Chapter 2: Atoms, Molecules, and Ions

Start Learning the Elemental Names and Symbols: Table 2.3 in Textbook (Table 2.3 Link Here)

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Early Ideas in Atomic Theory

An **element** is a substance that cannot be broken down into two or more simpler substances by any means. These are shown on the periodic table, with their symbols.

John Dalton (1803-1807) Dalton's Atomic Theory

1) All matter is composed of tiny particles called atoms. An **atom** is the smallest unit of an element that can participate in a chemical change.
2) All atoms of a given element have identical chemical properties that are characteristic of that element.
3) Atoms of one element differ in properties from atoms of all other elements.
4) A compound consists of atoms of 2 or more elements combined in a small, whole-number ratio.
5) Atoms can change how they are combined, but they are neither created nor destroyed in chemical reactions.

Law of Constant Composition/Law of Definite Proportion

All samples of a pure compound contain the same elements in the same proportion by mass (experiments of French chemist Joseph Proust)

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The Atomic Theory

Dalton used the data from Proust (who developed the Law of definite proportions) to formulate the Law of Multiple Proportions.

Law of Multiple Proportions (Dalton's Law):

- When two elements react, the mass of one element will react with masses of a second element so compounds always equals a ratio of small, whole number.

Compound A
32 grams O reacts with 24 grams C ⟶ ratio = $\frac{\text{g O}}{\text{g C}} = \frac{32}{24} = \frac{1.33\text{ g O}}{1\text{ g C}}$

Compound B
64 grams O reacts with 24 grams C ⟶ ratio = $\frac{\text{g O}}{\text{g C}} = \frac{64}{24} = \frac{2.66\text{ g O}}{1\text{ g C}}$

$$\frac{\text{Compound B}}{\text{Compound A}} = \frac{\left(\frac{2.66\text{ grams O}}{1\text{ gram C}}\right)}{\left(\frac{1.33\text{ grams O}}{1\text{ gram C}}\right)} = \frac{\left(\frac{2\text{ grams O}}{1\text{ gram C}}\right)}{\left(\frac{1\text{ gram O}}{1\text{ gram C}}\right)} = \frac{2}{1}$$

Therefore, compound B has 2 grams O for every 1-gram carbon; compound A has 1-gram oxygen for every 1-gram carbon. This means compound B has twice the amount of oxygen per carbon.

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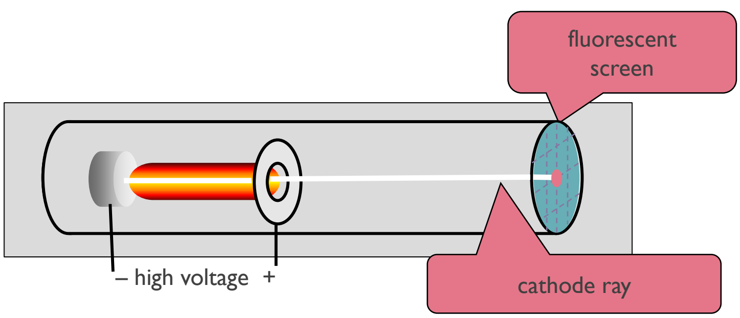
Subatomic Particles and Atomic Structure

In the late 1800's, many scientists were doing research involving **radiation**, the emission and transmission of energy in the form of waves.

They commonly used a **cathode ray tube**, which consists of two metal plates sealed inside a glass tube from which most of the air has been evacuated.



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Subatomic Particles and Atomic Structure

Researchers discovered that like charges repel each other, and opposite charges attract one another. J. J. Thomson (1856–1940) noted the rays were repelled by a plate bearing a negative charge and attracted to a plate bearing a positive charge.

Cathode Ray Tube Experiment Video
J.J. Thomson Talking about an electron

Thomson's contributions:

- He proposed the rays were a stream of negatively charged particles and these are called **electrons**.
- He determined the charge–to–mass ratio of electrons to be **1.76×10^8 C/g.**
- Thomson proposed the "plum pudding" model of an atom ("+" and "–" charges are squished together like a chocolate chip cookie).

Electric and magnetic fields deflect the beam.

- Gives charge/mass of e^- = 1.76×10^8 C/g
- Coulomb (C) = SI unit of charge
- 1 C = (A·s)

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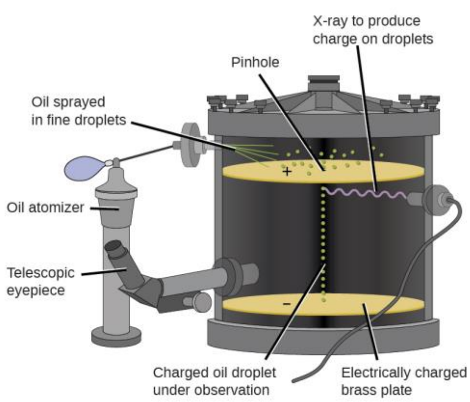
Subatomic Particles and Atomic Structure

Millikan (1911) studied electrically–charged oil drops.

R.A. Millikan (1868–1953) determined the charge on an electron by examining the motion of tiny oil drops.

The charge of an electron was determined to be **-1.6022×10^{-19} C**.



Oil Drop Experiment Video

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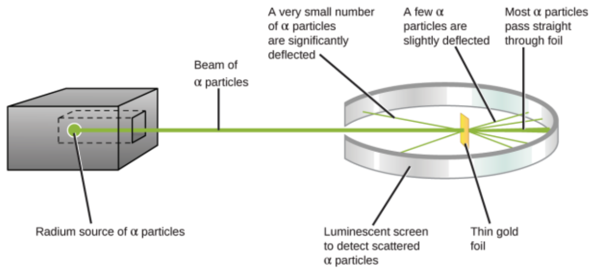
Subatomic Particles and Atomic Structure

Ernest Rutherford used α particles to find the structure of atoms. Most particles penetrated the gold foil undeflected, but some were deflected at a large angle, while others bounced back in the direction from which they had come.

Gold Foil Experiment Video 1
Gold Foil Experiment Video 2
Gold Foil Textbook Link



Rutherford proposed the nuclear model where the positive charge is concentrated in the **nucleus**.

The nucleus accounts for most of an atom's mass (extremely dense central core in the atom).
- A typical atomic radius is about 100 pm
- A typical nucleus has a radius of about 5×10^{-3} pm
- 1 pm = 1×10^{-12} m

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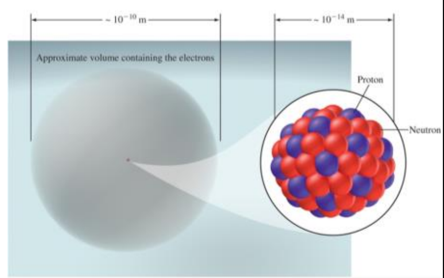
Subatomic Particles and Atomic Structure

Protons are positively charged particles found in the nucleus.
Neutrons are electronically neutral particles found in the nucleus.
Neutrons are slightly larger than protons.
Electrons are negatively charged particles distributed around the nucleus.



Particle	Mass (g)	Mass (amu)	Charge (C)	Charge Unit
Electron*	9.10938×10^{-28}	5.534×10^{-4}	-1.6022×10^{-19}	–1
Proton	1.67262×10^{-24}	1.0073	$+1.6022 \times 10^{-19}$	+1
Neutron	1.67493×10^{-24}	1.0086	0	0

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Subatomic Particles and Atomic Structure

Which of the following is true about protons and neutrons?

A. They both have a charge.

B. They have approximately the same mass.

C. They are unstable with respect to radioactive decay.

D. They are found in the same region of space around the nucleus as the electron.

Which experiment can be linked to Millikan?

A. Photographic plate fogged with Uranium – discovering radioactivity

B. Oil drop experiment – charge of an electron

C. Shooting alpha particles at gold foil – structure of the atom

D. Cathode ray tube – mass of an electron

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Atomic Number, Mass Number, and Isotopes

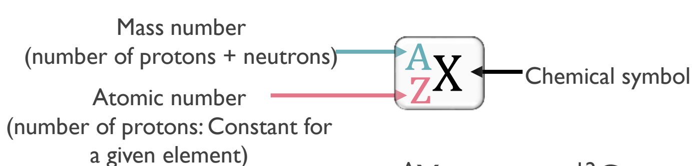
All atoms can be identified by the number of protons and neutrons they contain.

The **atomic number (Z)** is the number of protons in the nucleus.

- Atoms are neutral, so it's also the number of electrons.
- Protons determine the identity of an element.
  - For example, nitrogen's atomic number is 7, so every nitrogen has 7 protons.

The **mass number (A)** is the total number of protons and neutrons.

- Protons and neutrons are collectively referred to as **nucleons**.



or ^{A}X e.g. ^{12}C

or X–A e.g. carbon–12

For atoms: we use small units of measurement

1 amu = 1 atomic mass unit = 1/12 the mass of 1 carbon 12 atom

1 amu = 1.6605×10^{-24} g

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Examples

Determine the numbers of protons, neutrons, and electrons in each of the following species: (a) ^{28}Si, (b) ^{19}F, (c) ^{57}Fe, and (d) carbon–13.

Identify the # of protons, neutrons and electrons given the following information:

^{197}Au

^{48}Ti

^{55}Mn

Which neutral element are the following?

26 protons, 56 neutrons

20 neutrons, 18 electrons, no charge

76 neutrons, 52 protons

Pick the correct description for ^{59}Co.

A. 59 protons, 27 neutrons
B. 27 protons, 59 neutrons
C. 32 protons, 27 neutrons
D. 27 protons, 32 neutrons

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Atomic Number, Mass Number, and Isotopes

Most elements have two or more **isotopes**, atoms that have the same atomic number (Z) but different mass numbers (A). This is due to each isotope having a different number of neutrons.

Isotopes are atoms of the same element with different A (mass).

- **equal** numbers of p^+
- **different** numbers of n^0

Hydrogen isotopes:

1 proton 0 neutrons	1 proton 1 neutron	1 proton 2 neutrons
The mass of H is 1: 1 proton and zero neutrons	The mass is increased: 1 neutron/1 proton giving a mass of 2	The mass is increased: 2 neutrons/1 proton giving a mass of 2

Isotopes of the same element typically exhibit very similar chemical properties same types of compounds and similar reactivities.

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Isotopes and Atomic Weight

Most elements occur as a mixture of isotopes.

Magnesium is a mixture of:

	^{24}Mg	^{25}Mg	^{26}Mg
Number of p^{+}	12	12	12
Number of n°	12	13	14
Mass/amu	23.985	24.986	25.982

Atomic mass is the mass of an atom in atomic mass units (amu). The average atomic mass on the periodic table represents the average mass of the naturally occurring mixture of isotopes.

Isotope	Isotopic Mass (amu)	Natural abundance (%)
^{12}C	12.00000	98.93
^{13}C	13.003355	1.07

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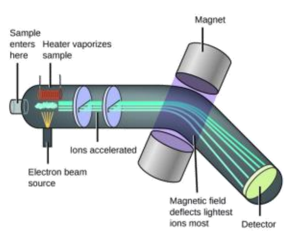
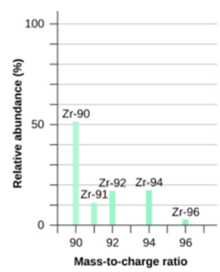
Average Atomic Mass

Measuring Atomic Mass: The most direct and most accurate method for determining atomic and molecular masses is mass spectrometry, using a mass spectrometer.





Mass Spectrometry Video

Isotopes and Mass Spectrometry Description

Chlorine has two stable isotopes, ^{35}Cl (75.78 percent) and ^{37}Cl (24.22 percent) which have masses of 34.9689 and 36.9659 amu, respectively. Calculate the average atomic mass of chlorine.

Strategy: Multiplying the mass of each isotope by its fractional abundance (percent value divided by 100) will give its contribution to the average atomic mass.

Solution

(0.7578)(34.9689 amu) + (0.2422)(36.9659 amu) = 35.44518022 amu

= 35.45 amu

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Example

Natural lithium is: 97.42% ^{6}Li (6.015 amu) and 2.58% ^{7}Li (7.016 amu), what is the average atomic mass of lithium?

Argon has three naturally occurring isotopes: 0.337% Ar-36 (35.968 amu), 0.063% Ar-38 (37.963 amu) and 99.60% Ar-40 (39.962 amu). What is the atomic weight of natural Argon?

The average atomic mass of nitrogen is 14.0067. The atomic masses of the two stable isotopes of nitrogen, ^{14}N and ^{15}N are 14.003074002 and 15.00010897 amu, respectively. Determine the percent abundance of each nitrogen isotope.

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Example

Silver has two stable isotopes; Ag-107 (106.90509 amu) and Ag-109 (108.90476 amu). Determine the percentage abundance of these two isotopes of silver.

Rubidium has two naturally occurring isotopes, ^{85}Rb (relative mass 84.9118 amu) and ^{87}Rb (relative mass 86.9092 amu). If rubidium has an average atomic mass of 85.47 amu, what is the percent abundance of each isotope?

A. 52.45% ^{85}Rb, 47.55% ^{87}Rb
B. 26.81% ^{85}Rb, 73.19% ^{87}Rb
C. ^{85}Rb :72.05%, ^{87}Rb: 27.95%
D. Aaaggh!

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Covalent Bonding and Molecules

A **molecule** is a combination of at least two atoms in a specific arrangement held together by chemical forces (chemical bonds).
A molecule may be an element or a compound. Some elements that naturally exist as molecules include:

Br_2, I_2, N_2, Cl_2, H_2, O_2, F_2, and C_{60} and S_8

BrINClHOF (**The super 7**)

Different samples of a given compound always contain the same elements in the same ratio. This is known as the **law of definite proportions**.

Sample	Mass of O (g)	Mass of C (g)	Ratio (O(*g*):C(*g*)
123 g carbon dioxide	89.4	33.6	2.66:1
50.5 g carbon dioxide	36.7	13.8	2.66:1
88.6 g carbon dioxide	64.4	24.2	2.66:1

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Covalent Bonding and Molecules

If two elements can form a series of different compounds, the **law of multiple proportions** tells us that the ratio of masses of one element that combine with a fixed mass of the other element can be expressed in small whole numbers.

Sample	Mass of O (g)	Mass of C (g)	Ratio (O(g):C(g))
123 g CO_2	89.4	33.6	2.66:1
50.5 g CO_2	36.7	13.8	2.66:1
88.6 g CO_2	64.4	24.2	2.66:1

In addition to carbon dioxide, carbon also combines with oxygen to form carbon monoxide.

Sample	Mass of O (g)	Mass of C (g)	Ratio (O(g):C(g)
16.3 g CO	9.31	6.99	1.33:1
25.9 g CO	14.8	11.1	1.33:1
88.4 g CO	50.5	37.9	1.33:1

$$\frac{2.66\ \text{O}}{1.33\ \text{O}} = \frac{2\ \text{O in } CO_2}{1\ \ \text{O in CO}}$$

law of multiple proportions tells us 2O in CO_2:1O in CO

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Molecular Formulas

Diatomic molecules contain two atoms and may be either ***heteronuclear*** or ***homonuclear*** (***BrINClHOF***). ***Polyatomic molecules*** contain more than two atoms.

Homonuclear diatomic: H_2 *Heteronuclear diatomic: CO* *Polyatomic: PO_4^{3-}*

A chemical formula denotes the composition of the substance through element symbols and numbers, as well as parentheses, dashes, brackets, commas, plus, and minus signs.

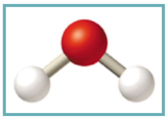
A **molecular formula** shows the exact number of atoms of each element in a molecule.

H_2O

Some elements have two or more distinct forms known as **allotropes**.
For example, oxygen (O_2) and ozone (O_3) are allotropes of oxygen.
C, S, and P also have allotropic forms

A **structural formula** shows not only the elemental composition, but also the general arrangements.

H—O—H



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Empirical Formulas

Molecular substances can also be represented using **empirical formulas**, the whole–number ratio of elements.

While the **molecular formulas** tell us the actual number of atoms in the molecule (the true formula), the **empirical formula** gives the lowest whole–number ratio of elements (the simplest formula).

Molecular formula: N_2H_4
Empirical formula: NH_2

The molecular and empirical formulas are often the same.

Compound	Molecular Formula	Empirical Formula
Water	H_2O	H_2O
Hydrogen peroxide	H_2O_2	HO
Ethane	C_2H_6	CH_3
Propane	C_3H_8	C_3H_8
Acetylene	C_2H_2	CH
Benzene	C_6H_6	CH

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Examples

Write the empirical formulas for the following molecules: (a) glucose ($C_6H_{12}O_6$), a substance known as blood sugar; (b) adenine ($C_5H_5N_5$), also known as vitamin B_4; and (c) nitrous oxide (N_2O), a gas that is used as an anesthetic ("laughing gas") and as an aerosol propellant for whipped cream.

Strategy To write the empirical formula, the subscripts in the molecular formula must be reduced to the smallest possible whole numbers

True or False:
If any coefficient (subscript) in the molecular formula is 1, then the molecular formula and empirical formula are the same.

A. True

B. False

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Molecular and Formula Mass

The **molecular mass** is the mass in atomic mass units (amu) of an individual molecule.

To calculate molecular mass, multiply the atomic mass of each element in a molecule by the number of atoms of that element and then total the masses

Molecular mass of H_2O:

2(atomic mass units of H) +(1)(atomic mass units of O)

2(1.008 amu) + (1)(16.00 amu) = 18.02 amu

Because the atomic masses on the periodic table are average atomic masses, the result of such a determination is an average molecular mass, sometimes referred to as the **molecular weight**.

Although an ionic compound does not have a molecular mass, we can use its empirical formula to determine its **formula mass** (the mass of a "formula unit"), sometimes called the **formula weight**.

The process is the same:
To calculate formula mass, multiply the atomic mass for each element in a formula unit by the number of atoms of that element and then total the masses

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Example Problem

Calculate the molecular mass or the formula mass, as appropriate, for each of the following:
(a) ethane, C_2H_6, (b) lithium hydroxide, (c) $CaCl_2$

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Practice Problem

Determine the Formula/Molecular weight for the following compounds:

Li_2NO_3

Potassium cyanide

Al_2O_3

Iron(III) sulfate

Carbon dioxide

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The Mole and Molar Mass

The **mole** is defined as the amount of a substance that contains as many elementary entities as there are atoms in exactly 12 g of carbon–12. This experimentally determined number is called **Avogadro's number (N_A)**

1 mole = 6.022×10^{23}

Molar Mass is the mass (grams) in 1 mole of a compound or element given in units of g mol^{-1}.

Example Problem 3:
What is the molar mass of nitrogen dioxide (NO_2)?
Hint: Use a similar strategy you would use to calculate the molecular mass of NO_2.

Molecular mass of NO_2 = atomic mass of N + 2(atomic mass of O)
=14.01 amu +(2)(16.00 amu) = 46.01 amu

Molar Mass of NO_2 = molar mass of N + 2(molar mass of O)

$$= \frac{14.01\ \text{g}}{\text{mol}} + (2)\left(\frac{16.00\ \text{g}}{\text{mol}}\right) = \frac{46.01\ \text{g}}{\text{mol}}$$

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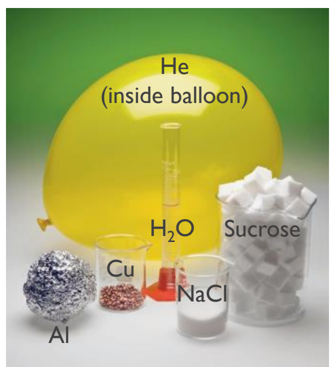
The Mole

One mole of some familiar substances:
- 4.0 g of He
- ~ 18.0 g of H_2O
- ~ 26.98 g Al
- ~ 63.55 g Cu
- ~ 58.44 g salt (NaCl)
- ~ 180.16 g sugar ($C_6H_{12}O_6$)



Mole Video
Mole Textbook Link

1 mole of marshmallows would be enough marshmallows to make a 19 km (12 mi) thick layer of marshmallows covering the entire face of the Earth. (Marshmallows would be in the clouds)

1 mole of cells would be approximately equivalent to the number of cells found composing every human on earth.
(100 trillion cells per person x 7.4 billion people = 7.4×10^{23} cells)

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Example Problems

Determine (a) the number of C atoms in 12.00 moles of carbon and (b) the number of moles of sodium in a sample containing 3.23×10^{10} Na atoms.

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Molar Mass

The **molar mass** of a substance is the mass in grams of one mole of the substance.

The mass of 1 mole of carbon–12 is exactly 12 g.
The mass of 1 carbon–12 atom is exactly 12 amu

We usually express molar masses in units of grams per mole (g mol^{-1}) to facilitate cancellation of units in calculations. Molar masses can be found on the periodic table. For example, in 1 mole of carbon–12, we have:

$$\frac{12 \text{ g carbon}}{1 \text{ mole carbon}} \quad \text{OR} \quad \frac{1 \text{ mole carbon}}{12 \text{ g carbon}}$$

How many moles are in 62.8 grams of Fe?

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Molecular Mass and Molar Mass

Molecular mass (or molecular weight) is the sum of the atomic masses (in amu) in a molecule.

1 Sulfur	32.07 amu
2 Oxygen	+ (2 x 16.00 amu)
SO_2	64.07 amu

For any molecule

molecular mass (amu) = molar mass (grams)

1 molecule SO_2 = 64.07 amu

1 mole SO_2 = 64.07 g SO_2

How many moles of $Ca_3(PO_4)_2$ are in 10.0 g of the compound?

Formula mass (Molar mass)

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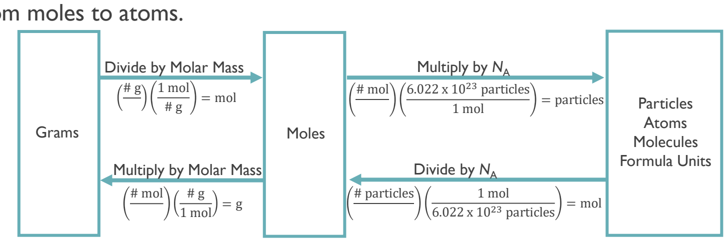
Conversion between Mass, Moles, Number of Atoms

Molar mass is the conversion factor from mass to moles, and vice versa. Avogadro's constant converts from moles to atoms.



Determine (a) the number of Ca atoms in 0.515 g of Ca, and (b) the mass of H that contains 6.89×10^{18} H atoms.

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Practice

(a) Determine the number of water molecules and the numbers of H and O atoms in 3.26 g of water. (b) Determine the mass of 7.92×10^{19} carbon dioxide molecules.

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Molar Mass

The **molar mass (M)** of a substance is the mass in grams of one mole of the substance. The molar mass of a compound is the sum of molar masses of the elements it contains.

$$1 \text{ mol } H_2O = 2 \times 1.008 \text{ g} + 16.00 \text{ g} = 18.02 \text{ g}$$

(a) How many moles are in 73.7 g of calcium nitrite?

(b) How many atoms are in 0.551 g of potassium (K) ?

(c) What is the number of moles of CO_2 in 10.00 g of carbon dioxide?

(d) How many H atoms are in 72.5 g of C_3H_8O ?

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Molar Mass

How many moles of Na are in 50.4 g of sodium?

How many atoms are in 0.0034 g Platinum?

The moles in 143.5 g of $Zn(NO_3)_2$

What is the mass in grams of 7.70×10^{20} molecules of caffeine, $C_8H_{10}N_4O_2$?

What is the molar mass of cholesterol if 0.00105 mol has a mass of 0.406 g?

The most important beryllium-containing compound is beryl, which occurs mostly as blue-green crystals with the formula $Be_3Al_2(SiO_3)_6$. How many moles of beryl are there in a 0.25 g crystal? How many molecules? How many Be atoms?

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