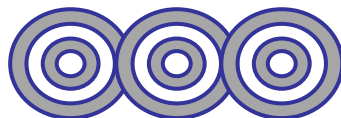




*Chemistry, Raymond Chang*  
*10th edition, 2010*  
*McGraw-Hill*



# Chapter 5

## Gases

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## **5.4**

# **The ideal gas equation**

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## Gas laws...

Boyle's law:  $V \propto \frac{1}{P}$  (at constant  $n$  and  $T$ )

Charles's law:  $V \propto T$  (at constant  $n$  and  $P$ )

Avogadro's law:  $V \propto n$  (at constant  $P$  and  $T$ )

We can combine all three expressions to form a single master equation for the behavior of gases:

$$V \propto \frac{nT}{P} \qquad V = R \frac{nT}{P}$$

$PV = nRT$  the ***ideal gas equation***, describes the relationship among the four variables  $P$ ,  $V$ ,  $T$ , and  $n$ .

where **R** , the proportionality constant, is called the **gas constant**.

An **ideal gas** is a hypothetical gas whose pressure-volume-temperature behavior can be completely accounted for by the ideal gas equation.

The molecules of an ideal gas do not attract or repel one another, and their volume is negligible compared with the volume of the container.

Although there is no such thing in nature as an ideal gas, the ideal gas approximation works rather well for most reasonable ***T*** and ***P*** ranges. Thus, we can safely use the ideal gas equation to solve many gas problems.

Before we can apply the ideal gas equation to a real system, we must evaluate the gas constant ***R***.

$$R = \frac{PV}{nT}$$

Units	Numerical Value
L-atm/mol-K	0.08206
J/mol-K*	8.314
cal/mol-K	1.987
m <sup>3</sup> -Pa/mol-K*	8.314
L-torr/mol-K	62.36

\*SI unit

In working problems with the ideal gas equation, the units of ***P***, ***T***, ***n*** and ***V*** must agree with the unit in the gas constant

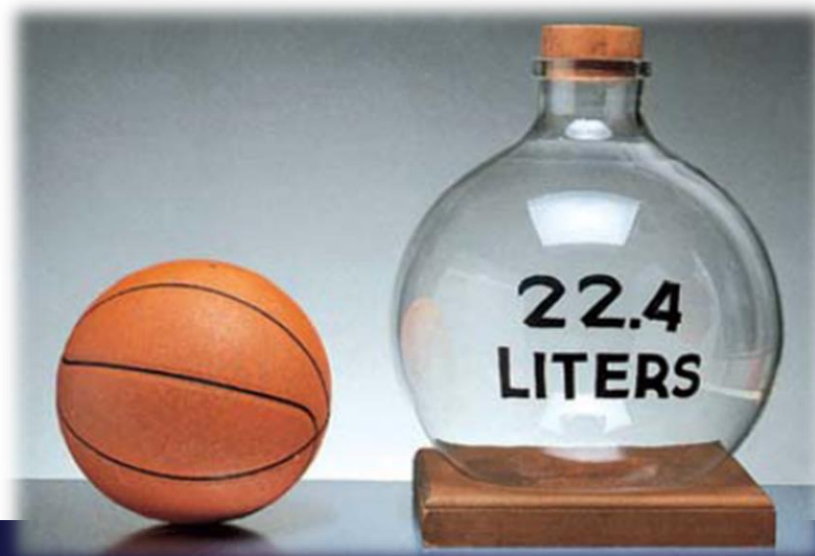
At 0°C (273.15 K) and 1 atm pressure, many real gases behave like an ideal gas. These conditions are called standard temperature and pressure (**STP**).

Suppose you have 1.00 mol of an ideal gas at **STP**. According to the ideal gas equation, the volume of the gas is:

$$V = \frac{nRT}{P} = \frac{(1.000 \text{ mol})(0.08206 \text{ L-atm/mol-K})(273.15 \text{ K})}{1.000 \text{ atm}} = 22.41 \text{ L}$$

The volume occupied by one mole of ideal gas at **STP** (22.41 L) is known as the **molar volume** of an ideal gas at **STP**.

1 mole of an ideal gas occupies 22.414 L, which is somewhat greater than the volume of a basketball.



## EXAMPLE

Sulfur hexafluoride ( $\text{SF}_6$ ) is a colorless, odorless, very unreactive gas. Calculate the pressure (in atm) exerted by 1.82 moles of the gas in a steel vessel of volume 5.43 L at  $69.5^\circ\text{C}$ .

$$\begin{aligned} P &= \frac{nRT}{V} \\ &= \frac{(1.82 \text{ mol})(0.0821 \text{ L} \cdot \text{atm}/\text{K} \cdot \text{mol})(69.5 + 273) \text{ K}}{5.43 \text{ L}} \\ &= 9.42 \text{ atm} \end{aligned}$$

## Practice Exercise

Calculate the volume (in liters) occupied by 2.12 moles of nitric oxide (NO) at 6.54 atm and  $76^\circ\text{C}$ .

## EXAMPLE

Calculate the volume (in liters) occupied by 7.40 g of  $\text{NH}_3$  at STP.

$$n = 7.4 / 17.03 = 0.4345 \text{ mol}$$

$$V = \frac{nRT}{P} = (0.4345 \times 0.0821 \times 273) / 1 = 9.74 \text{ L}$$

By using the fact that the molar volume of a gas occupies 22.41 L at STP, we can calculate the volume of a gas at STP without using the ideal gas equation.

$$\begin{aligned} V &= 7.40 \text{ g } \cancel{\text{NH}_3} \times \frac{1 \cancel{\text{ mol NH}_3}}{17.03 \text{ g } \cancel{\text{NH}_3}} \times \frac{22.41 \text{ L}}{1 \cancel{\text{ mol NH}_3}} \\ &= 9.74 \text{ L} \end{aligned}$$

## Practice Exercise

What is the volume (in liters) occupied by 49.8 g of  $\text{HCl}$  at STP?

If we need to deal with changes in pressure, volume, temperature, or the amount of gas, we must employ a modified form of the ideal gas equation that takes into account the initial and final conditions.

$$R = \frac{P_1 V_1}{n_1 T_1} \text{ (before change)}$$

$$R = \frac{P_2 V_2}{n_2 T_2} \text{ (after change)}$$

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

If  $n_1 = n_2$ , as is usually the case because the amount of gas normally does not change, the equation then becomes

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

The subscripts 1 and 2 denote the initial and final states of the gas, respectively.

## EXAMPLE

An inflated helium balloon with a volume of 0.55 L at sea level (1.0 atm) is allowed to rise to a height of 6.5 km, where the pressure is about 0.40 atm. Assuming that the temperature remains constant, what is the final volume of the balloon?

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

Because  $n_1 = n_2$  and  $T_1 = T_2$ ,

$$P_1 V_1 = P_2 V_2$$

Initial Conditions	Final Conditions
$P_1 = 1.0 \text{ atm}$	$P_2 = 0.40 \text{ atm}$
$V_1 = 0.55 \text{ L}$	$V_2 = ?$

Therefore,

$$\begin{aligned} V_2 &= V_1 \times \frac{P_1}{P_2} \\ &= 0.55 \text{ L} \times \frac{1.0 \text{ atm}}{0.40 \text{ atm}} \\ &= 1.4 \text{ L} \end{aligned}$$

## Practice Exercise

A sample of chlorine gas occupies a volume of 946 mL at a pressure of 726 mmHg. Calculate the pressure of the gas (in mmHg) if the volume is reduced at constant temperature to 154 mL.

## EXAMPLE

Argon is an inert gas used in lightbulbs to retard the vaporization of the tungsten filament. A certain lightbulb containing argon at 1.20 atm and 18°C is heated to 85°C at constant volume. Calculate its final pressure (in atm).

Because  $n_1 = n_2$  and  $V_1 = V_2$ , the Equation becomes

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

Initial Conditions	Final Conditions
$P_1 = 1.20 \text{ atm}$	$P_2 = ?$
$T_1 = (18 + 273) \text{ K} = 291 \text{ K}$	$T_2 = (85 + 273) \text{ K} = 358 \text{ K}$

$$\begin{aligned} P_2 &= P_1 \times \frac{T_2}{T_1} \\ &= 1.20 \text{ atm} \times \frac{358 \text{ K}}{291 \text{ K}} \\ &= 1.48 \text{ atm} \end{aligned}$$

## Practice Exercise

A sample of oxygen gas initially at 0.97 atm is cooled from 21°C to -68°C at constant volume. What is its final pressure (in atm)?

## EXAMPLE

A small bubble rises from the bottom of a lake, where the  $T$  and  $P$  are  $8^{\circ}\text{C}$  and  $6.4\text{ atm}$ , to the water's surface, where the temperature is  $25^{\circ}\text{C}$  and the pressure is  $1.0\text{ atm}$ . Calculate the final volume (in mL) of the bubble if its initial volume was  $2.1\text{ mL}$ .

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

the amount of air in the bubble remains constant, that is,  
 $n_1 = n_2$  so that

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

Initial Conditions	Final Conditions
$P_1 = 6.4\text{ atm}$	$P_2 = 1.0\text{ atm}$
$V_1 = 2.1\text{ mL}$	$V_2 = ?$
$T_1 = (8 + 273)\text{ K} = 281\text{ K}$	$T_2 = (25 + 273)\text{ K} = 298\text{ K}$

$$\begin{aligned} V_2 &= V_1 \times \frac{P_1}{P_2} \times \frac{T_2}{T_1} \\ &= 2.1\text{ mL} \times \frac{6.4\text{ atm}}{1.0\text{ atm}} \times \frac{298\text{ K}}{281\text{ K}} \\ &= 14\text{ mL} \end{aligned}$$

## Practice Exercise

A gas initially at  $4.0\text{ L}$ ,  $1.2\text{ atm}$ , and  $66^{\circ}\text{C}$  undergoes a change so that its final volume and temperature are  $1.7\text{ L}$  and  $42^{\circ}\text{C}$ . What is its final pressure? Assume the number of moles remains unchanged.

# Density Calculations

If we rearrange the ideal gas equation, we can calculate the density of a gas:

$$PV = nRT \longrightarrow \frac{n}{V} = \frac{P}{RT} \xrightarrow[\substack{n = \frac{m}{\mathcal{M}} \\ d = \frac{m}{V}}]{\quad} d = \frac{P \mathcal{M}}{RT}$$

Unlike molecules in condensed matter (that is, in liquids and solids), gaseous molecules are separated by distances that are large compared with their size.

Consequently, the density of gases is very low under atmospheric conditions. For this reason, gas densities are usually expressed in (g/L) rather than (g/mL).

## EXAMPLE

Calculate the density of carbon dioxide (CO<sub>2</sub>) in grams per liter (g/L) at 0.990 atm and 55°C.

$$T = 273 + 55 = 328 \text{ K}$$

$$\begin{aligned} d &= \frac{PM}{RT} \\ &= \frac{(0.990 \text{ atm})(44.01 \text{ g/mol})}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(328 \text{ K})} = 1.62 \text{ g/L} \end{aligned}$$

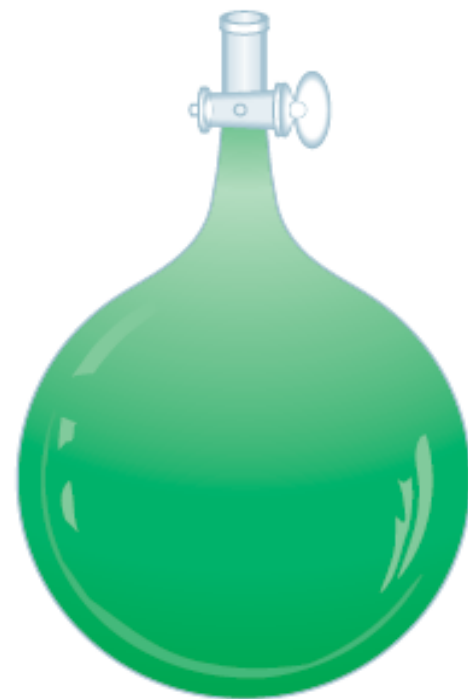
## Practice Exercise

What is the density (in g/L) of uranium hexafluoride (UF<sub>6</sub>) at 779 mmHg and 62°C?

## The Molar Mass of a Gaseous Substance

$$d = \frac{P \mathcal{M}}{RT} \longrightarrow \mathcal{M} = \frac{dRT}{P}$$

An apparatus for measuring the density of a gas. A **bulb** of known volume is filled with the gas under study at a certain ***T*** and ***P***. First the bulb is weighed, and then it is emptied (evacuated) and weighed again. The difference in masses gives the mass of the gas. Knowing the volume of the bulb, we can calculate the density of the gas. Once we know the density of a gas, we can calculate the molar mass of the substance.



## EXAMPLE

A chemist has synthesized a greenish-yellow gaseous compound of chlorine and oxygen and finds that its density is 7.71 g/L at 36°C and 2.88 atm. Calculate the molar mass of the compound and determine its molecular formula.

$$\begin{aligned}\mathcal{M} &= \frac{dRT}{P} \\ &= \frac{(7.71 \text{ g/L})(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(36 + 273) \text{ K}}{2.88 \text{ atm}} \\ &= 67.9 \text{ g/mol}\end{aligned}$$

Alternatively,

$$\text{molar mass of compound} = \frac{\text{mass of compound}}{\text{moles of compound}}$$

$$\begin{aligned}n &= \frac{PV}{RT} \\ &= \frac{(2.88 \text{ atm})(1.00 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(309 \text{ K})} \\ &= 0.1135 \text{ mol}\end{aligned}$$

$$\mathcal{M} = \frac{\text{mass}}{\text{number of moles}} = \frac{7.71 \text{ g}}{0.1135 \text{ mol}} = 67.9 \text{ g/mol}$$

## EXAMPLE

Chemical analysis of a gaseous compound showed that it contained 33.0 percent silicon (Si) and 67.0 percent fluorine (F) by mass. At 35°C, 0.210 L of the compound exerted a pressure of 1.70 atm. If the mass of 0.210 L of the compound was 2.38 g, calculate the molecular formula of the compound.

By assuming that we have 100 g of the compound, so the percentages are converted to grams. The number of moles of Si and F are:

$$n_{\text{Si}} = 33.0 \text{ g Si} \times \frac{1 \text{ mol Si}}{28.09 \text{ g Si}} = 1.17 \text{ mol Si}$$
$$n_{\text{F}} = 67.0 \text{ g F} \times \frac{1 \text{ mol F}}{19.00 \text{ g F}} = 3.53 \text{ mol F}$$

The empirical formula is  $\text{Si}_{1.17}\text{F}_{3.53}$ , dividing by the smaller subscript (1.17), we obtain  $\text{SiF}_3$ .

$$n = \frac{PV}{RT}$$
$$= \frac{(1.70 \text{ atm})(0.210 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(308 \text{ K})} = 0.0141 \text{ mol}$$

$$\mathcal{M} = \frac{2.38 \text{ g}}{0.0141 \text{ mol}} = 169 \text{ g/mol}$$

The molar mass of the empirical formula  $\text{SiF}_3$  is 85.09 g.

The ratio (molar mass/empirical molar mass) is always an integer ( $169/85.09 \sim 2$ ). Therefore, the molecular formula of the compound must be  $(\text{SiF}_3)_2$  or  $\text{Si}_2\text{F}_6$ .

## Practice Exercise

The density of a gaseous organic compound is 3.38 g/L at 40°C and 1.97 atm. What is its molar mass?

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## Practice Exercise

A gaseous compound is 78.14 percent boron and 21.86 percent hydrogen. At 27°C, 74.3 mL of the gas exerted a pressure of 1.12 atm. If the mass of the gas was 0.0934 g, what is its molecular formula?

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**5.5**

**Gas stoichiometry**

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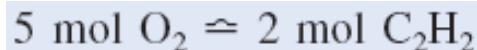
## EXAMPLE

Calculate the volume of  $\text{O}_2$  (in liters) required for the complete combustion of 7.64 L of acetylene ( $\text{C}_2\text{H}_2$ ) measured at the same temperature and pressure.

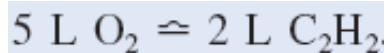


According to Avogadro's law, at the same  $T$  and  $P$ , the number of moles of gases are directly related to their volumes.

From the equation,



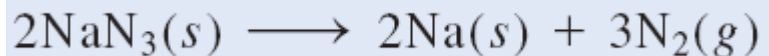
Therefore, we can also write



$$\begin{aligned} \text{volume of O}_2 &= 7.64 \text{ L C}_2\text{H}_2 \times \frac{5 \text{ L O}_2}{2 \text{ L C}_2\text{H}_2} \\ &= 19.1 \text{ L} \end{aligned}$$

## EXAMPLE

Sodium azide ( $\text{NaN}_3$ ) is used in some automobile air bags. The impact of a collision triggers the decomposition of  $\text{NaN}_3$  as follows:



The nitrogen gas produced quickly inflates the bag between the driver and the windshield and dashboard. Calculate the volume of  $\text{N}_2$  generated at  $80^\circ\text{C}$  and 823 mmHg by the decomposition of 60.0 g of  $\text{NaN}_3$ .

$$2 \text{ mol NaN}_3 \approx 3 \text{ mol N}_2$$

$$\frac{3 \text{ mol N}_2}{2 \text{ mol NaN}_3}$$

$$\begin{aligned} \text{moles of N}_2 &= 60.0 \text{ g NaN}_3 \times \frac{1 \text{ mol NaN}_3}{65.02 \text{ g NaN}_3} \times \frac{3 \text{ mol N}_2}{2 \text{ mol NaN}_3} \\ &= 1.38 \text{ mol N}_2 \end{aligned}$$

$$\begin{aligned} V &= \frac{nRT}{P} = \frac{(1.38 \text{ mol})(0.0821 \text{ L} \cdot \text{atm}/\text{K} \cdot \text{mol})(80 + 273 \text{ K})}{(823/760) \text{ atm}} \\ &= 36.9 \text{ L} \end{aligned}$$

## EXAMPLE

Aqueous lithium hydroxide solution is used to purify air in spacecrafts and submarines because it absorbs carbon dioxide, which is an end product of metabolism, according to the equation



The pressure of carbon dioxide inside the cabin of a submarine having a volume of  $2.4 \times 10^5 \text{ L}$  is  $7.9 \times 10^{-3} \text{ atm}$  at  $312 \text{ K}$ . A solution of lithium hydroxide ( $\text{LiOH}$ ) of negligible volume is introduced into the cabin. Eventually the pressure of  $\text{CO}_2$  falls to  $1.2 \times 10^{-4} \text{ atm}$ . How many grams of lithium carbonate are formed by this process?

The drop in  $\text{CO}_2$  pressure is  $(7.9 \times 10^{-3} \text{ atm}) - (1.2 \times 10^{-4} \text{ atm}) = 7.8 \times 10^{-3} \text{ atm}$ . Therefore, the number of moles of  $\text{CO}_2$  reacted is given by

$$n = P \times \left( \frac{V}{RT} \right)$$

$$\begin{aligned} \Delta n &= 7.8 \times 10^{-3} \text{ atm} \times \frac{2.4 \times 10^5 \text{ L}}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(312 \text{ K})} \\ &= 73 \text{ mol} \end{aligned}$$

$$1 \text{ mol CO}_2 \approx 1 \text{ mol Li}_2\text{CO}_3$$

so the amount of  $\text{Li}_2\text{CO}_3$  formed is also 73 moles.

$$\begin{aligned} \text{mass of Li}_2\text{CO}_3 \text{ formed} &= 73 \text{ mol Li}_2\text{CO}_3 \times \frac{73.89 \text{ g Li}_2\text{CO}_3}{1 \text{ mol Li}_2\text{CO}_3} \\ &= 5.4 \times 10^3 \text{ g Li}_2\text{CO}_3 \end{aligned}$$

### Practice Exercise

Assuming no change in temperature and pressure, calculate the volume of O<sub>2</sub> (in liters) required for the complete combustion of 14.9 L of butane (C<sub>4</sub>H<sub>10</sub>)



### Practice Exercise

The equation for the metabolic breakdown of glucose (C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>) is the same as the equation for the combustion of glucose in air:



Calculate the volume of CO<sub>2</sub> produced at 37°C and 1.00 atm when 5.60 g of glucose is used up in the reaction.

### Practice Exercise

A 2.14-L sample of hydrogen chloride (HCl) gas at 2.61 atm and 28°C is completely dissolved in 668 mL of water to form hydrochloric acid solution. Calculate the molarity of the acid solution. Assume no change in volume.

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**5.6**

**Dalton's law of partial  
pressures**

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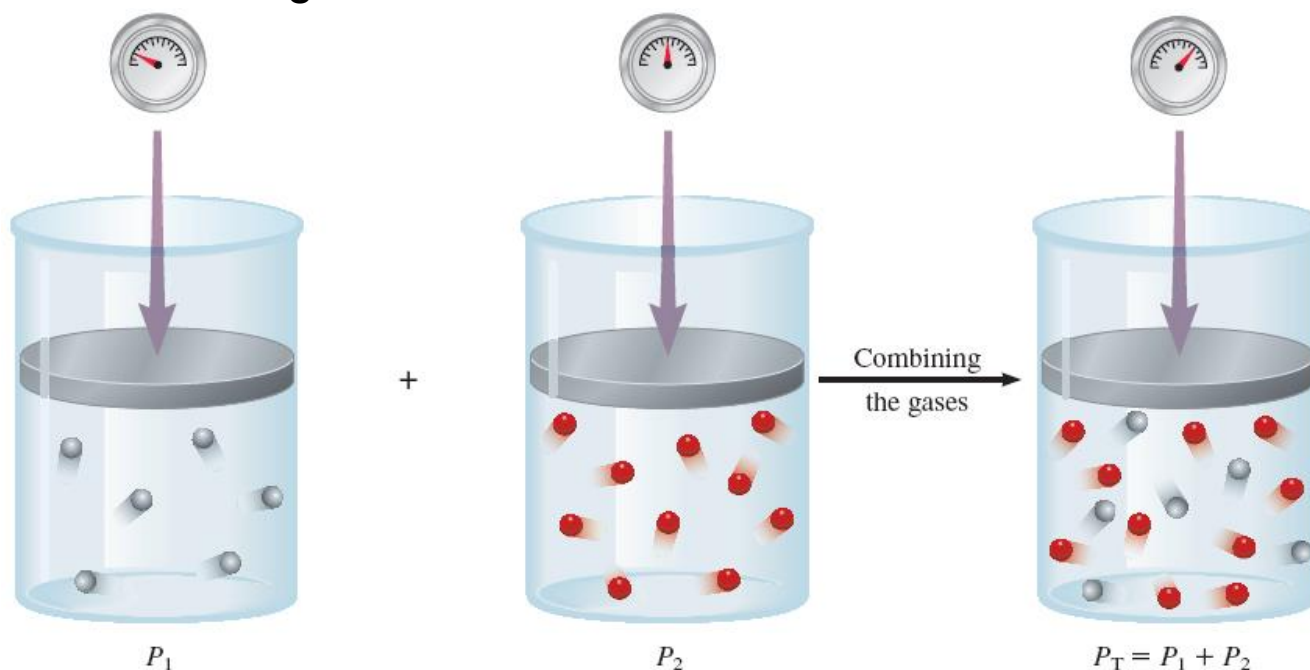
How do we deal with gases composed of a mixture of two or more substances ???

## - Dalton's Law of Partial Pressures

The total pressure of a mixture of gases equals the sum of the pressures that each would exert if it were present alone.

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

The pressure exerted by a particular component of a mixture of gases is called the **partial pressure** of that gas.



Consider a case in which two gases, A and B, are in a container of volume  $V$ . The pressure exerted by gas A, according to the ideal gas equation, is

$$P_A = \frac{n_A RT}{V} \quad P_B = \frac{n_B RT}{V}$$

In a mixture of gases A and B, the total pressure  $P_T$  is the result of the collisions of both types of molecules, A and B, with the walls of the container. Thus, according to Dalton's law,

$$\begin{aligned} P_T &= P_A + P_B \\ &= \frac{n_A RT}{V} + \frac{n_B RT}{V} \\ &= \frac{RT}{V}(n_A + n_B) \\ &= \frac{nRT}{V} \end{aligned}$$

where  $n$ , the total number of moles of gases present, is given by  $n = n_A + n_B$ , and  $P_A$  and  $P_B$  are the partial pressures of gases A and B, respectively.

For a mixture of gases, then,  $P_T$  depends only on the total number of moles of gas present, not on the nature of the gas molecules.

In general, the total pressure of a mixture of gases is given by

$$P_T = P_1 + P_2 + P_3 + \dots$$

where  $P_1, P_2, P_3, \dots$  are the partial pressures of components 1, 2, 3, . . . .

To see how each partial pressure is related to the total pressure, consider again the case of a mixture of two gases A and B. Dividing  $P_A$  by  $P_T$ , we obtain

$$\begin{aligned}\frac{P_A}{P_T} &= \frac{n_A RT/V}{(n_A + n_B) RT/V} && \text{where } X_A \text{ is called the mole fraction of A.} \\ &= \frac{n_A}{n_A + n_B} \\ &= X_A\end{aligned}$$

The **mole fraction** is a dimensionless quantity that expresses the ratio of the number of moles of one component to the number of moles of all components present.

$$X_i = \frac{n_i}{n_T}$$

where  $n_i$  and  $n_T$  are the number of moles of component  $i$  and the total number of moles present, respectively.

The mole fraction is always smaller than 1. We can now express the partial pressure as:

$$P_A = X_A P_T \qquad P_B = X_B P_T$$

The sum of the mole fractions for a mixture of gases must be unity (1). If only two components are present, then

$$X_A + X_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B} = 1$$

If a system contains more than two gases, then the partial pressure of the  $i$  th component is related to the total pressure by

$$P_i = X_i P_T$$

## EXAMPLE

A mixture of gases contains 4.46 moles of neon (Ne), 0.74 mole of argon (Ar), and 2.15 moles of xenon (Xe). Calculate the partial pressures of the gases if the total pressure is 2.00 atm at a certain temperature.

$$X_{\text{Ne}} = \frac{n_{\text{Ne}}}{n_{\text{Ne}} + n_{\text{Ar}} + n_{\text{Xe}}} = \frac{4.46 \text{ mol}}{4.46 \text{ mol} + 0.74 \text{ mol} + 2.15 \text{ mol}} = 0.607$$

$$\begin{aligned} P_{\text{Ne}} &= X_{\text{Ne}} P_{\text{T}} \\ &= 0.607 \times 2.00 \text{ atm} \\ &= 1.21 \text{ atm} \end{aligned}$$

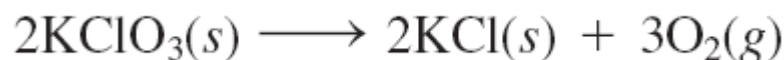
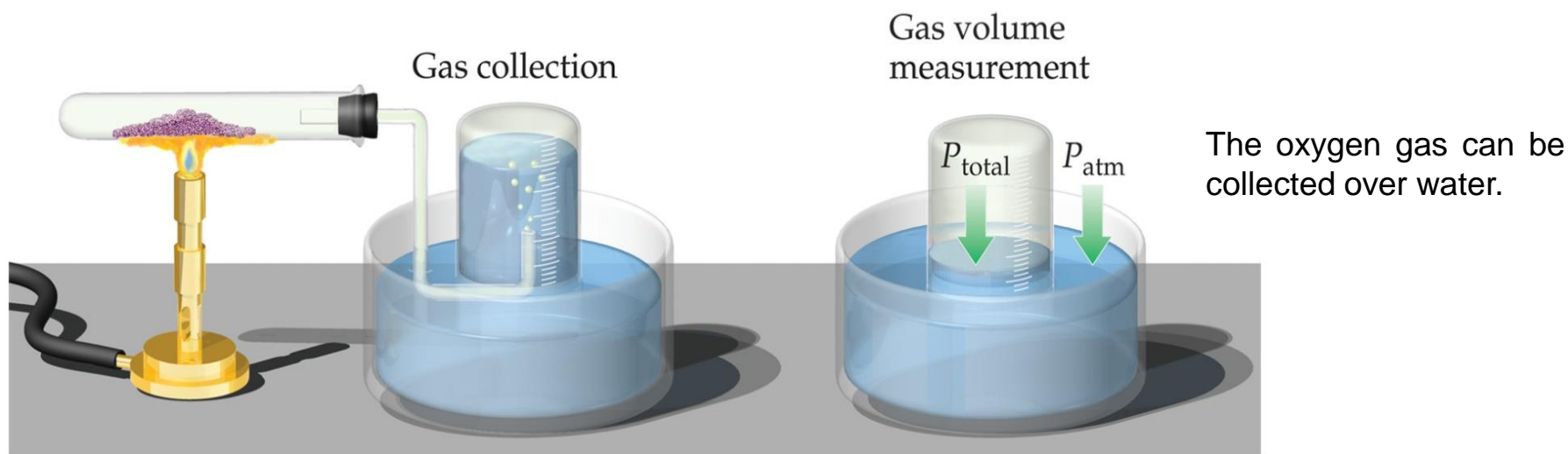
$$\begin{aligned} P_{\text{Ar}} &= X_{\text{Ar}} P_{\text{T}} \\ &= 0.10 \times 2.00 \text{ atm} \\ &= 0.20 \text{ atm} \end{aligned}$$

$$\begin{aligned} P_{\text{Xe}} &= X_{\text{Xe}} P_{\text{T}} \\ &= 0.293 \times 2.00 \text{ atm} \\ &= 0.586 \text{ atm} \end{aligned}$$

$$(1.21 + 0.20 + 0.586) \text{ atm} = 2.00 \text{ atm.}$$

# Collecting gases over water

An experiment that is often encountered in general chemistry laboratories involves determining the number of moles of gas collected from a chemical reaction. Sometimes this gas is collected over water.



When one collects a gas over water, there is water vapor mixed in with the gas.

- To find only the pressure of the desired gas, one must subtract the vapor pressure of water from the total pressure.

$$P_{\text{total}} = P_{\text{gas}} + P_{\text{H}_2\text{O}}$$

## EXAMPLE

Oxygen gas generated by the decomposition of potassium chlorate is collected over water. The volume of oxygen collected at 24°C and atmospheric pressure of 762 mmHg is 128 mL. Calculate the mass (in grams) of oxygen gas obtained. The pressure of the water vapor at 24°C is 22.4 mmHg.

$$P_T = P_{O_2} + P_{H_2O}$$

$$\begin{aligned} P_{O_2} &= P_T - P_{H_2O} \\ &= 762 \text{ mmHg} - 22.4 \text{ mmHg} \\ &= 740 \text{ mmHg} \end{aligned}$$

From the ideal gas equation

$$PV = nRT = \frac{m}{\mathcal{M}}RT$$

$$\begin{aligned} m &= \frac{PV\mathcal{M}}{RT} = \frac{(740/760)\text{atm}(0.128 \text{ L})(32.00 \text{ g/mol})}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(273 + 24) \text{ K}} \\ &= 0.164 \text{ g} \end{aligned}$$

## Practice Exercise

A sample of natural gas contains 8.24 moles of methane ( $\text{CH}_4$ ), 0.421 mole of ethane ( $\text{C}_2\text{H}_6$ ), and 0.116 mole of propane ( $\text{C}_3\text{H}_8$ ). If the total pressure of the gases is 1.37 atm, what are the partial pressures of the gases?

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## Practice Exercise

Hydrogen gas generated when calcium metal reacts with water is collected over water. The volume of gas collected at  $30^\circ\text{C}$  and pressure of 988 mmHg is 641 mL. What is the mass (in grams) of the hydrogen gas obtained? The pressure of water vapor at  $30^\circ\text{C}$  is 31.82 mmHg.

