10 Chemical Quantities

Big idea THE MOLE AND QUANTIFYING MATTER

10.1 The Mole: A Measurement of Matter

For students using the Foundation edition, assign problems 1–15.

Essential Understanding The mole represents a large number of very small particles.

Reading Strategy

Frayer Model The Frayer Model is a vocabulary development tool. The center of the diagram shows the concept being defined, while the quadrants around the concept are used for providing the details. Use this model when you want to understand a vocabulary term in more detail.

As you read Lesson 10.1, use the Frayer Model below. Complete the diagram for the term *mole*.

mole

Definition in your own words ***A mole is 6.02 × 10^{23} of anything.***	Facts/characteristics ***A mole contains Avogadro's number of anything. The mass of a mole of an element is the atomic mass of the element in grams.***
How moles are used ***Conversion factors can be used to change the number of particles to the number of moles or the number of moles to the number of particles.***	Examples ***1 mol Na = 23.0 g Na = 6.02 × 10^{23} Na atoms.*** ***1 mol H_2O = 18.0 g H_2O = 6.02 × 10^{23} H_2O molecules.***

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EXTENSION Use the Frayer Model to complete a diagram for the term *molar mass.*

molar mass

Definition in your own words *Molar mass is the mass in grams of 1 mole of an element or compound.*	**Facts/characteristics** *The molar mass of an element or compound contains 6.02×10^{23} particles. To find the molar mass of a compound, you must start with its formula.*
How molar mass is used *Chemists can measure samples in grams and use molar mass to find the number of particles in the sample.*	**Examples** *molar mass of Na = 23.0 g Na; molar mass of H_2O = 18.0 g H_2O*

Lesson Summary

Measuring Matter Matter can be measured in three ways: by count, by mass, by volume.

- Dimensional analysis is a tool for solving conversion problems.

What Is a Mole? A mole is Avogadro's number (6.02×10^{23}) of representative particles of something.

- A representative particle is the basic unit of the material, usually atoms, molecules, or formula units.
- The conversion factor *1 mol/6.02×10^{23} representative particles* can be used to find the number of moles when the number of representative particles is known.
- The conversion factor *6.02×10^{23} representative particles/1 mol* can be used to find the number of representative particles when the number of moles is known.

Molar Mass The molar mass of a substance is the mass in grams of a mole of that substance.

- For an element, molar mass is the atomic mass in grams.
- For a compound, molar mass is the sum of the atomic masses of each atom in a representative particle of the compound, in grams.

After reading Lesson 10.1, answer the following questions.

Measuring Matter

1. What do the questions “how much?” and “how many?” have in common?

They are questions about the amount of a substance and are similar to questions scientists ask.

2. List two or three ways to measure matter.

count the matter; measure the mass or weight; measure the volume

What Is a Mole?

3. Circle the letter of the term that is an SI unit for measuring the amount of a substance.

a. dozen **b.** ounce **c.** pair [✓ **d.**] mole

4. What is Avogadro’s number?

6.02×10^{23} representative particles of a substance

5. Circle the letter of the term that is NOT a representative particle of a substance.

a. molecule **b.** atom [✓ **c.**] grain **d.** formula unit

6. List the representative particle for each of the following types of substances.

a. molecular compounds *molecule*

b. ionic compounds *formula unit*

c. elements *atom*

7. Is the following sentence true or false? To determine the number of representative particles in a compound, you count the molecules by viewing them under a microscope. *false*

8. How can you determine the number of atoms in a mole of a molecular compound?

Use the chemical formula to find the number of atoms in one molecule and multiply this number by Avogadro’s number, the number of particles in one mole.

9. Complete the table about representative particles and moles.

Representative Particles and Moles			
	Representative Particle	**Chemical Formula**	**Representative Particles in 1.00 mol**
Atomic oxygen	*atom*	O	*6.02×10^{23}*
Oxygen gas	molecule	*O_2*	*6.02×10^{23}*
Sodium ion	*ion*	*Na^+*	*6.02×10^{23}*
Sodium chloride	*formula unit*	*$NaCl$*	*6.02×10^{23}*

Molar Mass

10. What is the atomic mass of an element?

The atomic mass of an element is the mass of a single atom in atomic mass units.

11. Circle the letter of the phrase that completes this sentence correctly.

The atomic masses of all elements

a. are the same.

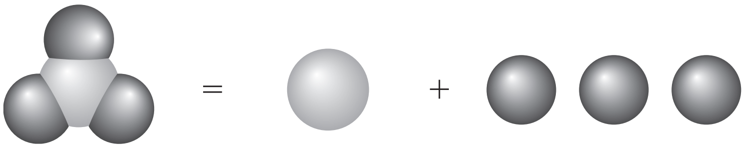
[✓ **b.**] are based on the mass of the carbon isotope C-12.

c. are based on the mass of a hydrogen atom.

12. How do you determine the mass of a mole of a compound?

The mass of a mole of a compound is determined by adding the atomic masses of the atoms making up the molecule.

13. Complete the labels on the diagram below.



SO_3 = 1 S atom + 3 O atoms

80.1 amu = 32.1 amu + *48.0* amu

14. What is the molar mass of a compound?

It is the mass of 1 mol of that compound.

15. Is the following sentence true or false? Molar masses can be calculated directly from atomic masses expressed in grams. *true*

10.2 Mole-Mass and Mole-Volume Relationships

For students using the Foundation edition, assign problems 1–12, 15–16.

Essential Understanding A mole always contains the same number of particles. But moles of different substances have different masses.

Lesson Summary

The Mole-Mass Relationship The molar mass of a substance can be used to convert between a sample's mass and the number of moles it contains.

- When a sample's mass is known, find the number of moles by multiplying the mass by the conversion factor *1 mol/molar mass.*
- When the number of moles is known, find the sample's mass by multiplying the number of moles by the conversion factor *molar mass/1 mol.*

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The Mole-Volume Relationship Mole-volume relationships are based on Avogadro's hypothesis, which says that equal volumes of gases at the same temperature and pressure contain equal numbers of particles.

- The volume of a gas is usually given at a standard temperature, 0°C, and a standard pressure, 1 atm or 101.3 kPa.
- At standard temperature and pressure (STP), a mole of gas occupies a volume of 22.4 liters.
- The quantities *1 mol* and *22.4 L* can be used in conversion factors that change moles to volume and volume to moles at STP.
- The molar mass of a gas can be found by multiplying its density at STP (in units of g/L) by *22.4 L/1 mol*.

BUILD Math Skills

Converting Between Mass and Moles When converting between mass and moles it is important to understand what a mole is. A mole is a representation of how many particles a sample has. The number of moles can be expressed as a coefficient in front of the compound or element in a chemical equation. For example, the chemical equation for the formation of aluminum oxide is:

$$4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$$

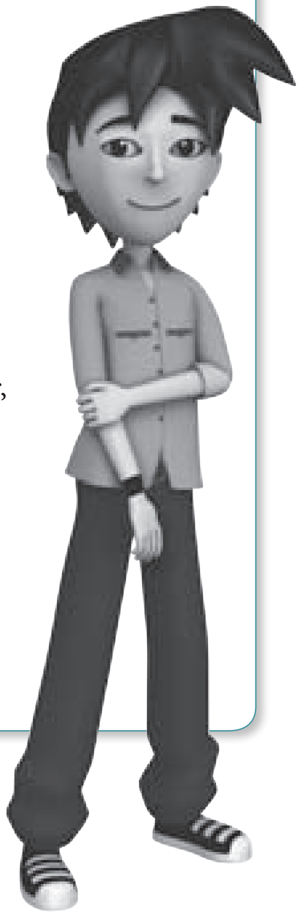
From this equation we can see that there are 4 moles of aluminum, 3 moles of oxygen, and 2 moles of aluminum oxide.

The mass of a mole of a compound is equal to the total mass of all the elements of the compound. Each element has an atomic mass that can be found in the periodic table. For example, the mass of one mole of NO_2 would be equal to the atomic mass of nitrogen plus twice the atomic mass of oxygen, or 46.01 g.

The conversion factor can be written as 1 mole $\frac{NO_2}{46.01\text{ g}}$. Remember, you multiply whatever you start with by any number on top, and you divide by any number on the bottom.

To convert from moles to mass or mass to moles, follow these simple steps:

- Determine the number of moles or grams in the given substance.
- Total the atomic masses for all the elements of any compound.
- Use the conversion process to get to the desired units.



Lesson 10.2 Reading and Study Workbook •

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Sample Problem Determine how many kilograms of water result from the following reaction, $2H_2 + O_2 \rightarrow 2H_2O$

Use the chemical equation to determine how many moles of water are present.

The coefficient in front of H_2O is 2, so 2 moles of water are present.

Total the mass for all elements that make up one mole of water.

Mass of H: 1.01 g

Mass of O: 16 g

Total Mass: $(1.01 \times 2) + 16 = 18.02$ g

Determine how many grams are present in 2 moles of water.

$$2 \text{ moles of } H_2O \times \frac{18.02 \text{ g}}{1 \text{ mole}} \text{ of } H_2O$$

$$= 36.04 \text{ g}$$

Convert 36.04 g of water to kilograms using the conversion method.

$$36.04 \text{ g } H_2O \times \frac{1 \text{ kg}}{1000 \text{ g}}$$

$$= 0.03604 \text{ kg of } H_2O$$

Hint: Remember that the atomic masses for the elements are given in grams, but the problem may require an answer in kilograms (kg) or other units. There are 1000 grams in 1 kilogram.

Now it's your turn to practice converting from moles to mass and mass to moles. Remember to multiply by numbers on the top and divide by numbers on the bottom of the conversion factor.

1. Determine how many moles are present in 0.23 kg of SO_2.

3.59 moles

2. How many grams of sodium chloride, NaCl, result from the reaction shown in the following equation, $FeCl_3 + 3NaOH \rightarrow Fe(OH)_3 + 3NaCl$?

175.33 g

3. Determine how many moles are present in 523.46 g of glucose, $C_6H_{12}O_6$.

2.91 moles

4. How many kilograms are in 4 moles of Na_2CO_3?

0.424 kg

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After reading Lesson 10.2, answer the following questions.

The Mole-Mass Relationship

5. What is the molar mass of a substance?

It is the mass (in grams) of one mole of the substance.

6. What is the molar mass of KI (potassium iodide)?

39.1 g K + 126.9 g I = 166.0 g KI

The Mole-Volume Relationship

7. Is the following sentence true or false? The volumes of one mole of different solid and liquid substances are the same. ***false***

8. Circle the letter of each term that can complete this sentence correctly. The volume of a gas varies with a change in

[✓ **a.**] temperature.

[✓ **b.**] the size of the container.

[✓ **c.**] pressure.

d. the amount of light in the container.

9. Circle the letter of the temperature that is defined as standard temperature.

a. 0 K

b. 100 K

[✓ **c.**] 0°C

d. 100°C

10. Is the following sentence true or false? Standard pressure is 101.3 kPa, or 1 atmosphere (atm). ***true***

11. What is the molar volume of a gas at standard temperature and pressure (STP)?
22.4 L

12. What units do you normally use to describe the density of a gas?
grams per liter (g/L)

13. What is Avogadro's hypothesis?

Avogadro's hypothesis says that equal volumes of gases at the same temperature and pressure contain equal numbers of particles.

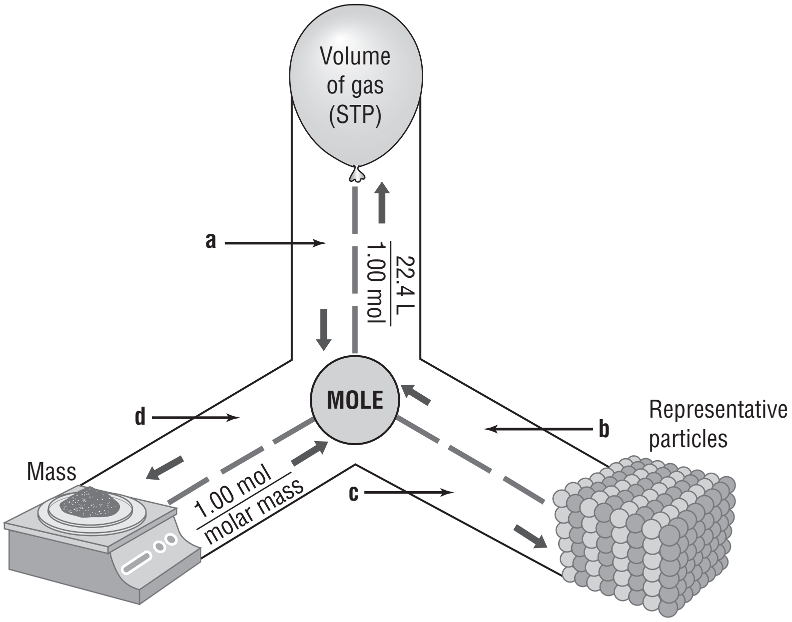
14. Look at Figure 10.7 to help you answer this question. Why is Avogadro's hypothesis reasonable?

As long as the gas particles are not tightly packed, there is a great deal of empty space between them. A container can easily accommodate the same number of relatively large or relatively small gas particles.

15. How many gas particles occupy a volume of 22.4 L at standard temperature and pressure?

6.02×10^{23} particles

16. The figure below shows how to convert from one unit to another unit. Write the missing conversion factors below.



a. $\frac{1.00 \text{ mol}}{22.4 \text{ L}}$

c. $\frac{6.02 \times 10^{23} \text{ particles}}{1.00 \text{ mol}}$

b. $\frac{1.00 \text{ mol}}{6.02 \times 10^{23} \text{ particles}}$

d. $\frac{\text{molar mass}}{1.00 \text{ mol}}$

10.3 Percent Composition and Chemical Formulas

For students using the Foundation edition, assign problems 1–6, 8–13.

Essential Understanding A molecular formula of a compound is a whole-number multiple of its empirical formula.

Lesson Summary

Percent Composition of a Compound Percent composition is the percent by mass of each element in a compound.

- To find the percent by mass of an element in a compound, use the formula:

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

- To find the mass of an element in a sample of a compound, use the formula:

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

Empirical Formulas The empirical formula of a compound is the formula with the smallest whole-number mole ratio of the elements.

- An empirical formula may or may not be the same as the actual molecular formula.

Molecular Formulas A molecular formula specifies the actual number of atoms in each element in one molecule or formula unit of the substance.

- ▶ To find a molecular formula, the molar mass of the compound must be determined.

After reading Lesson 10.3, answer the following questions.

Percent Composition of a Compound

1. How do you express relative amounts of each element in a compound?

Relative amounts are expressed by the percent composition or the percent by mass.

2. Circle the letter of the phrase that completes this sentence correctly. The number of percent values in the percent composition of a compound is

a. half as many as there are different elements in the compound.

[✓ **b.**] as many as there are different elements in the compound.

c. twice as many as there are different elements in the compound.

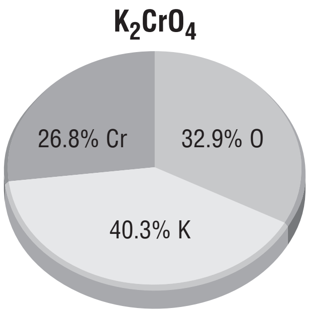
3. What is the formula for the percent by mass of an element in a compound?

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

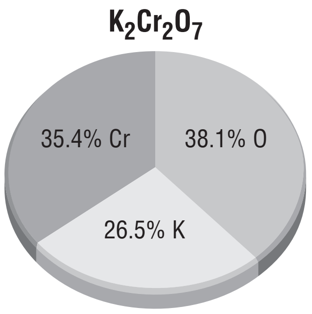
4. In the diagram below, which compound has a greater percent composition of chromium?

potassium dichromate, or $K_2Cr_2O_7$

How many more percentage points is this? *8.6%*



Potassium chromate, K_2CrO_4



Potassium dichromate, $K_2Cr_2O_7$

5. To calculate the percent composition of a known compound, start with the chemical formula of the compound and calculate the *molar mass*, which gives the mass of one mole of the compound.

6. Is the following sentence true or false? You can use percent composition to calculate the number of grams of an element in a given amount of a compound.

true

7. How do you calculate the grams of an element in a specific amount of a compound?

Multiply the mass of the compound by a conversion factor that is based on the percent composition.

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Empirical Formulas

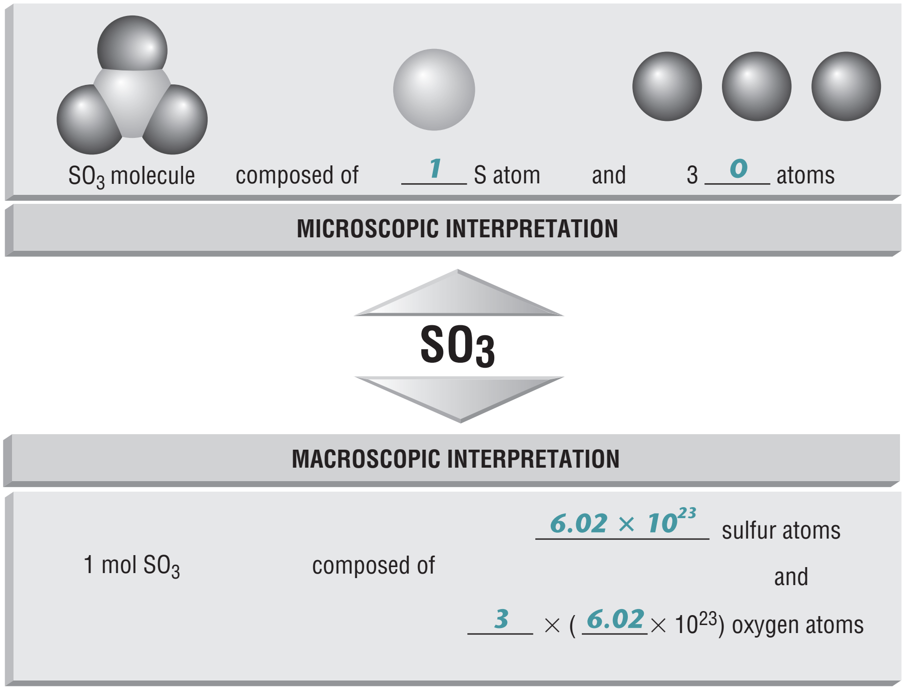
8. An empirical formula of a compound gives the ***lowest*** whole-number ratio of the atoms of the elements in a compound.

9. Is the following sentence true or false? The empirical formula of a compound is always the same as the molecular formula. ***false***

10. Look at Figure 10.11 and Table 10.3. Name three compounds that have an empirical formula of CH.

ethyne, styrene, benzene

11. Fill in the labels on the diagram below.



Molecular Formulas

12. The molecular formula of a compound is either the same as its empirical formula or a ***whole-number multiple*** of it.

13. What do you need to know to calculate the molecular formula of a compound?

You need the empirical formula of the compound and its molar mass.

14. If you divide the molar mass of a compound by the empirical formula mass, what is the result?

The result is the number of empirical formula units in a molecule of the compound.

15. What factor would you use to convert the empirical formula of a compound to a molecular formula?

the number of empirical formula units in a molecule of the compound

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Guided Practice Problems

Answer the following questions about Practice Problem 1.

If 0.20 bushel is 1 dozen apples and a dozen apples has a mass of 2.0 kg, what is the mass of 0.50 bushel of apples?

Analyze

Step 1. List the knowns and the unknown.

Knowns	**Unknown**
number of bushels = 0.5 bushel	*mass of 0.5 bushel of apples = ? kg*
0.20 bushel = 1 dozen apples	
1 dozen apples = 2.0 kg	

Use dimensional analysis to convert the number of bushels to the mass of apples, by following this sequence of conversions:

number of bushels → dozens of apples → mass of apples

Calculate

Step 2. Solve for the unknown.

The first conversion factor is $\frac{\text{1 dozen apples}}{\text{0.20 bushel}}$.

The second conversion factor is $\frac{\text{2.0 kg apples}}{\text{1 dozen apples}}$.

Multiplying the number of bushels by these two conversion factors gives the answer in kilograms.

$$\text{mass of apples} = 0.50\ \cancel{\text{bushel}} \times \frac{\cancel{\text{1 dozen apples}}}{0.20\ \cancel{\text{bushel}}} \times \frac{\text{2.0 kg apples}}{1\ \cancel{\text{dozen apples}}} = \text{5.0 kg apples}$$

The mass of 0.50 bushel of apples is *5.0 kg*.

Evaluate

Step 3. Does the result make sense?

Because a dozen apples is 2.0 kg and 0.5 bushel is more than two dozen but less than three dozen, the mass should be more than 4 kg (2 dozen × 2.0 kg) and less than 6 kg (3 dozen × 2.0 kg).

Answer the following questions about Practice Problem 3.

How many moles is 2.80×10^{24} atoms of silicon?

Step 1. List what you know.

2.80×10^{24} atoms of Si

$\boxed{\mathit{6.02 \times 10^{23}}}$ atoms in one mole

Step 2. Multiply the atoms of silicon by a mol/atoms conversion factor.

$$2.80 \times 10^{24} \text{ atoms Si} \times \frac{1 \text{ mol}}{\boxed{\mathit{6.02 \times 10^{23}}} \text{ atoms Si}}$$

Step 3. Divide.

$\boxed{\mathit{4.65}}$ mol

Answer the following questions about Practice Problem 5.

How many atoms are in 1.14 mol of sulfur trioxide (SO_3)?

Analyze

Step 1. List the knowns and the unknown.

Knowns

number of moles = 1.14 mol SO_3

1 mol SO_3 = 6.02×10^{23} molecules SO_3

1 molecule SO_3 = 4 atoms (1 S atom and 3 O atoms)

Unknown

1.14 mol SO_3 = ? atoms

Calculate

Step 2. Solve for the unknown.

The first conversion factor is $\dfrac{6.02 \times 10^{23} \text{ molecules of } SO_3}{1 \text{ mol } SO_3}$.

The second conversion factor is $\dfrac{4 \text{ atoms}}{1 \text{ molecule } SO_3}$.

Multiply moles of SO_3 by these conversion factors:

$$\text{number of atoms} = 1.14 \cancel{\text{mol } SO_3} \times \frac{6.02 \times 10^{23} \cancel{\text{molecules of } SO_3}}{\cancel{1 \text{ mol } SO_3}} \times \frac{4 \text{ atoms}}{\cancel{1 \text{ molecule } SO_3}}$$

$= \underline{\mathit{2.75 \times 10^{24} \text{ atoms}}}$

Evaluate

Step 3. Does the result make sense?

Because there are 4 atoms in a molecule of SO_3 and there is a little more than one mole of molecules, the answer should be more than 4 times Avogadro's number of atoms.

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Answer the following questions about Practice Problem 7.

Find the molar mass of PCl_3.

Analyze

Step 1. List the knowns and the unknown.

Knowns	**Unknown**
Molecular formula = PCl_3	*molar mass PCl_3 = ? g*
1 molar mass P = 31.0 g P	
1 molar mass Cl = 35.5 g Cl	

Calculate

Step 2. Solve for the unknown.

Convert moles of phosphorus and chlorine to grams of phosphorus and chlorine. Then add to get the results.

$$\cancel{1\text{ mol P}} \times \frac{31.0\text{ g P}}{\cancel{1\text{ mol P}}} = 31.0\text{ g P}$$

$$\cancel{3\text{ mol Cl}} \times \frac{35.5\text{ g Cl}}{\cancel{1\text{ mol Cl}}} = 106.5\text{ g Cl}$$

molar mass of PCl_3 = *137.5 g*

Evaluate

Step 3. Does the result make sense?

The answer is the sum of the molar mass of phosphorus and three times the molar mass of chlorine, expressed to the tenths decimal place.

Answer the following questions about Practice Problem 16.

Find the mass, in grams, of 4.52×10^{-3} mol $C_{20}H_{42}$.

Analyze

Step 1. List the known and the unknown.

Known	**Unknown**
number of moles = 4.52×10^{-3} mol $C_{20}H_{42}$	*mass = ? g $C_{20}H_{42}$*

Calculate

Step 2. Solve for the unknown.

Determine the molar mass of $C_{20}H_{42}$:

1 mol $C_{20}H_{42} = 20 \times 12.0\text{ g} + 42 \times 1.0\text{ g} = 282\text{ g}$

Multiply the given number of moles by the conversion factor:

$$\text{mass} = 4.52 \times 10^{-3}\ \cancel{\text{mol } C_{20}H_{42}} \times \frac{282\text{ g } C_{20}H_{42}}{\cancel{1\text{ mol } C_{20}H_{42}}} = \underline{\mathbf{1.27\ g\ C_{20}H_{42}}}$$

Evaluate

Step 3. Does the result make sense?

The amount of substance is a little more than four one-thousandths of a mole, so the mass should be only a small fraction of the molar mass.

Answer the following questions about Practice Problem 18.

Find the number of moles in 3.70×10^{-1} g of boron.

Analyze

Step 1. List the known and the unknown.

Known	**Unknown**
mass = 3.70×10^{-1} g boron	***number of moles = ? mol boron***

The unknown number of moles is calculated by converting the known mass to the number of moles using a conversion factor of mass → moles.

Calculate

Step 2. Solve for the unknown.

Determine the molar mass of boron:

1 mol B = 10.8 g B

Multiply the given mass by the conversion factor relating mass of boron to moles of boron:

$$\text{mass} = 3.70 \times 10^{-1}\ \cancel{\text{g B}} \times \frac{1\text{ mol B}}{10.8\ \cancel{\text{g B}}} = \underline{\mathbf{3.43 \times 10^{-2}\ mol\ B}}$$

Evaluate

Step 3. Does the result make sense?

Because the value of the conversion factor is about one-tenth, the numerical result should be about one-tenth of the given number of grams.

Answer the following questions about Practice Problems 20a and 20b.

What is the volume of these gases at STP?

a. 3.20×10^{-3} mol CO_2

b. 3.70 mol N_2

3.20×10^{-3} mol CO_2

Analyze

Step 1. List the knowns and the unknown.

Knowns	**Unknown**
number of moles = 3.20×10^{-3} mol CO_2	***volume = ? L CO_2***
1 mol CO_2 = 22.4 L CO_2	

To convert moles to liters, use the relationship 1 mol CO_2 = 22.4 L CO_2 (at STP).

Calculate

Step 2. Solve for the unknown.

Multiply the given number of moles of CO_2 by the conversion factor:

$$\text{volume} = 3.20 \times 10^{-3}\ \cancel{\text{mol CO}_2} \times \frac{22.4\ \text{L CO}_2}{1\ \cancel{\text{mol CO}_2}} = \underline{7.17 \times 10^{-2}\ \text{L CO}_2}$$

Evaluate

Step 3. Does the result make sense?

Because a mole of gas occupies a volume of a little more than 20 liters, the result should be a little larger than twenty times the given number of moles.

3.70 mol N_2

Analyze

Step 1. List the knowns and the unknown.

Knowns	**Unknown**
number of moles = 3.70 mol N_2	***volume = ? L N_2***
1 mol N_2 = 22.4 L N_2	

Use the relationship 1 mol N_2 = 22.4 L N_2 (at STP) to convert moles to liters.

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Calculate

Step 2. Solve for the unknown.

Multiply the given number of moles of N_2 by the conversion factor:

$$\text{volume} = 3.70\ \cancel{\text{mol N}_2} \times \frac{22.4\ \text{L N}_2}{1\ \cancel{\text{mol N}_2}} = \underline{\mathbf{82.9\ L\ N_2}}$$

Evaluate

Step 3. Does the result make sense?

Because the number of moles is slightly less than four, the result should be close to, but less than 88 L.

Answer the following questions about Practice Problem 22.

A gaseous compound composed of sulfur and oxygen has a density of 3.58 g/L at STP. What is the molar mass of this gas?

Analyze

Step 1. List the knowns and the unknown.

Knowns	**Unknown**
density = 3.58 g/L	***molar mass = ? g***
1 mol (gas at STP) = 22.4 L	

To convert density (g/L) to molar mass (g/mol), a conversion factor of L/mol is needed.

Calculate

Step 2. Solve for the unknown.

Multiply the density by the conversion factor relating liters and moles:

$$\text{molar mass} = \frac{3.58\ \text{g}}{1\ \text{L}} \times \frac{22.4\ \text{L}}{1\ \text{mol}} = \underline{\mathbf{80.2\ g/mol}}$$

Evaluate

Step 3. Does the result make sense?

Multiplying approximately 4 grams/liter by approximately 20 liters/mole yields about 80 grams/mole.

Answer the following questions about Practice Problem 33.

A compound is formed when 9.03 g Mg combines completely with 3.48 g N. What is the percent composition of this compound?

Analyze

Step 1. List the knowns and the unknowns.

Knowns

mass of Mg = 9.03 g Mg

mass of N = −3.48 g N

mass of compound = 9.03 g Mg + 3.48 g N = 12.51 g

Unknowns

percent Mg = ? %

percent N = ? %

The percent of an element in a compound is the mass of the element in the compound divided by the mass of the compound. To be expressed as a percentage, the ratio must be multiplied by 100%.

Calculate

Step 2. Solve for the unknown.

$$\text{percent Mg} = \frac{9.03\text{ g Mg}}{12.51\text{ g compound}} \times 100\% = \underline{72.2\%\text{ Mg}}$$

$$\text{percent N} = \frac{3.48\text{ g N}}{12.51\text{ g compound}} \times 100\% = \underline{27.8\%\text{ N}}$$

Evaluate

Step 3. Does the result make sense?

The percents of the elements of the compound add up to 100%.

72.2% + 27.8% = 100%

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Answer the following questions about Practice Problem 36.

Calculate the percent composition of these compounds.

a. ethane (C_2H_6)
b. sodium hydrogen sulfate ($NaHSO_4$)

Ethane (C_2H_6)

Analyze

Step 1. List the knowns and the unknowns.

Knowns

mass of C in one mole ethane = 2 × 12.0 g = 24.0 g

mass of H in one mole ethane = 6 × 1.0 g = 6.0 g

molar mass of C_2H_6 = 24.0 g + 6.0 g = 30.0 g

Unknowns

percent C = ? %

percent H = ? %

Because no masses are given, the percent composition can be determined based on the molar mass of the substance. The percent of an element in a compound is the mass of the element in the compound divided by the mass of the compound. To express the ratio as a percent, the ratio is multiplied by 100%.

Calculate

Step 2. Solve for the unknown.

$$\text{percent C} = \frac{24.0\text{ g C}}{30.0\text{ g compound}} \times 100\% = \underline{80.0\%\ \text{C}}$$

$$\text{percent H} = \frac{6.0\text{ g H}}{30.0\text{ g compound}} \times 100\% = \underline{20.0\%\ \text{H}}$$

Evaluate

Step 3. Does the result make sense?

The percents of the elements of the compound add up to 100%.

80.0% + 20.0% = 100%

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Sodium hydrogen sulfate ($NaHSO_4$)

Analyze

Step 1. List the knowns and the unknowns.

Knowns

mass of Na in one mole sodium hydrogen sulfate = 1 × 23.0 g = 23.0 g

mass of H in one mole sodium hydrogen sulfate = 1 × 1.0 g = 1.0 g

mass of S in one mole sodium hydrogen sulfate = 1 × 32.1 g = 32.1 g

mass of O in one mole sodium hydrogen sulfate = 4 × 16.0 g = 64.0 g

molar mass of $NaHSO_4$ = 23.0 g + 1.0 g + 32.1 g + 64.0 g = 120.1 g

Unknowns

percent Na = ? %

percent H = ? %

percent S = ? %

percent O = ? %

Because no masses are given, the percent composition can be determined based on the molar mass of the substance. The percent of an element in a compound is the mass of the element in the compound divided by the mass of the compound. To express the ratio as a percent, the ratio is multiplied by 100%.

Calculate

Step 2. Solve for the unknown.

$$\text{percent Na} = \frac{23.0\text{ g Na}}{120.1\text{ g compound}} \times 100\% = \underline{19.2\%\text{ Na}}$$

$$\text{percent H} = \frac{1.0\text{ g H}}{120.1\text{ g compound}} \times 100\% = \underline{0.83\%\text{ H}}$$

$$\text{percent S} = \frac{32.1\text{ g S}}{120.1\text{ g compound}} \times 100\% = \underline{26.7\%\text{ S}}$$

$$\text{percent O} = \frac{64.0\text{ g O}}{120.1\text{ g compound}} \times 100\% = \underline{53.3\%\text{ O}}$$

Evaluate

Step 3. Does the result make sense?

The percents of the elements of the compound add up to 100%.

19.2% + 0.83% + 26.7% + 53.3% = 100%

Answer the following questions about Practice Problem 39.

Calculate the empirical formula of each compound.

a. 94.1% O, 5.9% H

b. 67.6% Hg, 10.8% S, 21.6% O

94.1% O, 5.9% H

Analyze

Step 1. List the knowns and the unknown.

Knowns	**Unknown**
percent composition: 94.1% O, 5.9% H	***empirical formula =*** $H_?O_?$
molar mass O = 16.0 g/mol O	
molar mass H = 1.0 g/mol H	

Use the percent composition to convert to mass, recalling that *percent* means parts per hundred. Then use the molar mass to convert to number of moles. Finally, determine whole-number ratios based on the number of moles of each element per 100 grams of compound.

Calculate

Step 2. Solve for the unknown.

One hundred grams of compound has 5.9 g H and 94.1 g O.

Multiply by conversion factors relating moles of the elements to grams:

$$5.9\ \cancel{\text{g H}} \times \frac{1\ \text{mol H}}{1.0\ \cancel{\text{g H}}} = 5.9\ \text{mol H}$$

$$94.1\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.0\ \cancel{\text{g O}}} = 5.88\ \text{mol O}$$

So, the mole ratio for 100 g of the compound is $H_{5.9}O_{5.9}$. But formulas must have whole number subscripts. Divide each molar quantity by the smaller number of moles. This will give 1 mol for the element with the smaller number of moles. In this case, the ratio is 1:1, so the empirical formula is simply H_1O_1. However, a subscript of 1 is never written, so the answer is ***HO***.

Evaluate

Step 3. Does the result make sense?

The subscripts are whole numbers and the percent composition of this empirical formula equals the percent composition given in the original problem.

67.6% Hg, 10.8% S, 21.6% O

Analyze

Step 1. List the knowns and the unknown.

Knowns

percent composition: 67.6% Hg, 10.8% S, 21.6% O

molar mass Hg = 200.6 g/mol Hg

molar mass S = 32.1 g/mol S

molar mass O = 16.0 g/mol O

Unknown

empirical formula = $Hg_?S_?O_?$

Use the percent composition to convert to mass. Then use molar mass to convert to number of moles. Finally, determine whole-number ratios based on the number of moles of each element per 100 grams of compound.

Calculate

Step 2. Solve for the unknown.

One hundred grams of compound has 67.6 g Hg, 10.8 g S, and 21.6 g O.

Multiply by a conversion factor relating moles to grams:

$$67.6\ \cancel{\text{g Hg}} \times \frac{1\ \text{mol Hg}}{200.6\ \cancel{\text{g Hg}}} = 0.337\ \text{mol Hg}$$

$$10.8\ \cancel{\text{g S}} \times \frac{1\ \text{mol S}}{32.1\ \cancel{\text{g S}}} = 0.336\ \text{mol S}$$

$$21.6\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.0\ \cancel{\text{g O}}} = 1.35\ \text{mol O}$$

So, the mole ratio for 100 g of the compound is $Hg_{0.34}S_{0.34}O_{1.35}$.

Divide each molar quantity by the smaller number of moles:

$$\frac{0.34\ \text{mol Hg}}{0.34} = 1\ \text{mol Hg}$$

$$\frac{0.34\ \text{mol S}}{0.34} = 1\ \text{mol S}$$

$$\frac{1.35\ \text{mol O}}{1.35} = 4\ \text{mol O}$$

The empirical formula is $HgSO_4$.

Evaluate

Step 3. Does the result make sense?

The subscripts are whole numbers and the percent composition of this empirical formula equals the percent composition given in the original problem.

Answer the following questions about Practice Problem 42.

Find the molecular formula of ethylene glycol, which is used as antifreeze. The molar mass is 62.0 g/mol and the empirical formula is CH_3O.

Analyze

Step 1. List the knowns and the unknown.

Knowns	**Unknown**
molar mass = 62 g/mol	***molecular formula =*** $C_?H_?O_?$
empirical formula = CH_3O	

Calculate

Step 2. Solve for the unknown.

First, calculate the empirical formula mass (efm):

$$1\ \cancel{\text{mol C}} \times \frac{12\text{ g C}}{1\ \cancel{\text{mol C}}} = 12\text{ g C}$$

$$3\ \cancel{\text{mol H}} \times \frac{1.0\text{ g H}}{1\ \cancel{\text{mol H}}} = 3\text{ g H}$$

$$1\ \cancel{\text{mol O}} \times \frac{16\text{ g O}}{1\ \cancel{\text{mol O}}} = 16\text{ g O}$$

So, efm = 12 g + 3 g + 16 g = 31 g.

Divide the molar mass by the empirical formula mass:

Molar mass/efm = 62 g/31 g = 2

Multiply subscripts in the empirical formula by this value.

The molecular formula is $C_2H_6O_2$.

Evaluate

Step 3. Does the result make sense?

The molecular formula has the given molar mass, and can be reduced to the empirical formula.

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Apply the Big idea

A student has a sample of CO_2 gas that has a mass of 22.0 g.

a. Explain how she would find the volume of the sample at STP.

She would have to use the molar mass of CO_2 (44.0 g) and the given mass to find the number of moles of the gas. She could then use the number of moles and molar volume to find the volume of the sample.

b. What is the volume of 22.0 g of CO_2 at STP? Show your work.

$$22.0\,\cancel{\text{g}} \times \frac{1\text{ mol}}{44.0\,\cancel{\text{g}}} = 0.50\text{ mol}$$

$$0.50\,\cancel{\text{mol}} \times \frac{22.4\text{ L}}{1\,\cancel{\text{mol}}} = 11.2\text{ L}$$

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10 Self-Check Activity

For Questions 1–9, complete each statement by writing the correct word or words. If you need help, you can go online.

10.1 The Mole: A Measurement of Matter

1. Knowing how the ***count*** ______, mass, and volume of an item relate to a common unit allows you to convert among these units.

2. Counting the representative particles in a sample of a substance is based on the ***mole*** ______.

3. The molar mass of an element is its ***atomic mass*** ______ expressed in units of grams.

4. The molar mass of a(n) ***compound*** ______ is the sum of the molar mass of all the elements in the ***compound*** ______ .

10.2 Mole-Mass and Mole-Volume Relationships

5. The ***molar mass*** ______ of a substance is used to convert from mass to moles or moles to mass.

6. The ***molar volume*** ______ of a gas is used to convert moles to volume or volume to moles at STP.

10.3 Percent Composition and Chemical Formulas

7. To find the percent by mass of an element in a compound, divide the mass of the element by the ***mass of the compound*** ______, then multiply by 100%.

8. The ***empirical formula*** ______ of a compound gives the lowest whole-number ratio of the atoms or moles of the elements in a compound.

9. The ***molecular formula*** ______ of a molecular compound is the same as, or a whole-number multiple of, the empirical formula of the compound.

If You Have Trouble With...

Question	1	2	3	4	5	6	7	8	9
See Page	307	308	313	314	317	320	326	330	332

Review Key Equations

Sequence the following steps a student would use to find the percent composition of glucose, $C_6H_{12}O_6$.

a. Divide the mass of each element by the molar mass.

b. Find the mass of each element in one mole of glucose.

c. Multiply each quotient by 100%.

d. Find the molar mass of glucose.

b, d, a, c or d, b, a, c

EXTENSION **Use the sequence you wrote to find the percent composition of glucose.**

mass of C in 1 mol $C_6H_{12}O_6$: 6 mol × 12.0 g/mol = 72.0 g

mass of H in 1 mol $C_6H_{12}O_6$: 12 mol × 1.0 g/mol = 12.0 g

mass of O in 1 mole $C_6H_{12}O_6$: 6 mol × 16.0 g/mol = 96.0 g

molar mass of $C_6H_{12}O_6$ = 72.0 g + 12.0 g + 96.0 g = 180.0 g

$$\%C = \frac{72.0\ g}{180.0\ g} \times 100\% = 40.0\%\ C$$

$$\%H = \frac{12.0\ g}{180.0\ g} \times 100\% = 6.7\%\ H$$

$$\%O = \frac{96.0\ g}{180.0\ g} \times 100\% = 53.3\%\ O$$

Review Vocabulary

For each term in column 1, write the letter of the best match from column 2.

f **1.** mole

c **2.** Avogadro's number

e **3.** representative particle

a **4.** molar mass

d **5.** Avogadro's hypothesis

i **6.** standard temperature and pressure (STP)

b **7.** molar volume

h **8.** percent composition

g **9.** empirical formula

a. the mass of the Avogadro number of representative particles

b. 22.4 L

c. 6.02×10^{23}

d. relates volumes of gases with numbers of particles

e. atoms, molecules, or formula units

f. an SI unit used to measure amount

g. the lowest whole-number ratio of types of atoms in a compound

h. must equal 100

i. allows you to compare gases under the same physical conditions