SECTION 12.2

The Ideal Gas Law

Key Terms

ideal gas law
universal gas constant (R)
partial pressure
Dalton's law of partial pressures

You have related the combined gas law to Avogadro's volume-mole gas relationship using two sets of conditions. This enabled you to make calculations of pressure, temperature, volume, and amount of gas by holding two of the variables constant while you manipulated a third and calculated a fourth. The ideal gas law combines all four variables into a single relationship that lets you summarize mathematically the relationships expressed by the combined gas law and Avogadro's law for ideal gases. This relationship is expressed as

$$PV = nRT$$

The relationship $PV = nRT$ is the **ideal gas law**: the pressure of a gas multiplied by its volume is equal to its amount in moles multiplied by a proportionality constant, R, and the temperature. The proportionality constant, called the **universal gas constant**, is $8.314 \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}}$, and applies to all gases. This value for R is based on the molar volume of a gas at STP being 22.4 L/mol. Thus, using $n = 1.00$ mol, $T = 273.15$ K, $P = 101.325$ kPa, and $V = 22.4$ L, R can be calculated by using the ideal gas law equation.

This equation lets you calculate one of the four gas variables if you have data for the other three. Importantly, the equation does not require that you compare two sets of conditions for the same sample of gas, which you must do when using the other gas laws.

ideal gas law a gas law that describes the relationship among volume, pressure, temperature, and amount (in moles) of an ideal gas: $PV = nRT$

universal gas constant (R) a proportionality constant that relates the pressure on and the volume of an ideal gas to its amount and temperature: $R = 8.314 \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}}$

Guidelines for Using the Ideal Gas Law

To use the ideal gas law equation properly, remember the following guidelines:

- Always convert the temperature to kelvin units (K).
- Always convert the mass to moles (mol).
- Always convert the volume to litres (L).
- Calculations are easier if you always convert the pressure to kilopascals (kPa). Then you can remember just one value of R $\left(8.314 \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}}\right)$ for each calculation. If you forget the value of R, you can calculate it by finding R for 1 mol of gas at STP.

Go to **scienceontario** to find out more

Density and Molar Mass of Gases

Since the ideal gas law considers the amount of gas present, this allows you to determine other properties of the gas. As you learned in the last section, the molar volume of a gas is defined as the space that is occupied by one mole of the gas. It is always given in units of L/mol. The density of a gas is similar to the density of a solid or a liquid. Density is found by dividing mass by volume. The density of a gas is usually reported in units of g/L. The molar mass of a gas refers to the mass (in grams) of one mole of the gas. As you know, you can calculate the molar mass of a substance by adding the molar masses of its atoms from the periodic table of the elements. You can also calculate molar mass by dividing the mass of a sample by the number of moles that are present. Molar mass is always expressed in the units g/mol. **Table 12.3** at the top of the next page summarizes molar volume, density, and molar mass. These three properties are closely related, and any one of the properties can be calculated using the other two properties.

The Sample Problems and Practice Problems that follow on the next two pages will reinforce your understanding of how the ideal gas law can be used to determine various properties of gases.

Table 12.3 Units of Molar Volume, Density, and Molar Mass

	Unit	Meaning	Calculations
Molar volume	L/mol	volume/amount	molar volume $= \frac{\text{volume}}{\text{amount (in moles)}}$ $v = \frac{V}{n}$
Density	g/L	mass/volume	density $= \frac{\text{mass}}{\text{volume}}$ $D = \frac{m}{V}$
Molar mass	g/mol	mass/amount	molar mass = sum of the molar masses of the atoms in the substance, or molar mass $= \frac{\text{mass}}{\text{amount (in moles)}}$ $M = \frac{m}{n}$

Suggested Investigation

Inquiry Investigation 12-A, Using the Ideal Gas Law to Determine the Pressure to Make Popcorn

Sample Problem

Finding the Volume of a Gas Using the Ideal Gas Law

Problem

Find the volume of 100.0 g of oxygen gas at SATP.

What Is Required?

You need to find the volume, V, that 100.0 g of oxygen occupies at standard ambient temperature and pressure.

What Is Given?

You know the conditions of SATP:

$T = 298.15$ K

$P = 100.0$ kPa

You know the molar mass of O_2(g): 2×16.00 g/mol $= 32.00$ g/mol

Plan Your Strategy	Act on Your Strategy
Find the amount (in moles) in 100.0 g of oxygen by dividing the mass by the molar mass of oxygen.	$n = \frac{m}{M}$ $= \frac{100.0\ \cancel{\text{g}}}{32.00\ \cancel{\text{g}}/\text{mol}} = 3.125\ \text{mol}$
Use the ideal gas law.	$PV = nRT$
Isolate the variable V by dividing each side of the equation by P.	$PV = nRT$ $\cancel{P}V\left(\frac{1}{\cancel{P}}\right) = nRT\left(\frac{1}{P}\right)$ $V = \frac{nRT}{P}$
Substitute numbers and units for the known variables into the formula and solve for V.	$V = \frac{(3.125\ \cancel{\text{mol}})\left(8.314\ \frac{\cancel{\text{kPa}} \cdot \text{L}}{\cancel{\text{mol}} \cdot \cancel{\text{K}}}\right)(298.15\ \cancel{\text{K}})}{100.0\ \cancel{\text{kPa}}}$ $= 77.46$ L

Check Your Solution

One mole of an ideal gas at SATP occupies 24.8 L. Therefore, it makes sense that the volume of slightly more than 3 mol of gas should be in the range of 77 L.

Sample Problem

Finding the Temperature of a Gas Using the Ideal Gas Law

Problem
Find the temperature, in °C, of 2.50 mol of gas that occupies a volume of 56.5 L at a pressure of 1.20 atm.

What Is Required?
You need to find the Celsius temperature of 2.50 mol of gas, given its volume and pressure.

What Is Given?
You know the amount, pressure, and volume of the gas:
$n = 2.50$ mol
$P = 1.20$ atm
$V = 56.5$ L

Plan Your Strategy	Act on Your Strategy
Convert the units of pressure from atm to kPa.	$P = 1.20\ \cancel{\text{atm}}\left(\frac{101.325\ \text{kPa}}{\cancel{\text{atm}}}\right)$ $= 121.59$ kPa
Use the ideal gas law.	$PV = nRT$
Isolate the variable T by dividing each side of the equation by nR.	$PV = nRT$ $\frac{PV}{nR} = \frac{\cancel{nR}T}{\cancel{nR}}$ $\frac{PV}{nR} = T$
Substitute numbers and units for the known variables in the formula and solve for T.	$T = \frac{(121.59\ \cancel{\text{kPa}})(56.5\ \cancel{\text{L}})}{2.50\ \cancel{\text{mol}}\left(8.314\ \frac{\cancel{\text{kPa}} \cdot \cancel{\text{L}}}{\cancel{\text{mol}} \cdot \text{K}}\right)}$ $= 330.519$ K
Convert the temperature to degrees Celsius.	$T = 330.519\ \text{K} - 273.15$ $= 57.4°\text{C}$

Check Your Solution
The pressure has been converted to kPa and the correct value of R is used, given the units of the variables. The temperature has been converted from the Kelvin scale to the Celsius scale.

Sample Problem

Determining the Molecular Formula of a Gas Using Percentage Composition and the Ideal Gas Law

Problem
What is the molecular formula of an unknown gas that is composed of 80.0% carbon and 20.0% hydrogen if a 4.60 g sample occupies a 2.50 L volume at 25.00°C and 152 kPa?

What Is Required?
You need to determine the molecular formula of an unknown gas.

Continued on next page

What Is Given?

You know the following:

$T = 25.00°C$

$P = 152$ kPa

$V = 2.50$ L

$m = 4.60$ g

$R = 8.314 \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}}$

The percentage composition of the gas is 80.0% carbon and 20.0% hydrogen.

Plan Your Strategy	Act on Your Strategy
Find the empirical formula using the molar masses of carbon and hydrogen and the percentage compositions.	For a 100 g sample: carbon: 80.0% × 100 g = 80.0 g hydrogen: 20.0% × 100 g = 20.0 g Determine the moles of each element using the formula $n = \frac{m}{M}$ For carbon: $n = \frac{80.0 \cancel{\text{g}}}{12.01 \cancel{\text{g}}/\text{mol}} = 6.661116$ mol For hydrogen: $n = \frac{20.0 \cancel{\text{g}}}{1.01 \cancel{\text{g}}/\text{mol}} = 19.80198$ mol The simplest ratio of the two elements provides the empirical formula. $\frac{6.661116}{6.661116}$ mol of C : $\frac{19.80198}{6.661116}$ mol of H The mole ratio is 1 mol of C : 3 mol of H. The empirical formula is CH_3.
Convert the temperature to kelvin units.	$T = 25.00°C + 273.15$ $= 298.15$ K
Use the ideal gas law.	$PV = nRT$
Isolate the variable n by dividing both sides of the equation by RT and rearranging the equation. Substitute numbers and units for the known variables into the formula to solve for n.	$PV = nRT$ $n = \frac{PV}{RT}$ $= \frac{152 \cancel{\text{kPa}} \times 2.50 \cancel{\text{L}}}{8.314 \frac{\cancel{\text{kPa}} \cdot \cancel{\text{L}}}{\text{mol} \cdot \cancel{\text{K}}} \times 298.15 \cancel{\text{K}}}$ $= 0.153299$ mol
Find the molar mass (M) of the gas by dividing the mass (m) of the gas by the amount (n) of the gas.	$M = \frac{m}{n}$ $= \frac{4.60 \text{ g}}{0.153299 \text{ mol}}$ $= 30.007$ g/mol
Compare the molar mass of the unknown gas with the molar mass of the empirical formula. Multiplying the empirical formula by the ratio of the two molar masses provides the molecular formula.	Molar mass of the unknown gas = 30.007 g/mol Molar mass of the empirical formula = 12.01 g/mol + 3(1.01 g/mol) = 15.04 g/mol ratio of molar masses = $\frac{30.007 \cancel{\text{g/mol}}}{15.04 \cancel{\text{g/mol}}}$ Since the ratio of the molar masses is 2:1, the molecular formula is $CH_3 \times 2 = C_2H_6$

Check Your Solution

The simple integer ratio of the two molar masses makes the molecular formula a reasonable answer. When the ideal gas law is used, all units cancel out except for mol, which is appropriate since the amount of gas was being calculated.

Sample Problem

Finding the Density of a Gas Using the Ideal Gas Law

Problem

What is the density of nitrogen gas in grams per litre, at 25.00°C and 126.63 kPa?

What Is Required?

You need to determine the density of nitrogen gas at 25.00°C and 126.63 kPa.

What Is Given?

You know the temperature of the gas and its pressure:

$T = 25.00°C$

$P = 126.63\ kPa$

You also know the molar mass of $N_2(g)$: $2 \times 14.01\ g/mol = 28.02\ g/mol$

Plan Your Strategy	Act on Your Strategy
Convert the temperature to kelvin units.	$T = 25.00°C + 273.15$ $= 298.15\ K$
Isolate the variable n and rearrange the equation. Substitute the numbers and units for P, T, R, and V into the equation. Since the volume is not given, set it as 1.00 L and solve for n.	$n = \frac{PV}{RT}$ $= \frac{126.63\ \cancel{kPa} \times 1.00\ \cancel{L}}{8.314 \frac{\cancel{kPa} \cdot \cancel{L}}{mol \cdot \cancel{K}} \times 298.15\ \cancel{K}} = 5.1085 \times 10^{-2}\ mol$
Convert n to the mass of nitrogen (m) in 1.00 L by multiplying the amount by the molar mass (M) of nitrogen.	$m = n \times M$ $= 5.1085 \times 10^{-2}\ \cancel{mol} \times 28.02\ g/\cancel{mol} = 1.4314\ g$
Determine the density by dividing the mass by the 1.00 L that was used in the ideal gas law equation.	$D = \frac{m}{V}$ $= \frac{1.4314\ g}{1.00\ L}$ $= 1.431\ g/L$

Check Your Solution

When the units cancel out in the ideal gas equation, the unit "mol" remains. When units cancel out in the density equation, the unit "g/L" remains.

Sample Problem

Finding the Molar Mass of a Gas Using the Ideal Gas Law

Problem

A 1.58 g sample of gas occupies a volume of 500.0 mL at STP. Calculate the molar mass of the gas.

What Is Required?

You need to find the molar mass, M, of a gas, based on a 1.58 g sample occupying a volume of 500 mL at standard temperature and pressure.

What Is Given?

You know the volume of the gas, the temperature and pressure at STP, as well as the mass of the sample:

$V = 0.5000\ L$

$T = 273.15\ K$

$P = 101.325\ kPa$

$m = 1.58\ g$

Continued on next page

Plan Your Strategy	Act on Your Strategy
Use the ideal gas law.	$PV = nRT$
Isolate the variable n by dividing each side of the equation by RT and then rearranging the equation. Substitute numbers and units for the known variables into the formula and solve for n.	$n = \frac{PV}{RT}$ $= \frac{(101.325\ \cancel{\text{kPa}})(0.500\ \cancel{\text{L}})}{\left(8.314\frac{\cancel{\text{kPa}}\cdot\cancel{\text{L}}}{\text{mol}\cdot\cancel{\text{K}}}\right)((273.15\cancel{\text{K}}))}$ $= 0.02231$ mol
Calculate the molar mass (M) by dividing the mass (m) by the amount (in mol).	$M = \frac{m}{n}$ $= \frac{1.58\text{ g}}{0.02231\text{ mol}}$ $= 70.8$ g/mol

Check Your Solution

The molar volume of an ideal gas at STP is approximately 22.4 L/mol, which is about 45 times the volume occupied by 1.58 g of the gas. That mass, multiplied by 45, is consistent with the answer.

Practice Problems

21. What is the volume of 5.65 mol of helium gas at a pressure of 98 kPa and a temperature of 18.0°C?

22. Propane, C_3H_8, is a common gas used to supply energy for barbecue cookers as well as energy-requiring appliances in cabins and cottages, and heavy equipment such as the forklift shown in the photograph below. If a tank contains 20.00 kg of propane, what volume of propane gas could be supplied at 22°C and 100.5 kPa?

Forklift trucks that run on propane are alternatives to those that run on gasoline or diesel fuel.

23. Find the Celsius temperature of nitrogen gas if a 5.60 g sample occupies 2.40×10^3 mL at 3.00 atm of pressure.

24. What is the pressure of 3.25 mol of hydrogen gas that occupies a volume of 67.5 L at a temperature of 295 K?

25. A weather balloon filled with helium gas has a volume of 960 L at 101 kPa and 25°C. What mass of helium was required to fill the balloon?

26. Find the molar mass of 6.24 g of an unknown gas that occupies 2.5 L at 18.3°C and 100.5 kPa.

27. A scientist isolates 2.366 g of a gas. The sample occupies a volume of 8.00×10^2 mL at 78.0°C and 103 kPa. Calculate the molar mass of the gas. Is the gas most likely to be bromine, krypton, neon, or fluorine?

28. What is the density of carbon dioxide gas, in grams per litre, at SATP?

29. A hydrocarbon gas used for fuel contains the elements carbon and hydrogen in percentages of 82.66 percent and 17.34 percent. Some of the gas, 1.77 g, was trapped in a 750 mL round-bottom flask. The gas was collected at a temperature of 22.1°C and a pressure of 99.7 kPa.

a. Determine the empirical formula for this gas.

b. Calculate the molar mass of the gas.

c. Determine the molecular formula for this gas.

30. A 10.0 g sample of an unknown liquid is vaporized at 120.0°C and 5.0 atm. The volume of the vapour is found to be 568.0 mL. The liquid is determined to be made up of 84.2% carbon and 15.8% hydrogen. What is the molecular formula of the liquid?

Dalton's Law of Partial Pressures

For the calculations you have done so far, 1 mol of any gas will have the same volume when the temperature and pressure are constant. Because gases are completely miscible, any mixture of non-reacting gases will have the same molar volume as any pure gas. However, scientists often want to consider just one particular gas in a mixture of gases. Is there a way to describe the pressure of one particular type of gas when it is mixed with other gases?

This question was investigated by John Dalton when he was studying atmospheric humidity. It is common for water vapour from the surroundings to mix with gases that are being studied in the laboratory. Upon adding water vapour to dry gases, Dalton observed that the total gas pressure was equal to the sum of the pressure of the dry gas and the pressure of the water vapour on the walls of the container. Mathematically, this is presented as

$$P_{\text{total}} = P_{\text{dry air}} + P_{\text{water vapour}}$$

Further studies showed that this phenomenon holds true for the addition of any type of gas to a mixture of gases. When gases are mixed, the pressure that any one gas exerts on the walls of the container is called the **partial pressure** of that gas. **Dalton's law of partial pressures** states that, in a mixture of gases that do not react chemically, the total pressure is the sum of the partial pressures of each individual gas.

partial pressure the portion of the total pressure of a mixture of gases contributed by a single gas component

Dalton's law of partial pressures a gas law stating that the total pressure of a mixture of gases is the sum of the individual pressures of each gas

Figure 12.5 illustrates Dalton's law of partial pressures. When two separate gases originally at the same temperature are mixed, the temperature—and thus the average speed of the gases—does not change. Only the total number of molecules in the container increases. As a result, there are more molecules colliding with the walls of the container and, thus, the pressure is higher.

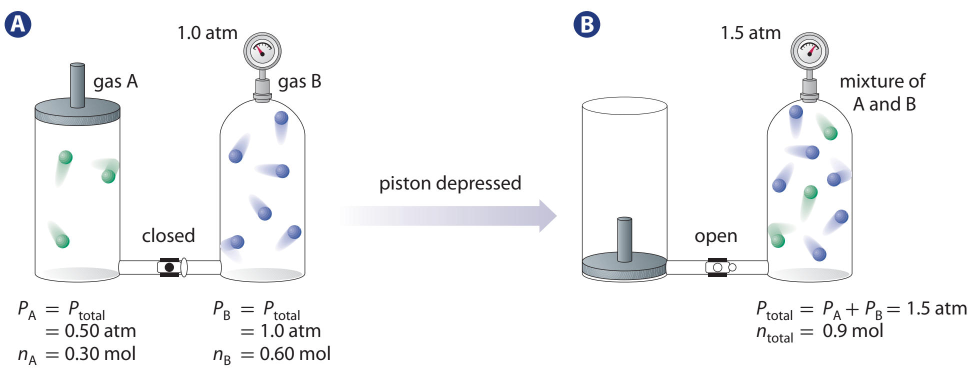


Figure 12.5 When two gases are in the same container, the molecules of each gas collide with the walls of the container as many times and with the same force as the molecules would do if the gases were in separate containers. Therefore, the pressure on the wall of the container with mixed gases is the sum of the pressures of the gases in separate containers.

Identify *How would the total pressure in the container with the mixture of A and B change if twice the amount (in moles) of gas A were added?*

Learning Check

7. Identify the two laws that are used to derive the ideal gas law.

8. Explain why the use of specific units for temperature, pressure, and volume is important when using the ideal gas law. What units should be used for each variable?

9. A gas mixture is composed of 40.0 percent xenon gas that has a partial pressure of 110 kPa. What is the total pressure of the gas mixture?

10. A gas mixture contains 12 percent helium, 25 percent argon, and 63 percent neon. If the total pressure is 2.0 atm, what is the partial pressure of each gas?

11. Why might the humidity of the air influence the total pressure of a gas in an open container?

12. Our atmosphere is approximately 80.0 percent nitrogen gas. At standard pressure, how much pressure is contributed by the nitrogen?

Collecting a Gas in the Laboratory

One of the safest ways to collect a gas in the laboratory is by downward displacement of water. A container, such as a graduated cylinder, is filled with water and inverted into a beaker of water. Care must be taken to avoid letting any air into the inverted cylinder. Tubing from the source of the gas is placed in the water and directed up into the inverted cylinder, as shown in **Figure 12.6**. The gas bubbles up through the water and collects in the closed end of the cylinder. The gas pushes the water down and out of the cylinder, leaving the gas sample trapped above the water. The cylinder can be adjusted so that the water level inside the cylinder is even with the water level in the beaker. The volume of the gas is read from the scale on the cylinder, and the pressure on the gas is the same as the atmospheric pressure. Many laboratories have a barometer in the room showing the barometric (atmospheric) pressure. The ambient temperature can be measured with a laboratory thermometer.

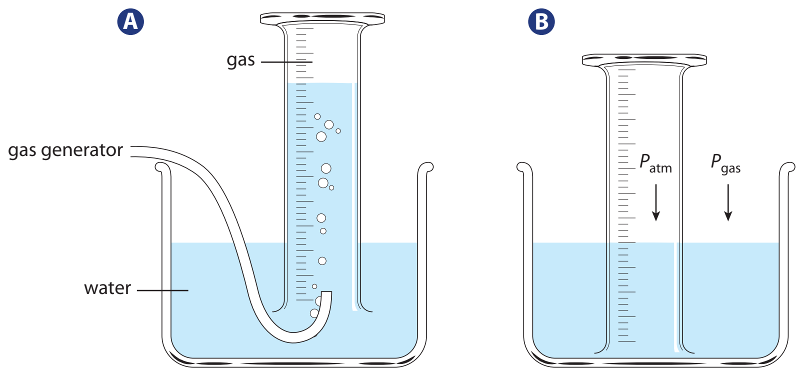


Figure 12.6 **(A)** Because gases have a much lower density than water does, gases float to the top of the container. **(B)** When the position of the container is adjusted so that the water levels inside and outside the container are the same, the gas pressures above the water are the same inside and outside the container.

Explain *Why must care be taken to avoid letting any air into the cylinder when it is inverted?*

When you collect gases over water, there is one factor that you must take into account, even if the gas is not soluble in water. The molecules of liquid water vaporize and mix with the molecules of the gas. Therefore, according to Dalton's law of partial pressures, the total pressure of the gases trapped in the cylinder is equal to the sum of the pressure exerted by each component of the gas mixture. Mathematically, this is represented by

$$P_{\text{total}} = P_{\text{gas}} + P_{\text{water vapour}}$$

Thus, to determine the pressure of the collected gas, often referred to as the pressure of the dry gas, the pressure of the water vapour must be subtracted from the total pressure. **Table 12.4** provides the partial pressures of water vapour at different temperatures.

Table 12.4 Partial Pressures of Water Vapour at Different Temperatures

Temperature (°C)	Pressure (kPa)	Temperature (°C)	Pressure (kPa)
15	1.71	22	2.81
16	1.81	23	2.99
17	1.93	24	3.17
18	2.07	25	3.36
19	2.20	26	3.36
20	2.33	27	3.56
21	2.49	28	3.37

Gas Stoichiometry

As you know, stoichiometry refers to the relationship between the amounts (in moles) of reactants and products in a chemical reaction. The stoichiometry of chemical reactions allows you to determine the quantity of one reactant or product if you know the quantity of another reactant or product.

When the volumes of gaseous reactants or products are not under the same conditions of temperature and pressure, you must use the ideal gas law to determine the quantities of reactant or products in a chemical reaction. A general set of steps to follow when using the ideal gas law to solve gas stoichiometry problems is shown below.

Suggested **Investigation**

Inquiry Investigation 12-B, Measuring the Molar Volume of a Gaseous Product in a Chemical Reaction

Steps for Solving Gas Stoichiometry Problems

1. Write a balanced equation for the reaction.
2. Convert all amounts to moles.
3. Compare molar amounts using stoichiometry ratios from the balanced equation. Solve for the unknown molar amount.
4. Convert the new molar amount into the units required, using a set of conditions with the ideal gas law, $PV = nRT$.

Also keep in mind that gaseous products of a chemical reaction are often collected over water in the laboratory. When using the ideal gas law for a gas collected over water, you must use Dalton's law of partial pressures to correct the pressure before substituting it into the gas law. To find the partial pressure of the dry gas, subtract the pressure of the water vapour from the total pressure: $P_{\text{dry gas}} = P_{\text{total}} - P_{\text{water vapour}}$

The following Sample Problem and Practice Problems will reinforce your understanding of the use of the ideal gas law in gas stoichiometry.

Sample Problem

Gas Stoichiometry Using the Ideal Gas Law

Problem

What volume of hydrogen gas is produced when excess sulfuric acid reacts with 40.0 g of iron at 18.0°C and 100.3 kPa?

What Is Required?

You need the volume of gas that is produced when sulfuric acid reacts with iron under specific temperature and pressure conditions.

What Is Given?

You know each of the following:
Reactants: sulfuric acid and iron
Products: an iron(II) compound and a gas
$m_{\text{iron}} = 40.0$ g
$T = 18.0$°C
$P = 100.3$ kPa
You also know the molar mass of iron: 55.85 g/mol

Continued on next page ❯

Plan Your Strategy	Act on Your Strategy
Write a balanced equation for the chemical reaction.	$Fe(s) + H_2SO_4(aq) \longrightarrow H_2(g) + FeSO_4(aq)$
Calculate the amount (in moles) of iron present by dividing the mass, m, by the molar mass, M. Use this value, along with the mole ratios from the balanced equation, to calculate the amount of gas produced.	$n = \frac{m}{M}$ $= \frac{40.0\text{ g}}{55.85\text{ g/mol}}$ $= 0.71620\text{ mol}$ The mole ratio is 1 mol of H_2 : 1 mol of Fe Therefore, the amount of hydrogen gas formed is $\frac{n\text{ mol }H_2}{0.71620\text{ mol Fe}} = \frac{1\text{ mol }H_2}{1\text{ mol Fe}}$ $0.71620\text{ mol Fe} \times \frac{n\text{ mol }H_2}{0.71620\text{ mol Fe}} = \frac{1\text{ mol }H_2}{1\ \cancel{\text{mol Fe}}} \times 0.71620\ \cancel{\text{mol Fe}}$ $n = 0.71620\text{ mol }H_2$
Isolate the variable V by dividing both sides of the ideal gas law equation by P. Rearrange the equation, substitute the values for the amount of gas, the temperature, and the pressure into the ideal gas law, and solve for the volume of the gas.	$PV = nRT$ $V = \frac{nRT}{P}$ $= \frac{0.71620\ \cancel{\text{mol}} \times 8.314\frac{\cancel{\text{kPa}}\cdot\text{L}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}} \times 291.15\ \cancel{\text{K}}}{100.3\ \cancel{\text{kPa}}}$ $= 17.3\text{ L}$

Check Your Solution

The answer is less than the molar volume of hydrogen gas at STP. Since less than 1 mol of hydrogen gas was formed, this seems reasonable.

Practice Problems

31. What volume of hydrogen gas will be produced at 93.0 kPa and 23°C from the reaction of 33 mg of magnesium with hydrochloric acid?

32. At STP, 0.72 g of hydrogen gas reacts with 8.0 L of chlorine gas. How many litres of hydrogen chloride gas are produced?

33. Determine the volume of nitrogen gas produced when 120 g of sodium azide, NaN(s), decomposes at 27°C and 100.5 kPa. Sodium metal is the other product.

34. When calcium carbonate, $CaCO_3(s)$, is heated, it decomposes to form calcium oxide, CaO(s), and carbon dioxide gas. How many liters of carbon dioxide will be produced at STP if 2.38 kg of calcium carbonate reacts completely?

35. When iron rusts, it undergoes a reaction with oxygen to form solid iron(III) oxide. Calculate the volume of oxygen gas at STP that is required to completely react with 52.0 g of iron.

36. Oxygen gas and magnesium react to form 2.43 g of magnesium oxide, MgO(s). What volume of oxygen gas at 94.9 kPa and 25.0°C would be consumed to produce this mass of magnesium oxide?

37. In the semiconductor industry, hexafluoroethane, $C_2F_6(g)$, is used to remove silicon dioxide, $SiO_2(s)$, according to the following chemical equation:

$$2SiO_2(s) + 2C_2F_6(g) + O_2(g) \longrightarrow 2SiF_4(g) + 2COF_2(g) + 2CO_2(g)$$

What mass of silicon dioxide reacts with 1.270 L of hexafluoroethane at 0.200 kPa and 400.0°C?

38. What mass of oxygen gas reacts with hydrogen gas to produce 0.62 L of water vapour at 100.0°C and 101.3 KPa?

39. One method of producing ammonia gas involves the reaction of ammonium chloride, $NH_4Cl(aq)$, with sodium hydroxide, NaOH(aq); water and aqueous sodium chloride are also products of the reaction. During an experiment, 98 mL of ammonia gas was collected using water displacement. If the gas was collected at 20.0°C and 780 mmHg, determine the amount of sodium hydroxide that must have reacted.

40. A student reacts 0.15 g of magnesium metal with excess dilute hydrochloric acid to produce hydrogen gas, which she collects over water. What volume of dry hydrogen gas does she collect over water at 28°C and 101.8 kPa?

Understanding Non-Ideal Gas Behaviour

The gas laws were developed to describe the behaviour and properties of ideal gases. How well does the behaviour of a *real* gas follow the laws for ideal gases? Look again at **Table 12.2** in the previous section. If you compare the measured molar volumes of gases in this table with the molar volume of an ideal gas (22.4 L/mol), a quick calculation shows that the maximum deviation of the molar volume of these gases is less than two percent from the ideal molar volume. In fact, real gases begin to deviate from ideal behaviour only when the temperature and pressure of their molecules diverge significantly from standard conditions. To understand why, recall what the kinetic molecular theory assumes about the of molecules of ideal gases:

- Particles move in straight lines at speeds determined by their temperature.
- Collisions are elastic and therefore do not use energy.
- Molecules are point masses, which have no volume.
- Molecules have no forces of attraction between each other and no forces of attraction with their container.

At STP, molecules of gases are moving very rapidly and are very far apart, which makes their interactions and volumes insignificant; thus, the assumptions for the ideal behaviour of gas molecules are valid. However, real gases begin to deviate from ideal behaviour when their molecules move slowly and are relatively close together. Low temperatures and high pressures are responsible for these conditions.

The Effects of Low Temperature

Intermolecular forces of attraction exist among real molecules of the same kind. These attractive forces do not significantly affect the behaviour of real gas molecules at standard temperatures, because the molecules are moving at high speeds and have a large amount of kinetic energy. When the molecules collide, their kinetic energy allows them to easily break their attractive interactions. The ability to break these short-range attractive forces decreases as the temperature of molecules decreases. At lower speeds and with reduced kinetic energy, molecules cannot easily break their attractive interactions. Eventually, these attractive interactions cause the gas to condense into a liquid.

The Effects of High Pressure: Reduced Collisions with the Container

Under standard atmospheric pressure, gas molecules are so far apart that interactions among them are very infrequent, so gases tend to behave ideally. When the external pressure on a gas increases substantially, however, the molecules are pushed closer together, and interactions are more frequent. As shown in **Figure 12.7**, when a gas molecule is about to collide with the walls of its container, nearby molecules exert attractive forces on the molecule, pulling it away. Therefore, the force of the collision with the wall is reduced, and the gas exerts less pressure than expected on the walls of the container. If you measured the pressure, you would find that it is lower than it would be for an ideal gas under the same conditions. Thus, calculations of other variables would be incorrect.

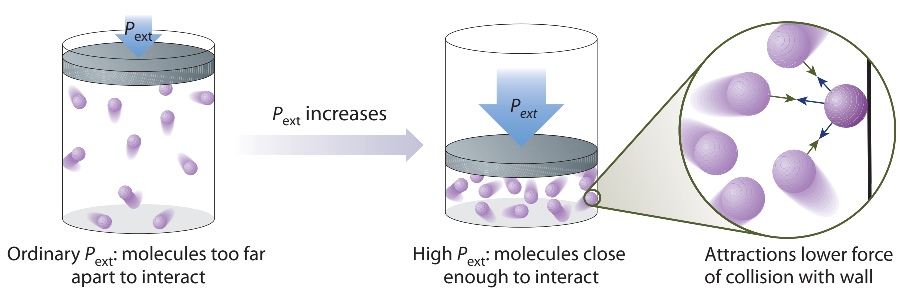


Figure 12.7 Under high pressures, gas molecules are close enough together to interact with one another when they are about to collide with the wall of the container. These interactions reduce the force of the collisions with the wall.

The Effects of High Pressure: Gas Molecule Volume Is Significant

Another effect of high pressure involves volume. Under standard atmospheric pressure, the total volume percent of a container taken up by gas molecules is insignificant, so gas molecules behave as if they were point masses. When high pressures reduce the volume of the container, the total volume percent taken up by the gas molecules becomes significant. The V in the ideal gas law represents the total empty space between molecules, as per the kinetic molecular theory. However, when you measure the volume of a gas at high pressures, the value of V is larger than the actual volume of empty space, as shown in **Figure 12.8**. If you used this measured volume to calculate other variables, the calculations of other variables would be incorrect.

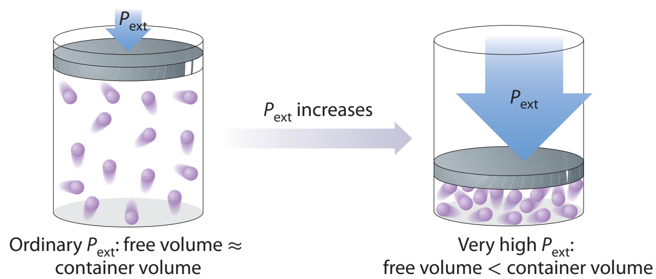


Figure 12.8 At standard atmospheric pressure, gas molecules are so far apart that they take up a very small percentage of the volume of the container. At high pressures, the actual volume of the gas molecules is a significant percentage of the volume of the container.

Summarizing Gas Laws and Properties

The graphic organizer in **Figure 12.9** summarizes the ideal gas law and concepts related to it.

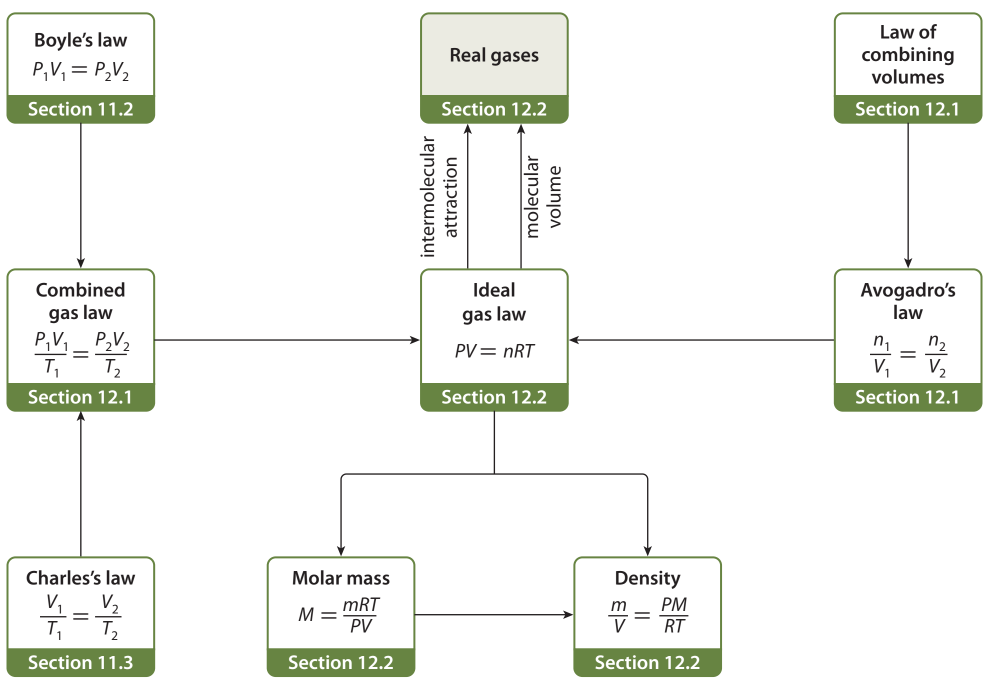


Figure 12.9 This graphic organizer relates various concepts you have learned in this unit to the ideal gas law.

Analyze *Which gas law is not included, and where would you place it in this organizer?*

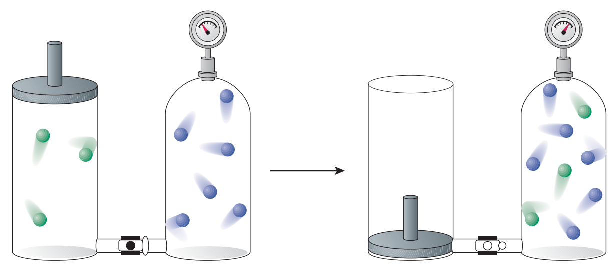
Section 12.2 REVIEW

Section Summary

- The ideal gas law is derived from the combined gas law and Avogadro's law. The mathematical equation for it is $PV = nRT$. R is the universal gas constant with a value of $8.314 \frac{\text{kPa} \cdot \text{L}}{\text{mol} \cdot \text{K}}$
- For gas stoichiometry, conditions of pressure, volume, and temperature must be considered. When the reactants and products are not under the same conditions the ideal gas law is used to calculate the amounts of gaseous reactants or products.
- Real gases behave similarly to ideal gases under standard conditions of temperature and pressure. However, this behaviour deviates under conditions of low temperature and high pressure.

Review Questions

1. **C** Using a graphic organizer, summarize how the ideal gas law is derived from existing gas law equations. Why is the ideal gas law a useful equation for a chemist?
2. **K/U** Rearrange the ideal gas law equation to solve for each of the following variables.
   a. volume of a gas
   b. pressure of a gas
   c. temperature of a gas
   d. amount of a gas
3. **T/I** A tank of chlorine gas contains 25.00 kg of pressurized liquid chlorine. If all the chlorine is released in its gaseous form at 17°C and 0.98 atm, what volume does it occupy?
4. **T/I** A 5.0×10^2 sample of sulfur dioxide, $SO_2(g)$, has a volume of 2.37×10^6 L. If the temperature of the gas is 30.0°C, determine the pressure of the gas in units of kPa.
5. **T/I** An 11.91 g sample of a gas occupies 3.20 L at 24.2°C and 102 kPa. Find the molar mass of the gas.
6. **T/I** A syringe was filled with halogen gas to the 60.0 mL mark at the pressure of the surrounding room. The mass of the empty syringe was 65.780 g and the mass of the syringe and gas was 65.875 g. The room had an atmospheric pressure of 101.8 kPa, and the temperature was 21.7°C. Calculate the molar mass and identify the gas.
7. **T/I** The molar mass of dry air is 28.57 g/mol. Determine the density of air at 25°C and 102 kPa.
8. **T/I** A gaseous compound contains 92.31% carbon and 7.69% hydrogen by mass. If 4.35 g of the gas occupies 4.16 L at 22.0°C and 738 torr, determine the molecular formula of the gas.
9. **T/I** A sample of hydrogen gas is collected by water displacement at 20.0°C when the atmospheric pressure is 99.8 kPa. What is the pressure of the "dry" hydrogen, if the partial pressure of water vapour is 2.33 kPa at that temperature?
10. **K/U** Use the following diagram to explain Dalton's law of partial pressures.



11. **T/I** Zinc metal reacts with nitric acid, $HNO_3(aq)$, to produce 34.0 L of dry hydrogen gas at 900.0 torr and 20.0°C. How many grams of zinc are consumed?
12. **A** The process of cellular respiration involves a chemical reaction in which glucose (a fuel) reacts with oxygen, forming carbon dioxide and water. The general equation for cellular respiration is
    $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(\ell)$
    If 10.0 g of glucose is reacted, calculate the total volume of carbon dioxide that is formed at 100.0 kPa and 37.0°C.
13. **T/I** When solid calcium is placed in water, it undergoes a vigorous reaction that produces solid calcium oxide, $CaO(s)$, and hydrogen gas. Determine the volume of dry hydrogen gas that is produced when 5.00 g of calcium reacts and the gas is collected at 42°C and 770 mmHg.
14. **C** Under what conditions does a real gas deviate from ideal gas behaviour? Use diagrams and kinetic molecular theory to illustrate your answer.