{ mol}}{? \text{ g}} \)

4. \( \frac{? \text{ g}}{1 \text{ mol}} \)

5. Use molar mass

6. Mass
Relating Mass to Numbers of Atoms, continued

Sample Problem B
What is the mass in grams of 3.50 mol of the element copper, Cu?
Relating Mass to Numbers of Atoms, continued

Sample Problem B Solution

**Given:** 3.50 mol Cu

**Unknown:** mass of Cu in grams

**Solution:** the mass of an element in grams can be calculated by multiplying the amount of the element in moles by the element’s molar mass.

\[
\text{moles Cu} \times \frac{\text{grams Cu}}{\text{moles Cu}} = \text{grams Cu}
\]
Relating Mass to Numbers of Atoms, continued

Sample Problem B Solution, continued

The molar mass of copper from the periodic table is rounded to 63.55 g/mol.

\[
3.50 \text{ mol Cu} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = 222 \text{ g Cu}
\]
Relating Mass to Numbers of Atoms, continued

Sample Problem C
A chemist produced 11.9 g of aluminum, Al. How many moles of aluminum were produced?
Relating Mass to Numbers of Atoms, continued

Sample Problem C Solution

**Given:** 11.9 g Al

**Unknown:** amount of Al in moles

**Solution:**

$$\text{grams Al} \times \frac{\text{moles Al}}{\text{grams Al}} = \text{moles Al}$$

The molar mass of aluminum from the periodic table is rounded to 26.98 g/mol.

$$11.9 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 0.441 \text{ mol Al}$$
Relating Mass to Numbers of Atoms, \textit{continued}

Sample Problem D

How many moles of silver, Ag, are in $3.01 \times 10^{23}$ atoms of silver?
Section 3 Counting Atoms

Relating Mass to Numbers of Atoms, continued

Sample Problem D Solution

**Given:** $3.01 \times 10^{23}$ atoms of Ag

**Unknown:** amount of Ag in moles

**Solution:**

$$\text{Ag atoms} \times \frac{\text{moles Ag}}{\text{Avogadro's number of Ag atoms}} = \text{moles Ag}$$

$$3.01 \times 10^{23} \text{ Ag atoms} \times \frac{1 \text{ mol Ag}}{6.022 \times 10^{23} \text{ Ag atoms}} = 0.500 \text{ mol Ag}$$
Relating Mass to Numbers of Atoms, continued

Sample Problem E

What is the mass in grams of $1.20 \times 10^8$ atoms of copper, Cu?
Relating Mass to Numbers of Atoms, continued

Sample Problem E Solution

**Given:** $1.20 \times 10^8$ atoms of Cu

**Unknown:** mass of Cu in grams

**Solution:**

$$\text{Cu atoms} \times \frac{\text{moles Cu}}{\text{Avogadro's number of Cu atoms}} \times \frac{\text{grams Cu}}{\text{moles Cu}} = \text{grams Cu}$$

The molar mass of copper from the periodic table is rounded to 63.55 g/mol.

$$1.20 \times 10^8 \text{ Cu atoms} \times \frac{1 \text{ mol Cu}}{6.022 \times 10^{23} \text{ Cu atoms}} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = 1.27 \times 10^{-14} \text{ g Cu}$$
End of Chapter 3 Show