



تم تحميل هذا الملف من موقع المناهج الإماراتية

موقع المناهج ← المناهج الإماراتية ← الصف الحادي عشر المتقدم ← رياضيات ← الفصل الثالث ← ملفات متنوعة ← الملف

تاريخ إضافة الملف على موقع المناهج: 07:30:57 2026-06-05

ملفات اكتب للمعلم اكتب للطالب ا اختبارات الكترونية ا اختبارات ا حلول ا عروض بوربوينت ا أوراق عمل
منهج انجليزي ا ملخصات وتقارير ا مذكرات وبنوك ا الامتحان النهائي للمدرس

المزيد من مادة
رياضيات:

التواصل الاجتماعي بحسب الصف الحادي عشر المتقدم



صفحة المناهج
الإماراتية على
فيسبوك

الرياضيات

اللغة الانجليزية

اللغة العربية

التربية الاسلامية

المواد على تلغرام

المزيد من الملفات بحسب الصف الحادي عشر المتقدم والمادة رياضيات في الفصل الثالث

تحليل ورقة تدريب امتحان نهاية الفصل في الخلايا الجلفانية

1

منهج ريفيل حلول مراجعة امتحانية تدريبات وأسئلة

2

منهج ريفيل مراجعة امتحانية تدريبات وأسئلة غير محلول

3

ملزمة هيكل واختياري وأسئلة

4

هيكل اختبار نهاية الفصل 2026م

5

EOT EXAM PRACTICE

10 ADVANCE

TERM 3

2025-2026

BY MS.SHAMMA EBRAHIM ALMAZROUEI
SCHOOL PRINCEPLE : MOUZA RASHED KASEB



Particle and Mole Relationships

When a colorless sodium hydrogen sulfite (NaHSO_3) solution is added to a purple potassium permanganate (KMnO_4) solution, the purple color eventually disappears. Why do you think this happens? If you concluded that the potassium permanganate had been used up and the reaction had stopped, you are right. Chemical reactions stop when one of the reactants is used up. When planning the reaction of potassium permanganate and sodium hydrogen sulfite, a chemist might ask, "How many grams of potassium permanganate are needed to react completely with a known mass of sodium hydrogen sulfite?" Or, when analyzing a photosynthesis reaction, you might ask, "How much oxygen and carbon dioxide are needed to form a known mass of sugar?"

Stoichiometry

The study of quantitative relationships between the amounts of reactants used and amounts of products formed by a chemical reaction is called **stoichiometry**. Stoichiometry is based on the law of conservation of mass. Recall that the law states that matter is neither created nor destroyed in a chemical reaction, but can change form. In any chemical reaction, the amount of matter present at the end of the reaction is the same as the amount of matter present at the beginning, given that atoms are conserved during a chemical reaction. Therefore, the mass of the reactants equals the mass of the products. Note the reaction of powdered iron (Fe) with oxygen (O_2) shown in **Figure 1**. Although iron reacts with oxygen to form a new compound, iron(III) oxide (Fe_2O_3), the total mass is unchanged.



Figure 1 The balanced chemical equation for this reaction between iron and oxygen provides the relationships between the amounts of reactants and products.

According to the law of conservation of mass, the total mass of reactants in a chemical reaction is:

- A) Greater than the mass of products
- B) Less than the mass of products
- C) Equal to the mass of products
- D) Unrelated to the mass of products

The disappearance of the purple color of KMnO_4 when NaHSO_3 is added indicates that:

- A) KMnO_4 concentration increased
- B) One of the reactants has been completely used up
- C) No reaction occurred
- D) A gas was produced

The quantitative relationship between reactants and products in a chemical reaction is called:

- A) Thermodynamics
- B) Organic chemistry
- C) Stoichiometry
- D) Analytical chemistry

Stoichiometry is based on which scientific law?

- A) Law of definite proportions
- B) Ideal gas law
- C) Law of multiple proportions
- D) Law of conservation of mass

Which of the following is an application of stoichiometry?

- A) Determining the color of the product
- B) Measuring the temperature of the reaction
- C) Calculating the amount of oxygen needed to produce a known mass of sugar in photosynthesis
- D) Identifying the shape of a molecule

Table 1 Relationships Derived from a Balanced Chemical Equation

4Fe(s)	+	3O ₂ (g)	→	2Fe ₂ O ₃ (s)
Iron	+	oxygen	→	iron (III) oxide
4 atoms of Fe	+	3 molecules O ₂	→	2 formula units of Fe ₂ O ₃
4 mol Fe	+	3 mol O ₂	→	2 mol Fe ₂ O ₃
223.4 g Fe	+	96.00 g O ₂	→	319.4 g Fe ₂ O ₃
319.4 g reactants			→	319.4 g products

The balanced chemical equation for the chemical reaction shown in **Figure 1** is as follows.



You can interpret this equation in terms of representative particles by saying that four atoms of iron react with three molecules of oxygen to produce two formula units of iron(III) oxide. Remember that coefficients in an equation represent not only numbers of individual particles but also numbers of moles of particles. Therefore, you can also say that four moles of iron react with three moles of oxygen to produce two moles of iron(III) oxide.

The chemical equation does not directly tell you anything about the masses of the reactants and products. However, by converting the known mole quantities to mass, the mass relationships become obvious. Recall that moles are converted to mass by multiplying by the molar mass. The masses of the reactants are as follows.

$$4 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 223.4 \text{ g Fe}$$

$$3 \text{ mol O}_2 \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = 96.00 \text{ g O}_2$$

The total mass of the reactants is: (223.4 g + 96.00 g) = 319.4 g. Similarly, the mass of the product is calculated as follows:

$$2 \text{ mol Fe}_2\text{O}_3 \times \frac{159.7 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = 319.4 \text{ g Fe}_2\text{O}_3$$

Note that the mass of the reactants equals the mass of the product.

$$\text{mass of reactants} = \text{mass of products}$$

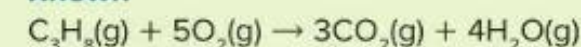
$$319.4 \text{ g} = 319.4 \text{ g}$$

EXAMPLE Problem 1

INTERPRETING CHEMICAL EQUATIONS The combustion of propane (C₃H₈) provides energy for heating homes, cooking food, and soldering metal parts. Interpret the equation for the combustion of propane in terms of representative particles, moles, and mass. Show that the law of conservation of mass is observed.

1 ANALYZE THE PROBLEM

The coefficients in the balanced chemical equation shown below represent both moles and representative particles, in this case molecules. Therefore, the equation can be interpreted in terms of molecules and moles. The law of conservation of mass will be verified if the masses of the reactants and products are equal.

Known**Unknown**

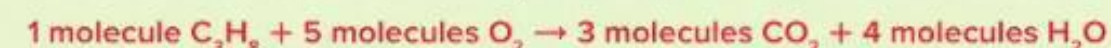
Equation interpreted in terms of molecules = ?

Equation interpreted in terms of moles = ?

Equation interpreted in terms of mass = ?

2 SOLVE FOR THE UNKNOWN

The coefficients in the chemical equation indicate the number of molecules.



The coefficients in the chemical equation also indicate the number of moles.



To verify that mass is conserved, first convert moles of reactant and product to mass by multiplying by a conversion factor—the molar mass—that relates grams to moles.

$$\text{moles of reactant or product} \times \frac{\text{grams reactant or product}}{1 \text{ mol reactant or product}} = \text{grams of reactant or product}$$

$$1 \text{ mol C}_3\text{H}_8 \times \frac{44.09 \text{ g C}_3\text{H}_8}{1 \text{ mol C}_3\text{H}_8} = 44.09 \text{ g C}_3\text{H}_8$$
 Calculate the mass of the reactant C₃H₈.

$$5 \text{ mol O}_2 \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = 160.0 \text{ g O}_2$$
 Calculate the mass of the reactant O₂.

$$3 \text{ mol CO}_2 \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 132.0 \text{ g CO}_2$$
 Calculate the mass of the product CO₂.

$$4 \text{ mol H}_2\text{O} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 72.08 \text{ g H}_2\text{O}$$
 Calculate the mass of the product H₂O.

$$44.09 \text{ g C}_3\text{H}_8 + 160.0 \text{ g O}_2 = \mathbf{204.1 \text{ g reactants}}$$
 Add the masses of the reactants.

$$132.0 \text{ g CO}_2 + 72.08 \text{ g H}_2\text{O} = \mathbf{204.1 \text{ g products}}$$
 Add the masses of the products.

$$\mathbf{204.1 \text{ g reactants} = 204.1 \text{ g products}}$$

The law of conservation of mass is observed.

PRACTICE Problems



ADDITIONAL PRACTICE

1. Interpret the following balanced chemical equations in terms of particles, moles, and mass. Show that the law of conservation of mass is observed.

- a. $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$
 b. $\text{HCl}(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow \text{KCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 c. $2\text{Mg}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2\text{MgO}(\text{s})$

2. **CHALLENGE** For each of the following, balance the chemical equation; interpret the equation in terms of particles, moles, and mass; and show that the law of conservation of mass is observed.

- a. $\text{Na}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{NaOH}(\text{aq}) + \text{H}_2(\text{g})$
 b. $\text{Zn}(\text{s}) + \text{HNO}_3(\text{aq}) \rightarrow \text{Zn}(\text{NO}_3)_2(\text{aq}) + \text{N}_2\text{O}(\text{g}) + \text{H}_2\text{O}(\text{l})$



2 atoms Mg + 1 molecule

$\text{O}_2 \rightarrow 2$ formula units MgO

2 moles Mg + 1 mole $\text{O}_2 \rightarrow 2$ moles MgO

Mass:

$2\text{Mg}: (2 \text{ mol})(24.305 \text{ g/mol}) = 48.610 \text{ g}$

$\text{O}_2: (2 \text{ mol})(15.999 \text{ g/mol}) = 31.998 \text{ g}$

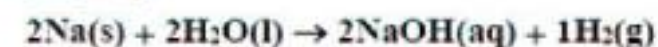
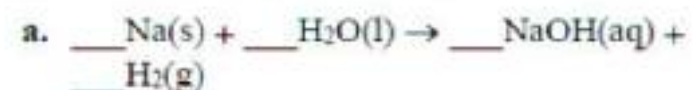
$2\text{MgO}: (2 \text{ mol})(24.305 \text{ g/mol}) +$
 $(2 \text{ mol})(15.999 \text{ g/mol}) = 80.608 \text{ g}$

$48.610 \text{ g Mg} + 31.998 \text{ g O}_2 \rightarrow$

80.608 g MgO

$80.608 \text{ g reactants} = 80.608 \text{ g products}$

2. **Challenge** For each of the following, balance the chemical equation; interpret the equation in terms of particles, moles, and mass; and show that the law of conservation of mass is observed.



2 atoms Na + 2 molecules $\text{H}_2\text{O} \rightarrow$

2 formula units NaOH + 1 molecule H_2

2 mol Na + 2 mol $\text{H}_2\text{O} \rightarrow$

2 mol NaOH + 1 mol H_2

Mass:

$2\text{Na}: (2 \text{ mol})(22.990 \text{ g/mol}) = 45.980 \text{ g}$

$2\text{H}_2\text{O}: (4 \text{ mol})(1.008 \text{ g/mol}) +$
 $(2 \text{ mol})(15.999 \text{ g/mol}) = 36.030 \text{ g}$

$2\text{NaOH}: (2 \text{ mol})(22.990 \text{ g/mol}) +$
 $(2 \text{ mol})(15.999 \text{ g/mol}) + (2 \text{ mol})(1.008 \text{ g/mol}) =$
 79.994 g

1. Interpret the following balanced chemical equations in terms of particles, moles, and mass. Show that the law of conservation of mass is observed.



1 molecule N_2 + 3 molecules

$\text{H}_2 \rightarrow 2$ molecules NH_3

1 mole N_2 + 3 moles $\text{H}_2 \rightarrow 2$ moles NH_3

Mass:

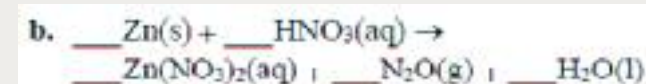
$\text{N}_2: (2 \text{ mol})(14.007 \text{ g/mol}) = 28.014 \text{ g}$

$3\text{H}_2: (6 \text{ mol})(1.008 \text{ g/mol}) = 6.048 \text{ g}$

$2\text{NH}_3: (2 \text{ mol})(14.007 \text{ g/mol}) +$
 $(6 \text{ mol})(1.008 \text{ g/mol}) = 34.062 \text{ g}$

$28.014 \text{ g N}_2 + 6.048 \text{ g H}_2 \rightarrow 34.062 \text{ g NH}_3$

$34.062 \text{ g reactants} = 34.062 \text{ g products}$



4 atoms Zn + 10 molecules $\text{HNO}_3 \rightarrow$

4 formula units $\text{Zn}(\text{NO}_3)_2$ + 1 molecule N_2O + 5 molecules H_2O

4 mol Zn + 10 mol $\text{HNO}_3 \rightarrow$

4 mol $\text{Zn}(\text{NO}_3)_2$ + 1 mol N_2O + 5 mol H_2O

Mass:

$4\text{Zn}: (4 \text{ mol})(65.39 \text{ g/mol}) = 261.56 \text{ g}$

$10\text{HNO}_3: (10 \text{ mol})(1.008 \text{ g/mol}) +$
 $(10 \text{ mol})(14.007 \text{ g/mol}) +$
 $(30 \text{ mol})(15.999 \text{ g/mol}) = 630.12 \text{ g}$

$4\text{Zn}(\text{NO}_3)_2: (4 \text{ mol})(65.39 \text{ g/mol}) +$
 $(8 \text{ mol})(14.007 \text{ g/mol}) +$
 $(24 \text{ mol})(15.999 \text{ g/mol}) = 757.592 \text{ g}$

$\text{N}_2\text{O}: (2 \text{ mol})(14.007 \text{ g/mol}) +$
 $(1 \text{ mol})(15.999 \text{ g/mol}) = 44.013 \text{ g}$

$5\text{H}_2\text{O}: (10 \text{ mol})(1.008 \text{ g/mol}) +$
 $(5 \text{ mol})(15.999 \text{ g/mol}) = 90.075 \text{ g}$

$261.56 \text{ g Zn} + 630.12 \text{ g HNO}_3 \rightarrow$
 $757.592 \text{ g Zn}(\text{NO}_3)_2 + 44.013 \text{ g N}_2\text{O} +$
 $90.075 \text{ g H}_2\text{O}$

$891.68 \text{ g reactants} = 891.68 \text{ g products}$



1 molecule HCl + 1 formula unit KOH $\rightarrow$

1 formula unit KCl + 1 molecule H_2O

1 mole HCl + 1 mole KOH $\rightarrow$

1 mole KCl + 1 mole H_2O

Mass:

$\text{HCl}: (1 \text{ mol})(1.008 \text{ g/mol}) +$
 $(1 \text{ mol})(35.453 \text{ g/mol}) = 36.461 \text{ g}$

$\text{KOH}: (1 \text{ mol})(39.098 \text{ g/mol}) +$
 $(1 \text{ mol})(15.999 \text{ g/mol}) +$
 $(1 \text{ mol})(1.008 \text{ g/mol}) = 56.105 \text{ g}$

$\text{KCl}: (1 \text{ mol})(39.098 \text{ g/mol}) +$
 $(1 \text{ mol})(35.453 \text{ g/mol}) = 74.551 \text{ g}$

$\text{H}_2\text{O}: (2 \text{ mol})(1.008 \text{ g/mol}) +$
 $(1 \text{ mol})(15.999 \text{ g/mol}) = 18.015 \text{ g}$

$36.461 \text{ g HCl} + 56.105 \text{ g KOH} \rightarrow$
 $74.551 \text{ g KCl} + 18.015 \text{ g H}_2\text{O}$

$92.566 \text{ g reactants} = 92.566 \text{ g products}$

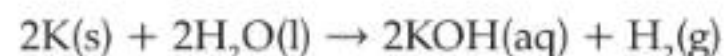
3	CHM.5.3.01.012.02 Calculate the number of moles of a reactant or a product given the number of moles of another reactant or product	Textbook + Example problem 2 + Practice Problems	215, 216, 217
4	CHM.5.3.01.012.01 List the steps included in stoichiometric calculations	Textbook + Problem-Solving Strategy	216

Using Stoichiometry

All stoichiometric calculations use mole ratios based on a balanced chemical equation. Mass-to-mole conversions are also needed.

Stoichiometric mole-to-mole conversion

The reaction between potassium and water is shown in **Figure 3**. The equation is:



Here, you know that two moles of potassium yield one mole of hydrogen. But how much hydrogen would be produced if only 0.0400 mol of potassium were used? To answer this question, identify the given substance and the substance that you need to determine. The given substance is 0.0400 mol of potassium. The unknown is the number of moles of hydrogen.

To solve the problem, you need to know how the unknown moles of hydrogen are related to the known moles of potassium. Mole ratios are used as conversion factors to convert the known number of moles of one substance to the unknown number of moles of another substance in the same reaction. Several mole ratios can be written from the equation, but how do you choose the correct one?

As shown below, the correct mole ratio, 1 mol H₂ to 2 mol K, has moles of unknown in the numerator and moles of known in the denominator. Using this mole ratio converts the moles of potassium to the unknown number of moles of hydrogen.

$$\text{moles of known} \times \frac{\text{moles of unknown}}{\text{moles of known}} = \text{moles of unknown}$$

$$0.0400 \text{ mol K} \times \frac{1 \text{ mol H}_2}{2 \text{ mol K}} = 0.0200 \text{ mol H}_2$$



Figure 3 Potassium metal reacts vigorously with water, releasing so much heat that the hydrogen gas formed in the reaction catches fire.

PROBLEM-SOLVING STRATEGY

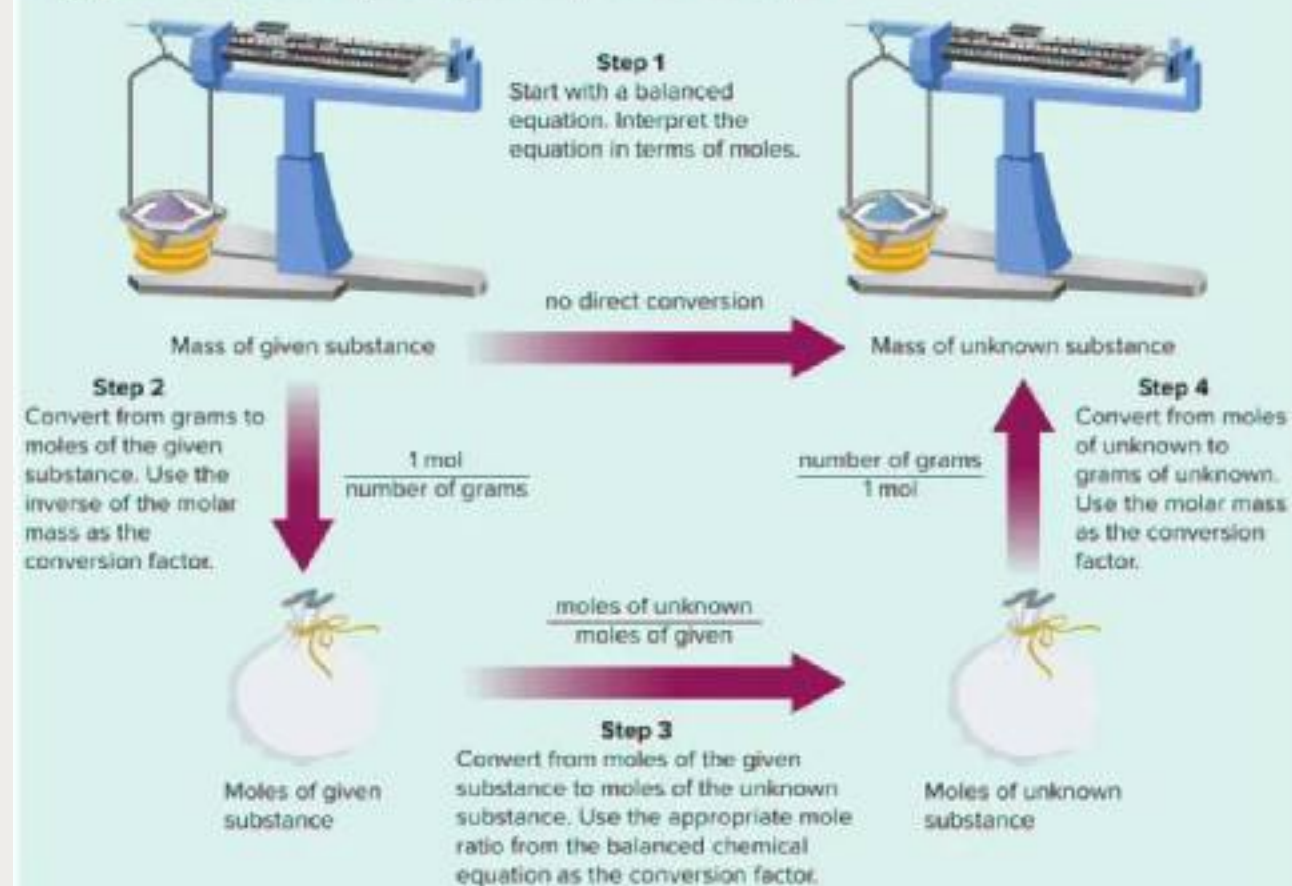
Mastering Stoichiometry

The flowchart below outlines the steps used to solve mole-to-mole, mole-to-mass, and mass-to-mass stoichiometric problems.

- Complete Step 1 by writing the balanced chemical equation for the reaction.
- To determine where to start your calculations, note the unit of the given substance.
 - If mass (in grams) of the given substance is the starting unit, begin your calculations with Step 2.
 - If amount (in moles) of the given substance is the starting unit, skip Step 2 and begin your calculations with Step 3.
- The end point of the calculation depends on the desired unit of the unknown substance.
 - If the answer must be in moles, stop after completing Step 3.
 - If the answer must be in grams, stop after completing Step 4.

Apply the Strategy

Apply the Problem-Solving Strategy to Example Problems 2, 3, and 4.



3	CHM.5.3.01.012.02 Calculate the number of moles of a reactant or a product given the number of moles of another reactant or product	Textbook + Example problem 2 + Practice Problems	215, 216, 217
4	CHM.5.3.01.012.01 List the steps included in stoichiometric calculations	Textbook + Problem-Solving Strategy	216

EXAMPLE Problem 2

MOLE-TO-MOLE STOICHIOMETRY One disadvantage of burning propane (C_3H_8) is that carbon dioxide (CO_2) is one of the products. The released carbon dioxide increases the concentration of CO_2 in the atmosphere. How many moles of CO_2 are produced when 10.0 mol of C_3H_8 are burned in excess oxygen in a gas grill?

1 ANALYZE THE PROBLEM

You are given moles of the reactant, C_3H_8 and must find the moles of the product, CO_2 . First write the balanced chemical equation, then convert from moles of C_3H_8 to moles of CO_2 . The correct mole ratio has moles of unknown substance in the numerator and moles of known substance in the denominator.

Known

moles $C_3H_8 = 10.0 \text{ mol } C_3H_8$

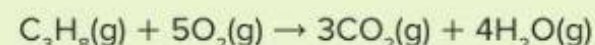
Unknown

moles $CO_2 = ? \text{ mol } CO_2$

2 SOLVE FOR THE UNKNOWN

Write the balanced chemical equation for the combustion of C_3H_8 . Use the correct mole ratio to convert moles of known (C_3H_8) to moles of unknown (CO_2).

10.0 mol ? mol



Mole ratio: $\frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8}$

$$10.0 \text{ mol } C_3H_8 \times \frac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8} = \text{30.0 mol } CO_2$$

Burning 10.0 moles of C_3H_8 produces 30.0 moles CO_2 .

3 EVALUATE THE ANSWER

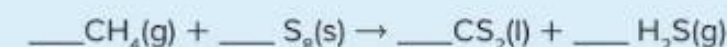
Because the given number of moles has three significant figures, the answer also has three figures. The balanced chemical equation indicates that 1 mol of C_3H_8 produces 3 mol of CO_2 . Thus, 10.0 mol of C_3H_8 produces three times as many moles of CO_2 , or 30.0 mol.

PRACTICE Problems



ADDITIONAL PRACTICE

11. Methane and sulfur react to produce carbon disulfide (CS_2), a liquid often used in the production of cellophane.



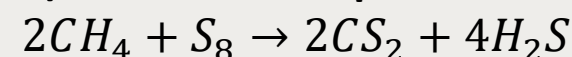
- Balance the equation.
- Calculate the moles of CS_2 produced when 1.50 mol S_8 is used.
- How many moles of H_2S are produced?

12. CHALLENGE Sulfuric acid (H_2SO_4) is formed when sulfur dioxide (SO_2) reacts with oxygen and water.

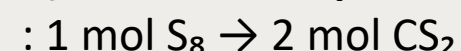
- Write the balanced chemical equation for the reaction.
- How many moles of H_2SO_4 are produced from 12.5 moles of SO_2 ?
- How many moles of O_2 are needed?

11. Methane + Sulfur $\rightarrow$ CS_2 + H_2S

a) Balance the equation



b) Moles of CS_2 produced from 1.50 mol S_8



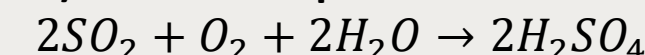
$$!1.50 \text{ mol } S_8 \times 2 = 3.00 \text{ mol } CS_2$$

c) Moles of H_2S produced

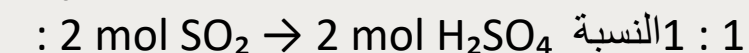


$$1.50 \text{ mol } S_8 \times 4 = 6.00 \text{ mol } H_2S$$

a) Balanced equation

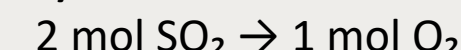


b) Moles of H_2SO_4 from 12.5 mol SO_2



$$!12.5 \text{ mol } SO_2 \rightarrow 12.5 \text{ mol } H_2SO_4$$

c) Moles of O_2 needed



$$12.5 \div 2 = 6.25 \text{ mol } O_2$$

EXAMPLE Problem 3

MOLE-TO-MASS STOICHIOMETRY Determine the mass of sodium chloride (NaCl), commonly called table salt, produced when 1.25 mol of chlorine gas (Cl₂) reacts vigorously with excess sodium.

1 ANALYZE THE PROBLEM

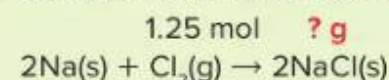
You are given the moles of the reactant, Cl₂, and must determine the mass of the product, NaCl. You must convert from moles of Cl₂ to moles of NaCl using the mole ratio from the equation. Then, you need to convert moles of NaCl to grams of NaCl using the molar mass as the conversion factor.

Known

moles of chlorine = 1.25 mol Cl₂

Unknown

mass of sodium chloride = ? g NaCl

2 SOLVE FOR THE UNKNOWN

Write the balanced chemical equation, and identify the known and the unknown values.

$$\text{Mole ratio: } \frac{2 \text{ mol NaCl}}{1 \text{ mol Cl}_2}$$

$$1.25 \text{ mol Cl}_2 \times \frac{2 \text{ mol NaCl}}{1 \text{ mol Cl}_2} = 2.50 \text{ mol NaCl}$$

Multiply moles of Cl₂ by the mole ratio to get moles of NaCl.

$$2.50 \text{ mol NaCl} \times \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = \mathbf{146 \text{ g NaCl}}$$

Multiply moles of NaCl by the molar mass to get grams of NaCl.

3 EVALUATE THE ANSWER

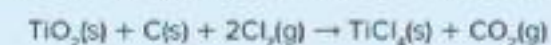
Because the given number of moles has three significant figures, the mass of NaCl also has three. To quickly assess whether the calculated mass value for NaCl is correct, perform the calculations in reverse: divide the mass of NaCl by the molar mass of NaCl, and then divide the result by 2. You will obtain the given number of moles of Cl₂.

PRACTICE Problems**ADDITIONAL PRACTICE**

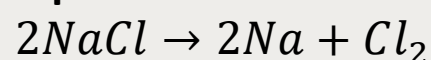
13. Sodium chloride is decomposed into the elements sodium and chlorine by means of electrical energy. How much chlorine gas, in grams, is obtained from the process diagrammed at right?



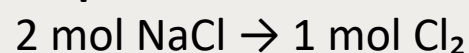
14. CHALLENGE Titanium is a transition metal used in many alloys because it is extremely strong and lightweight. Titanium tetrachloride (TiCl₄) is extracted from titanium oxide (TiO₂) using chlorine and coke (carbon).



- What mass of Cl₂ gas is needed to react with 1.25 mol of TiO₂?
- What mass of C is needed to react with 1.25 mol of TiO₂?
- What is the mass of all of the products formed by reaction with 1.25 mol of TiO₂?

Equation:**Given:**

NaCl = 2.50 mol

Step 1: Find moles of Cl₂ produced

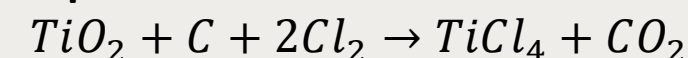
$$2.50 \div 2 = 1.25 \text{ mol Cl}_2$$

Step 2: Convert moles of Cl₂ to grams

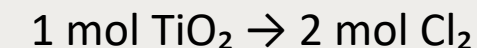
Mass = *moles* × *molar mass*

Molar mass of Cl₂ = 70.90 g/mol

$$1.25 \times 70.90 = 88.6 \text{ g Cl}_2$$

14. Titanium Tetrachloride Formation**Equation:****Given:**

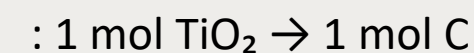
TiO₂ = 1.25 mol

a) Mass of Cl₂ needed

$$1.25 \times 2 = 2.50 \text{ mol Cl}_2$$

Molar mass of Cl₂ = 70.90 g/mol

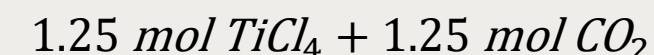
$$2.50 \times 70.90 = 177.3 \text{ g Cl}_2$$

b) Mass of C needed

$$1.25 \times 12.01 = 15.0 \text{ g C}$$

c) Mass of all products

Products: TiCl₄ + CO₂



Molar masses: TiCl₄ = 189.7 g/mol CO₂ = 44.01 g/mol

$$(1.25 \times 189.7) + (1.25 \times 44.01) = 236.9 + 55.0 = 291.9 \text{ g total}$$

Stoichiometric mass-to-mass calculations

If you were preparing to carry out a chemical reaction in the laboratory, you would need to know how much of each reactant to use in order to produce the mass of product you required. You can use the balanced chemical reaction to determine this amount. You could also use the balanced chemical reaction and mole ratios to determine how much of a product you would obtain when starting with a known mass of a reactant.

Example Problem 4 demonstrates how you can use a measured mass of the known substance, the balanced chemical equation, and mole ratios from the equation to find the mass of the unknown substance. Notice, as always, a balanced chemical equation is essential in the calculations associated with any chemical reaction. Without the ratios between reactants and products, it would not be possible to convert between moles of a reactant and the moles of a product.

EXAMPLE Problem 4

MASS-TO-MASS STOICHIOMETRY Ammonium nitrate (NH_4NO_3), an important fertilizer, produces dinitrogen monoxide (N_2O) gas and H_2O when it decomposes. Determine the mass of H_2O produced from the decomposition of 25.0 g of solid NH_4NO_3 .

1 ANALYZE THE PROBLEM

Write the balanced equation and convert the known mass of the reactant to moles of the reactant. Next, use a mole ratio to relate moles of the reactant to moles of the product. Then, use the molar mass to convert from moles of the product to the mass of the product.

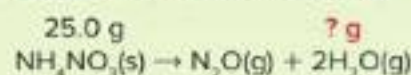
Known

mass of ammonium nitrate = 25.0 g NH_4NO_3

Unknown

mass of water = ? g H_2O

2 SOLVE FOR THE UNKNOWN



Write the balanced chemical equation, and identify the known and unknown values.

$$25.0 \text{ g } \text{NH}_4\text{NO}_3 \times \frac{1 \text{ mol } \text{NH}_4\text{NO}_3}{80.04 \text{ g } \text{NH}_4\text{NO}_3} = 0.312 \text{ mol } \text{NH}_4\text{NO}_3$$

Multiply grams of NH_4NO_3 by the inverse of molar mass to get moles of NH_4NO_3 .

$$\text{Mole ratio: } \frac{2 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \text{NH}_4\text{NO}_3}$$

$$0.312 \text{ mol } \text{NH}_4\text{NO}_3 \times \frac{2 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \text{NH}_4\text{NO}_3} = 0.624 \text{ mol } \text{H}_2\text{O}$$

Multiply moles of NH_4NO_3 by the mole ratio to get moles of H_2O .

$$0.624 \text{ mol } \text{H}_2\text{O} \times \frac{18.02 \text{ g } \text{H}_2\text{O}}{1 \text{ mol } \text{H}_2\text{O}} = 11.2 \text{ g } \text{H}_2\text{O}$$

Multiply moles of H_2O by the molar mass to get grams.

PRACTICE Problems

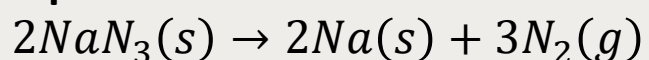
ADDITIONAL PRACTICE

15. One of the reactions used to inflate automobile air bags involves sodium azide (NaN_3): $2\text{NaN}_3(\text{s}) \rightarrow 2\text{Na}(\text{s}) + 3\text{N}_2(\text{g})$. Determine the mass of N_2 produced from the decomposition of NaN_3 shown below.



16. **CHALLENGE** In the formation of acid rain, sulfur dioxide (SO_2) reacts with oxygen and water in the air to form sulfuric acid (H_2SO_4). Write the balanced chemical equation for the reaction. If 2.50 g of SO_2 reacts with excess oxygen and water, how much H_2SO_4 , in grams, is produced?

15. Sodium Azide Decomposition Equation:



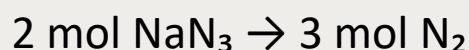
Given:



Step 1: Moles of NaN_3

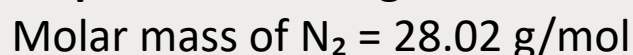
$$\text{Moles} = \frac{100.0}{65.02} = 1.54 \text{ mol } \text{NaN}_3$$

Step 2: Moles of N_2 produced



$$1.54 \times \frac{3}{2} = 2.31 \text{ mol } \text{N}_2$$

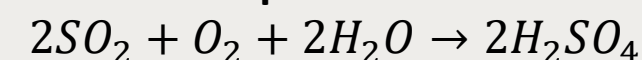
Step 3: Convert to grams



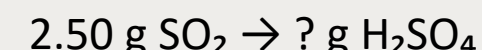
$$2.31 \times 28.02 = 64.7 \text{ g } \text{N}_2$$

16. Formation of Acid Rain (Sulfuric Acid)

Balanced equation:



Given:



Step 1: Moles of SO_2

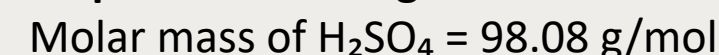
$$\frac{2.50}{64.07} = 0.0390 \text{ mol } \text{SO}_2$$

Step 2: Moles of H_2SO_4 produced



$$0.0390 \text{ mol } \text{H}_2\text{SO}_4$$

Step 3: Convert to grams



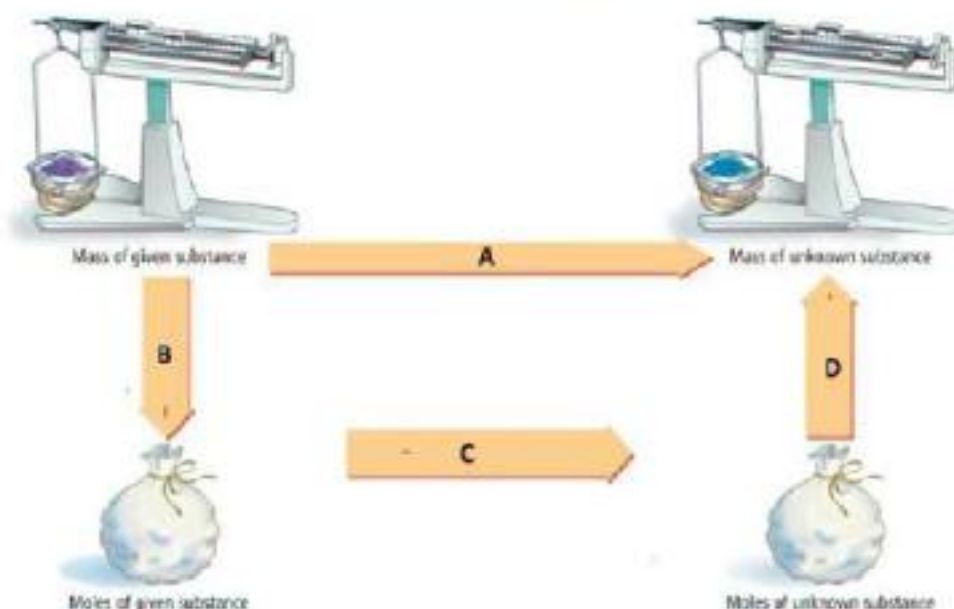
$$0.0390 \times 98.08 = 3.83 \text{ g } \text{H}_2\text{SO}_4$$

Summary

- Chemists use stoichiometric calculations to predict the amounts of reactants used and products formed in specific reactions.
- The first step in solving stoichiometric problems is writing the balanced chemical equation.
- Mole ratios derived from the balanced chemical equation are used in stoichiometric calculations.
- Stoichiometric problems make use of mole ratios to convert between mass and moles.

What is the **correct** conversion factor for describing the arrow with the letter **D** in the chart below ?

ما معامل التحويل **الصحيح** لوصف السهم الذي يحمل الحرف **D** في المخطط أدناه؟



A. Balanced chemical equation

B. molar mass

C. Inverse Molar mass

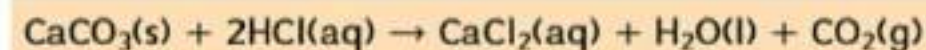
D. Mole ratio

What is the mass of CaCl_2 produced from the reaction

175 g ما كتلة كلوريد الكالسيوم CaCl_2 التي تنتج من تفاعل

of 175 g of CaCO_3 ?

من كربونات الكالسيوم CaCO_3 ؟



Molar Mass $\text{CaCO}_3 = 100 \text{ g/mol}$

الكتلة المولية $\text{CaCO}_3 = 100 \text{ g/mol}$

Molar Mass $\text{CaCl}_2 = 110.9 \text{ g/mol}$

الكتلة المولية $\text{CaCl}_2 = 110.9 \text{ g/mol}$

What is the **first** step in solving stoichiometry problems?

ما الخطوة **الأولى** في حل مسائل الحسابات الكيميائية؟

A. Writing the unit of the given substance

كتابة وحدة المادة المعطاة

B. Writing the unit of the unknown substance

كتابة وحدة المادة غير المعروفة

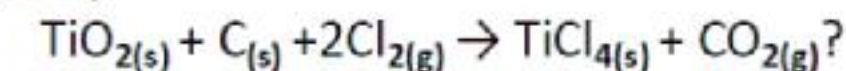
C. Writing the balanced chemical equation

كتابة المعادلة الكيميائية الموزونة

D. Writing the mole ratios

كتابة النسب المولية

What is the mass of Cl_2 gas is needed to reacted with 1.25 mol of $\text{TiO}_2(\text{s})$ (molar mass of $\text{Cl}_2 = 71 \text{ g/mol}$)



A. 88.75 g

B. 44.4 g

C. 177.5 g

D. 1 g13.

Why do reactions stop?

Rarely in nature are the reactants present in the exact ratios specified by the balanced chemical equation. Generally, one or more reactants are in excess and the reaction proceeds until all of one reactant is used up.

When a reaction is carried out in the laboratory, the same principle applies. Usually, one or more of the reactants are in excess, while one is limited. The amount of product depends on the reactant that is limited.

Limiting and excess reactants

Recall the reaction of sodium hydrogen sulfite with potassium permanganate from Lesson 1. After a colorless solution forms, adding more sodium hydrogen sulfite has no effect because there is no potassium permanganate left to react. Potassium permanganate is a limiting reactant. As the name implies, the **limiting reactant** limits the extent of the reaction and thereby determines the amount of product formed. A portion of all the other reactants remains after the reaction stops. Reactants leftover when a reaction stops are **excess reactants**.

To help you understand limiting and excess reactants, consider the analogy in **Figure 4**. From the available tools, four complete sets consisting of a pair of pliers, a hammer, and two screwdrivers can be assembled. The number of sets is limited by the number of available hammers. Pliers and screwdrivers remain in excess.

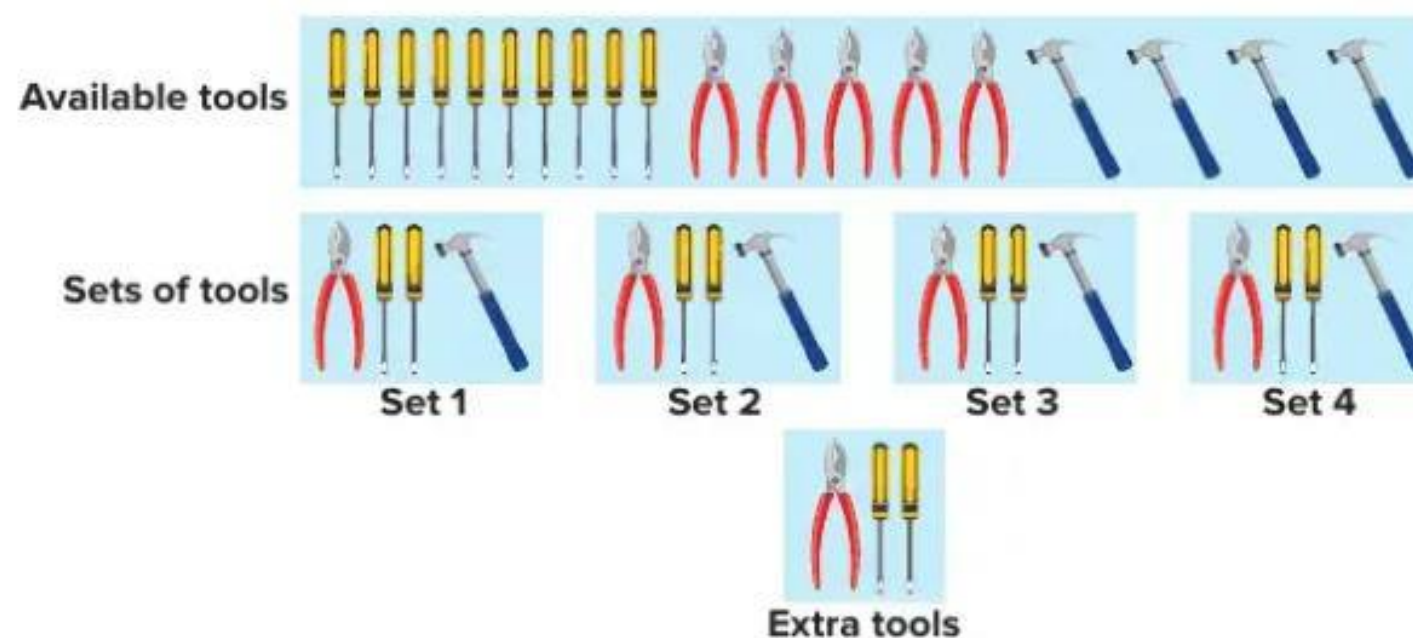


Figure 4 Each tool set must have one hammer, so only four sets can be assembled.

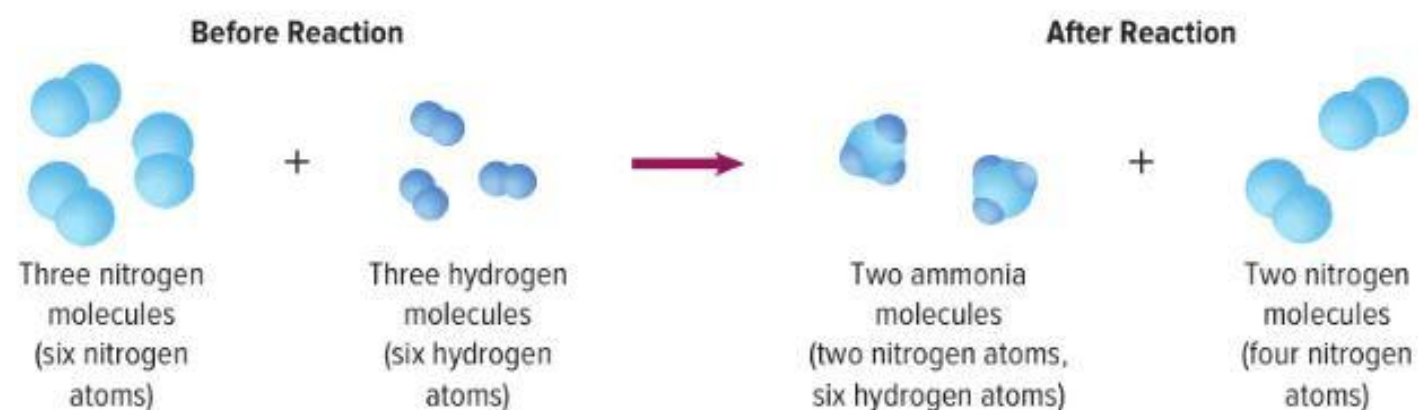


Figure 5 If you check all of the atoms present before and after the reaction, you will find that some of the nitrogen molecules are unchanged. These nitrogen molecules are the excess reactant.

Determining the limiting reactant

The calculations you did in the previous section were based on having the reactants present in the ratio described by the balanced chemical equation. When this is not the case, the first thing you must do is determine which reactant is limiting.

Consider the reaction shown in **Figure 5**, in which three molecules of nitrogen (N_2) and three molecules of hydrogen (H_2) react to form ammonia (NH_3). First, all the nitrogen and hydrogen molecules are separated into individual atoms. These atoms are available for reassembling into ammonia molecules, just as the tools in **Figure 4** can be assembled into tool kits. How many molecules of ammonia can be produced from the available atoms? Two ammonia molecules can be assembled from the hydrogen atoms and nitrogen atoms because only six hydrogen atoms are available—three for each ammonia molecule. When the hydrogen is gone, two unreacted molecules of nitrogen remain. Thus, hydrogen is the limiting reactant and nitrogen is the excess reactant. It is important to know which reactant is limiting because the amount of product formed depends on this reactant.

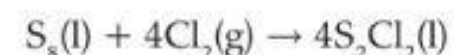


Get It?

Extend How many more hydrogen molecules would be needed to completely react with excess nitrogen molecules shown in **Figure 5**?

Calculating the Amount of Product When a Reactant Is Limiting

How can you calculate the amount of product formed when one of the reactants is limiting? Consider the formation of disulfur dichloride (S_2Cl_2), which is used to vulcanize rubber. As shown in **Figure 6**, the properties of vulcanized rubber make it useful for many products. In the production of disulfur dichloride, molten sulfur reacts with chlorine gas according to the following equation.



If 200.0 g of sulfur reacts with 100.0 g of chlorine, what mass of disulfur dichloride is produced?



Figure 6 Natural rubber, which is soft and sticky, is hardened in a chemical process called vulcanization. During vulcanization, molecules become linked together, forming a durable material that is harder, smoother, and less sticky, making it ideal for many products, such as this caster.

8	CHM.5.3.01.013.04 Identify limiting reactant and excess reactant in a chemical reaction given the number of moles of reactants	Textbook + Example problem 5 + Practice Problems	222, 223, 224, 225
9	CHM.5.3.01.013.05 Identify limiting reactant and excess reactant in a chemical reaction given the mass of reactants	Textbook + Example problem 5 + Practice Problems	224, 225
10	CHM.5.3.01.013.08 Calculate the mass of product(s) produced in a chemical reaction given a particulate diagram, mass or mole of the reactants	Textbook + Example problem 5 + Practice Problems	224, 225

Calculating the limiting reactant

The masses of both reactants are given. First, determine which is the limiting reactant, since the reaction stops forming product when the limiting reactant is used up.

Moles of reactants Identifying the limiting reactant involves finding the number of moles of each reactant. You do this by converting the masses of chlorine and sulfur to moles. Multiply each mass by the inverse of molar mass.

$$100.0\text{ g Cl}_2 \times \frac{1\text{ mol Cl}_2}{70.91\text{ g Cl}_2} = 1.410\text{ mol Cl}_2$$

$$200.0\text{ g S}_8 \times \frac{1\text{ mol S}_8}{256.5\text{ g S}_8} = 0.7797\text{ mol S}_8$$

Using mole ratios Next, determine whether the two reactants are in the correct mole ratio, derived from the balanced equation. The coefficients in the equation indicate that 4 mol of chlorine is needed to react with 1 mol of sulfur. This 4:1 ratio must be compared with the calculated ratio of the moles of available reactants above. To determine the actual ratio of moles, divide the number of available moles of chlorine by the number of available moles of sulfur.

$$\frac{1.410\text{ mol Cl}_2\text{ available}}{0.7797\text{ mol S}_8\text{ available}} = \frac{1.808\text{ mol Cl}_2\text{ available}}{1\text{ mol S}_8\text{ available}}$$

Only 1.808 mol of chlorine is available for every 1 mol of sulfur, instead of the 4 mol required by the balanced chemical equation. Therefore, chlorine is the limiting reactant.

Analyzing the excess reactant

Now that you have determined the limiting reactant and the amount of product formed, what about the excess reactant, sulfur? How much of it reacted?

Moles reacted You need to make a mole-to-mass calculation to determine the mass of sulfur needed to react completely with 1.410 mol of chlorine. First, obtain the number of moles of sulfur by multiplying the moles of chlorine by the S₈-to-Cl₂ mole ratio.

$$1.410\text{ mol Cl}_2 \times \frac{1\text{ mol S}_8}{4\text{ mol Cl}_2} = 0.3525\text{ mol S}_8$$

Mass reacted Next, to obtain the mass of sulfur needed, multiply by its molar mass.

$$0.3525\text{ mol S}_8 \times \frac{256.5\text{ g S}_8}{1\text{ mol S}_8} = 90.42\text{ g S}_8\text{ needed}$$

EXAMPLE Problem 5

DETERMINING THE LIMITING REACTANT The reaction between solid white phosphorus (P₄) and oxygen produces solid tetraphosphorus decoxide (P₄O₁₀). This compound is often called diphosphorus pentoxide because its empirical formula is P₂O₅.

- Determine the mass of P₄O₁₀ formed if 25.0 g of P₄ and 50.0 g of oxygen are combined.
- How much of the excess reactant remains after the reaction stops?

1 ANALYZE THE PROBLEM

You are given the masses of both reactants, so you must identify the limiting reactant and find the mass of the product. From the limiting reactant, the moles of the excess reactant used in the reaction can be determined. This can be converted to mass and subtracted from the given mass to find the amount in excess.

Known

mass of phosphorus = 25.0 g P₄
mass of oxygen = 50.0 g O₂

Unknown

mass of tetraphosphorus decoxide = ? g P₄O₁₀
mass of excess reactant = ? g excess reactant

2 SOLVE FOR THE UNKNOWN

Determine the limiting reactant.

25.0 g P₄(s) + 50.0 g O₂(g) → P₄O₁₀(s)

Write the balanced chemical equation, and identify the known and the unknown.

Determine the number of moles of the reactants by using the inverse of molar mass.

25.0 g P₄ × $\frac{1\text{ mol P}_4}{123.9\text{ g P}_4}$ = 0.202 mol P₄

Calculate the moles of P₄.

50.0 g O₂ × $\frac{1\text{ mol O}_2}{32.00\text{ g O}_2}$ = 1.56 mol O₂

Calculate the moles of O₂.

Calculate the actual ratio of available moles of O₂ and available moles of P₄.

$\frac{1.56\text{ mol O}_2}{0.202\text{ mol P}_4} = \frac{7.72\text{ mol O}_2}{1\text{ mol P}_4}$

Calculate the ratio of moles of O₂ to moles of P₄.

Determine the mole ratio of the two reactants from the balanced chemical equation.

Mole ratio: $\frac{5\text{ mol O}_2}{1\text{ mol P}_4}$

Because 7.72 mol of O₂ is available but only 5 mol is needed to react with 1 mol of P₄, O₂ is in excess and P₄ is the limiting reactant. Use the moles of P₄ to determine the moles of P₄O₁₀ that will be produced. Multiply the number of moles of P₄ by the mole ratio of P₄O₁₀ to P₄.

0.202 mol P₄ × $\frac{1\text{ mol P}_4\text{O}_{10}}{1\text{ mol P}_4}$ = 0.202 mol P₄O₁₀ Calculate the moles of product (P₄O₁₀) formed.

To calculate the mass of P₄O₁₀, multiply moles of P₄O₁₀ by the molar mass.

0.202 mol P₄O₁₀ × $\frac{283.9\text{ g P}_4\text{O}_{10}}{1\text{ mol P}_4\text{O}_{10}}$ = 57.3 g P₄O₁₀ Calculate the mass of the product P₄O₁₀.

8	CHM.5.3.01.013.04 Identify limiting reactant and excess reactant in a chemical reaction given the number of moles of reactants	Textbook + Example problem 5 + Practice Problems	222, 223, 224, 225
9	CHM.5.3.01.013.05 Identify limiting reactant and excess reactant in a chemical reaction given the mass of reactants	Textbook + Example problem 5 + Practice Problems	224, 225
10	CHM.5.3.01.013.08 Calculate the mass of product(s) produced in a chemical reaction given a particulate diagram, mass or mole of the reactants	Textbook + Example problem 5 + Practice Problems	224, 225

EXAMPLE Problem 5 (continued)

Because O_2 is in excess, only part of the available O_2 is consumed. Use the limiting reactant, P_4 , to determine the moles and mass of O_2 used.

$0.202 \text{ mol } P_4 \times \frac{5 \text{ mol } O_2}{1 \text{ mol } P_4} = 1.01 \text{ mol } O_2$ Multiply the moles of limiting reactant by the mole ratio to find moles of excess reactant

Convert moles of O_2 consumed to mass of O_2 consumed.

$1.01 \text{ mol } O_2 \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} = 32.3 \text{ g } O_2$ Multiply the moles of O_2 by the molar mass.

Calculate the amount of excess O_2 .

$50.0 \text{ g } O_2 \text{ available} - 32.3 \text{ g } O_2 \text{ consumed} = \mathbf{17.7 \text{ g } O_2 \text{ in excess}}$ Subtract mass used from the initial mass.

3 EVALUATE THE ANSWER

All values have a minimum of three significant figures, so the mass of P_4O_{10} is correctly stated with three digits. The mass of excess O_2 (17.7 g) is found by subtracting two numbers that are accurate to the first decimal place, so the mass of excess O_2 correctly shows one decimal place. The sum of the O_2 consumed (32.3 g) and the given mass of P_4 (25.0 g) is 57.3 g, the calculated mass of the product P_4O_{10} .

PRACTICE Problems

23. The reaction between solid sodium and iron(III) oxide is one in a series of reactions that inflates an automobile airbag: $6Na(s) + Fe_2O_3(s) \rightarrow 3Na_2O(s) + 2Fe(s)$. If 100.0 g of Na and 100.0 g of Fe_2O_3 are used in this reaction, determine the following.
a. limiting reactant
b. reactant in excess
c. mass of solid iron produced
d. mass of excess reactant that remains after the reaction is complete

24. CHALLENGE Photosynthesis reactions in green plants use carbon dioxide and water to produce glucose ($C_6H_{12}O_6$) and oxygen. A plant has 88.0 g of carbon dioxide and 64.0 g of water available for photosynthesis. Be sure to report the correct level of accuracy based on measurements given in the question.
a. Write the balanced chemical equation for the reaction.
b. Determine the limiting reactant and the excess reactant.
c. Determine the mass in excess.
d. Determine the mass of glucose produced.

24. Challenge Photosynthesis reactions in green plants use carbon dioxide and water to produce glucose ($C_6H_{12}O_6$) and oxygen. A plant has 88.0 g of carbon dioxide and 64.0 g of water available for photosynthesis.

a. Write the balanced chemical equation for the reaction.

$$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g)$$

b. Determine the limiting reactant.
Make mass $\rightarrow$ mole conversion.
 $88.0 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} = 2.00 \text{ mol } CO_2$
 $64.0 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} = 3.55 \text{ mol } H_2O$
Make mole ratio comparison.
 $\frac{2.00 \text{ mol } CO_2}{3.55 \text{ mol } H_2O}$ compared to $\frac{6 \text{ mol } CO_2}{6 \text{ mol } H_2O}$;
0.563 compared to 1.00
The actual ratio is less than the needed ratio
carbon dioxide is the limiting reactant.

c. Determine the excess reactant.
Water is the excess reactant.
d. Determine the mass in excess.
Make mole $\rightarrow$ mole conversion.
 $2.00 \text{ mol } CO_2 \times \frac{6 \text{ mol } H_2O}{6 \text{ mol } CO_2} = 2.00 \text{ mol } H_2O$
Make mole $\rightarrow$ mass conversion.
 $2.00 \text{ mol } H_2O \times \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O}$
 $= 36.0 \text{ g } H_2O \text{ needed}$
 $64.0 \text{ g } H_2O \text{ given} - 36.0 \text{ g } H_2O \text{ needed}$
 $= 28.0 \text{ g } H_2O \text{ in excess}$

11	CHM.5.3.01.013.13 Define the percent yield of a chemical reaction	Textbook + Figure 8	227, 228
12	CHM.5.3.01.013.11 Distinguish between theoretical yield and actual yield in stoichiometric calculations	Textbook + Example problem 6 + Practice Problems	227, 228, 229
13	CHM.5.3.01.013.14 Calculate percent yield and theoretical in a chemical reaction	Textbook + Example problem 6 + Practice Problems	227, 228, 229

How much product?

While solving stoichiometric problems, you might have concluded that chemical reactions always proceed according to the balanced equation and produce the calculated amount of product. This, however, is not the case. Most reactions never succeed in producing the predicted amount of product.

Reactions do not go to completion or yield as expected for many reasons. Reactants and products might adhere to the surfaces of their containers or evaporate. In some cases, products other than the intended ones might be formed by competing reactions, thus reducing the yield of the desired product. Or, as shown in **Figure 8**, some amount of a solid product can be left behind on filter paper or lost in the purification process. Because of this, chemists need to know how to gauge the yield of a reaction.



Figure 8 Silver chromate is formed when potassium chromate is added to silver nitrate. Note that not all of the precipitate can be removed from the filter paper. Still more of the precipitate is lost because it adheres to the sides of the beaker.

Theoretical and actual yields

In many of the calculations you have performed, you have calculated the amount of product given an amount of reactant. This value is the **theoretical yield** of the reaction, and represents the maximum amount of product that can be produced from a given amount of reactant.

A chemical reaction rarely produces the theoretical yield. A chemist determines the actual yield of a reaction through an experiment in which the mass of the product is measured. The **actual yield** is the amount produced when the experiment is performed.

Percent yield

Percent yield of product is the ratio of the actual yield to the theoretical yield expressed as a percent. Chemists use this to measure the efficiency of the reaction.

Percent Yield

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

The actual yield divided by the theoretical yield multiplied by 100 is the percent yield.

EXAMPLE Problem 6

PERCENT YIELD Solid silver chromate (Ag_2CrO_4) forms when excess potassium chromate (K_2CrO_4) is added to a solution containing 0.500 g of silver nitrate (AgNO_3). Determine the theoretical yield of Ag_2CrO_4 . Find the percent yield if the reaction yields 0.455 g of Ag_2CrO_4 .

1 ANALYZE THE PROBLEM

You know the mass of a reactant and the actual yield of the product. Write the balanced chemical equation, and calculate theoretical yield by converting grams of AgNO_3 to moles of AgNO_3 , moles of AgNO_3 to moles of Ag_2CrO_4 , and moles of Ag_2CrO_4 to grams of Ag_2CrO_4 . Calculate the percent yield from the actual yield and the theoretical yield.

Known

mass of silver nitrate = 0.500 g AgNO_3

actual yield = 0.455 g Ag_2CrO_4

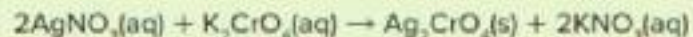
Unknown

theoretical yield = ? g Ag_2CrO_4

percent yield = ? % Ag_2CrO_4

2 SOLVE FOR THE UNKNOWN

0.500 g ? g



$$0.500 \text{ g } \text{AgNO}_3 \times \frac{1 \text{ mol } \text{AgNO}_3}{169.9 \text{ g } \text{AgNO}_3} = 2.94 \times 10^{-3} \text{ mol } \text{AgNO}_3$$

$$2.94 \times 10^{-3} \text{ mol } \text{AgNO}_3 \times \frac{1 \text{ mol } \text{Ag}_2\text{CrO}_4}{2 \text{ mol } \text{AgNO}_3} = 1.47 \times 10^{-3} \text{ mol } \text{Ag}_2\text{CrO}_4$$

$$1.47 \times 10^{-3} \text{ mol } \text{Ag}_2\text{CrO}_4 \times \frac{331.7 \text{ g } \text{Ag}_2\text{CrO}_4}{1 \text{ mol } \text{Ag}_2\text{CrO}_4} = 0.488 \text{ g } \text{Ag}_2\text{CrO}_4$$

$$\frac{0.455 \text{ g } \text{Ag}_2\text{CrO}_4}{0.488 \text{ g } \text{Ag}_2\text{CrO}_4} \times 100 = 93.2\% \text{ } \text{Ag}_2\text{CrO}_4$$

Write the balanced chemical equation, and identify the known and the unknown.

Convert grams of AgNO_3 to moles.

Use the mole ratio to convert moles of AgNO_3 to moles of Ag_2CrO_4 .

Calculate the theoretical yield.

Calculate the percent yield.

PRACTICE Problems



ADDITIONAL PRACTICE

- 28.** Aluminum hydroxide ($\text{Al}(\text{OH})_3$) is often present in antacids to neutralize stomach acid (HCl). The reaction occurs as follows: $\text{Al}(\text{OH})_3(\text{s}) + 3\text{HCl}(\text{aq}) \rightarrow \text{AlCl}_3(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$. If 14.0 g of $\text{Al}(\text{OH})_3$ is present in an antacid tablet, determine the theoretical yield of AlCl_3 produced when the tablet reacts with HCl .
- 29.** Zinc reacts with iodine in a synthesis reaction: $\text{Zn} + \text{I}_2 \rightarrow \text{ZnI}_2$.
- Determine the theoretical yield if 1.912 mol of zinc is used.
 - Determine the percent yield if 515.6 g of product is recovered.
- 30. CHALLENGE** When copper wire is placed into a silver nitrate solution (AgNO_3), silver crystals and copper(II) nitrate ($\text{Cu}(\text{NO}_3)_2$) solution form.
- Write the balanced chemical equation for the reaction.
 - If a 20.0-g sample of copper is used, determine the theoretical yield of silver.
 - If 60.0 g of silver is recovered from the reaction, determine the percent yield of the reaction.

11	CHM.5.3.01.013.13 Define the percent yield of a chemical reaction	Textbook + Figure 8	227, 228
12	CHM.5.3.01.013.11 Distinguish between theoretical yield and actual yield in stoichiometric calculations	Textbook + Example problem 6 + Practice Problems	227, 228, 229
13	CHM.5.3.01.013.14 Calculate percent yield and theoretical in a chemical reaction	Textbook + Example problem 6 + Practice Problems	227, 228, 229

Zinc reacts with iodine in a synthesis reaction

يتفاعل الخارصين مع اليود في تفاعل تكوين طبقاً للمعادلة التالية:

as the following equation: If 3.50 mol of Zinc

إذا تم استخدام 3.50 mol من الخارصين Zn . ما المردود النظري؟

is used. What is theoretical yield?

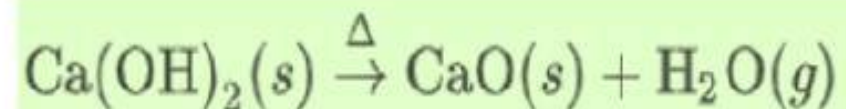


Molar Mass of $\text{ZnI}_2 = 319 \text{ g/mol}$

الكتلة المولية لـ $\text{ZnI}_2 = 319 \text{ g/mol}$

Solid calcium hydroxide decomposes by heat into solid calcium oxide and water vapor as in the equation below. If **59.2 g** of Ca(OH)_2 is used at the beginning of the reaction and **40.32 g** of CaO is produced, what is the percent yield?

يتفكك هيدروكسيد الكالسيوم الصلب بالحرارة إلى أكسيد الكالسيوم الصلب وبخار الماء كما في المعادلة أدناه. إذا تم استخدام **59.2 g** من Ca(OH)_2 في بداية التفاعل ونُتج **40.32 g** من CaO . فما النسبة المئوية للمردود؟



Molar Mass of $\text{Ca(OH)}_2 = 74.0 \text{ g/mol}$

الكتلة المولية لـ $\text{Ca(OH)}_2 = 74.0 \text{ g/mol}$

Molar Mass of $\text{CaO} = 56.0 \text{ g/mol}$

الكتلة المولية لـ $\text{CaO} = 56.0 \text{ g/mol}$

The Kinetic-Molecular Theory

You have learned that composition (the types of atoms present) and structure (their arrangement) determine the chemical properties of matter. Composition and structure also affect the physical properties of matter. Based solely on physical appearance, you can distinguish between the solids and liquids, as shown in **Figure 1**. By contrast, substances that are gases at room temperature usually display similar physical properties despite their different compositions. Why is there so little variation in behavior among gases? Why are the physical properties of gases different from those of liquids and solids?



Figure 1 You can distinguish some materials by looking at them, but this is not true for many gases.

By the eighteenth century, scientists knew how to collect gaseous products by displacing water. Now, they could observe and measure properties of individual gases. About 1860, chemists Ludwig Boltzmann and James Maxwell, who were working in different countries, each proposed a model to explain the properties of gases. That model is the kinetic-molecular theory. Because all of the gases known to Boltzmann and Maxwell contained molecules, the name of the model refers to molecules. The word *kinetic* comes from a Greek word meaning *to move*. Objects in motion have energy called kinetic energy. The **kinetic-molecular theory** describes the behavior of matter in terms of particles in motion. The model makes several assumptions about the size, motion, and energy of gas particles.

Particle size

Gases consist of small particles that are separated from one another by empty space. The volume of the particles is small compared with the volume of the empty space. Because gas particles are far apart, they experience no significant attractive or repulsive forces.

Particle motion

Gas particles are in constant, random motion. Particles move in a straight line until they collide with other particles or with the walls of their container, as shown in **Figure 2**. Collisions between gas particles are elastic. An **elastic collision** is one in which no kinetic energy is lost. Kinetic energy can be transferred between colliding particles, but the total kinetic energy of the two particles does not change.

Particle energy

Two factors determine the kinetic energy of a particle: mass and velocity. The kinetic energy of a particle can be represented by the following equation.

$$KE = \frac{1}{2} mv^2$$

KE is kinetic energy, m is the mass of the particle, and v is its velocity. Velocity reflects both the speed and the direction of motion. In a sample of a single gas, all particles have the same mass, but all particles do not have the same velocity. Therefore, all particles do not have the same kinetic energy. **Temperature** is a measure of the average kinetic energy of the particles in a sample of matter.

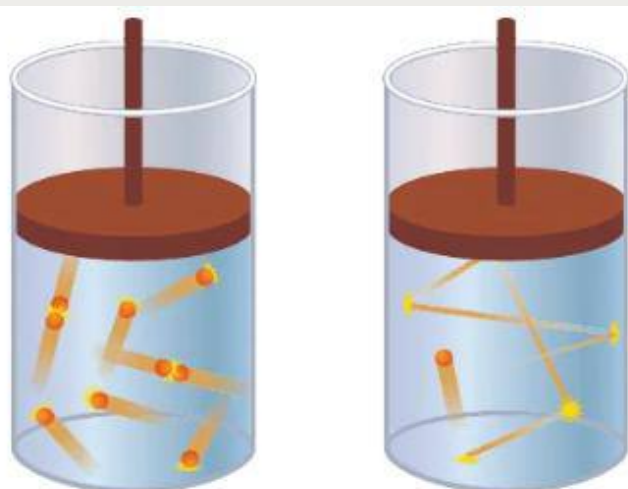


Figure 2 Kinetic energy can be transferred between gas particles during an elastic collision. Between collisions, the particles move in straight lines.

Explain the influence that gas particles have on each other, both in terms of collisions and what happens to particles between collisions.

Diffusion and effusion

According to the kinetic-molecular theory, there are no significant forces of attraction between gas particles. Thus, gas particles can flow easily past each other. Often, the space into which a gas flows is already occupied by another gas. The random motion of the gas particles causes the gases to mix until they are evenly distributed. **Diffusion** is the term used to describe the movement of one material through another. The term might be new, but you are probably familiar with the process.

If food is cooking in the kitchen, you can smell it throughout the house because the gas particles diffuse. Particles diffuse from an area of high concentration (the kitchen) to one of low concentration (the other rooms in the house).

Effusion is a process related to diffusion. During effusion, a gas escapes through a tiny opening. What happens when you puncture a container, such as a balloon or a tire?

In 1846, Thomas Graham conducted experiments to measure the rates of effusion for different gases at the same temperature. Graham designed his experiments so that the gases effused into a vacuum—space containing no matter. He discovered an inverse relationship between effusion rates and molar mass. **Graham's law of effusion** states that the rate of effusion for a gas is inversely proportional to the square root of its molar mass.

Graham's Law

$$\text{Rate of effusion} \propto \frac{1}{\sqrt{\text{molar mass}}}$$

The rate of diffusion or effusion of a gas is inversely proportional to the square root of its molar mass.

The rate of diffusion depends mainly on the mass of the particles involved. Lighter particles diffuse more rapidly than heavier particles. Recall that different gases at the same temperature have the same average kinetic energy as described by the equation $KE = \frac{1}{2}mv^2$. However, the mass of gas particles varies from gas to gas. For lighter particles to have the same average kinetic energy as heavier particles, they must have, on average, a greater velocity.

Graham's law also applies to rates of diffusion, which is logical because heavier particles diffuse more slowly than lighter particles at the same temperature. Using Graham's law, you can set up a proportion to compare the diffusion rates for two gases.

$$\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$$

EXAMPLE Problem 1

GRAHAM'S LAW Ammonia has a molar mass of 17.0 g/mol; hydrogen chloride has a molar mass of 36.5 g/mol. What is the ratio of their diffusion rates?

1 ANALYZE THE PROBLEM

You are given the molar masses for ammonia and hydrogen chloride. To find the ratio of the diffusion rates for ammonia and hydrogen chloride, use the equation for Graham's law of effusion.

Known

molar mass_{HCl} = 36.5 g/mol

molar mass_{NH₃} = 17.0 g/mol

Unknown

ratio of diffusion rates = ?

2 SOLVE FOR THE UNKNOWN

$$\frac{\text{Rate}_{\text{NH}_3}}{\text{Rate}_{\text{HCl}}} = \sqrt{\frac{\text{molar mass}_{\text{HCl}}}{\text{molar mass}_{\text{NH}_3}}}$$

$$= \sqrt{\frac{36.5 \text{ g/mol}}{17.0 \text{ g/mol}}} = 1.47$$

State the ratio derived from Graham's law.

Substitute molar mass_{HCl} = 36.5 g/mol and molar mass_{NH₃} = 17.0 g/mol.

The ratio of diffusion rates is 1.47.

3 EVALUATE THE ANSWER

A ratio of roughly 1.5 is logical because molecules of ammonia are about half as massive as molecules of hydrogen chloride. Because the molar masses have three significant figures, the answer also does. Note that the units cancel, and the answer is stated correctly without any units.

PRACTICE Problems



ADDITIONAL PRACTICE

1. Calculate the ratio of effusion rates for nitrogen (N₂) and neon (Ne).
2. Calculate the ratio of diffusion rates for carbon monoxide and carbon dioxide.
3. **CHALLENGE** What is the rate of effusion for a gas that has a molar mass twice that of a gas that effuses at a rate of 3.6 mol/min?

EXAMPLE Problem 1

GRAHAM'S LAW Ammonia has a molar mass of 17.0 g/mol; hydrogen chloride has a molar mass of 36.5 g/mol. What is the ratio of their diffusion rates?

1 ANALYZE THE PROBLEM

You are given the molar masses for ammonia and hydrogen chloride. To find the ratio of the diffusion rates for ammonia and hydrogen chloride, use the equation for Graham's law of effusion.

Known

molar mass_{HCl} = 36.5 g/mol

molar mass_{NH₃} = 17.0 g/mol

Unknown

ratio of diffusion rates = ?

2 SOLVE FOR THE UNKNOWN

$$\frac{\text{Rate}_{\text{NH}_3}}{\text{Rate}_{\text{HCl}}} = \sqrt{\frac{\text{molar mass}_{\text{HCl}}}{\text{molar mass}_{\text{NH}_3}}}$$

State the ratio derived from Graham's law.

$$= \sqrt{\frac{36.5 \text{ g/mol}}{17.0 \text{ g/mol}}} = 1.47$$

Substitute molar mass_{HCl} = 36.5 g/mol and molar mass_{NH₃} = 17.0 g/mol.

The ratio of diffusion rates is 1.47.

3 EVALUATE THE ANSWER

A ratio of roughly 1.5 is logical because molecules of ammonia are about half as massive as molecules of hydrogen chloride. Because the molar masses have three significant figures, the answer also does. Note that the units cancel, and the answer is stated correctly without any units.

PRACTICE Problems**ADDITIONAL PRACTICE**

1. Calculate the ratio of effusion rates for nitrogen (N₂) and neon (Ne).
2. Calculate the ratio of diffusion rates for carbon monoxide and carbon dioxide.
3. **CHALLENGE** What is the rate of effusion for a gas that has a molar mass twice that of a gas that effuses at a rate of 3.6 mol/min?

1. Calculate the ratio of effusion rates for nitrogen (N₂) and neon (Ne).

$$\text{Rate}_{\text{N}_2} / \text{Rate}_{\text{Ne}} = 0.8488$$

$$\frac{\text{Rate}_{\text{N}_2}}{\text{Rate}_{\text{Ne}}} = \sqrt{\frac{\text{molar mass}_{\text{Ne}}}{\text{molar mass}_{\text{N}_2}}} = \sqrt{\frac{20.18}{28.01}} = 0.8488$$

2. Calculate the ratio of diffusion rates for carbon monoxide and carbon dioxide.

$$1.253$$

$$\frac{\text{Rate}_{\text{CO}}}{\text{Rate}_{\text{CO}_2}} = \sqrt{\frac{\text{molar mass}_{\text{CO}_2}}{\text{molar mass}_{\text{CO}}}} = \sqrt{\frac{44.01}{28.01}} = 1.253$$

3. **Challenge** What is the rate of effusion for a gas that has a molar mass twice that of a gas that effuses at a rate of 3.6 mol/min?

$$2.5 \text{ mol/min}$$

$$\frac{\text{Rate}_x}{3.6 \text{ mol/min}} = \sqrt{\frac{1}{2}} \text{ Rate}_x = 2.5 \text{ mol/min}$$



Figure 4 High-heeled shoes increase the pressure on a surface because the area touching the floor is reduced. In flatter-heeled shoes, such as boots, the force is applied over a larger area.

Infer where the highest pressure is located between the floor and high-heel shoe.

Gas particles also exert pressure when they collide with the walls of their container. Because an individual gas particle has little mass, it can exert little pressure. However, a liter-sized container could hold 10^{22} gas particles. With this many particles colliding, the pressure can be high.

Air pressure

Earth is surrounded by an atmosphere that extends into space for hundreds of kilometers. Because the particles in air move in every direction, they exert pressure in all directions. This pressure is called atmospheric pressure, or air pressure. Air pressure varies at different points on Earth. Because gravity is greater at the surface of Earth, there are more particles than at higher altitudes where the force of gravity is less. Fewer particles at higher elevations exert less force than the greater concentration of particles at lower altitudes. Therefore, air pressure is less at higher altitudes than it is at sea level. At sea level, atmospheric pressure is about one-kilogram per square centimeter.

Measuring air pressure Italian physicist Evangelista Torricelli (1608–1647) was the first to demonstrate that air exerted pressure. He noticed that water pumps were unable to pump water higher than about 10 m. He hypothesized that the height of a column of liquid would vary with the density of the liquid. To test this idea, Torricelli designed the equipment shown in **Figure 5**. He filled a thin glass tube that was closed at one end with mercury. While covering the open end so that air could not enter, he inverted the tube and placed it (open end down) in a dish of mercury. The height of the mercury column fell to about one-fourteenth of a similar water column. This validated Torricelli's hypothesis because mercury is approximately fourteen times more dense than water.

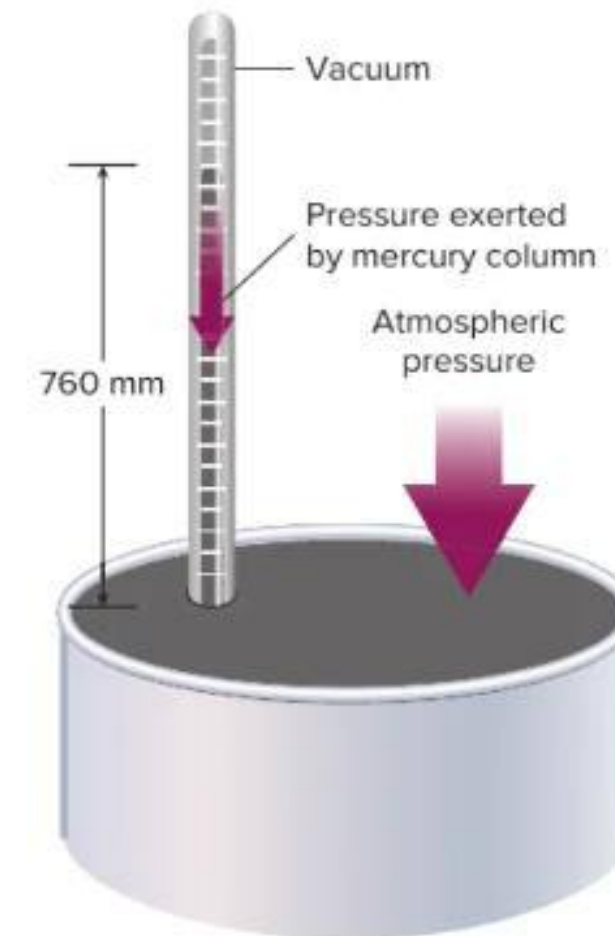


Figure 5 Torricelli was the first to design equipment to show that the atmosphere exerted pressure.

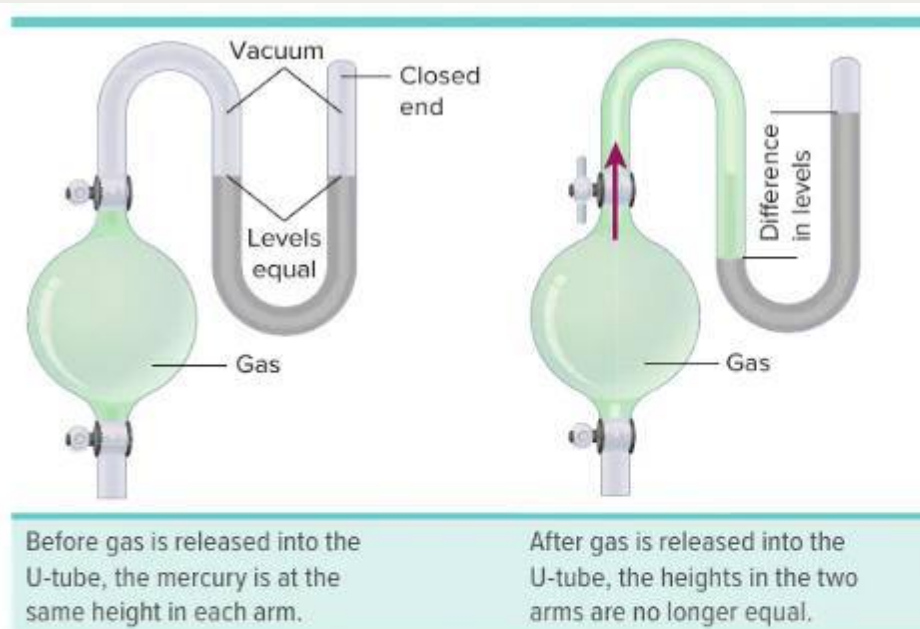


Figure 6 A manometer measures the pressure of an enclosed gas.

Barometers The device that Torricelli invented is called a barometer. A **barometer** is an instrument used to measure atmospheric pressure. As Torricelli demonstrated, at sea level the height of the mercury in a barometer is usually about 760 mm.

The exact height of the mercury is determined by two forces. Gravity exerts a constant downward force on the mercury. This force is opposed by an upward force exerted by air pressing down on the surface of the mercury. Changes in air temperature or humidity cause air pressure to vary.

Manometers A manometer is an instrument used to measure gas pressure in a closed container. In a manometer, a flask is connected to a U-tube that contains mercury, as shown in **Figure 6**. When the valve between the flask and the U-tube is opened, gas particles diffuse out of the flask into the U-tube. The released gas particles push down on the mercury in the tube. The difference in the height of the mercury in the two arms is used to calculate the pressure of the gas in the flask.

Units of pressure

The SI unit of pressure is the pascal (Pa). It is named for Blaise Pascal (1623–1662), a French mathematician and philosopher. The pascal is derived from the SI unit of force, the newton (N). One **pascal** is equal to a force of one newton per square meter: 1 Pa equals 1 N/m².

Many fields of science still use more traditional units of pressure. For example, engineers often report pressure as pounds per square inch (psi). You might see this unit of pressure listed on the tires for bicycles and cars. The pressures measured by barometers and manometers can be reported in millimeters of mercury (mmHg). However, if you listen to your local weather report, you might hear the meteorologist report the pressure in inches, which is understood to mean inches of mercury. It is similar to the unit millimeters of mercury as it describes the height of a mercury column that could be supported by the atmospheric pressure.

Unit	Number Equivalent to 1 atm	Number Equivalent to 1 kPa
Kilopascal (kPa)	101.3 kPa	—
Atmosphere (atm)	—	0.009869 atm
Millimeters of mercury (mmHg)	760 mmHg	7.501 mmHg
Torr	760 torr	7.501 torr
Pounds per square inch (psi or lb/in ²)	14.7 psi	0.145 psi
Bar	1.01 bar	0.01 bar

At sea level, the average air pressure is 101.3 kPa when the temperature is 0°C. Air pressure is often reported in a unit called an atmosphere (atm). One **atmosphere** is equal to 760 mmHg or 760 torr or 101.3 kilopascals (kPa). **Table 1** compares different units of pressure. Because the units 1 atm, 760 mmHg, and 760 torr are defined units, they should have as many significant figures as needed when used in calculations.

Dalton's law of partial pressures

When Dalton studied the properties of gases, he found that each gas in a mixture exerts pressure independently of the other gases present. Illustrated in **Figure 7**, **Dalton's law of partial pressures** states that the total pressure of a mixture of gases is equal to the sum of the pressures of all the gases in the mixture. The portion of the total pressure contributed by a single gas is called its partial pressure. The partial pressure of a gas depends on the number of moles of gas, the size of the container, and the temperature of the mixture. It does not depend on the identity of the gas. At a given temperature and pressure, the partial pressure of 1 mol of any gas is the same.

Look again at **Figure 7**. What happens when 1 mol of helium and 1 mol of nitrogen are combined in a single closed container? Because neither the volume nor the number of particles changes, the total pressure equals the sum of the two partial pressures.

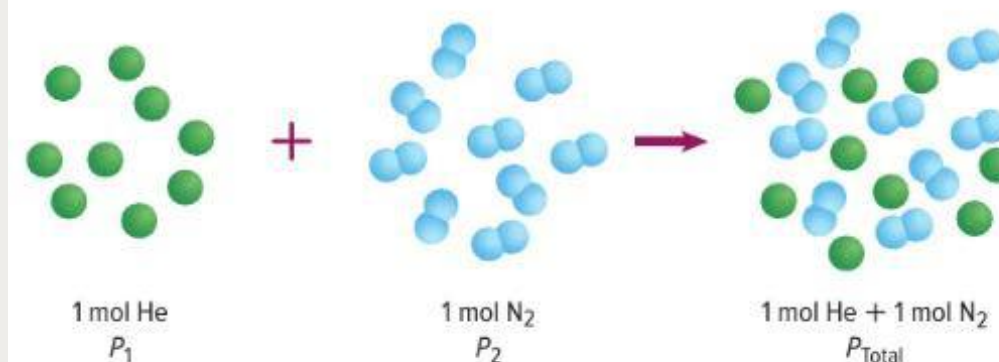


Figure 7 When gases mix, the total pressure of the mixture is equal to the sum of the partial pressures of the individual gases.

Determine How do the partial pressures of nitrogen gas and helium gas compare when a mole of nitrogen gas and a mole of helium gas are in the same closed container?

Unit	Number Equivalent to 1 atm	Number Equivalent to 1 kPa
Kilopascal (kPa)	101.3 kPa	—
Atmosphere (atm)	—	0.009869 atm
Millimeters of mercury (mmHg)	760 mmHg	7.501 mmHg
Torr	760 torr	7.501 torr
Pounds per square inch (psi or lb/in ²)	14.7 psi	0.145 psi
Bar	1.01 bar	0.01 bar

At sea level, the average air pressure is 101.3 kPa when the temperature is 0°C. Air pressure is often reported in a unit called an atmosphere (atm). One **atmosphere** is equal to 760 mmHg or 760 torr or 101.3 kilopascals (kPa). **Table 1** compares different units of pressure. Because the units 1 atm, 760 mmHg, and 760 torr are defined units, they should have as many significant figures as needed when used in calculations.

Dalton's law of partial pressures

When Dalton studied the properties of gases, he found that each gas in a mixture exerts pressure independently of the other gases present. Illustrated in **Figure 7**, **Dalton's law of partial pressures** states that the total pressure of a mixture of gases is equal to the sum of the pressures of all the gases in the mixture. The portion of the total pressure contributed by a single gas is called its partial pressure. The partial pressure of a gas depends on the number of moles of gas, the size of the container, and the temperature of the mixture. It does not depend on the identity of the gas. At a given temperature and pressure, the partial pressure of 1 mol of any gas is the same.

Look again at **Figure 7**. What happens when 1 mol of helium and 1 mol of nitrogen are combined in a single closed container? Because neither the volume nor the number of particles changes, the total pressure equals the sum of the two partial pressures.

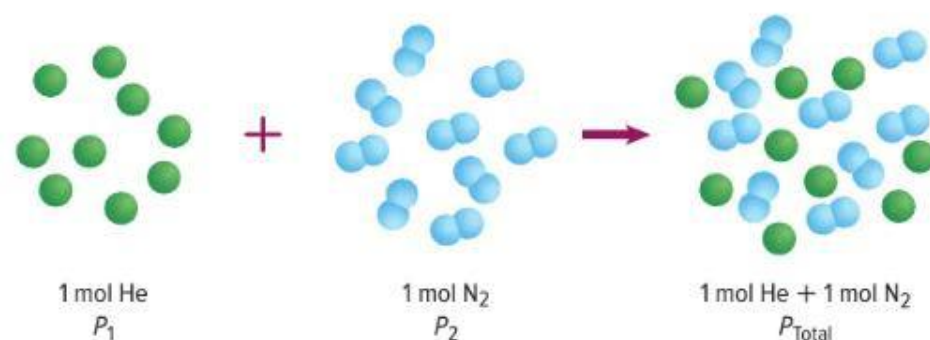


Figure 7 When gases mix, the total pressure of the mixture is equal to the sum of the partial pressures of the individual gases.

Determine How do the partial pressures of nitrogen gas and helium gas compare when a mole of nitrogen gas and a mole of helium gas are in the same closed container?

Dalton's Law of Partial Pressures

Dalton's law of partial pressures can be summarized by the following equation.

$$P_{\text{total}} = P_1 + P_2 + P_3 + \dots + P_n$$

P_{total} represents total pressure. P_1 , P_2 , and P_3 represent the partial pressures of each gas up to the final gas, P_n .

To calculate the total pressure of a mixture of gases, add the partial pressures of each of the gases in the mixture.

EXAMPLE Problem 2

THE PARTIAL PRESSURE OF A GAS A mixture of oxygen (O_2), carbon dioxide (CO_2), and nitrogen (N_2) has a total pressure of 0.97 atm. What is the partial pressure of O_2 if the partial pressure of CO_2 is 0.70 atm and the partial pressure of N_2 is 0.12 atm?

1 ANALYZE THE PROBLEM

You are given the total pressure of a mixture and the partial pressure of two gases in the mixture. To find the partial pressure of the third gas, use the equation that relates partial pressures to total pressure.

Known

$$P_{\text{N}_2} = 0.12 \text{ atm}$$

$$P_{\text{CO}_2} = 0.70 \text{ atm}$$

$$P_{\text{total}} = 0.97 \text{ atm}$$

Unknown

$$P_{\text{O}_2} = ? \text{ atm}$$

2 SOLVE FOR THE UNKNOWN

$$P_{\text{total}} = P_{\text{N}_2} + P_{\text{CO}_2} + P_{\text{O}_2}$$

State Dalton's law of partial pressures.

$$P_{\text{O}_2} = P_{\text{total}} - P_{\text{CO}_2} - P_{\text{N}_2}$$

Solve for P_{O_2} .

$$P_{\text{O}_2} = 0.97 \text{ atm} - 0.70 \text{ atm} - 0.12 \text{ atm}$$

Substitute $P_{\text{N}_2} = 0.12 \text{ atm}$, $P_{\text{CO}_2} = 0.70 \text{ atm}$, and $P_{\text{total}} = 0.97 \text{ atm}$.

$$P_{\text{O}_2} = 0.15 \text{ atm}$$

3 EVALUATE THE ANSWER

Adding the calculated value for the partial pressure of oxygen to the known partial pressures gives the total pressure, 0.97 atm. The answer has two significant figures to match the data.

PRACTICE Problems

ADDITIONAL PRACTICE

- What is the partial pressure of hydrogen gas in a mixture of hydrogen and helium if the total pressure is 600 mmHg and the partial pressure of helium is 439 mmHg?
- Find the total pressure for a mixture that contains four gases with partial pressures of 5.00 kPa, 4.56 kPa, 3.02 kPa, and 1.20 kPa.
- Find the partial pressure of carbon dioxide in a gas mixture with a total pressure of 30.4 kPa if the partial pressures of the other two gases in the mixture are 16.5 kPa and 3.7 kPa.
- CHALLENGE** Air is a mixture of gases. By percentage, it is roughly 78 percent nitrogen, 21 percent oxygen, and 1 percent argon. (There are trace amounts of many other gases in air.) If the atmospheric pressure is 760 mmHg, what are the partial pressures of nitrogen, oxygen, and argon in the atmosphere?

17	CHM.5.2.01.004.07 Use the mathematical formula of Dalton's law of partial pressures to calculate partial pressures and total pressure of a mixture of gases	Textbook + Figure 7 + Example problem 2 + Practice Problems	244, 245
----	---	---	----------

Dalton's Law of Partial Pressures

Dalton's law of partial pressures can be summarized by the following equation.

$$P_{\text{total}} = P_1 + P_2 + P_3 + \dots + P_n$$

P_{total} represents total pressure. P_1 , P_2 , and P_3 represent the partial pressures of each gas up to the final gas, P_n .

To calculate the total pressure of a mixture of gases, add the partial pressures of each of the gases in the mixture.

EXAMPLE Problem 2

THE PARTIAL PRESSURE OF A GAS A mixture of oxygen (O_2), carbon dioxide (CO_2), and nitrogen (N_2) has a total pressure of 0.97 atm. What is the partial pressure of O_2 if the partial pressure of CO_2 is 0.70 atm and the partial pressure of N_2 is 0.12 atm?

1 ANALYZE THE PROBLEM

You are given the total pressure of a mixture and the partial pressure of two gases in the mixture. To find the partial pressure of the third gas, use the equation that relates partial pressures to total pressure.

Known

$$P_{\text{N}_2} = 0.12 \text{ atm}$$

$$P_{\text{CO}_2} = 0.70 \text{ atm}$$

$$P_{\text{total}} = 0.97 \text{ atm}$$

Unknown

$$P_{\text{O}_2} = ? \text{ atm}$$

2 SOLVE FOR THE UNKNOWN

$$P_{\text{total}} = P_{\text{N}_2} + P_{\text{CO}_2} + P_{\text{O}_2}$$

State Dalton's law of partial pressures.

$$P_{\text{O}_2} = P_{\text{total}} - P_{\text{CO}_2} - P_{\text{N}_2}$$

Solve for P_{O_2} .

$$P_{\text{O}_2} = 0.97 \text{ atm} - 0.70 \text{ atm} - 0.12 \text{ atm}$$

Substitute $P_{\text{N}_2} = 0.12 \text{ atm}$, $P_{\text{CO}_2} = 0.70 \text{ atm}$, and $P_{\text{total}} = 0.97 \text{ atm}$.

$$P_{\text{O}_2} = 0.15 \text{ atm}$$

3 EVALUATE THE ANSWER

Adding the calculated value for the partial pressure of oxygen to the known partial pressures gives the total pressure, 0.97 atm. The answer has two significant figures to match the data.

4. What is the partial pressure of hydrogen gas in a mixture of hydrogen and helium if the total pressure is 600 mmHg and the partial pressure of helium is 439 mmHg?

$$161 \text{ mmHg}$$

$$600 \text{ mmHg} - 439 \text{ mmHg} = 161 \text{ mmHg}$$

5. Find the total pressure for a mixture that contains four gases with partial pressures of 5.00 kPa, 4.56 kPa, 3.02 kPa, and 1.20 kPa.

$$5.00 \text{ kPa} + 4.56 \text{ kPa} + 3.02 \text{ kPa} + 1.20 \text{ kPa} = 13.78 \text{ kPa}$$

6. Find the partial pressure of carbon dioxide in a gas mixture with a total pressure of 30.4 kPa if the partial pressures of the other two gases in the mixture are 16.5 kPa and 3.7 kPa.

$$30.4 \text{ kPa} - 16.5 \text{ kPa} - 3.7 \text{ kPa} = 10.2 \text{ kPa}$$

7. **Challenge** Air is a mixture of gases. By percentage, it is roughly 78 percent nitrogen, 21 percent oxygen, and 1 percent argon. (There are trace amounts of many other gases in air.) If the atmospheric pressure is 760 mmHg, what are the partial pressures of nitrogen, oxygen, and argon in the atmosphere?

$$760 \text{ mmHg} \times 0.78 = \text{N}_2 = 590 \text{ mmHg};$$

$$760 \text{ mmHg} \times 0.21 = \text{O}_2 = 160 \text{ mmHg};$$

$$760 \text{ mmHg} \times 0.01 = \text{Ar} = 8 \text{ mmHg}$$

PRACTICE Problems



ADDITIONAL PRACTICE

- What is the partial pressure of hydrogen gas in a mixture of hydrogen and helium if the total pressure is 600 mmHg and the partial pressure of helium is 439 mmHg?
- Find the total pressure for a mixture that contains four gases with partial pressures of 5.00 kPa, 4.56 kPa, 3.02 kPa, and 1.20 kPa.
- Find the partial pressure of carbon dioxide in a gas mixture with a total pressure of 30.4 kPa if the partial pressures of the other two gases in the mixture are 16.5 kPa and 3.7 kPa.
- CHALLENGE** Air is a mixture of gases. By percentage, it is roughly 78 percent nitrogen, 21 percent oxygen, and 1 percent argon. (There are trace amounts of many other gases in air.) If the atmospheric pressure is 760 mmHg, what are the partial pressures of nitrogen, oxygen, and argon in the atmosphere?

What is the partial pressure of **oxygen** in an air sample when the total pressure is **1.10 atm**, the partial pressure of nitrogen is **0.75 atm**, the partial pressure of water vapor is **0.0056 atm**, and the partial pressure of all other gases in air is **0.0044 atm**?

ما هو الضغط الجزئي **للأكسجين** في عينة من الهواء عندما يكون الضغط الكلي **1.10 atm** ، والضغط الجزئي للنيتروجين **0.75 atm** والضغط الجزئي لبخار الماء **0.0056 atm** والضغط الجزئي لجميع الغازات الأخرى في الهواء **0.0044 atm** ؟

0.20 atm

0.21 atm

0.34 atm

0.36 atm

A sealed flask contains oxygen, helium, and nitrogen.

If the total pressure in the flask is 4.711 atm, the partial pressure for O_2 is 2.592 atm, and the partial pressure for He is 0.836 atm, what is the partial pressure of N_2 ?

وعاء مغلق يحتوي خليط من غازات الأكسجين والهيليوم والنيتروجين. إذا كان الضغط الكلي في الوعاء 4.711 atm ، والضغط الجزئي لـ O_2 هو 2.592 atm ، والضغط الجزئي لـ He هو 0.836 atm ، ما هو الضغط الجزئي لـ N_2 ؟



2.955 atm

8.139 atm

0.467 atm

1.283 atm

What is the total pressure for a mixture that contains three gases with partial pressures of 1.35 kPa, 3.81 kPa, and 5.22 kPa?

ما الضغط الكلي لخليط يحتوي على ثلاث غازات بضغطها الجزئية كالتالي 1.35 kPa ، 3.81 kPa ، 5.22 kPa ؟



7.68 kPa

10.38 kPa

12.76 kPa

6.57 kPa

What is the partial pressure of water vapor in an air sample when the total pressure is **1.00 atm**, the partial pressure of nitrogen is **0.79 atm**, the partial pressure of oxygen is **0.20 atm**, and the partial pressure of all other gases in air is **0.0044 atm**?

ما الضغط الجزئي لبخار الماء في عينة من الهواء عندما يكون الضغط الكلي **1.00 atm** والضغط الجزئي للنيتروجين **0.79 atm** والضغط الجزئي للأكسجين **0.20 atm** والضغط الجزئي لجميع الغازات الأخرى **0.0044 atm** ؟

0.0056 atm

0.2100 atm

0.80 atm

0.9956 atm

Boyle's Law

Robert Boyle (1627–1691), an Irish chemist, described the relationship between the pressure and the volume of a gas, such as the air in a balloon.

How are pressure and volume related?

Boyle designed experiments like the one shown in **Figure 1**. He showed that if the temperature and the amount of gas are constant, doubling the pressure decreases the volume by one-half. On the other hand, reducing the pressure by one-half doubles the volume. A relationship in which one variable increases proportionally as the other variable decreases is known as an inversely proportional relationship.

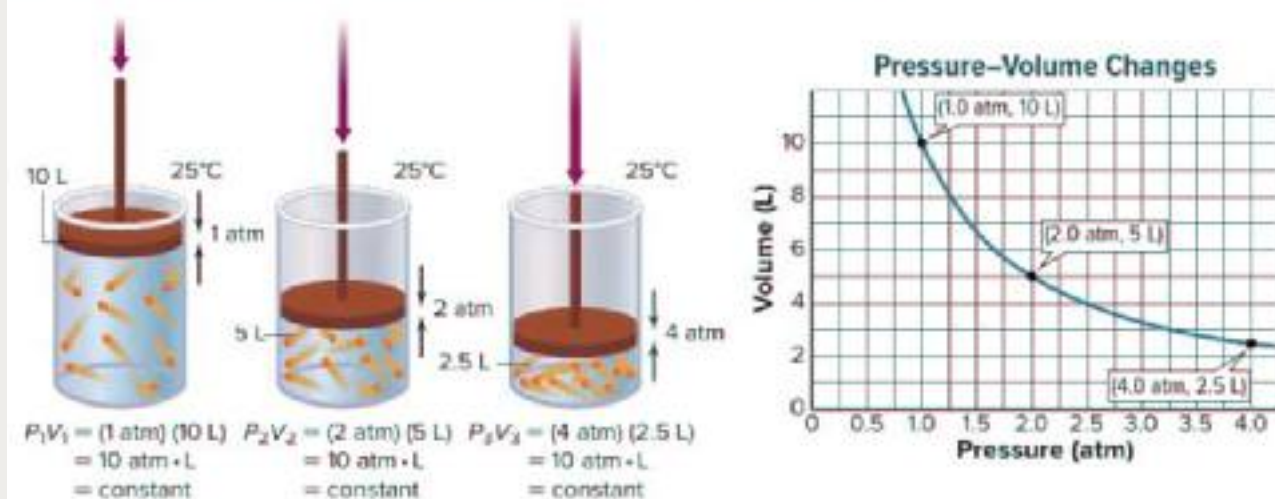


Figure 1 As the external pressure on the cylinder's piston increases, the volume inside the cylinder decreases. The graph shows the inverse relationship between pressure and volume.

Boyle's law states that the volume of a fixed amount of gas held at a constant temperature varies inversely with the pressure. Look at the graph in **Figure 1**, on the previous page, in which volume versus pressure is plotted for a gas. The plot of an inversely proportional relationship results in a downward curve.

Note that the product of the pressure and the volume for each point in **Figure 1** is 10 atm·L. Boyle's law can be expressed mathematically as follows.

Boyle's Law

$$P_1V_1 = P_2V_2 \quad P \text{ represents pressure. } V \text{ represents volume.}$$

For a given amount of gas held at constant temperature, the product of pressure and volume is constant.

P_1 and V_1 represent the initial conditions, and P_2 and V_2 represent new conditions. If you know any three of these values, you can solve for the fourth by rearranging the equation.

EXAMPLE Problem 1

BOYLE'S LAW A diver blows a 0.75-L air bubble 10 m under water. As it rises to the surface, the pressure goes from 2.25 atm to 1.03 atm. What will be the volume of air in the bubble at the surface?

1 ANALYZE THE PROBLEM

According to Boyle's law, the decrease in pressure on the bubble will result in an increase in volume, so the initial volume should be multiplied by a pressure ratio greater than 1.

Known

$$\begin{aligned} V_1 &= 0.75 \text{ L} \\ P_1 &= 2.25 \text{ atm} \\ P_2 &= 1.03 \text{ atm} \end{aligned}$$

Unknown

$$V_2 = ? \text{ L}$$

2 SOLVE FOR THE UNKNOWN

Use Boyle's law. Solve for V_2 , and calculate the new volume.

$$P_1V_1 = P_2V_2$$

State Boyle's law.

$$V_2 = V_1 \left(\frac{P_1}{P_2} \right)$$

Solve for V_2 .

$$V_2 = 0.75 \text{ L} \left(\frac{2.25 \text{ atm}}{1.03 \text{ atm}} \right)$$

Substitute $V_1 = 0.75 \text{ L}$, $P_1 = 2.25 \text{ atm}$, and $P_2 = 1.03 \text{ atm}$.

$$V_2 = 0.75 \text{ L} \left(\frac{2.25 \text{ atm}}{1.03 \text{ atm}} \right) = 1.6 \text{ L}$$

Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

The pressure decreases by roughly half, so the volume should roughly double. The answer is expressed in liters, a unit of volume, and correctly contains two significant figures.

PRACTICE Problems

ADDITIONAL PRACTICE

Assume that the temperature and the amount of gas are constant in the following problems.

1. The volume of a gas at 99.0 kPa is 300.0 mL. If the pressure is increased to 188 kPa, what will be the new volume?
2. The pressure of a sample of helium in a 1.00-L container is 0.988 atm. What is the new pressure if the sample is placed in a 2.00-L container?
3. **CHALLENGE** Air trapped in a cylinder fitted with a piston occupies 145.7 mL at 1.08 atm pressure. What is the new volume when the piston is depressed, increasing the pressure by 25%?

EXAMPLE Problem 1

BOYLE'S LAW A diver blows a 0.75-L air bubble 10 m under water. As it rises to the surface, the pressure goes from 2.25 atm to 1.03 atm. What will be the volume of air in the bubble at the surface?

1 ANALYZE THE PROBLEM

According to Boyle's law, the decrease in pressure on the bubble will result in an increase in volume, so the initial volume should be multiplied by a pressure ratio greater than 1.

Known

$$V_1 = 0.75 \text{ L}$$

$$P_1 = 2.25 \text{ atm}$$

$$P_2 = 1.03 \text{ atm}$$

Unknown

$$V_2 = ? \text{ L}$$

2 SOLVE FOR THE UNKNOWN

Use Boyle's law. Solve for V_2 , and calculate the new volume.

$$P_1 V_1 = P_2 V_2$$

State Boyle's law.

$$V_2 = V_1 \left(\frac{P_1}{P_2} \right)$$

Solve for V_2 .

$$V_2 = 0.75 \text{ L} \left(\frac{2.25 \text{ atm}}{1.03 \text{ atm}} \right)$$

Substitute $V_1 = 0.75 \text{ L}$, $P_1 = 2.25 \text{ atm}$, and $P_2 = 1.03 \text{ atm}$.

$$V_2 = 0.75 \text{ L} \left(\frac{2.25 \text{ atm}}{1.03 \text{ atm}} \right) = 1.6 \text{ L}$$

Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

The pressure decreases by roughly half, so the volume should roughly double. The answer is expressed in liters, a unit of volume, and correctly contains two significant figures.

Assume that the temperature and the amount of gas are constant in the following problems.

1. The volume of a gas at 99.0 kPa is 300.0 mL. If the pressure is increased to 188 kPa, what will be the new volume?

$$158 \text{ mL}$$

$$V_2 = \frac{V_1 P_1}{P_2} = \frac{(300.0 \text{ mL})(99.0 \text{ kPa})}{188 \text{ kPa}} = 158 \text{ mL}$$

2. The pressure of a sample of helium in a 1.00-L container is 0.988 atm. What is the new pressure if the sample is placed in a 2.00-L container?

$$0.494 \text{ atm}$$

$$P_2 = \frac{V_1 P_1}{V_2} = \frac{(1.00 \text{ L})(0.988 \text{ atm})}{2.00 \text{ L}} = 0.494 \text{ atm}$$

3. **Challenge** Air trapped in a cylinder fitted with a piston occupies 145.7 mL at 1.08 atm pressure. What is the new volume when the piston is depressed, increasing the pressure by 25%?

$$117 \text{ mL}$$

$$P_2 = 1.08 \text{ atm} + (1.08 \text{ atm} \times 0.25) = 1.35 \text{ atm}$$

$$V_2 = \frac{V_1 P_1}{P_2} = \frac{(145.7 \text{ mL})(1.08 \text{ atm})}{1.35 \text{ atm}} = 117 \text{ mL}$$

PRACTICE Problems**ADDITIONAL PRACTICE**

Assume that the temperature and the amount of gas are constant in the following problems.

1. The volume of a gas at 99.0 kPa is 300.0 mL. If the pressure is increased to 188 kPa, what will be the new volume?
2. The pressure of a sample of helium in a 1.00-L container is 0.988 atm. What is the new pressure if the sample is placed in a 2.00-L container?
3. **CHALLENGE** Air trapped in a cylinder fitted with a piston occupies 145.7 mL at 1.08 atm pressure. What is the new volume when the piston is depressed, increasing the pressure by 25%?

The volume of a gas is 400.0 mL, and the pressure is 1.00 atm. When the volume of the gas is 2.0 L, what is the pressure, if the temperature remains the same?

حجم غاز 400.0 ml وعند ضغط 1.00 atm
إذا أصبح حجم الغاز 2.0 L ، ما ضغط الغاز عند نفس درجة الحرارة؟

0.20 atm

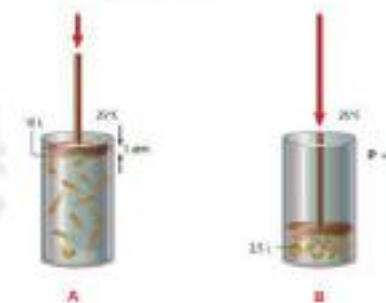
0.5 atm

5.0 atm

0.80 atm

As the external pressure on the cylinder's piston increases in the figure below. The gas volume changes from 10L in A to 2.5 L in B. What is the new pressure on the cylinder's piston in B?

عند زيادة الضغط الخارجي على مكبس الأسطوانة في الشكل أدناه، تغير حجم الغاز من 10 L في A إلى 2.5 L في B. ما الضغط الجديد على مكبس الأسطوانة في B؟



2 atm

4 atm

6 atm

8 atm

Air trapped in a cylinder fitted with a piston occupies 365.5 mL at 0.985 atm pressure. What is the new volume (mL) when the piston is depressed, increasing the pressure by 50%?

هواء محصور في أسطوانة مغلقة بمكبس يشغل 365.5 mL عند ضغط 0.985 atm ، ما الحجم الجديد (mL) إذا تم ضغط المكبس بحيث يزداد الضغط بمقدار 50%؟

244

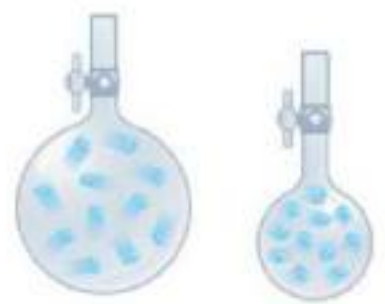
354

198

455

The pressure of a sample of helium in a 1.0 L container is 0.857 atm. What is the pressure if the same sample is placed in a 0.50 L container? (Assume that the temperature is constant.)

ضغط عينة من الهيليوم في حاوية سعة 1.0 L هو 0.857 atm. ما الضغط إذا تم وضع نفس العينة في حاوية سعة 0.50 L؟ (افترض أن درجة الحرارة ثابتة)



0.44 atm

1.4 atm

1.7 atm

0.14 atm

Charles's Law

After a cool evening, a rubber pool raft can appear partially inflated. During a sunny afternoon, the same raft can appear fully inflated. Why does the appearance of the raft change? This can be answered by applying a second gas law—Charles's law.

How are temperature and volume related?

Jacques Charles (1746–1823), a French physicist, studied the relationship between volume and temperature. He observed that as temperature increases, so does the volume of a gas sample when the amount of gas and the pressure remain constant. The cylinders in **Figure 2** show how the volume of a fixed amount of gas changes as the gas is heated. Unlike **Figure 1**, where pressure in addition to that of the atmosphere was applied to the piston, the piston in **Figure 2** is free to float. The piston is supported by the gas in the cylinder at a level where the pressure of the gas exactly matches that of the atmosphere.

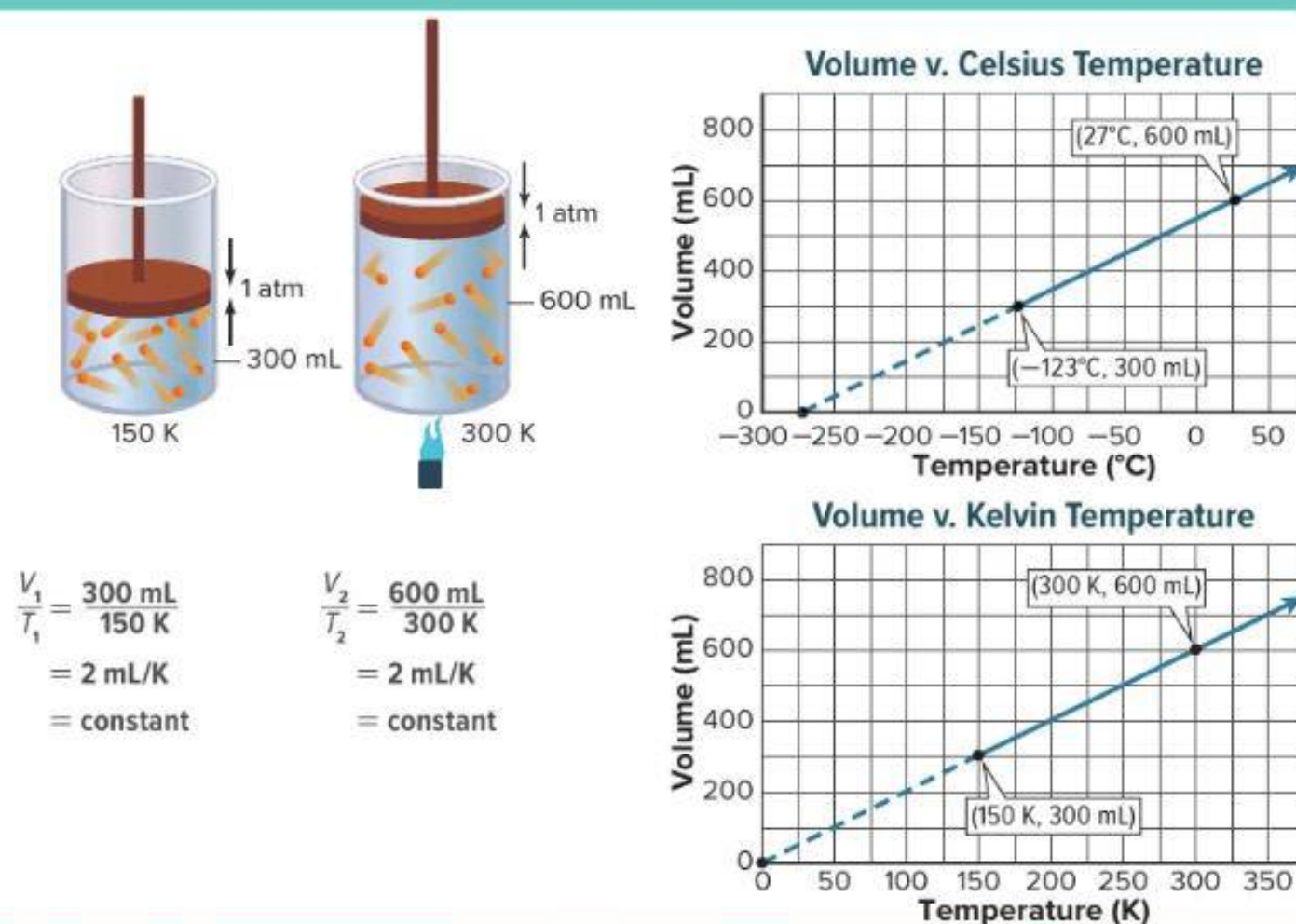


Figure 2 When the cylinder is heated, the kinetic energy of the gas particles increases, causing them to push the piston outward. The graphs show the relationship of volume to Celsius and Kelvin temperatures.

19	CHM.5.2.01.004.15 Use the mathematical formula of Charles's law to calculate volume–temperature changes at constant pressure	Textbook + Figure 2 + Example problem 2 + Practice Problems	277, 278, 279
20	CHM.5.2.01.004.19 Use the mathematical formula of Gay-Lussac's law to calculate pressure–temperature changes at constant volume	Textbook + Figure 3 + Example problem 3 + Practice Problems	280, 281

Charles's Law

After a cool evening, a rubber pool raft can appear partially inflated. During a sunny afternoon, the same raft can appear fully inflated. Why does the appearance of the raft change? This can be answered by applying a second gas law—Charles's law.

How are temperature and volume related?

Jacques Charles (1746–1823), a French physicist, studied the relationship between volume and temperature. He observed that as temperature increases, so does the volume of a gas sample when the amount of gas and the pressure remain constant. The cylinders in **Figure 2** show how the volume of a fixed amount of gas changes as the gas is heated. Unlike **Figure 1**, where pressure in addition to that of the atmosphere was applied to the piston, the piston in **Figure 2** is free to float. The piston is supported by the gas in the cylinder at a level where the pressure of the gas exactly matches that of the atmosphere.

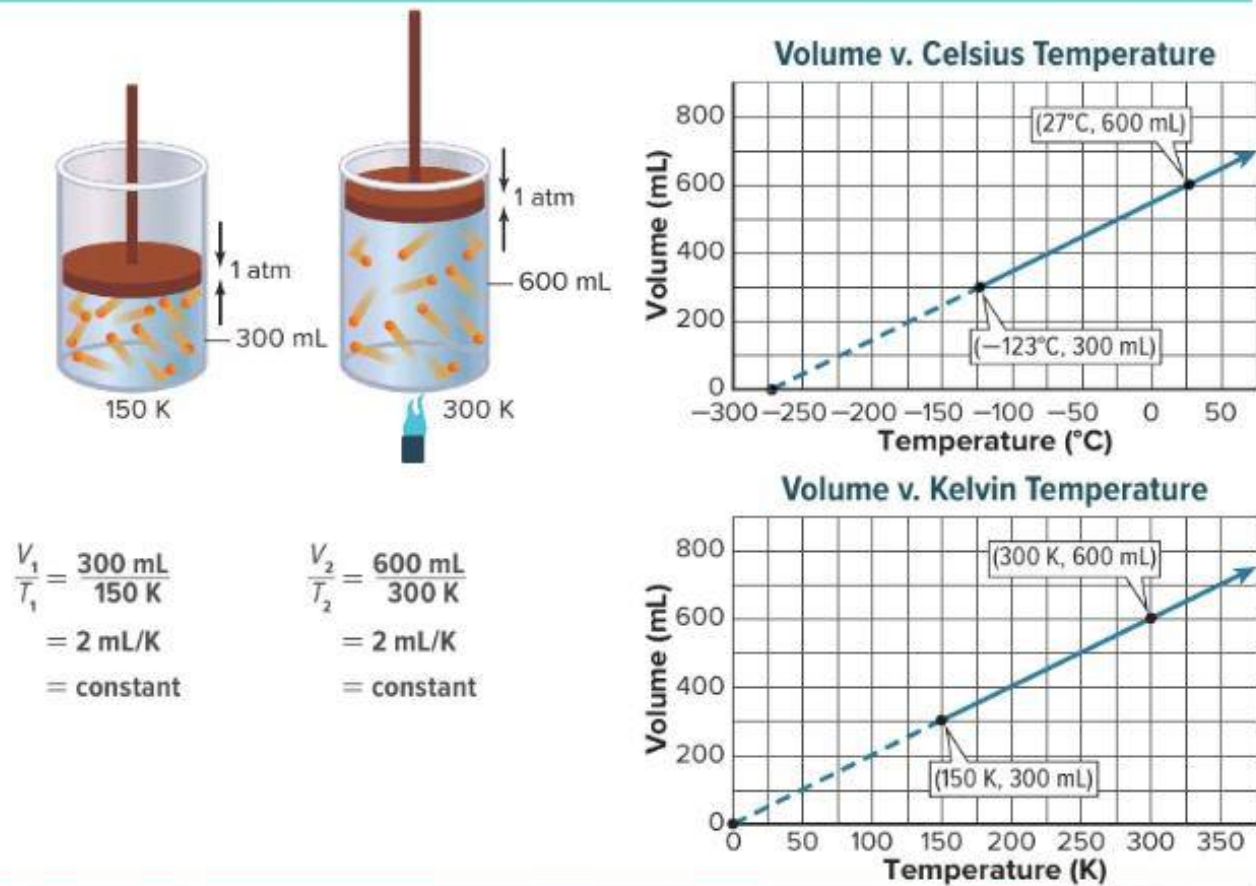


Figure 2 When the cylinder is heated, the kinetic energy of the gas particles increases, causing them to push the piston outward. The graphs show the relationship of volume to Celsius and Kelvin temperatures.

Graphing the relationship of temperature and volume

Figure 2 also shows graphs of the relationship between the temperature and the volume of a fixed amount of gas at constant pressure. The plot of volume versus temperature is a straight line. Note that you can predict the temperature at which the volume will reach 0 L by extrapolating the line to temperatures below the values that were measured.

In the first graph, the temperature that corresponds to 0 L is -273.15°C . This relationship is linear, but it is not a direct proportion. For example, you can see that the graph of the line does not pass through the origin and that doubling the temperature from 25°C to 50°C does not double the volume.

The second graph in **Figure 2**, which plots the Kelvin (K) temperature against volume, does show a direct proportion. A temperature of 0 K corresponds to 0 mL, and doubling the temperature doubles the volume. Zero on the Kelvin scale is also known as **absolute zero**. Absolute zero represents the lowest possible theoretical temperature. At absolute zero, the atoms are all in the lowest possible energy state.



Get It?

Explain why the second graph in **Figure 2** shows a direct proportion, but the first graph does not.

Using Charles's law

Charles's law states that the volume of a given amount of gas is directly proportional to its Kelvin temperature at constant pressure. Charles's law can be expressed as follows.

Charles's law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

V represents volume.
T represents Temperature.

For a given amount of gas at constant pressure, the quotient of the volume and Kelvin temperature is a constant.

V_1 and T_1 represent initial conditions, while V_2 and T_2 are new conditions. As with Boyle's law, if you know three of the values, you can calculate the fourth. The temperature must be expressed in kelvin when using the equation for Charles's law. To convert a temperature from Celsius degrees to kelvin, add 273 to the Celsius temperature:

$$T_K = 273 + T_C$$

19	CHM.5.2.01.004.15 Use the mathematical formula of Charles's law to calculate volume–temperature changes at constant pressure	Textbook + Figure 2 + Example problem 2 + Practice Problems	277, 278, 279
20	CHM.5.2.01.004.19 Use the mathematical formula of Gay–Lussac's law to calculate pressure–temperature changes at constant volume	Textbook + Figure 3 + Example problem 3 + Practice Problems	280, 281

EXAMPLE Problem 2

CHARLES'S LAW A helium balloon in a closed car occupies a volume of 2.32 L at 40.0°C. If the car is parked on a hot day and the temperature inside rises to 75.0°C, what is the new volume of the balloon, assuming the pressure remains constant?

1 ANALYZE THE PROBLEM

Charles's law states that as the temperature of a fixed amount of gas increases, so does its volume, assuming constant pressure. Therefore, the volume of the balloon will increase. The initial volume should be multiplied by a temperature ratio greater than 1.

Known
 $T_1 = 40.0^\circ\text{C}$
 $V_1 = 2.32\text{ L}$
 $T_2 = 75.0^\circ\text{C}$

Unknown
 $V_2 = ?\text{ L}$

2 SOLVE FOR THE UNKNOWN

Convert degrees Celsius to kelvin.

$T_1 = 273 + T_c$ Apply the conversion factor.
 $T_1 = 273 + 40.0^\circ\text{C} = 313.0\text{ K}$ Substitute $T_c = 40.0^\circ\text{C}$.
 $T_2 = 273 + 75.0^\circ\text{C} = 348.0\text{ K}$ Substitute $T_c = 75.0^\circ\text{C}$.

Use Charles's law. Solve for V_2 , and substitute the known values into the rearranged equation.

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ State Charles's law.

$V_2 = V_1 \left(\frac{T_2}{T_1} \right)$ Solve for V_2 .

$V_2 = 2.32\text{ L} \left(\frac{348.0\text{ K}}{313.0\text{ K}} \right)$ Substitute $V_1 = 2.32\text{ L}$, $T_1 = 313.0\text{ K}$, and $T_2 = 348.0\text{ K}$.

$V_2 = 2.32\text{ L} \left(\frac{348.0\text{ K}}{313.0\text{ K}} \right) = 2.58\text{ L}$ Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

The increase in temperature is relatively small, so the volume should show a small increase. The unit of the answer is liters, a volume unit, and there are three significant figures.

PRACTICE Problems

Assume that the pressure and the amount of gas remain constant in the following problems.

- What volume will the gas in the balloon at right occupy at 250 K?
- A gas at 89°C occupies a volume of 0.67 L. At what Celsius temperature will the volume increase to 1.12 L?
- The Celsius temperature of a 3.00-L sample of gas is lowered from 80.0°C to 30.0°C. What will be the resulting volume of this gas?
- CHALLENGE** A gas occupies 0.67 L at 350 K. What temperature is required to reduce the volume by 45%?



Assume that the pressure and the amount of gas are constant in the following problems.

- What volume will the gas in the balloon below occupy at 250 K?



3.1 L

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{(4.3\text{ L})(250\text{ K})}{350\text{ K}} = 3.1\text{ L}$$

- A gas at 89°C occupies a volume of 0.67 L. At what Celsius temperature will the volume increase to 1.12 L?

330°C

$$T_1 = 89^\circ\text{C} + 273 = 362\text{ K}$$

$$T_2 = \frac{T_1 V_2}{V_1} = \frac{(362\text{ K})(1.12\text{ L})}{0.67\text{ L}} = 605\text{ K}$$

$$605 - 273 = 332^\circ\text{C} = 330^\circ\text{C}$$

- The Celsius temperature of a 3.00-L sample of gas is lowered from 80.0°C to 30.0°C. What will be the resulting volume of this gas?

2.58 L

$$T_1 = 80.0^\circ\text{C} + 273 = 353\text{ K}$$

$$T_2 = 30.0^\circ\text{C} + 273 = 303\text{ K}$$

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{(3.00\text{ L})(303\text{ K})}{353\text{ K}} = 2.58\text{ L}$$

- Challenge** A gas occupies 0.67 L at 350 K. What temperature is required to reduce the volume by 45%?

190 K

$$V_2 = 0.67\text{ L} - (0.67\text{ L} \times 0.45) = 0.37\text{ L}$$

$$T_2 = \frac{T_1 V_2}{V_1} = \frac{(350\text{ K})(0.37\text{ L})}{0.67\text{ L}} = 190\text{ K}$$

19	CHM.5.2.01.004.15 Use the mathematical formula of Charles's law to calculate volume–temperature changes at constant pressure	Textbook + Figure 2 + Example problem 2 + Practice Problems	277, 278, 279
20	CHM.5.2.01.004.19 Use the mathematical formula of Gay–Lussac's law to calculate pressure–temperature changes at constant volume	Textbook + Figure 3 + Example problem 3 + Practice Problems	280, 281

A gas occupies a volume of **0.85 L** at **375 K**. What temperature (**K**) is required to reduce the volume to **60%** of the original volume increase?

يشغل غاز حجمًا مقداره **0.85 L** عند درجة حرارة **375 K**. فما درجة الحرارة بوحدة (**K**) المطلوبة لخفض الحجم إلى **60 %** من الحجم الأصلي؟

938

625

225

150

A sample of gas occupies a volume of **6.50 L** at **95.0 °C**. What is the Celsius temperature (**°C**) at which the volume of the gas sample will become **1.63 L**?

تشغل عينة من غاز حجمًا **6.50 L** عند درجة حرارة **95.0 °C** فما درجة الحرارة (**°C**) التي يُصبح عندها حجم عينة الغاز **1.63 L**؟

-181

-92

365

418

What is the volume for the gas in the below balloon when temperature changes to 348 K?

ما حجم الغاز الموجود في البالون أدناه عندما تتغير درجة الحرارة إلى 348 K؟



4.01 L

3.84 L

2.73 L

2.31 L

19	CHM.5.2.01.004.15 Use the mathematical formula of Charles's law to calculate volume–temperature changes at constant pressure	Textbook + Figure 2 + Example problem 2 + Practice Problems	277, 278, 279
20	CHM.5.2.01.004.19 Use the mathematical formula of Gay-Lussac's law to calculate pressure–temperature changes at constant volume	Textbook + Figure 3 + Example problem 3 + Practice Problems	280, 281

Gay-Lussac's Law

If volume is constant, is there a relationship between temperature and pressure? The answer to that question is found in Gay-Lussac's law.

How are temperature and pressure of a gas related?

Pressure is a direct result of collisions between gas particles and the walls of their container. An increase in temperature increases collision frequency and energy, so raising the temperature should also raise the pressure if the volume is not changed. Joseph Gay-Lussac (1778–1850) found that a direct proportion exists between Kelvin temperature and pressure, as illustrated in Figure 3. **Gay-Lussac's law** states that the pressure of a fixed amount of gas varies directly with the Kelvin temperature when the volume remains constant. It can be expressed mathematically as follows.

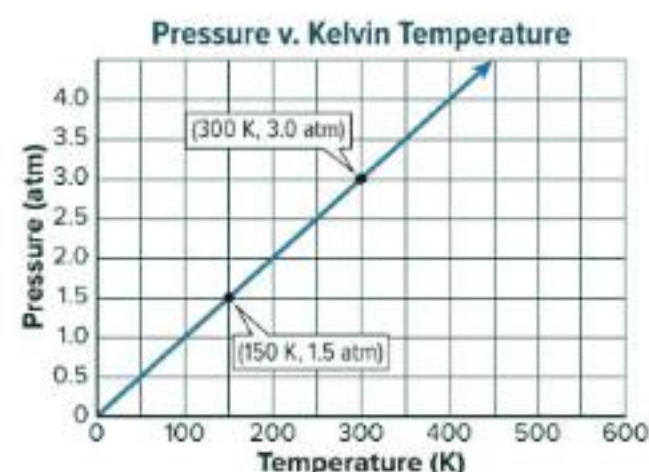
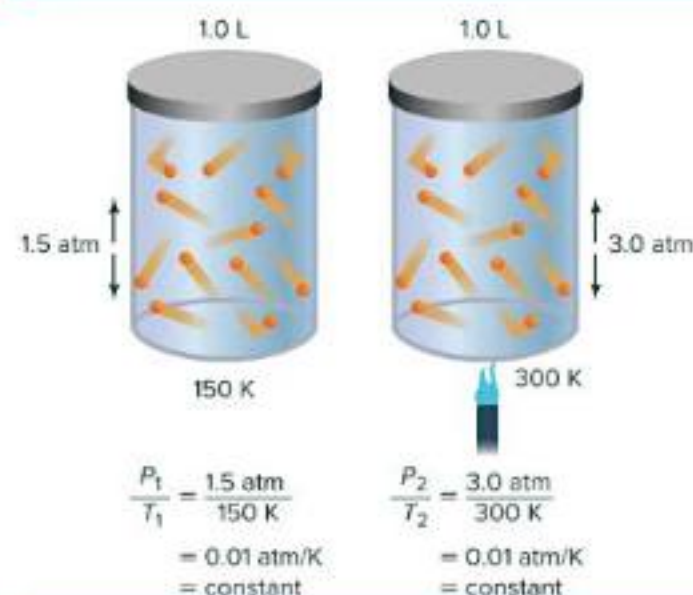
Gay-Lussac's Law

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

P represents pressure.
 T represents temperature.

For a given amount of gas held at constant volume, the quotient of the pressure and the Kelvin temperature is a constant.

As with Boyle's and Charles's laws, if you know any three of the four variables, you can calculate the fourth using this equation. Remember that temperature must be in kelvin whenever it is used in a gas law equation.



EXAMPLE Problem 3

GAY-LUSSAC'S LAW The pressure of the oxygen gas inside a canister is 5.00 atm at 25.0°C. The canister is located at a camp high on Mount Everest. If the temperature there falls to −10.0°C, what is the new pressure inside the canister?

1 ANALYZE THE PROBLEM

Gay-Lussac's law states that if the temperature of a gas decreases, so does its pressure when volume is constant. Therefore, the pressure in the oxygen canister will decrease. The initial pressure should be multiplied by a temperature ratio less than 1.

Known
 $P_1 = 5.00 \text{ atm}$
 $T_1 = 25.0^\circ\text{C}$
 $T_2 = -10.0^\circ\text{C}$

Unknown
 $P_2 = ? \text{ atm}$

2 SOLVE FOR THE UNKNOWN

Convert degrees Celsius to kelvin.

$$T_2 = 273 + T_c$$

Apply the conversion factor.

$$T_1 = 273 + 25.0^\circ\text{C} = 298.0 \text{ K}$$

Substitute $T_1 = 25.0^\circ\text{C}$.

$$T_2 = 273 + (-10.0^\circ\text{C}) = 263.0 \text{ K}$$

Substitute $T_2 = -10.0^\circ\text{C}$.

Use Gay-Lussac's law. Solve for P_2 , and substitute the known values into the rearranged equation.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

State Gay-Lussac's law.

$$P_2 = P_1 \left(\frac{T_2}{T_1} \right)$$

Solve for P_2 .

$$P_2 = 5.00 \text{ atm} \left(\frac{263.0 \text{ K}}{298.0 \text{ K}} \right)$$

Substitute $P_1 = 5.00 \text{ atm}$,
 $T_1 = 298.0 \text{ K}$, and $T_2 = 263.0 \text{ K}$.

$$P_2 = 5.00 \text{ atm} \left(\frac{263.0 \text{ K}}{298.0 \text{ K}} \right) = 4.41 \text{ atm}$$

Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

Kelvin temperature decreases, so the pressure should decrease. The unit is atm, a pressure unit, and there are three significant figures.

Real-World Chemistry Gay-Lussac's Law



PRESSURE COOKERS A pressure cooker is a pot with a lid that locks into place. This seals the container, which keeps its volume constant. Heating the pot increases the pressure in the cooker. As pressure increases, the temperature continues to increase and foods cook faster.

PRACTICE Problems

Assume that the volume and the amount of gas are constant in the following problems.

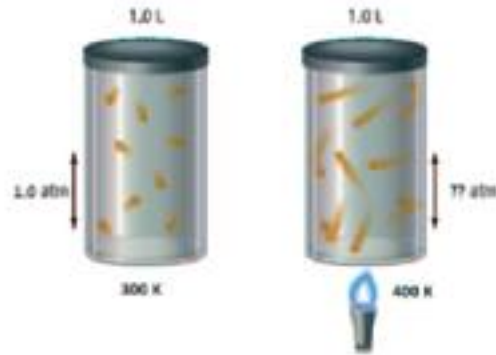
- The pressure in an automobile tire is 1.88 atm at 25.0°C. What will be the pressure if the temperature increases to 37.0°C?
- Helium gas in a 2.00-L cylinder is under 112 atm pressure. At 36.5°C, that same gas sample has a pressure of 2.56 atm. What was the initial temperature in degrees Celsius of the gas in the cylinder?
- CHALLENGE** If a gas sample has a pressure of 30.7 kPa at 0.00°C, by how many degrees Celsius does the temperature have to increase to cause the pressure to double?

ADDITIONAL PRACTICE

19	CHM.5.2.01.004.15 Use the mathematical formula of Charles's law to calculate volume–temperature changes at constant pressure	Textbook + Figure 2 + Example problem 2 + Practice Problems	277, 278, 279
20	CHM.5.2.01.004.19 Use the mathematical formula of Gay–Lussac's law to calculate pressure–temperature changes at constant volume	Textbook + Figure 3 + Example problem 3 + Practice Problems	280, 281

The pressure for a gas in a cylinder is 1.00 atm at 300 K.
What will be the pressure if the temperature increases to 400 K?

الضغط لغاز في أسطوانة 1.00 atm عند 300 K.
كم سيصبح الضغط إذا زادت درجة الحرارة إلى 400 K؟



0.75 atm

2.67 atm

2.44 atm

1.30 atm

If the of pressure of a sample of hydrogen gas equals

إذا كان ضغط عينة من غاز الهيدروجين يُساوي **3.33 kPa** عند

3.33 kPa at **20.0°C**. What is the Celsius temperature

20.0°C. ما درجة الحرارة السيليزية التي يُصبح عندها ضغط الغاز

at which the gas pressure becomes **3.44 kPa**?

3.44 kPa؟

29.7 °C

35. 0 °C

27.5 °C

18.5 °C

21	CHM.5.2.01.003.13 List conditions of a gas at STP and RTP	Textbook + Example problem 5 + Practice Problems	285, 286, 287
22	CHM.5.2.01.004.28 Use the ideal gas law to calculate pressure, volume, temperature, mass of a gas, when three quantities are given	Textbook + Table 2 + Example problem 6 + Practice Problems	287, 288, 289

Avogadro's Principle

The particles that make up different gases can vary greatly in size. However, kinetic-molecular theory assumes that the particles in a gas sample are far enough apart that size has very little influence on the volume occupied by a gas. For example, 1000 relatively large krypton gas particles occupy the same volume as 1000 smaller helium gas particles at the same temperature and pressure. It was Avogadro who first proposed this idea in 1811. **Avogadro's principle** states that equal volumes of gases at the same temperature and pressure contain equal numbers of particles. **Figure 5** shows equal volumes of carbon dioxide, helium, and oxygen.



Figure 5 Gas tanks of equal volume that are at the same pressure and temperature contain equal numbers of gas particles, regardless of which gas they contain.

Infer Why doesn't Avogadro's principle apply to liquids and solids?

Volume and moles

Recall that one mole of a substance contains 6.02×10^{23} particles. The **molar volume** of a gas is the volume that 1 mol occupies at 0.00°C and 1.00 atm pressure. The conditions of 0.00°C and 1.00 atm are known as **standard temperature and pressure (STP)**. Avogadro showed experimentally that 1 mol of any gas occupies a volume of 22.4 L at STP.

Because the volume of 1 mol of a gas at STP is 22.4 L, you can use 22.4 L/mol as a conversion factor whenever a gas is at STP. For example, suppose you want to find the number of moles in a sample of gas that has a volume of 3.72 L at STP. Use the molar volume to convert from volume to moles.

$$3.72 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 0.166 \text{ mol}$$

EXAMPLE Problem 5

MOLAR VOLUME The main component of natural gas used for home heating and cooking is methane (CH_4). Calculate the volume that 2.00 kg of methane gas will occupy at STP.

1 ANALYZE THE PROBLEM

The number of moles can be calculated by dividing the mass of the sample, m , by its molar mass, M . The gas is at STP (0.00°C and 1.00 atm pressure), so you can use the molar volume to convert from the number of moles to the volume.

Known

$$m = 2.00 \text{ kg}$$

$$T = 0.00^\circ\text{C}$$

$$P = 1.00 \text{ atm}$$

Unknown

$$V = ? \text{ L}$$

2 SOLVE FOR THE UNKNOWN

Determine the molar mass for methane.

$$\begin{aligned} M &= 1 \text{ C atom} \left(\frac{12.01 \text{ amu}}{1 \text{ C atom}} \right) + 4 \text{ H atoms} \left(\frac{1.01 \text{ amu}}{1 \text{ H atom}} \right) \\ &= 12.01 \text{ amu} + 4.04 \text{ amu} = 16.05 \text{ amu} \\ &= 16.05 \text{ g/mol} \end{aligned}$$

Determine the molecular mass.

Express the molecular mass as g/mol to arrive at the molar mass.

Determine the number of moles of methane.

$$2.00 \text{ kg} \left(\frac{1000 \text{ g}}{1 \text{ kg}} \right) = 2.00 \times 10^3 \text{ g}$$

Convert the mass from kg to g.

$$\frac{m}{M} = \frac{2.00 \times 10^3 \text{ g}}{16.05 \text{ g/mol}} = 125 \text{ mol}$$

Divide mass by molar mass to determine the number of moles.

Use the molar volume to determine the volume of methane at STP.

$$V = 125 \text{ mol} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 2.80 \times 10^3 \text{ L}$$

Use the molar volume, 22.4 L/mol, to convert from moles to the volume.

3 EVALUATE THE ANSWER

The amount of methane present is much more than 1 mol, so you should expect a large volume, which is in agreement with the answer. The unit is liters, a volume unit, and there are three significant figures.

PRACTICE Problems

ADDITIONAL PRACTICE

- What size container do you need to hold 0.0459 mol of N_2 gas at STP?
- How much carbon dioxide gas, in grams, is in a 1.0-L balloon at STP?
- What volume in milliliters will 0.00922 g of H_2 gas occupy at STP?
- What volume will 0.416 g of krypton gas occupy at STP?
- Calculate the volume that 4.5 kg of ethylene gas (C_2H_4) will occupy at STP.
- CHALLENGE** A flexible plastic container contains 0.860 g of helium gas in a volume of 19.2 L. If 0.205 g of helium is removed at constant pressure and temperature, what will be the new volume?

21	CHM.5.2.01.003.13 List conditions of a gas at STP and RTP	Textbook + Example problem 5 + Practice Problems	285, 286, 287
22	CHM.5.2.01.004.28 Use the ideal gas law to calculate pressure, volume, temperature, mass of a gas, when three quantities are given	Textbook + Table 2 + Example problem 6 + Practice Problems	287, 288, 289

The Ideal Gas Law

Avogadro's principle and the laws of Boyle, Charles, and Gay-Lussac can be combined into a single mathematical statement that describes the relationships between pressure, volume, temperature, and number of moles of a gas. This formula works best for gases that obey the assumptions of the kinetic-molecular theory. Known as ideal gases, their particles occupy a negligible volume and are far enough apart that they exert minimal attractive or repulsive forces on one another.

From the combined gas law to the ideal gas law

The combined gas law relates the variables of pressure, volume, and temperature for a given amount of gas.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

For a specific sample of gas, this relationship of pressure, volume, and temperature is always the same. You could rewrite the relationship represented in the combined gas law as follows.

$$\frac{PV}{T} = \text{constant}$$

As **Figure 6** illustrates, increasing the amount of gas present in a sample will raise the pressure if temperature and volume are constant. Likewise, if pressure and temperature remain constant, the volume will increase as more particles of a gas are added. In fact, we know that both volume and pressure are directly proportional to the number of moles, n , so n can be incorporated into the combined gas law as follows.

$$\frac{PV}{nT} = \text{constant}$$

Experiments using known values of P , T , V , and n have determined the value of this constant. It is called the **ideal gas constant**, and it is represented by the symbol R . If pressure is in atmospheres, the value of R is 0.0821 L·atm/mol·K. Note that the units for R are simply the combined units for each of the four variables. **Table 2** shows the numerical values for R in different units of pressure.



Figure 6 The volume and temperature of this tire stay the same as air is added. However, the pressure in the tire increases as the amount of air present increases.

Table 2 Values of R

Value of R	Units of R
0.0821	$\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$
8.314	$\frac{\text{L} \cdot \text{kPa}}{\text{mol} \cdot \text{K}}$
62.4	$\frac{\text{L} \cdot \text{mm Hg}}{\text{mol} \cdot \text{K}}$

Substituting R for the constant in the equation on the previous page and rearranging the variables gives the most familiar form of the ideal gas law. The **ideal gas law** describes the physical behavior of an ideal gas in terms of the pressure, volume, temperature, and number of moles of gas present.

The Ideal Gas Law

$$PV = nRT$$

P represents pressure. V represents volume.
 n represents number of moles. R is the ideal gas constant.
 T represents temperature.

For a given amount of gas held at constant temperature, the product of pressure and volume is a constant.

If you know any three of the four variables, you can rearrange the equation to solve for the unknown.

EXAMPLE Problem 6

THE IDEAL GAS LAW Calculate the number of moles of ammonia gas (NH_3) contained in a 3.0-L vessel at $3.00 \times 10^2 \text{ K}$ with a pressure of 1.50 atm.

1 ANALYZE THE PROBLEM

You are given the volume, temperature, and pressure of a gas sample. Use the ideal gas law, and select the value of R that contains the pressure units given in the problem. Because the pressure and temperature are close to STP, but the volume is much smaller than 22.4 L, it would make sense if the calculated answer were much smaller than 1 mol.

Known
 $V = 3.0 \text{ L}$
 $T = 3.00 \times 10^2 \text{ K}$
 $P = 1.50 \text{ atm}$
 $R = 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$

Unknown
 $n = ? \text{ mol}$

2 SOLVE FOR THE UNKNOWN

Use the ideal gas law. Solve for n , and substitute the known values.

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

$$n = \frac{(1.50 \text{ atm})(3.0 \text{ L})}{(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(3.00 \times 10^2 \text{ K})}$$

$$n = \frac{(1.50 \text{ atm})(3.0 \text{ L})}{(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}})(3.00 \times 10^2 \text{ K})} = 0.18 \text{ mol}$$

State the ideal gas law.

Solve for n .

Substitute $V = 3.0 \text{ L}$, $T = 3.00 \times 10^2 \text{ K}$, $P = 1.50 \text{ atm}$, and $R = 0.0821 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K}$.

Multiply and divide numbers and units.

3 EVALUATE THE ANSWER

The answer agrees with the prediction that the number of moles present will be significantly less than 1 mol. The unit of the answer is the mole, and there are two significant figures.

21	CHM.5.2.01.003.13 List conditions of a gas at STP and RTP	Textbook + Example problem 5 + Practice Problems	285, 286, 287
22	CHM.5.2.01.004.28 Use the ideal gas law to calculate pressure, volume, temperature, mass of a gas, when three quantities are given	Textbook + Table 2 + Example problem 6 + Practice Problems	287, 288, 289

PRACTICE Problems



ADDITIONAL PRACTICE

26. Determine the Celsius temperature of 2.49 mol of a gas contained in a 1.00-L vessel at a pressure of 143 kPa.
27. Calculate the volume of a 0.323-mol sample of a gas at 265 K and 0.900 atm.
28. What is the pressure, in atmospheres, of a 0.108-mol sample of helium gas at a temperature of 20.0°C if its volume is 0.505 L?
29. If the pressure exerted by a gas at 25°C in a volume of 0.044 L is 3.81 atm, how many moles of gas are present?
30. **CHALLENGE** An ideal gas has a volume of 3.0 L. If the number of moles of gas and the temperature are doubled, while the pressure remains constant, what is the new volume?

The Ideal Gas Law—Molar Mass and Density

The ideal gas law can be used to solve for the value of any one of the four variables P , V , T , or n if the values of the other three are known. However, you can also rearrange the $PV = nRT$ equation to calculate the molar mass and density of a gas sample.

Molar mass and the ideal gas law

To find the molar mass of a gas sample, the mass, temperature, pressure, and volume of the gas must be known. Recall that the number of moles of a gas (n) is equal to the mass (m) divided by the molar mass (M). Therefore, the n in the equation can be replaced by m/M .

$$PV = nRT \quad \text{substitute } n = \frac{m}{M} \quad PV = \frac{mRT}{M}$$

You can rearrange the new equation to solve for the molar mass.

$$M = \frac{mRT}{PV}$$

Density and the ideal gas law

Recall that the density (D) of a substance is defined as mass (m) per unit volume (V). After rearranging the ideal gas equation to solve for molar mass, you can substitute D for m/V .

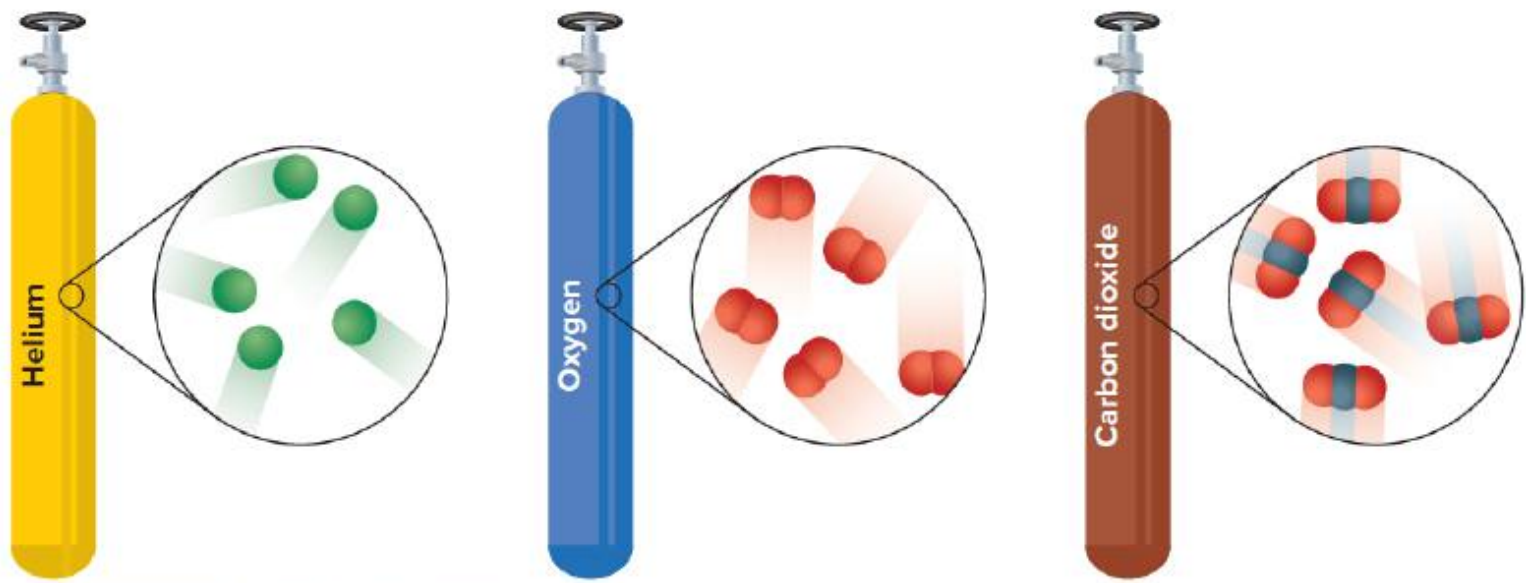
$$M = \frac{mRT}{PV} \quad \text{substitute } \frac{m}{V} = D \quad M = \frac{DRT}{P}$$

You can rearrange the new equation to solve for density.

$$D = \frac{MP}{RT}$$

21	CHM.5.2.01.003.13 List conditions of a gas at STP and RTP	Textbook + Example problem 5 + Practice Problems	285, 286, 287
22	CHM.5.2.01.004.28 Use the ideal gas law to calculate pressure, volume, temperature, mass of a gas, when three quantities are given	Textbook + Table 2 + Example problem 6 + Practice Problems	287, 288, 289

Avogadro's principle states that **equal volumes of gases** contain **equal numbers of particles**.



1 mol of any gas occupies a volume of 22.4 L at STP.

According to Avogadro's principle, 1 mol of any gas at STP occupies a volume of _____.

حسب مبدأ أفوجادرو، 1 مول (mol) من أي غاز عند الضغط ودرجة الحرارة القياسيين STP يشغل حجماً قدره _____.

22.4 L	☒
3.72 L	☐
1.00 L	☐
6.02 L	☐

What are the standard temperature and pressure (STP) for a gas?

ما هي درجة الحرارة والضغط القياسيان (STP) لغاز؟

Temperature = 0 °C and pressure = 1 kPa	درجة الحرارة = 0 °C والضغط = 1 kPa	☐
Temperature = 273 °C and pressure = 1 atm	درجة الحرارة = 273 °C والضغط = 1 atm	☐
Temperature = 1 °C and pressure = 0 atm	درجة الحرارة = 1 °C والضغط = 0 atm	☐
Temperature = 0 °C and pressure = 1 atm	درجة الحرارة = 0 °C والضغط = 1 atm	☒

21	CHM.5.2.01.003.13 List conditions of a gas at STP and RTP	Textbook + Example problem 5 + Practice Problems	285, 286, 287
22	CHM.5.2.01.004.28 Use the ideal gas law to calculate pressure, volume, temperature, mass of a gas, when three quantities are given	Textbook + Table 2 + Example problem 6 + Practice Problems	287, 288, 289

How much carbon dioxide gas, in grams, is in a **2.75 L** balloon at **STP**?

Molar mass of carbon dioxide – **44.0 g/mol**

$$R = 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

7.25 g

6.35 g

5.40 g

3.65 g

How many neon atoms are there in **1.86 L** sample at standard temperature and pressure (**STP**)?

ما عدد ذرات النيون في **1.86 L** منه عند الضغط ودرجة الحرارة القياسيين (**STP**)؟

$$6.02 \times 10^{23}$$

ثابت (عدد) أفوجادرو

Avogadro's constant(number)

$$R = 0.0821 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K}$$

$$5.00 \times 10^{22}$$

$$2.70 \times 10^{23}$$

$$3.50 \times 10^{24}$$

$$1.40 \times 10^{25}$$

21	CHM.5.2.01.003.13 List conditions of a gas at STP and RTP	Textbook + Example problem 5 + Practice Problems	285, 286, 287
22	CHM.5.2.01.004.28 Use the ideal gas law to calculate pressure, volume, temperature, mass of a gas, when three quantities are given	Textbook + Table 2 + Example problem 6 + Practice Problems	287, 288, 289

A **4.25 L** flask is filled with butane gas (C_4H_{10}) at a pressure of **1.5 atm** and a temperature of **-20 °C**. What is the mass of butane in the flask?

دورق حجمه **4.25 L** مملوء بغاز البيوتان (C_4H_{10}) عند ضغط **1.5 atm** ودرجة حرارة **-20 °C** فما كتلة البيوتان في الدورق؟

R	الكتلة المولية للبيوتان C_4H_{10} Molar mass of butane C_4H_{10}
0.0821 L.atm/mol.K	58.1 g/mol

17.8 g

8.9 g

26.7 g

13.5 g

What is the volume of a 0.323 mol sample of a gas at 12°C and 0.900 atm?

ما حجم عينة من غاز عدد مولاتها 0.323 mol عند 12°C و 0.900 atm؟

$$R = 0.0821 \text{ L.atm/mol.K}$$

7.26 L

8.40 L

3.53 L

6.52 L

21	CHM.5.2.01.003.13 List conditions of a gas at STP and RTP	Textbook + Example problem 5 + Practice Problems	285, 286, 287
22	CHM.5.2.01.004.28 Use the ideal gas law to calculate pressure, volume, temperature, mass of a gas, when three quantities are given	Textbook + Table 2 + Example problem 6 + Practice Problems	287, 288, 289

A **4.50 L** flask is filled with butane gas (C_4H_{10}) at a pressure of **1.20 atm** and a temperature of **(-10.0°C)**.

What is the density of the butane?

دورق حجمه **4.50 L** مملوء بـغاز البيوتان (C_4H_{10}) عند ضغط **1.20 atm** ودرجة حرارة **(-10.0 °C)** فما كثافة البيوتان؟

$$R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$$

0.82 g/L

1.35 g/L

3.22 g/L

3.40 g/L

If the pressure exerted by a gas at **35°C** in a volume of **0.044 L** is **7.62 atm**

How many moles of gas are present?

What is the volume(L) of **0.216 mol** sample of helium gas at a temperature of **30.0 °C** and a pressure of **7.16 atm**?

ما حجم (L) عينة من غاز الهيليوم مقدارها **0.216 mol** عند درجة حرارة **30.0 °C** وضغط **7.16 atm**؟

$$R = 0.0821 L \cdot atm/mol \cdot K$$

0.750 L

1.40 L

0.375 L

2.85 L

إذا كان الضغط المبذول من غاز عند درجة حرارة **35 °C** يساوي **7.62 atm** في حجم مقداره **0.044 L**

فكم عدد مولات الغاز الموجودة؟

$$R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$$

4.25 mol

0.068 mol

2.45 mol

0.013 mol

21	CHM.5.2.01.003.13 List conditions of a gas at STP and RTP	Textbook + Example problem 5 + Practice Problems	285, 286, 287
22	CHM.5.2.01.004.28 Use the ideal gas law to calculate pressure, volume, temperature, mass of a gas, when three quantities are given	Textbook + Table 2 + Example problem 6 + Practice Problems	287, 288, 289

What is the molar mass of unknown gas at STP, if its density was 1.70 g/L?
 ما الكتلة المولية لغاز مجهول عند درجة الحرارة والضغط القياسيين STP، إذا كانت كثافة الغاز 1.70 g/L ؟

$R = 0.0821 \text{ L.atm/mol.K}$

87.3 g/mol

25.6 g/mol

38.1 g/mol

5.11 g/mol

What is the volume of a 0.323 mol sample of a gas at 12°C and 0.900 atm?
 ما حجم عينة من غاز عدد مولاتها 0.323 mol عند 12°C و 0.900 atm ؟

$R = 0.0821 \text{ L.atm/mol.K}$

7.26 L

8.40 L

3.53 L

6.52 L

21	CHM.5.2.01.003.13 List conditions of a gas at STP and RTP	Textbook + Example problem 5 + Practice Problems	285, 286, 287
22	CHM.5.2.01.004.28 Use the ideal gas law to calculate pressure, volume, temperature, mass of a gas, when three quantities are given	Textbook + Table 2 + Example problem 6 + Practice Problems	287, 288, 289
23	CHM.5.2.01.003.15 Predict the conditions under which a real gas might deviate from ideal behavior while explaining its effect	Textbook + Figure 8	291

A **4.25 L** flask is filled with butane gas (C_4H_{10}) at a pressure of **1.5atm** and a temperature of **-20 °C**. What is the mass of butane in the flask?

دورق حجمه **4.25 L** مملوء بغاز البيوتان (C_4H_{10}) عند ضغط **1.5atm** ودرجة حرارة **-20 °C** فما كتلة البيوتان في الدورق؟

R	الكتلة المولية للبيوتان C_4H_{10} Molar mass of butane C_4H_{10}	17.8 g	8.9 g	26.7 g	13.5 g
0.0821 L.atm/mol.K	58.1 g/mol				

	Ideal gases(follow kinetic molecular theory)	Real gases(deviate from ideal)
Particles behaviour	Particles take up no space	Particles have volume
	No attraction or repulsion	There are intermolecular interactions
	Collisions between particles are elastic	Collisions are not elastic
Conditions	High temperature	Low temperature
	Low pressure	High pressure
	Nonpolar	Polar
	Small Particles (Low Molar Mass)	Large Particles (high Molar Mass)

When real gases deviate most from ideal gas behavior?

- A – At high pressures and low temperatures
- B – At low pressures and high temperatures
- C – At high pressures and high temperatures
- D - At low pressures and low temperatures

Which of the following is a characteristic of the ideal gas?

أي مما يلي من خصائص الغاز المثالي؟

Its particles move at variable velocities and on winding (zigzag) lines

تتحرك جسيماته بسرعات متغيرة وبمسارات متعرجة



Its particles take up space and measured in volume units (L)

تشغل جسيماته حيزاً من الفراغ ويعبر عنها بوحدة الحجم (L)



Its particles collide with each other or with the wall surface in perfectly elastic way

تتصادم جسيماته ببعضها أو مع جدران الوعاء تصادمات مرنة بشكل مثالي



Its particles experience intermolecular attractive forces

تتعرض جسيماته لقوى تجاذب بينها



21	CHM.5.2.01.003.13 List conditions of a gas at STP and RTP	Textbook + Example problem 5 + Practice Problems	285, 286, 287
22	CHM.5.2.01.004.28 Use the ideal gas law to calculate pressure, volume, temperature, mass of a gas, when three quantities are given	Textbook + Table 2 + Example problem 6 + Practice Problems	287, 288, 289
23	CHM.5.2.01.003.15 Predict the conditions under which a real gas might deviate from ideal behavior while explaining its effect	Textbook + Figure 8	291

. When does a real gas behave like an ideal gas?

- When the particles are far apart and attractive forces decreases
- When the particles are closer together and attractive forces increases
- At high pressure and low temperature
- When the gas is liquefied if enough pressure is applied

When does a real gas behave like an ideal gas?

متى يسلك الغاز الحقيقي مثل الغاز المثالي؟

At high pressure and low temperature

عند الضغط العالي ودرجة الحرارة المنخفضة

When high pressure is applied and the gas changes to the liquid phase

عندما تتحول حالة الغاز إلى سائل، عند التأثير عليه بضغط مرتفع

When the particles are close to each other, and attractive forces are high

عندما تقرب الجزيئات عن بعضها البعض وتزداد قوى التجاذب

When the particles are far apart, and the attractive forces are low

عندما تبعد الجزيئات عن بعضها البعض وتقل قوى التجاذب

Real Versus Ideal Gases

What does the term **ideal gas** mean? Ideal gases follow the assumptions of the kinetic-molecular theory. According to this theory, an ideal gas is one whose particles take up no space. Ideal gases experience no intermolecular attractive (electrical) forces, nor are they attracted or repelled by the walls of their containers. The particles of an ideal gas are in constant, random motion, moving in straight lines until they collide with each other or with the walls of the container. Additionally, these collisions are perfectly elastic, which means that the kinetic energy of the system does not change. An ideal gas follows the gas laws under all conditions of temperature and pressure.

In reality, no gas is truly ideal. All gas particles have some volume, however small, and are subject to intermolecular interactions. Also, the collisions that particles make with each other and with the container are not perfectly elastic. Despite that, most gases behave like ideal gases at a wide range of temperatures and pressures. Calculations made using the ideal gas law often closely approximate experimental measurements.



Get It?

Explain how the interactions of real gases are determined by attractive (electrical) forces between particles.

Extreme pressure and temperature

When is the ideal gas law not likely to work for a real gas? Real gases deviate most from ideal gas behavior at high pressures and low temperatures. The nitrogen gas shown in **Figure 8 (left)** behaves as a real gas. Lowering the temperature of nitrogen gas results in less kinetic energy of the gas particles, which means their intermolecular attractive (electric) forces are strong enough to affect their behavior. When the temperature is low enough, this real gas condenses to form a liquid.

The propane gas in the tanks shown in **Figure 8 (right)** also behaves as a real gas. Increasing the pressure on a gas forces the gas particles closer together until the volume occupied by the gas particles themselves is no longer negligible. Real gases such as propane will liquefy if enough pressure is applied.



Nitrogen gas turns to liquid at -196°C . At this temperature, scientists can preserve biological specimens, such as body tissues, for future research or medical procedures.



About 270 times more propane can be stored as a liquid than as a gas in the same amount of space. Your family might use small tanks of liquid propane as fuel for your barbecue grill or larger tanks for heating and cooking.

	Ideal gases(follow kinetic molecular theory)	Real gases(deviate from ideal)
Particles behaviour	Particles take up no space No attraction or repulsion Collisions between particles are elastic	Particles have volume There are intermolecular interactions Collisions are not elastic
Conditions	High temperature	Low temperature
	Low pressure	High pressure
	Nonpolar	Polar
	Small particles(low molar mass)	Large particles(high molar mass)

When does a real gas behave like an ideal gas?

متى يسلك الغاز الحقيقي مثل الغاز المثالي؟

At high pressure and low temperature

عند الضغط العالي ودرجة الحرارة المنخفضة

When high pressure is applied and the gas changes to the liquid phase

عندما تتحول حالة الغاز إلى سائل، عند التأثير عليه بضغط مرتفع

When the particles are close to each other, and attractive forces are high

عندما تقترب الجزيئات عن بعضها البعض وتزداد قوى التجاذب

When the particles are far apart, and the attractive forces are low

عندما تبتعد الجزيئات عن بعضها البعض وتقل قوى التجاذب

أي مما يلي من خصائص الغاز المثالي؟ Which of the following is a characteristic of the ideal gas?

- ☐ تتحرك جسيماته بسرعات متغيرة وبمسارات متعرجة
Its particles move at variable velocities and on winding (zigzag) lines
- ☐ تشغل جسيماته حيزاً من الفراغ ويعبر عنها بوحدة الحجم (L)
Its particles take up space and measured in volume units (L)
- ☐ تتصادم جسيماته ببعضها أو مع جدران الوعاء تصادمات مرنة بشكل مثالي
Its particles collide with each other or with the wall surface in perfectly elastic way
- ☐ تتعرض جسيماته لقوى تجاذب بينها
Its particles experience intermolecular attractive forces

فيما يتعلق بكل من الغازات المثالية والغازات الحقيقية. Regarding both ideal and real gases.
أي مما يأتي **صحيح**؟ Which of the following is **correct**?

- ☐ تسلك الغازات القطبية مثل بخار الماء سلوك الغاز المثالي
Polar gases like water vapor behave as ideal gases
- ☐ تسلك الغازات ذات الجزيئات الأكبر حجماً مثل البيوتان (C₄H₁₀) سلوك الغاز المثالي
Gases of larger particles like butane (C₄H₁₀) behave as ideal gases
- ☐ تنحرف الغازات الحقيقية كثيراً عن سلوك الغاز المثالي عند الضغوط العالية ودرجات الحرارة المنخفضة
Real gases deviate most from ideal gas behavior at high pressures and low temperatures
- ☐ الغازات المثالية لديها قوى تجاذب بين جسيماتها وتنجذب وتتنافر مع جدران الأوعية الموجودة فيها
Ideal gases experience intermolecular attractive forces, and they attracted, repelled by the walls of their containers

24	CHM.5.2.03.002.12 Calculate percent by mass of a solution	Textbook +Table 3+ Example problem 1 + Practice Problems	308, 309, 310
25	CHM.5.2.03.002.05 Calculate molarity when the moles or the mass of solute and volume of solution are given and vice versa	Textbook + Example problem 2 + Practice Problems	311, 312
26	CHM.5.2.03.002.16 Perform calculation involving diluting a concentrated solution	Textbook + Figure 9 + Example problem 3+ Practice Problems	314, 315, 316
27	CHM.5.2.03.002.07 Calculate molality when the moles or the mass of solute and mass of solvent are given and vice versa	Textbook + Example problem 4 + Practice Problems	316, 317

Expressing Concentration

The **concentration** of a solution is a measure of how much solute is dissolved in a specific amount of solvent or solution. Visual observations can be used to describe solutions qualitatively by using the words concentrated or dilute. Notice the containers of tea in **Figure 6**. One of the tea solutions is more concentrated than the other. In general, a concentrated solution contains a large amount of solute. The darker tea has more tea particles than the lighter tea. Conversely, a dilute solution contains a small amount of solute. The lighter tea in **Figure 6** is dilute and contains fewer tea particles than the darker tea.

Although qualitative descriptions of concentration can be useful, solutions are more often described quantitatively. Some commonly used quantitative descriptions are percent by mass, percent by volume, molarity, and molality. These descriptions express concentration as a ratio of measured amounts of solute and solvent or solution.



Figure 6 The strength of the tea corresponds to its concentration. The lighter tea is less concentrated than the darker tea.

Table 3 Concentration Ratios

Concentration Description	Ratio
Percent by mass	$\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$
Percent by volume	$\frac{\text{volume of solute}}{\text{volume of solution}} \times 100$
Molarity	$\frac{\text{moles of solute}}{\text{liter of solution}}$
Molality	$\frac{\text{moles of solute}}{\text{kilogram of solvent}}$
Mole fraction	$\frac{\text{mole of solute}}{\text{moles of solute} + \text{moles of solvent}}$

Table 3 shows ratios for common quantitative descriptions. Which quantitative description should be used? The description used depends on the type of solution analyzed and the reason for describing it. For example, a chemist working with a reaction in an aqueous solution most likely refers to the molarity of the solution because he or she needs to know the number of particles involved in the reaction.

24	CHM.5.2.03.002.12 Calculate percent by mass of a solution	Textbook +Table 3+ Example problem 1 + Practice Problems	308, 309, 310
25	CHM.5.2.03.002.05 Calculate molarity when the moles or the mass of solute and volume of solution are given and vice versa	Textbook + Example problem 2 + Practice Problems	311, 312
26	CHM.5.2.03.002.16 Perform calculation involving diluting a concentrated solution	Textbook + Figure 9 + Example problem 3+ Practice Problems	314, 315, 316
27	CHM.5.2.03.002.07 Calculate molality when the moles or the mass of solute and mass of solvent are given and vice versa	Textbook + Example problem 4 + Practice Problems	316, 317

Percent by mass

The percent by mass is the ratio of the solute's mass to the solution's mass expressed as a percent. The mass of the solution is the sum of the masses of the solute and solvent.

Percent by Mass

$$\text{percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

Percent by mass equals the mass of the solute divided by the mass of the whole solution, multiplied by 100.

EXAMPLE Problem 1

CALCULATE PERCENT BY MASS In order to maintain a sodium chloride (NaCl) concentration similar to ocean water, an aquarium must contain 3.6 g NaCl per 100.0 g of water. What is the percent by mass of NaCl in the solution?

1 ANALYZE THE PROBLEM

You are given the amount of sodium chloride dissolved in 100.0 g of water. The percent by mass of a solute is the ratio of the solute's mass to the solution's mass, which is the sum of the masses of the solute and the solvent.

Known

mass of solute = 3.6 g NaCl

mass of solvent = 100.0 g H₂O

Unknown

percent by mass = ?

EXAMPLE Problem 1 (continued)

2 SOLVE FOR THE UNKNOWN

Find the mass of the solution.

mass of solution = grams of solute + grams of solvent

mass of solution = 3.6 g + 100.0 g = 103.6 g

Substitute mass of solute = 3.6 g, and mass of solvent = 100.0 g.

Calculate the percent by mass.

$$\text{percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100$$

State the equation for percent by mass.

$$\text{percent by mass} = \frac{3.6 \text{ g}}{103.6 \text{ g}} \times 100 = 3.5\%$$

Substitute mass of solute = 3.6 g, and mass of solution = 103.6 g.

3 EVALUATE THE ANSWER

Because only a small mass of sodium chloride is dissolved per 100.0 g of water, the percent by mass should be a small value, which it is. The mass of sodium chloride was given with two significant figures; therefore, the answer is also expressed with two significant figures.

PRACTICE Problems

ADDITIONAL PRACTICE

- What is the percent by mass of NaHCO₃ in a solution containing 20.0 g of NaHCO₃ dissolved in 600.0 mL of H₂O?
- You have 1500.0 g of a bleach solution. The percent by mass of the solute sodium hypochlorite (NaOCl) is 3.62%. How many grams of NaOCl are in the solution?
- In Question 10, how many grams of solvent are in the solution?
- CHALLENGE** The percent by mass of calcium chloride in a solution is found to be 2.65%. If 50.0 g of calcium chloride is used, what is the mass of the solution?

24	CHM.5.2.03.002.12 Calculate percent by mass of a solution	Textbook +Table 3+ Example problem 1 + Practice Problems	308, 309, 310
25	CHM.5.2.03.002.05 Calculate molarity when the moles or the mass of solute and volume of solution are given and vice versa	Textbook + Example problem 2 + Practice Problems	311, 312
26	CHM.5.2.03.002.16 Perform calculation involving diluting a concentrated solution	Textbook + Figure 9 + Example problem 3+ Practice Problems	314, 315, 316
27	CHM.5.2.03.002.07 Calculate molality when the moles or the mass of solute and mass of solvent are given and vice versa	Textbook + Example problem 4 + Practice Problems	316, 317

Percent by volume

Percent by volume usually describes solutions in which both solute and solvent are liquids. The percent by volume is the ratio of the volume of the solute to the volume of the solution, expressed as a percent. The volume of the solution is the sum of the volumes of the solute and the solvent. Calculations of percent by volume are similar to those involving percent by mass.

Percent by Volume

$$\text{percent by volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100$$

Percent by volume equals the volume of solute divided by the volume of the solution, multiplied by 100.

Biodiesel, shown in **Figure 7**, is an alternative fuel that is produced from renewable resources. Biodiesel can be used in diesel engines with little or no modifications. Biodiesel is simple to use, biodegradable, nontoxic, and it does not contain some of the pollutants found in regular gasoline.



Figure 7 B20 is 20% by volume biodiesel and 80% by volume petroleum diesel. Biodiesel is a clean-burning alternative fuel that can be produced from renewable resources, such as vegetable oil.

Molarity

Percent by volume and percent by mass are only two of the commonly used ways to quantitatively describe the concentrations of liquid solutions. One of the most common units of solution concentration is molarity. **Molarity** (M) is the number of moles of solute dissolved per liter of solution. Molarity is also known as molar concentration, and the unit M is read as molar. A liter of solution containing 1 mol of solute is a 1M solution, which is read as a one-molar solution. A liter of solution containing 0.1 mol of solute is a 0.1M solution. To calculate a solution's molarity, you must know the volume of the solution in liters and the amount of dissolved solute in moles.

Molarity

$$\text{molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

The molarity of a solution equals the moles of solute divided by the liters of solution.

Molarity is the measure of concentration that you will likely use most in your studies of chemistry. The molarity of a solution tells you the amount of solute present in a solution, and this is the information that is important to know when you are planning and analyzing reactions in solution. Knowing the amounts of different solutes in solution allows you to predict how they will react based on a balanced chemical equation.

When performing calculations involving concentrations, you can think of molarity as a conversion factor that you can use to convert between volume of solution and amount of solute in moles. (Note that sometimes you will see molarity expressed using units of mol/L instead of M .)

PRACTICE Problems

- What is the percent by volume of ethanol in a solution that contains 35 mL of ethanol dissolved in 155 mL of water?
- What is the percent by volume of isopropyl alcohol in a solution that contains 24 mL of isopropyl alcohol in 1.1 L of water?
- CHALLENGE** If 18 mL of methanol is used to make an aqueous solution that is 15% methanol by volume, how many milliliters of solution is produced?



ADDITIONAL PRACTICE

24	CHM.5.2.03.002.12 Calculate percent by mass of a solution	Textbook +Table 3+ Example problem 1 + Practice Problems	308, 309, 310
25	CHM.5.2.03.002.05 Calculate molarity when the moles or the mass of solute and volume of solution are given and vice versa	Textbook + Example problem 2 + Practice Problems	311, 312
26	CHM.5.2.03.002.16 Perform calculation involving diluting a concentrated solution	Textbook + Figure 9 + Example problem 3+ Practice Problems	314, 315, 316
27	CHM.5.2.03.002.07 Calculate molality when the moles or the mass of solute and mass of solvent are given and vice versa	Textbook + Example problem 4 + Practice Problems	316, 317

EXAMPLE Problem 2

CALCULATING MOLARITY A 100.5-mL intravenous (IV) solution contains 5.10 g of glucose ($C_6H_{12}O_6$). What is the molarity of this solution? The molar mass of glucose is 180.16 g/mol.

1 ANALYZE THE PROBLEM

You are given the mass of glucose dissolved in a volume of water. The molarity of the solution is the ratio of moles of solute per liter of solution.

Known

mass of solute = 5.10 g $C_6H_{12}O_6$
molar mass of $C_6H_{12}O_6$ = 180.16 g/mol
volume of solution = 100.5 mL

Unknown

solution concentration = ? M

2 SOLVE FOR THE UNKNOWN

Calculate the number of moles of $C_6H_{12}O_6$.

$$(5.10 \text{ g } C_6H_{12}O_6) \left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.16 \text{ g } C_6H_{12}O_6} \right)$$

$$= 0.0283 \text{ mol } C_6H_{12}O_6$$

Convert the volume of H_2O to liters.

$$(100.5 \text{ mL solution}) \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) = 0.1005 \text{ L solution}$$

Solve for the molarity.

$$M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$M = \left(\frac{0.0283 \text{ mol } C_6H_{12}O_6}{0.1005 \text{ L solution}} \right)$$

$$M = \left(\frac{0.0283 \text{ mol } C_6H_{12}O_6}{1 \text{ L solution}} \right) = 0.282M$$

Multiply grams of $C_6H_{12}O_6$ by the molar mass of $C_6H_{12}O_6$.

Use the conversion factor 1 L/1000 mL.

State the molarity equation.

Substitute moles of $C_6H_{12}O_6$ = 0.0283 and volume of solution = liters of solution = 0.1005 L.

Divide numbers and units.

3 EVALUATE THE ANSWER

The molarity value will be small because only a small mass of glucose was dissolved in the solution. The mass of glucose used in the problem has three significant figures; therefore, the value of the molarity also has three significant figures.

PRACTICE Problems



ADDITIONAL PRACTICE

- What is the molarity of an aqueous solution containing 40.0 g of glucose ($C_6H_{12}O_6$) in 1.5 L of solution?
- Calculate the molarity of 1.60 L of a solution containing 1.55 g of dissolved KBr.
- What is the molarity of a bleach solution containing 9.5 g of NaOCl per liter of bleach?
- CHALLENGE** How much calcium hydroxide ($Ca(OH)_2$), in grams, is needed to produce 1.5 L of a 0.25M solution?

Diluting molar solutions In the laboratory, you might use concentrated solutions of standard molarities, called stock solutions. For example, concentrated hydrochloric acid (HCl) is 12M. Recall that a concentrated solution has a large amount of solute. You can prepare a less-concentrated solution by diluting the stock solution with additional solvent. When you add solvent, you increase the number of solvent particles among which the solute particles move, as shown in **Figure 9**, thereby decreasing the solution's concentration.

How do you determine the volume of stock solution you must dilute? You can rearrange the expression of molarity to solve for moles of solute.

$$\text{molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$\text{moles of solute} = \text{molarity} \times \text{liters of solution}$$

Because the total number of moles of solute does not change during dilution,

$$\text{moles of solute in the stock solution} = \text{moles of solute after dilution.}$$

Substituting moles of solute with molarity times liters of solution, the relationship can be expressed in the dilution equation on the following page.

24	CHM.5.2.03.002.12 Calculate percent by mass of a solution	Textbook +Table 3+ Example problem 1 + Practice Problems	308, 309, 310
25	CHM.5.2.03.002.05 Calculate molarity when the moles or the mass of solute and volume of solution are given and vice versa	Textbook + Example problem 2 + Practice Problems	311, 312
26	CHM.5.2.03.002.16 Perform calculation involving diluting a concentrated solution	Textbook + Figure 9 + Example problem 3+ Practice Problems	314, 315, 316
27	CHM.5.2.03.002.07 Calculate molality when the moles or the mass of solute and mass of solvent are given and vice versa	Textbook + Example problem 4 + Practice Problems	316, 317

Dilution Equation

$M_1V_1 = M_2V_2$

M represents molarity.

V represents volume.

For a given amount of solute, the product of the molarity and volume of the stock solution equals the product of the molarity and the volume of the dilute solution.

In the dilution equation, M_1 and V_1 represent the molarity and volume of the stock solution, and M_2 and V_2 represent the molarity and volume of the dilute solution. Before dilution, a concentrated solution contains a fairly high ratio of solute particles to solvent particles. After adding more solvent, however, the ratio of solute particles to solvent particles has decreased. However, because the total number of solute particles would not change during dilution, the total number of moles of solute would remain the same.

EXAMPLE Problem 3

DILUTING STOCK SOLUTIONS

If you know the concentration and volume of the solution you want to prepare, you can calculate the volume of stock solution you will need. What volume, in milliliters, of 2.00M calcium chloride (CaCl₂) stock solution would you use to make 0.50 L of 0.300M calcium chloride solution?

1 ANALYZE THE PROBLEM

You are given the molarity of a stock solution of CaCl₂ and the volume and molarity of a dilute solution of CaCl₂. Use the relationship between molarities and volumes to find the volume, in liters, of the stock solution required. Then, convert the volume to milliliters.

Known

$M_1 = 2.00M$ CaCl₂

$M_2 = 0.300M$

$V_2 = 0.50$ L

Unknown

$V_1 = ?$ mL 2.00M CaCl₂

2 SOLVE FOR THE UNKNOWN

Solve the molarity-volume relationship for the volume of the stock solution V_1 .

$M_1V_1 = M_2V_2$

State the dilution equation.

$V_1 = V_2 \left(\frac{M_2}{M_1} \right)$

Solve for V_1 .

$V_1 = (0.50 \text{ L}) \left(\frac{0.300M}{2.00M} \right)$

Substitute $M_1 = 2.00M$, $M_2 = 0.300M$, and $V_2 = 0.50$ L.

$V_1 = (0.50 \text{ L}) \left(\frac{0.300M}{2.00M} \right) = 0.075 \text{ L}$

Multiply and divide numbers and units.

$V_1 = (0.075 \text{ L}) \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right) = 75 \text{ mL}$

Convert to milliliters using the conversion factor 1000 mL/1 L.

3 EVALUATE THE ANSWER

The volume V_1 was calculated, and then its value was converted to milliliters. This volume should be less than the final volume of the dilute solution, and it is. Of the given information, V_2 had the fewest number of significant figures, with two. Thus, the volume V_1 should also have two significant figures, and it does.

PRACTICE Problems

24. What volume of a 3.00M KI stock solution would you use to make 0.300 L of a 1.25M KI solution?

25. How many milliliters of a 5.0M H₂SO₄ stock solution would you need to prepare 100.0 mL of 0.25M H₂SO₄?

26. CHALLENGE If 0.50 L of 5.00M stock solution of HCl is diluted to make 2.0 L of solution, how much HCl, in grams, is in the solution?

Molality

The volume of a solution changes with temperature as it expands or contracts. This change in volume alters the molarity of the solution. Masses, however, do not change with temperature. It is sometimes more useful to describe solutions in terms of how many moles of solute are dissolved in a specific mass of solvent. Such a description is called **molality**—the ratio of the number of moles of solute dissolved in 1 kg of solvent. The unit m is read as molal. A solution containing 1 mol of solute per kilogram of solvent is a one-molal solution.

Molality

$$\text{molality } (m) = \frac{\text{moles of solute}}{\text{kg of solvent}}$$

The molality of a solution equals the moles of solute divided by kilograms of solvent.

EXAMPLE Problem 4

CALCULATING MOLALITY

In the lab, a student adds 4.5 g of sodium chloride (NaCl) to 100.0 g of water. Calculate the molality of the solution.

1 ANALYZE THE PROBLEM

You are given the mass of solute and solvent. Determine the number of moles of solute. Then, you can calculate the molality.

Known

mass of water (H₂O) = 100.0 g

mass of sodium chloride (NaCl) = 4.5 g

Unknown

$m = ?$ mol/kg

2 SOLVE FOR THE UNKNOWN

$4.5 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 0.077 \text{ mol NaCl}$

Calculate the number of moles of solute.

$100.0 \text{ g H}_2\text{O} \times \frac{1 \text{ kg H}_2\text{O}}{1000 \text{ g H}_2\text{O}} = 0.1000 \text{ kg H}_2\text{O}$

Convert the mass of H₂O from grams to kilograms using the factor 1 kg/1000 g.

Substitute the known values into the expression for molality, and solve.

$m = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$

Write the equation for molality.

$m = \frac{0.077 \text{ mol NaCl}}{0.1000 \text{ kg H}_2\text{O}} = 0.77 \text{ mol/kg}$

Substitute moles of solute = 0.077 mol NaCl, kilograms of solvent = 0.1000 kg H₂O.

3 EVALUATE THE ANSWER

Because there was less than one-tenth of a mole of solute present in one-tenth of a kilogram of water, the molality should be less than one, and it is. The mass of sodium chloride was given with two significant figures; therefore, the molality is also expressed with two significant figures.

almanahj.com/ae
2026 2025
موقع المناهج الإلكترونية

24	CHM.5.2.03.002.12 Calculate percent by mass of a solution	Textbook +Table 3+ Example problem 1 + Practice Problems	308, 309, 310
25	CHM.5.2.03.002.05 Calculate molarity when the moles or the mass of solute and volume of solution are given and vice versa	Textbook + Example problem 2 + Practice Problems	311, 312
26	CHM.5.2.03.002.16 Perform calculation involving diluting a concentrated solution	Textbook + Figure 9 + Example problem 3+ Practice Problems	314, 315, 316
27	CHM.5.2.03.002.07 Calculate molality when the moles or the mass of solute and mass of solvent are given and vice versa	Textbook + Example problem 4 + Practice Problems	316, 317

What is the percent by mass of NaHCO_3 in a solution containing 40.0 g of NaHCO_3 dissolved in 760.0 mL of water H_2O? (density of water = 1 g / mL) A – 8.40% B – 6.70 % C – 5.30 % D - 5.00 %	ما هي النسبة المئوية بالكتلة لـ NaHCO_3 في محلول يحتوي على 40.0 g من NaHCO_3 مذابة في 760.0 mL من الماء H_2O ؟ (كثافة الماء = 1 g / mL) 8.40% – A 6.70 % – B 5.30 % – C 5.00 % – D
--	---

What is the percent by mass of sodium chloride NaCl in a solution containing 4.0 g of NaCl dissolved in 100.0 g of water H_2O? 3.8% 4.0% 3.3% 4.8%	ما النسبة المئوية بالكتلة لكلوريد الصوديوم NaCl في محلول يحتوي على 4.0 g من NaCl مذابة في 100.0 g من الماء H_2O ؟
--	---

24	CHM.5.2.03.002.12 Calculate percent by mass of a solution	Textbook +Table 3+ Example problem 1 + Practice Problems	308, 309, 310
25	CHM.5.2.03.002.05 Calculate molarity when the moles or the mass of solute and volume of solution are given and vice versa	Textbook + Example problem 2 + Practice Problems	311, 312
26	CHM.5.2.03.002.16 Perform calculation involving diluting a concentrated solution	Textbook + Figure 9 + Example problem 3+ Practice Problems	314, 315, 316
27	CHM.5.2.03.002.07 Calculate molality when the moles or the mass of solute and mass of solvent are given and vice versa	Textbook + Example problem 4 + Practice Problems	316, 317

What is the molarity of a 50.0 mL solution containing 10.0 g of table sugar ($C_{12}H_{22}O_{11}$)?
(molar mass of table sugar = 342.3 g/mol)

ما مولارية 50.0 mL من محلول يحتوي 10.0 g من سكر المائدة ($C_{12}H_{22}O_{11}$) ؟
(الكتلة المولية لسكر المائدة = 342.3 g/mol)

0.584 M

200 M

5.84×10^{-4} M

2.00 M

What is the molarity of **2.5 L** of a solution containing **(5.95 g)** of KBr?

(Molar mass of KBr =119 g/mol)

ما مولارية محلول حجمه **2.5 L** مذاب فيه **(5.95 g)** من KBr؟

(الكتلة المولية لـ KBr = 119 g / mol)

A – 0.02 M

B – 0.04 M

C – 0.05 M

D - 0.08 M

0.02 M – A

0.04 M – B

0.05 M – C

0.08 M – D

24	CHM.5.2.03.002.12 Calculate percent by mass of a solution	Textbook +Table 3+ Example problem 1 + Practice Problems	308, 309, 310
25	CHM.5.2.03.002.05 Calculate molarity when the moles or the mass of solute and volume of solution are given and vice versa	Textbook + Example problem 2 + Practice Problems	311, 312
26	CHM.5.2.03.002.16 Perform calculation involving diluting a concentrated solution	Textbook + Figure 9 + Example problem 3+ Practice Problems	314, 315, 316
27	CHM.5.2.03.002.07 Calculate molality when the moles or the mass of solute and mass of solvent are given and vice versa	Textbook + Example problem 4 + Practice Problems	316, 317

What mass of calcium hydroxide Ca(OH)_2 is needed to produce 1500 mL of a 0.025M solution?

(Molar mass of $\text{Ca(OH)}_2 = 74.09 \text{ g/mol}$)

ما كتلة هيدروكسيد الكالسيوم Ca(OH)_2 اللازمة لتحضير محلول حجمه 1500 mL وتركيزه 0.025 M؟

(الكتلة المولية لـ $\text{Ca(OH)}_2 = 74.09 \text{ g/mol}$)

2.34 g

2.78 g

1.85 g

3.17 g

24	CHM.5.2.03.002.12 Calculate percent by mass of a solution	Textbook +Table 3+ Example problem 1 + Practice Problems	308, 309, 310
25	CHM.5.2.03.002.05 Calculate molarity when the moles or the mass of solute and volume of solution are given and vice versa	Textbook + Example problem 2 + Practice Problems	311, 312
26	CHM.5.2.03.002.16 Perform calculation involving diluting a concentrated solution	Textbook + Figure 9 + Example problem 3+ Practice Problems	314, 315, 316
27	CHM.5.2.03.002.07 Calculate molality when the moles or the mass of solute and mass of solvent are given and vice versa	Textbook + Example problem 4 + Practice Problems	316, 317

What is the molality (m) of a solution containing 7.10 g of Na_2SO_4 dissolved in 500.0 g of water? (Molar mass of Na_2SO_4 = 142 g/mol)	ما مولالية (m) لمحلول يحتوي على 7.10 g من Na_2SO_4 ذائبة في 500.0 g من الماء؟ (الكتلة المولية لـ Na_2SO_4 = 142 g/mol)
A – 0.500	0.500 – A
B – 0.250	0.250 – B
C – 0.100	0.100 – C
D - 0.050	0.050 – D

What is the molality of a solution that contain 3.5 mol of KCl dissolved in 1.5 kg of water?	ما مولالية محلول يحتوي على 3.5 mol من KCl المذابة في 1.5 kg من الماء؟
1.13 m	
4.28 m	
2.33 m	
7.78 m	

What is the mass in (g) of Na_2CO_3 in a 0.20 m solution dissolved in 750.0 g of water? (Molar mass of Na_2CO_3 = 106g/mol)	ما كتلة Na_2CO_3 بوحدة (g) في محلول تركيزه 0.20 m ذائبة في 750.0 g من الماء؟ (الكتلة المولية لـ Na_2CO_3 = 106 g / mol)
A – 12.5	12.5 – A
B – 15.9	15.9 – B
C – 24.6	24.6 – C
D - 28.5	28.5 – D

28	CHM.5.2.02.002.03 Explain three factors that affect the rate at which a solid solute dissolves in a liquid while explaining its effect	Textbook + Figure 14	322
29	CHM.5.2.02.002.04 Interpret temperature – solubility graph for solids	Textbook + Figure 16 + Table 4	324
30	CHM.5.2.02.002.03 Explain three factors that affect the rate at which a solid solute dissolves in a liquid while explaining its effect	Textbook + Figure 17	323, 324, 325

Factors That Affect Solvation

Solvation occurs only when the solute and solvent particles collide with each other. There are three common ways to increase the collisions between solute and solvent particles and thus increase the rate at which the solute dissolves. These three common ways, shown in **Figure 14**, are agitation, increasing the surface area of the solute, and increasing the temperature of the solvent.

Agitation

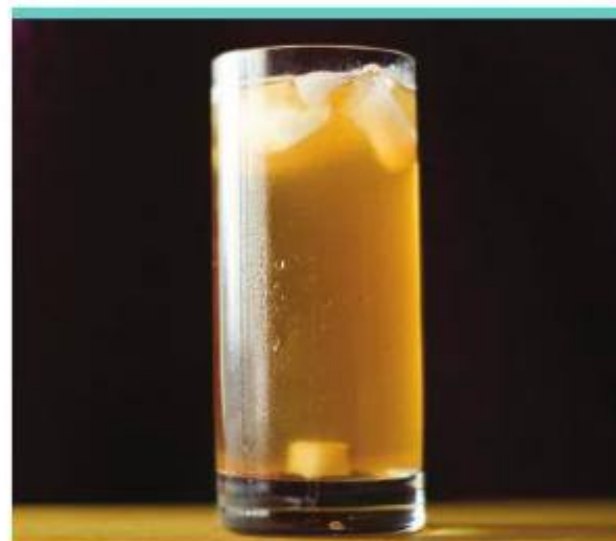
Stirring or shaking—agitation of the mixture—moves dissolved solute particles away from the contact surfaces more quickly and thereby allows new collisions between solute and solvent particles to occur. Without agitation, the solvated solute particles move away from the contact areas slowly.

Surface area

Breaking the solute into small pieces increases its surface area. A greater surface area allows more collisions to occur between the solute particles and the solvent particles. This is why a teaspoon of granulated sugar dissolves more quickly than an equal amount of sugar in cube form.

Temperature

The rate of solvation is affected by temperature. For example, sugar dissolves more quickly in hot tea, shown in **Figure 14**, than it does in iced tea. Additionally, hotter solvents generally can dissolve more solid solute. Hot tea can hold more dissolved sugar than the iced tea.



A sugar cube in iced tea will dissolve slowly, but stirring will make the sugar cube dissolve more quickly.



Granulated sugar dissolves faster in iced tea than a sugar cube. Stirring speeds the process.

Most solids act in the same way as sugar—as temperature increases, the rate of solvation also increases. Solvation of other substances, such as gases, decreases at higher temperatures. For example, a carbonated soft drink will lose its fizz (carbon dioxide) much faster when it is at room temperature than when it is cold.



Get It?

Explain how the rate of solvation can be increased.



Granulated sugar dissolves very quickly in hot tea.

Figure 14 Agitation, surface area, and temperature affect the rate of solvation.

28	CHM.5.2.02.002.03 Explain three factors that affect the rate at which a solid solute dissolves in a liquid while explaining its effect	Textbook + Figure 14	322
29	CHM.5.2.02.002.04 Interpret temperature – solubility graph for solids	Textbook + Figure 16 + Table 4	324
30	CHM.5.2.02.002.03 Explain three factors that affect the rate at which a solid solute dissolves in a liquid while explaining its effect	Textbook + Figure 17	323, 324, 325

Solubility

Solubility is a property of a solute that describes its ability to dissolve in a solvent. Just as solvation can be understood at the particle level, so can solubility. When a solute comes in contact with water molecules and solvation is initiated, at first, the solute particles are carried away from the solute. As the number of dissolved particles increases, collisions between dissolved solute particles and the remaining solute also increases.

However, solvation is a reversible process, and some colliding solute particles rejoin the solid solute. As long as the solvation rate is greater than this crystallization rate, solvation continues. Eventually, the rate of solvation and crystallization may become equal, forming a state of dynamic equilibrium, as shown in **Figure 15**, as long as the temperature remains constant. During dynamic equilibrium, solute particles return to the surface of the solid solute at the same rate as they are leaving.

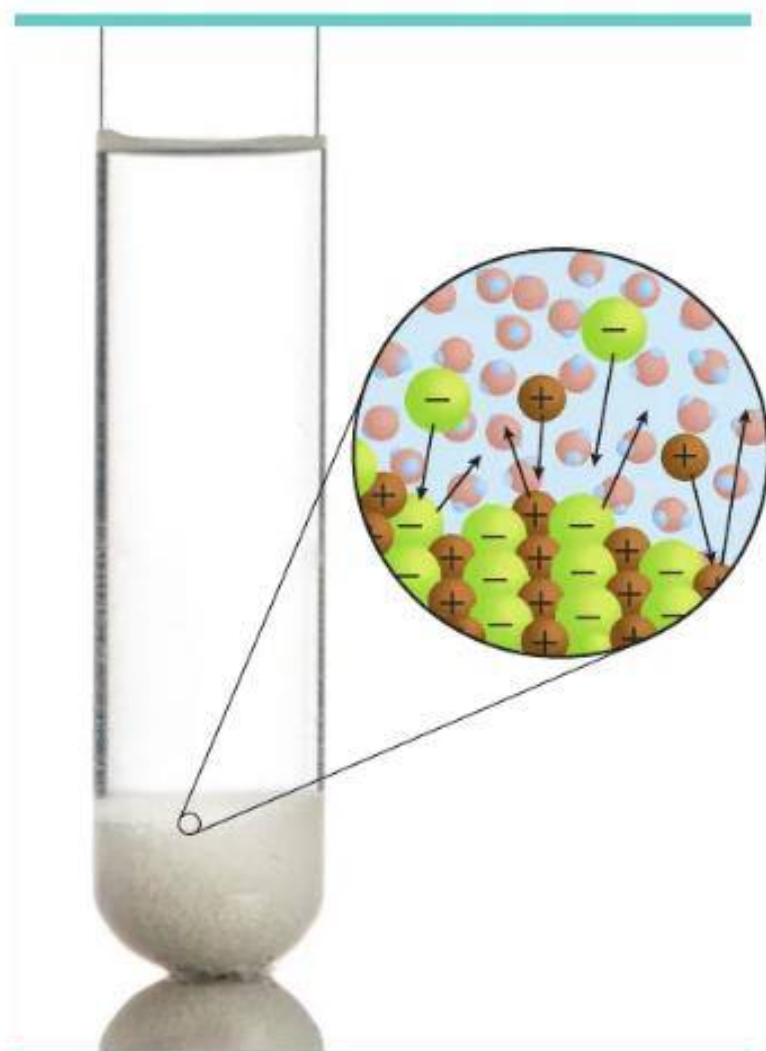


Figure 15 During dynamic equilibrium, the rate of solvation equals the rate of crystallization and the amount of dissolved solute remains unchanged.

Saturated solutions

Saturated solutions occur when no more solute can be dissolved. The solution contains the maximum amount of dissolved solute for a given amount of solvent at a specific temperature and pressure.

Unsaturated solutions

Unsaturated solutions contain less dissolved solute for a given temperature and pressure than a saturated solution. This means that more solute can be dissolved in an unsaturated solution.

28	CHM.5.2.02.002.03 Explain three factors that affect the rate at which a solid solute dissolves in a liquid while explaining its effect	Textbook + Figure 14	322
29	CHM.5.2.02.002.04 Interpret temperature – solubility graph for solids	Textbook + Figure 16 + Table 4	324
30	CHM.5.2.02.002.03 Explain three factors that affect the rate at which a solid solute dissolves in a liquid while explaining its effect	Textbook + Figure 17	323, 324, 325

Temperature and supersaturated solutions

Solubility is affected by raising the temperature of the solvent because the kinetic energy of its particles is increased, resulting in more-frequent collisions and collisions with greater energy than those that occur at lower temperatures. The fact that many substances are more soluble at high temperatures is demonstrated in **Figure 16**. For example, calcium chloride (CaCl_2) has a solubility of about 64 g CaCl_2 per 100 g H_2O at 10°C . Increasing the temperature to approximately 27°C increases the solubility by almost 50%, to 100 g CaCl_2 per 100 g H_2O . For other substances, such as cerium sulfate, $\text{Ce}_2(\text{SO}_4)_3$, solubility initially decreases rapidly as temperature increases, but then levels off and remains constant.

The effect of temperature on solubility is also illustrated by the data in **Table 4**. Notice in **Table 4** that at 20°C , 203.9 g of sucrose ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) dissolves in 100 g of water. At 100°C , 487.2 g of sucrose dissolves in 100 g of water, a nearly 140% increase in solubility.

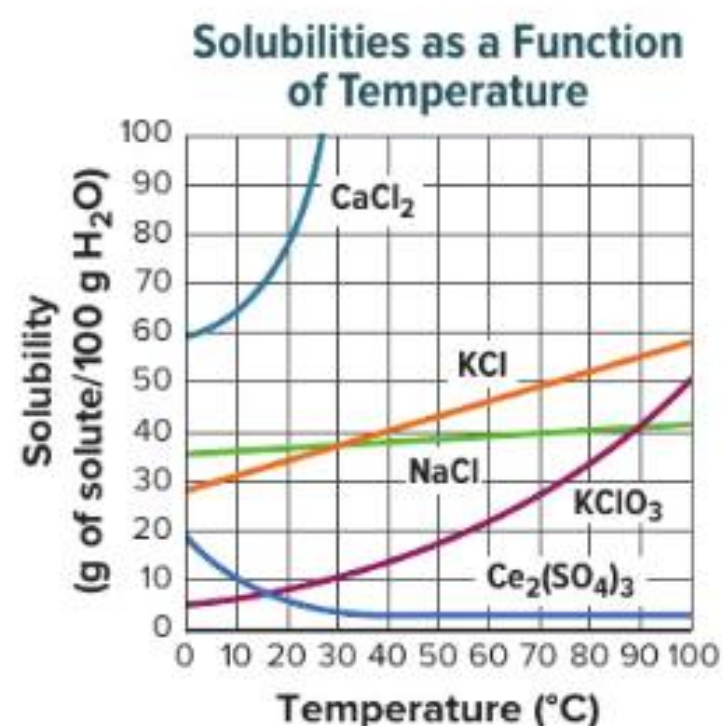


Figure 16 The solubilities of several substances as a function of temperature are shown in this graph.

Table 4 Solubilities of Solutes in Water at Various Temperatures

Substance	Formula	Solubility (g/100 g H_2O)*			
		0°C	20°C	60°C	100°C
Aluminum sulfate	$\text{Al}_2(\text{SO}_4)_3$	31.2	36.4	59.2	89.0
Barium hydroxide	$\text{Ba}(\text{OH})_2$	1.67	3.89	20.94	—
Calcium hydroxide	$\text{Ca}(\text{OH})_2$	0.189	0.173	0.121	0.076
Lithium sulfate	Li_2SO_4	36.1	34.8	32.6	—
Potassium chloride	KCl	28.0	34.2	45.8	56.3
Sodium chloride	NaCl	35.7	35.9	37.1	39.2
Silver nitrate	AgNO_3	122	216	440	733
Sucrose	$\text{C}_{12}\text{H}_{22}\text{O}_{11}$	179.2	203.9	287.3	487.2
Ammonia*	NH_3	1130	680	200	—
Carbon dioxide*	CO_2	1.713	0.878	0.359	—
Oxygen*	O_2	0.048	0.031	0.019	—



Get It?

Interpret According to the graph in **Figure 16**, what is the solubility of NaCl at 80°C ?

28	CHM.5.2.02.002.03 Explain three factors that affect the rate at which a solid solute dissolves in a liquid while explaining its effect	Textbook + Figure 14	322
29	CHM.5.2.02.002.04 Interpret temperature – solubility graph for solids	Textbook + Figure 16 + Table 4	324
30	CHM.5.2.02.002.03 Explain three factors that affect the rate at which a solid solute dissolves in a liquid while explaining its effect	Textbook + Figure 17	323, 324, 325

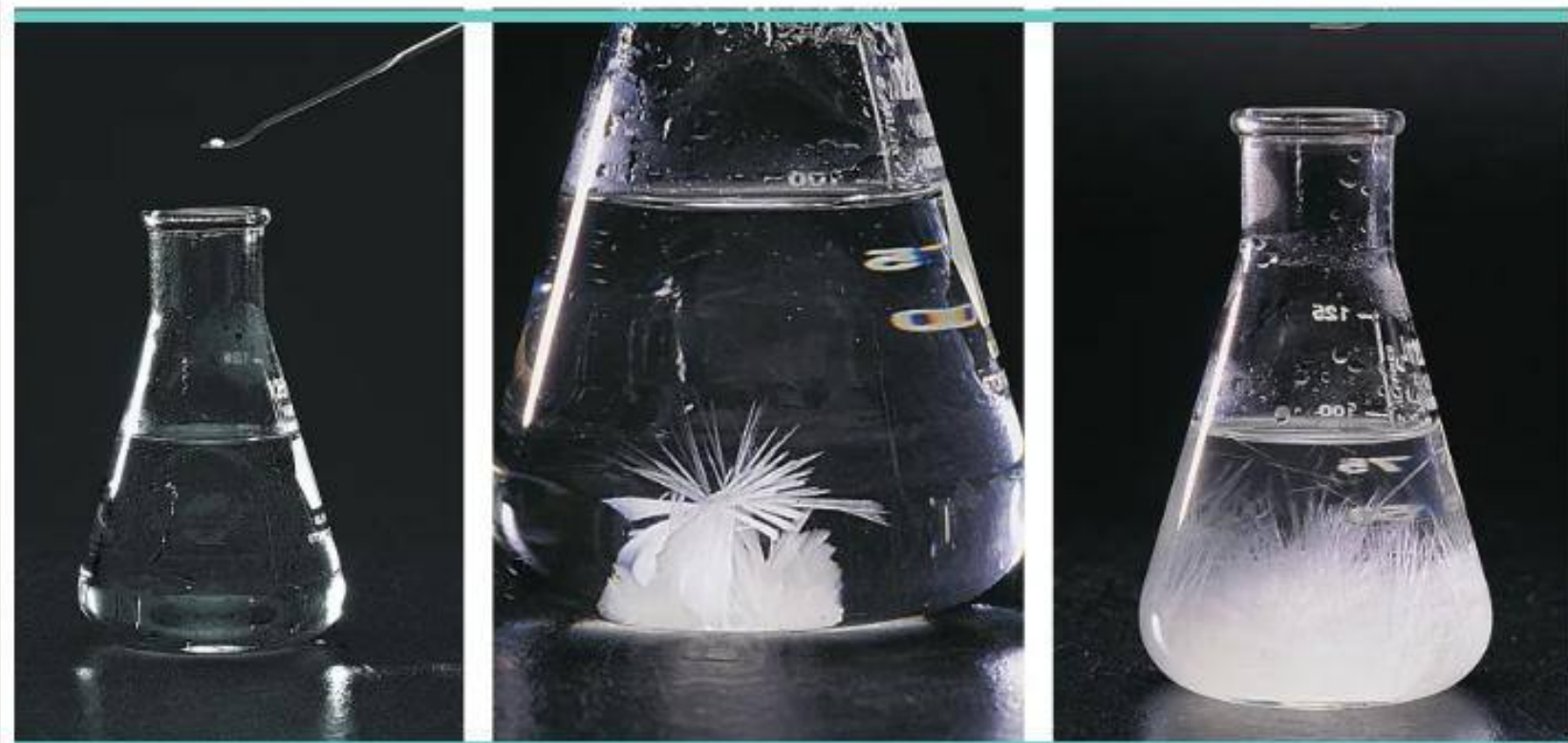


Figure 17 When a seed crystal is added to a supersaturated solution, the excess solute crystallizes out of the solution.

The fact that solubility changes with temperature and that some substances become more soluble with increasing temperature is the key to the formation of **supersaturated solutions**. A supersaturated solution contains more dissolved solute than a saturated solution at the same temperature. To make a supersaturated solution, a saturated solution is formed at a high temperature and then cooled slowly. The slow cooling allows the excess solute to remain dissolved in solution at the lower temperature.

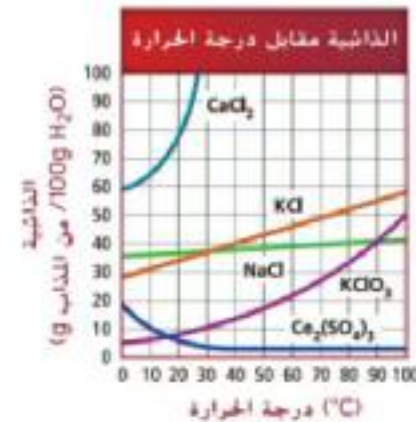
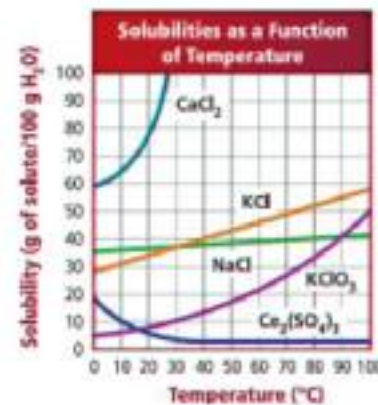
Supersaturated solutions are unstable. If a tiny amount of solute, called a seed crystal, is added to a supersaturated solution, the excess solute crystallizes quickly, as illustrated in **Figure 17**. Crystallization can also occur if the inside of the container is scratched or the supersaturated solution undergoes a physical shock, such as stirring or tapping the container. Rock candy and mineral deposits at the edges of mineral springs, such as those shown in **Figure 18**, are both formed from supersaturated solutions.



Figure 18 Hot spring mineral deposits are an example of crystals that formed from supersaturated solutions.

28	CHM.5.2.02.002.03 Explain three factors that affect the rate at which a solid solute dissolves in a liquid while explaining its effect	Textbook + Figure 14	322
29	CHM.5.2.02.002.04 Interpret temperature – solubility graph for solids	Textbook + Figure 16 + Table 4	324
30	CHM.5.2.02.002.03 Explain three factors that affect the rate at which a solid solute dissolves in a liquid while explaining its effect	Textbook + Figure 17	323, 324, 325

Using the graph below, which of the following statements is **correct**? باستخدام الرسم البياني أدناه، أي من العبارات التالية **صحيحة**؟



CaCl₂ has a solubility equals 64 g per 100 g of H₂O at 10°C

ذائبية CaCl₂ تساوي 64 g لكل 100 g H₂O عند 10°C

NaCl has the highest increase in solubility with increasing temperature

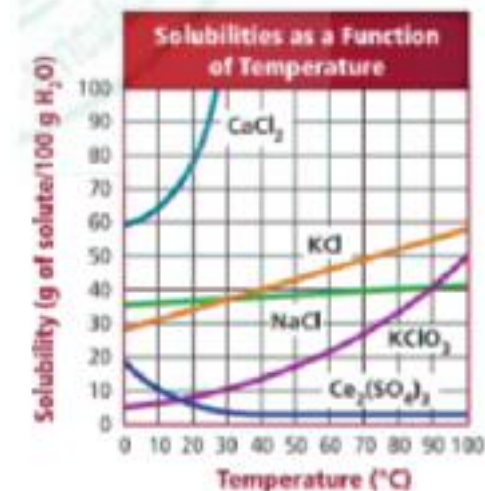
له NaCl أعلى زيادة في الذائبية مع زيادة درجة الحرارة

KClO₃ solubility decreases as temperature increases

ذائبية KClO₃ تنخفض إذا ما ارتفعت درجة الحرارة

Ce₂(SO₄)₃ solubility increases rapidly as temperature increases

ذائبية Ce₂(SO₄)₃ تزداد بسرعة عند زيادة درجة الحرارة



Which of the following substance present initially a fast solubility decreases when the temperature increase?

a. Ce(SO₄)₃

b. NaCl

c. KCl

d. CaCl₂