Midterm Study Guide

Mid-Second Semester Examination

Academic Year 2025-2026

Subject: Chemistry

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Content: Experience Chemistry

Topics Covered

Chapter	Lesson	Pages
Physical Properties of Materials	1- States of Matter	112-123
	2- Modeling Phase Changes	125-134
Chemical Quantities	1- The Mole Concept	174-181
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	3- Percent Composition and Empirical Formulas	192-201

Please study the material listed in the table above with a focus on the points below.

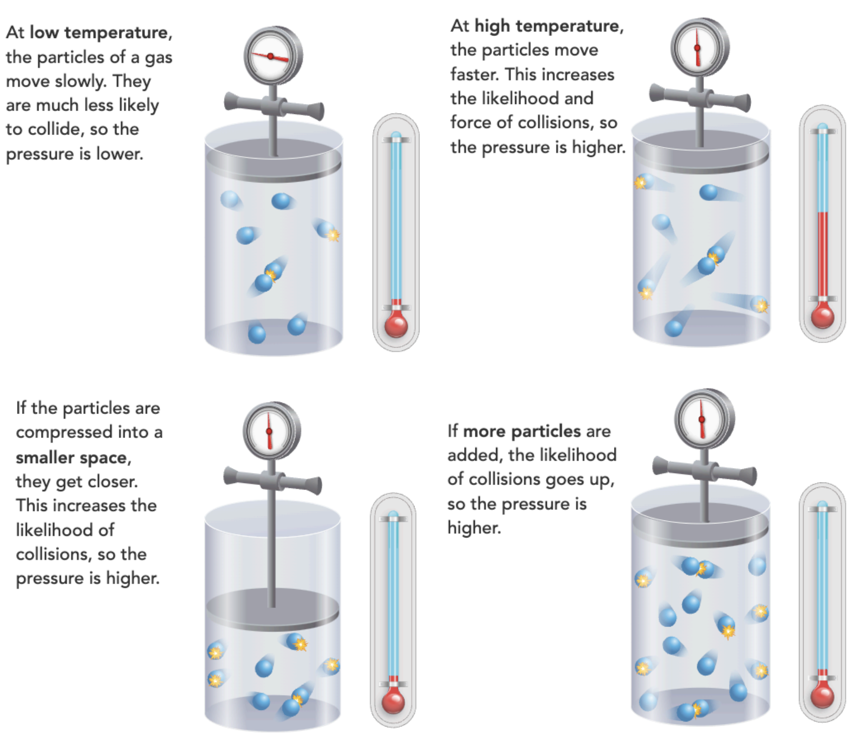
Key Topics and Concepts

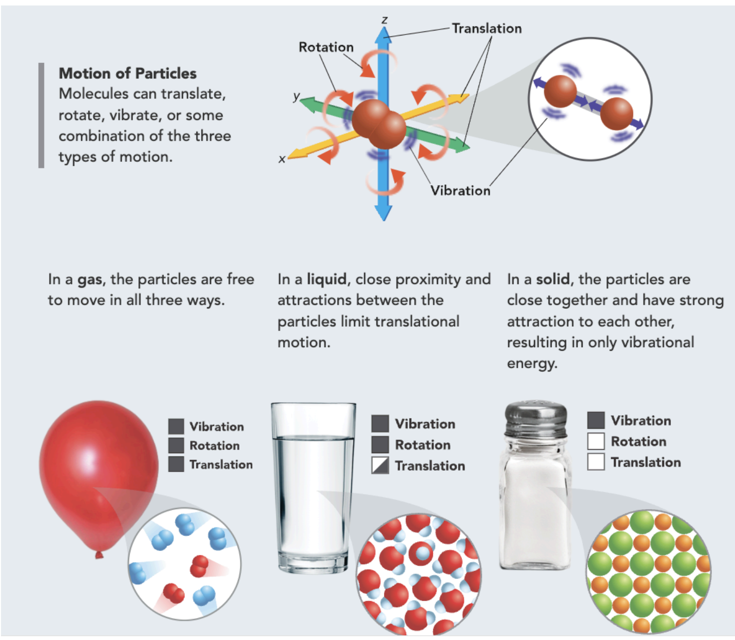
Chapter 1: Physical Properties of Materials

- Describe three states of matter (solid, liquid, gas)
- Understand common gases at STP **[Temperature of 0℃ and Atmospheric Pressure of 100 kPa]**
- Explain behaviors of gases with changes in temperature, pressure, and volume
- Converting between units of pressure
- Explain particle motion in solids, liquids, and gases
- Understand intermolecular forces to explain states of matter
- Describe solids and attractive forces **(FROM HANDOUT)**
- Explain the crystal structure of solids **(FROM HANDOUT)**

- Explain the relationship between energy and phase change
- Define the types of phase changes and describe how variations in kinetic energy among particles result in evaporation and condensation, even without heating or cooling
- Summarize the mechanism by which gas bubbles in a heated liquid exert enough pressure to break free of the surface, resulting in boiling
- Explain the relationship between vapor pressure and boiling
- Relate the melting point of different solids to the strength of intermolecular forces, to predict the bond energy of a compound **[The higher the melting point, the higher the bond energy]**
- Interpret phase diagrams to determine the state of water at at various pressure and temperature points

Temperature, Volume, and Pressure Changes in temperature and volume affect the speed of the particles and the number of collisions between particles.





Comparing units Conversion factors are used to convert between different units of pressure.

$1 \text{ atm} = 760 \text{ mm Hg}$

$1 \text{ atm} = 101.3 \text{ kPa}$

A pressure gauge records a pressure of 450 kPa. Convert this measurement to **a.** atmospheres and **b.** millimeters of mercury.

ANALYZE List the knowns and unknowns.

Knowns
pressure = 450 kPa
1 atm = 101.3 kPa
1 atm = 760 mm Hg

Unknowns
pressure = ? atm
pressure = ? mm Hg

CALCULATE Solve for the unknowns.

Identify the appropriate conversion factor to convert kPa to atm.

The units in the known should be in the denominator so that the units will cancel.

a. $\frac{1 \text{ atm}}{101.3 \text{ kPa}}$

Multiply the given pressure by the conversion factor.

$$450 \text{ kPa} \times \frac{1 \text{ atm}}{101.3 \text{ kPa}} = 4.4 \text{ atm}$$

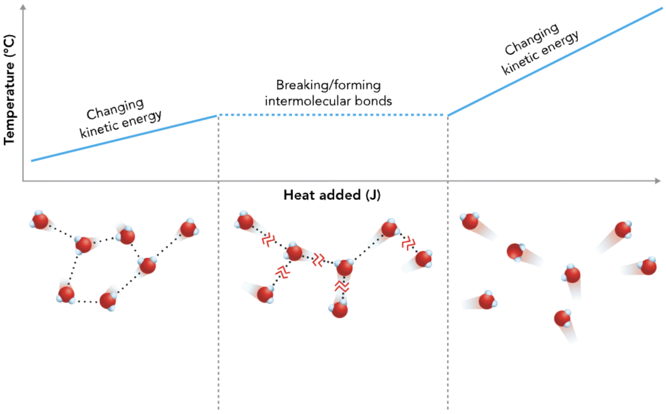
Identify the appropriate conversion factor to convert kPa to mm Hg.

b. $\frac{760 \text{ mm Hg}}{101.3 \text{ kPa}}$

Multiply the given pressure by the conversion factor.

Round the answer to two significant figures because 450 kPa has only two significant figures.

$$450 \text{ kPa} \times \frac{760 \text{ mm Hg}}{101.3 \text{ kPa}} = 3376 \text{ mm Hg} = 3.4 \times 10^3 \text{ mm Hg}$$

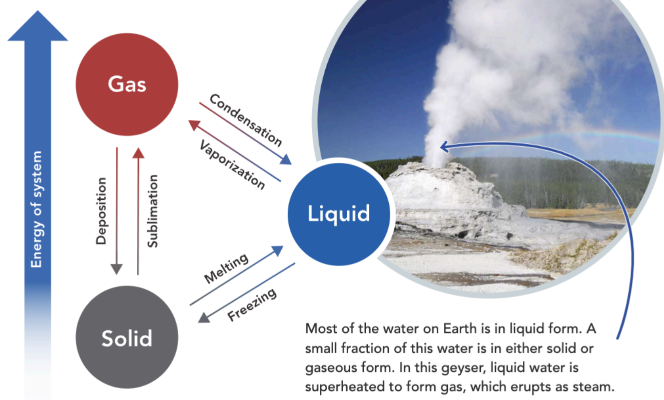


Increasing Kinetic Energy of Liquid Water Molecules Added energy is absorbed by the water molecules and turned into increased molecular motion, increasing the temperature.

Breaking Intermolecular Forces Added energy goes into breaking the hydrogen bonds between water molecules, and not into increasing molecular motion. The temperature remains constant.

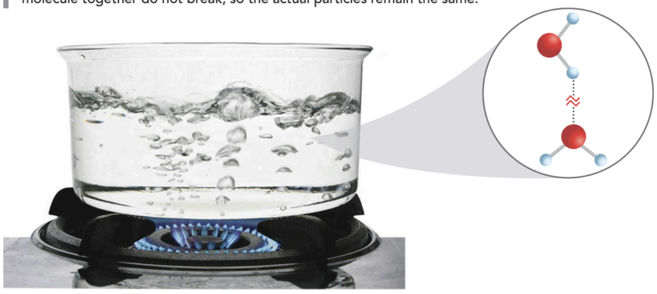
Increasing Kinetic Energy of Gaseous Water Molecules Once all hydrogen bonds are broken, added energy goes into molecular motion, increasing temperature.

Heat Energy in Materials The heat energy in a material increases as it transitions from solid to liquid, and liquid to gas.



Most of the water on Earth is in liquid form. A small fraction of this water is in either solid or gaseous form. In this geyser, liquid water is superheated to form gas, which erupts as steam.

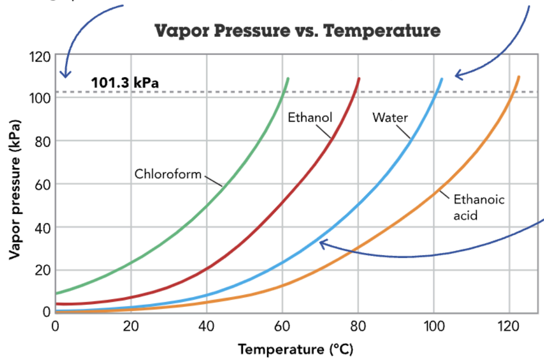
Breaking Bonds The intermolecular hydrogen bonds in water are broken during boiling. But the intramolecular covalent bonds that hold a molecule together do not break, so the actual particles remain the same.



Boiling Points The normal boiling point is defined as the boiling point of a liquid at 101.3 kPa, which is approximately atmospheric pressure at sea level. If the surrounding external pressure is greater, then the boiling point will also be higher.

The dashed line marks 101.3 kPa which is where the normal boiling point is read on the graph for each substance.

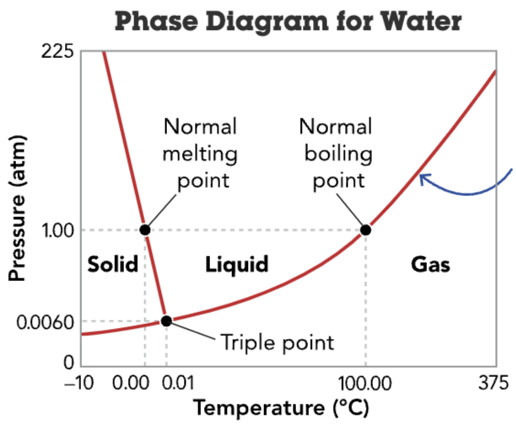
At 100°C, the vapor pressure of water equals 101.3 kPa, so the normal boiling point of water is 100°C.



If the atmospheric pressure dropped to 25 kPa, then the boiling point of water would drop to approximately 60°C.

The relative strength of intermolecular forces can be used to predict differences in bulk properties like boiling point. Stronger intermolecular forces result in higher boiling points. The strength of intermolecular forces is related to the type of force and the size and weight of the molecules.

Reading a Phase Diagram A phase diagram shows the pressure as a function of the temperature and outlines in what phase a material will be found at a specific temperature and pressure.



The lines in a phase diagram show the temperatures and pressures at which transitions between phases occur.

The **normal melting point** is the temperature at which water melts at 101.3 kPa (1 atm). The normal melting point of water is 0°C.

The **triple point** describes the only set of conditions under which all three phases can exist in equilibrium with each other.

The **normal boiling point** is the temperature at which water boils at 101.3 kPa. For water, this temperature is 100°C.

Chapter 2: Chemical Quantities

- Know the three methods used to measure matter —count, mass, and volume
- Explain the relationship between the mole and Avogadro's number **[6.02x10²³ - memorize]**
- Use Avogadro's number to convert from moles to particles and particles to moles
- Use the periodic table to find the molar mass of elements and compounds
- Convert mole quantities to masses, and mass quantities to moles
- Explain the relationship between moles and volume of gases at STP
- Calculate the density of a gas at STP, given the molar mass of the gas
- Be able to find the percent composition of a compound
- Describe the law of definite proportion and law of constant composition
- Find the empirical and molecular formulas for a compound
- Use subscripts within a chemical formula to represent a mole ratio

What is the mass of 90 apples if 1 dozen of the apples has a mass of 2.0 kg?

ANALYZE List the knowns and the unknown.

Knowns
number of apples = 90
12 apples = 1 dozen apples
1 dozen apples = 2.0 kg apples

Unknown
mass of 90 apples = ? kg

CALCULATE Solve for the unknown.

Identify the steps to convert from number, or count, to mass.

number of apples → dozens of apples → mass of apples

Multiply the number of apples by the two conversion factors needed to convert from number of apples to mass of apples.

$$90 \text{ apples} \times \frac{1 \text{ dozen apples}}{12 \text{ apples}} \times \frac{2.0 \text{ kg apples}}{1 \text{ dozen apples}} = 15 \text{ kg apples}$$

Magnesium is a light metal used in the manufacture of aircraft and tools. How many moles of magnesium is 1.25×10^{23} atoms of magnesium?

ANALYZE List the knowns and the unknown.

Knowns
number of atoms = 1.25×10^{23} atoms Mg

Unknown
moles = ? mol Mg

CALCULATE Solve for the unknown.

Identify the relationship between moles and number of representative particles.

$$1 \text{ mol Mg} = 6.02 \times 10^{23} \text{ atoms Mg}$$

Write the conversion factors based on this relationship.

$$\frac{1 \text{ mol Mg}}{6.02 \times 10^{23} \text{ atoms Mg}} \text{ and } \frac{6.02 \times 10^{23} \text{ atoms Mg}}{1 \text{ mol Mg}}$$

Identify the conversion factor needed to convert from atoms to moles.

$$\frac{1 \text{ mol Mg}}{6.02 \times 10^{23} \text{ atoms Mg}}$$

The units in the known should be in the denominator so that the units will cancel.

Multiply the number of atoms of Mg by the conversion factor.

$$1.25 \times 10^{23} \text{ atoms Mg} \times \frac{1 \text{ mol Mg}}{6.02 \times 10^{23} \text{ atoms Mg}}$$

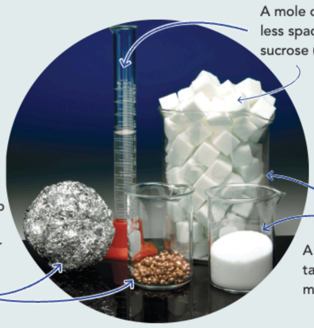
$$= 0.208 \text{ mol Mg}$$

A Mole of Common Substances Common substances may have atoms, molecules, or formula units as their representative particles. A water molecule is the representative unit of water. Aluminum and copper both have atoms as their representative particles. The volume of these substances depends on the properties of these particles.

A water molecule is composed of one oxygen atom and two hydrogen atoms.

Therefore, 1 mole of water has 3 moles of atoms.

$1 \text{ mole of } H_2O = 1 \text{ mole of O} + 2 \text{ moles of H}$

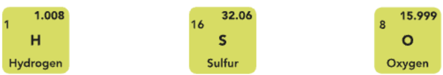


Molar Mass of Sulfuric Acid Adding the masses of 2 moles of hydrogen, 1 mole of sulfur, and 4 moles of oxygen gives the molar mass of H_2SO_4.

2 H atoms + 1 S atom + 4 O atoms = 1 H_2SO_4 molecule

The moles of H, S, and O in 1 mol H_2SO_4 are the same as the number of atoms of each element in one molecule of H_2SO_4.

2 mol H + 1 mol S + 4 mol O = 1 mol 1 H_2SO_4



Round the values in the periodic table to one decimal place to simplify the molar mass calculation.

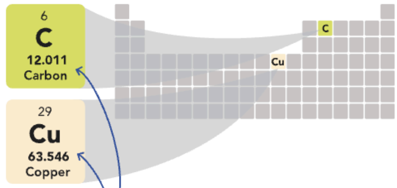
Multiply the moles of each element by its molar mass.

$$\left(2 \text{ mol H} \times \frac{1.0 \text{ g H}}{1 \text{ mol H}}\right) + \left(1 \text{ mol S} \times \frac{32.1 \text{ g S}}{1 \text{ mol S}}\right) + \left(4 \text{ mol O} \times \frac{16.0 \text{ g O}}{1 \text{ mol O}}\right) = 98.1 \text{ g/mol}$$

Molar Mass

How do you obtain the **molar mass from the periodic table?**

Atomic Weight The periodic table gives the atomic weight of each element. Atomic weight is a dimensionless quantity that represents the **average mass of atoms of an element divided by 1/12 the atomic mass of a carbon-12 atom.**

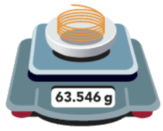


Molar Mass The **molar mass is the mass of 1 mole of an element.** It is the atomic weight multiplied by 1 g/mol. To obtain the molar mass, read the atomic weight from the periodic table and add the unit g/mol.

1 mole of carbon

Molar mass of carbon = 12.011 g/mol

1 mole of copper



Molar mass of copper = 63.546 g/mol

Molar Mass as Conversion Factor The **molar mass can be represented as a fraction and used as a conversion factor** to convert a known mass into moles or known moles into mass. The molar mass is often rounded to 1 decimal place to simplify calculations.

$$\frac{12.0 \text{ g C}}{1 \text{ mol C}} \text{ or } \frac{1 \text{ mol C}}{12.0 \text{ g C}}$$

$$\frac{63.5 \text{ g Cu}}{1 \text{ mol Cu}} \text{ or } \frac{1 \text{ mol Cu}}{63.5 \text{ g Cu}}$$

The decomposition reaction of hydrogen peroxide (H_2O_2) provides sufficient energy to launch a rocket. What is the molar mass of hydrogen peroxide?

ANALYZE List the knowns and the unknown.

Knowns
molecular formula = H_2O_2
mass of 1 mol H = 1.0 g H
mass of 1 mol O = 16.0 g O

Unknown
molar mass = ? g/mol

CALCULATE Solve for the unknown.

Convert moles of atoms to grams using conversion factors (g/mol) based on the molar mass of each element.

Multiply the molar mass of each element by its subscript in the chemical formula to find the total mass of the element.

$$2 \text{ mol H} \times \frac{1.0 \text{ g H}}{1 \text{ mol H}} = 2.0 \text{ g H}$$

$$2 \text{ mol O} \times \frac{16.0 \text{ g O}}{1 \text{ mol O}} = 32.0 \text{ g O}$$

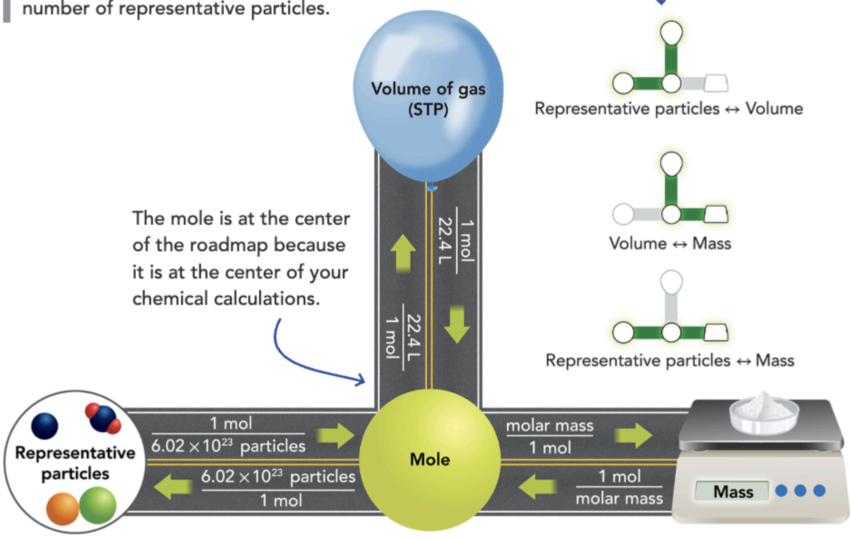
Add the masses of the elements to determine the molar mass of the compound.

$\text{mass of 1 mol } H_2O_2 = 2.0 \text{ g H} + 32.0 \text{ g O} = 34.0 \text{ g}$

$\text{molar mass of } H_2O_2 = 34.0 \text{ g/mol}$

$$\% \text{ by mass of element} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100$$

The Mole Roadmap Follow the arrows on the roadmap to determine which conversion factors you need to use to convert between mass, volume, and number of representative particles.



When iron is exposed to air, it undergoes a corrosion reaction to form a redbrown rust. Rust is iron(III) oxide (Fe_2O_3). How many moles of iron(III) oxide are contained in 92.2 g of pure Fe_2O_3?

ANALYZE List the knowns and the unknown.

Knowns
mass = 92.2 g Fe_2O_3

Unknown
number of moles = ? mol Fe_2O_3

CALCULATE Solve for the unknown.

Identify the necessary conversion.

mass → moles

Use the subscripts to determine the moles of each element in the compound. Multiply the moles by the molar masses to determine the mass of each element in the compound.

$$2\ \text{mol Fe} \times \frac{55.8\ \text{g Fe}}{1\ \text{mol Fe}} = 111.6\ \text{g Fe}$$

$$3\ \text{mol O} \times \frac{16.0\ \text{g O}}{1\ \text{mol O}} = 48.0\ \text{g O}$$

Add the mass of the elements to find the mass of 1 mole of the compound.

$$1\ \text{mol}\ Fe_2O_3 = 111.6\ \text{g Al} + 48.0\ \text{g O} = 159.6\ \text{g}\ Fe_2O_3$$

Identify the conversion factor for converting mass to moles.

$$\frac{1\ \text{mol}\ Fe_2O_3}{159.6\ \text{g}\ Fe_2O_3}$$

Note that the known unit (g) is in the denominator and the unknown unit (mol) is in the numerator.

Multiply the given mass by the conversion factor.

$$92.2\ \text{g}\ Fe_2O_3 \times \frac{1\ \text{mol}\ Fe_2O_3}{159.6\ \text{g}\ Fe_2O_3} = 0.578\ \text{mol}\ Fe_2O_3$$

Items made of aluminum, such as aircraft parts and cookware, are resistant to corrosion because the aluminum reacts with oxygen in the air to form a coating of aluminum oxide (Al_2O_3). This tough, resistant coating prevents any further corrosion. What is the mass, in grams, of 9.45 mol aluminum oxide?

ANALYZE List the knowns and the unknown.

Knowns
number of moles = 9.45 mol Al_2O_3

Unknown
mass = ? g Al_2O_3

CALCULATE Solve for the unknown.

Identify the necessary conversion.

moles → mass

Multiply the moles by the molar masses to determine the mass of each element in the compound.

$$2\ \text{mol Al} \times \frac{27.0\ \text{g Al}}{1\ \text{mol Al}} = 54.0\ \text{g Al}$$

$$3\ \text{mol O} \times \frac{16.0\ \text{g O}}{1\ \text{mol O}} = 48.0\ \text{g O}$$

Use the subscripts to determine the moles of each element in the compound.

Add the mass of the elements to find the mass of 1 mole of the compound.

$$1\ \text{mol}\ Al_2O_3 = 54.0\ \text{g Al} + 48.0\ \text{g O} = 102.0\ \text{g}\ Al_2O_3$$

Identify the conversion factor for converting from moles to mass

$$\frac{102.0\ \text{g}\ Al_2O_3}{1\ \text{mol}\ Al_2O_3}$$

The known unit (mol) is in the denominator and the unknown unit (g) is in the numerator.

Multiply the given number of moles by the conversion factor.

$$9.45\ \text{mol}\ Al_2O_3 \times \frac{102.0\ \text{g}\ Al_2O_3}{1\ \text{mol}\ Al_2O_3} = 964\ \text{g}\ Al_2O_3$$

EVALUATE Does the result make sense?

The number of moles of Al_2O_3 is approximately 10, and each mole has a mass of approximately 100 g. The answer should be close to 1000 g. The answer has been rounded to the correct number of significant figures.

Sulfur dioxide (SO_2) is a gas produced by burning coal. It is an air pollutant and one of the causes of acid rain. Determine the volume, in liters, of 0.60 mol SO_2 gas at STP.

ANALYZE List the knowns and unknown.

Knowns
number of moles = 0.60 mol SO_2
1 mol SO_2 = 22.4 L SO_2 at STP

Unknown
volume = ? L SO_2

CALCULATE Solve for the unknown.

Identify the conversion factor relating moles of SO_2 gas to volume of SO_2 gas at STP.

$$\frac{22.4\ \text{L}\ SO_2}{1\ \text{mol}\ SO_2}$$

Multiply the given number of moles by the conversion factor.

$$0.60\ \text{mol}\ SO_2 \times \frac{22.4\ \text{L}\ SO_2}{1\ \text{mol}\ SO_2} = 13\ \text{L}\ SO_2$$

Applying the Mole-to-Mass Roadmap What is the mass, in grams, of 3.2 mol nitrogen dioxide (NO_2)?

Use the mole roadmap to determine how to convert moles to mass.

$$\text{moles of } NO_2 \times \frac{\text{molar mass } NO_2}{1 \text{ mol } NO_2} \rightarrow \text{mass of } NO_2$$

Use the subscripts from the chemical formula to find the mass of each element in the compound.

$$1 \text{ mol N} \times \frac{14.0 \text{ g N}}{1 \text{ mol N}} \rightarrow 14.0 \text{ g N}$$

$$2 \text{ mol O} \times \frac{16.0 \text{ g O}}{1 \text{ mol O}} \rightarrow 32.0 \text{ g O}$$

Add the mass of the elements to find the mass of 1 mole of the compound.

$$1 \text{ mol } NO_2 = 14.0 \text{ g N} + 32.0 \text{ g O} = 46.0 \text{ g } NO_2$$

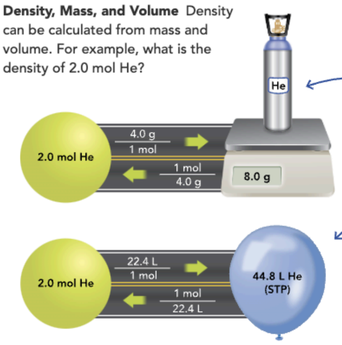
Multiply the given number of moles by the conversion factor.

$$3.2 \text{ mol } NO_2 \times \frac{46.0 \text{ g } NO_2}{1 \text{ mol } NO_2} \rightarrow 147.2 \text{ g } NO_2$$

The density of a gas at STP can be determined using one of the following expressions:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad \text{or} \quad \text{density} = \frac{\text{molar mass}}{\text{molar volume}}$$

Density, Mass, and Volume Density can be calculated from mass and volume. For example, what is the density of 2.0 mol He?



$$\text{density He} = \frac{\text{mass}}{\text{volume}} = \frac{8.0 \text{ g He}}{44.8 \text{ L He}} = 0.18 \text{ g/L He}$$

Helium has a density of 0.18 g/L at STP, which is less than the density of air (1.23 g/L). Therefore, a helium-filled balloon floats.

$$\text{density He} = \frac{\text{molar mass He}}{\text{molar volume He}} = \frac{\frac{4.0 \text{ g He}}{1 \text{ mol He}}}{\frac{22.4 \text{ L He}}{1 \text{ mol He}}} = 0.18 \text{ g/L He}$$

Since density is a ratio, you can also calculate it directly from the molar mass and molar volume. The moles cancel!

When a 13.60-g sample of a compound containing only magnesium and oxygen is decomposed, 5.40 g of oxygen is obtained. What is the percent composition of this compound?

ANALYZE List the knowns and the unknowns.

Knowns
mass of compound = 13.60 g
mass of oxygen = 5.40 g O
mass of magnesium = 13.6 g − 5.40 g O = 8.20 g Mg

Unknowns
percent by mass of Mg = ?% Mg
percent by mass of O = ?% O

CALCULATE Solve for the unknowns.

Determine the percent by mass of Mg in the compound.

$$\%Mg = \frac{\text{mass of Mg}}{\text{mass of compound}} \times 100 = \frac{8.20 \text{ g}}{13.60 \text{ g}} \times 100 = 60.3\% \text{ Mg}$$

Determine the percent by mass of O in the compound.

$$\%O = \frac{\text{mass of O}}{\text{mass of compound}} \times 100 = \frac{5.40 \text{ g}}{13.60 \text{ g}} \times 100 = 39.7\% \text{O}$$

$$\% \text{ by mass of element} = \frac{\text{mass of element in 1 mol of compound}}{\text{molar mass of compound}} \times 100$$

Percent Composition of Water Any sample of pure water will have the same proportions of hydrogen and oxygen. To determine the percent composition of each element in water, calculate its proportion by mass.

STEP 1 Use the subscripts to determine how many moles of each element are in the compound.

$$1 \text{ mol } H_2O = 2 \text{ mol H} + 1 \text{ mol O}$$

STEP 2 Convert moles to mass for each element in the compound.

$$2 \text{ mol H} \times \frac{1.0 \text{ g H}}{1 \text{ mol H}} \rightarrow 2.0 \text{ g H} \qquad 1 \text{ mol O} \times \frac{16.0 \text{ g O}}{1 \text{ mol O}} \rightarrow 16.0 \text{ g O}$$

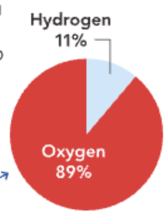
STEP 3 Calculate the percent composition of each element.

$$\% \text{ by mass of element} \rightarrow \frac{\text{mass of element in 1 mol of compound}}{\text{molar mass of compound}} \times 100$$

$$\% \text{ by mass of H} \rightarrow \frac{2.0 \text{ g}}{18.0 \text{ g}} \times 100 \rightarrow 11\% \text{ H}$$

$$\% \text{ by mass of O} \rightarrow \frac{16.0 \text{ g}}{18.0 \text{ g}} \times 100 \rightarrow 89\% \text{ O}$$

Any sample of pure water is 11% hydrogen to 89% oxygen by mass.



Propane (C_3H_8), the fuel commonly used in gas grills, is one of the compounds obtained from petroleum. Calculate the percent composition of propane.

ANALYZE List the knowns and the unknowns.

Knowns
mass of C in 1 mol C_3H_8 = 3 mol × 12.0 g/mol = 36.0 g
mass of H in 1 mol C_3H_8 = 8 mol × 1.0 g/mol = 8.0 g
molar mass of C_3H_8 = 36.0 g/mol + 8.0 g/mol = 44.0 g/mol

Unknowns
percent by mass of C = ?% C
percent by mass of H = ?% H

CALCULATE Solve for the unknown.

Determine the percent by mass of C in the C_3H_8 by dividing the mass of C in 1 mole of the compound by the molar mass of the compound and multiplying by 100.

$$\%\,C = \frac{\text{mass of C in 1 mol } C_3H_8}{\text{molar mass of } C_3H_8} \times 100 = \frac{36.0\ \cancel{g}}{44.0\ \cancel{g}} \times 100 = 81.8\%\ C$$

Determine the percent by mass of H in C_3H_8.

$$\%\,H = \frac{\text{mass of H in 1 mol } C_3H_8}{\text{molar mass of } C_3H_8} \times 100 = \frac{8.0\ \cancel{g}}{44.0\ \cancel{g}} \times 100 = 18\%\ H$$

An unknown compound is analyzed and found to contain 25.9% nitrogen and 74.1% oxygen. What is the empirical formula of the compound?

ANALYZE List the knowns and the unknown.

Knowns
percent by mass of N = 25.9% N
percent by mass of O = 74.1% O

Unknown
empirical formula = $N_?O_?$

CALCULATE Solve for the unknown.

Convert the percent by mass of each element to moles to determine the mole ratio.

$$25.9\ \cancel{g\,N} \times \frac{1\ \text{mol N}}{14.0\ \cancel{g\,N}} = 1.85\ \text{mol N}$$

$$74.1\ \cancel{g\,O} \times \frac{1\ \text{mol O}}{16.0\ \cancel{g\,O}} = 4.63\ \text{mol O}$$

The mole ratio is $N_{1.85}$:$O_{4.63}$.

Divide each molar quantity by the smaller number of moles to get 1 mol from the element with the smaller number of moles.

$$\frac{1.85\ \text{mol N}}{1.85} = 1\ \text{mol N}$$

$$\frac{4.63\ \text{mol O}}{1.85} = 2.5\ \text{mol O}$$

The mole ratio becomes N_1:$O_{2.5}$.

Reduce the mole ratio to the lowest whole-number ratio by multiplying each part of the ratio by the smallest whole number that will convert both subscripts to a whole number.

1 mol N × 2 = 2 mol N
2.5 mol O × 2 = 5 mol N
The empirical formula is N_2O_5.

Calculate the mass of carbon and the mass of hydrogen in 82.0 g of propane (C_3H_8).

ANALYZE List the known and the unknowns.

Knowns
mass of C_3H_8 = 82.0 g

Unknowns
mass of carbon = ? g C
mass of hydrogen = ? g H

CALCULATE Solve for the unknowns.

Write the conversion factor to convert from mass of C_3H_8 to mass of C. The percent by mass of C in C_3H_8 is 81.8%.

$$\frac{81.8\ \text{g C}}{100\ \text{g } C_3H_8}$$

Multiply the mass of C_3H_8 by the conversion factor.

$$82.0\ \cancel{g\,C_3H_8} \times \frac{81.8\ \text{g C}}{100\ \cancel{g\,C_3H_8}} = 67.1\ \text{g C}$$

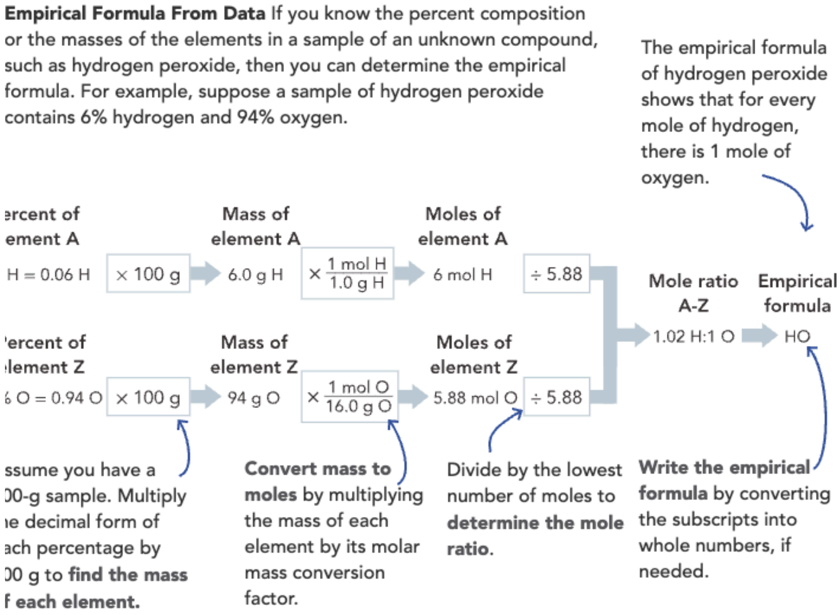
Write the conversion factor to convert from mass of C_3H_8 to mass of H. The percent by mass of H in C_3H_8 is 18.2%.

$$\frac{18.2\ \text{g H}}{100\ \text{g } C_3H_8}$$

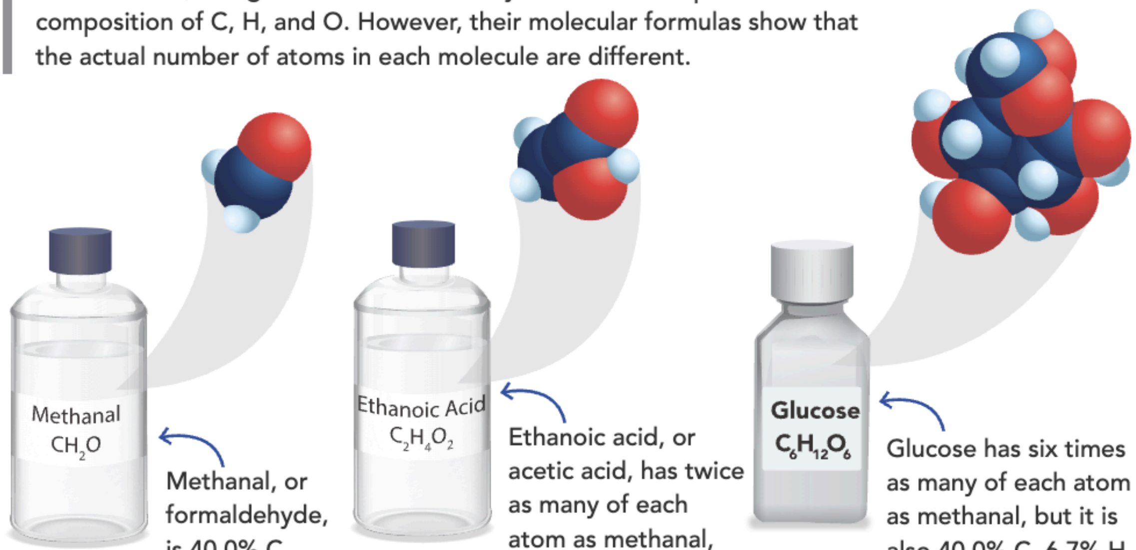
Multiply the mass of C_3H_8 by the conversion factor.

$$82.0\ \cancel{g\,C_3H_8} \times \frac{18.2\ \text{g H}}{100\ \cancel{g\,C_3H_8}} = 14.9\ \text{g H}$$

Empirical Formula From Data If you know the percent composition or the masses of the elements in a sample of an unknown compound, such as hydrogen peroxide, then you can determine the empirical formula. For example, suppose a sample of hydrogen peroxide contains 6% hydrogen and 94% oxygen.



Empirical vs. Molecular Formula The empirical formula of methanal, ethanoic acid, and glucose shows that they have the same percent composition of C, H, and O. However, their molecular formulas show that the actual number of atoms in each molecule are different.



Methanal, or formaldehyde, is 40.0% C, 6.7% H, and 53.3% O.

Ethanoic acid, or acetic acid, has twice as many of each atom as methanal, but it is also 40.0% C, 6.7% H, and 53.3% O.

Glucose has six times as many of each atom as methanal, but it is also 40.0% C, 6.7% H, and 53.3% O.

Calculate the molecular formula of a compound whose molar mass is 60.0 g/mol and empirical formula is CH_4N.

ANALYZE List the knowns and the unknown.

Knowns
empirical formula = CH_4N
molar mass = 60.0g/ mol

Unknown
molecular formula = $C_?H_?N_?$

CALCULATE Solve for the unknown.

Calculate the empirical formula mass (EFM). It is the molar mass of the empirical formula.

EFM of CH_4N =
$12.0\text{g/mol} + 4(1.0\text{g/ mol}) + 14.0\text{g/mol} = 30.0\text{g/ mol}$

Divide the molar mass by the empirical formula mass to obtain a whole number.

$$\frac{\text{molar mass}}{\text{EFM}} = \frac{60.0\ \cancel{\text{g/mol}}}{30.0\ \cancel{\text{g/mol}}} = 2$$

Multiply the empirical formula subscripts by this value to get the molecular formula.

$(CH_4N) \times 2 = C_2H_8N_2$