



General Chemistry

CHEM 101
(3+1+0)

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Chapter 4

Gases

Substances That Exist as Gases

Elements that exist as **gases** at 25°C and 1 atmosphere

1A H																	8A He
2A Li	Be											3A B	4A C	5A N	6A O	7A F	Ne
Na	Mg	3B	4B	5B	6B	7B	8B		1B	2B	Zn	3A Al	4A Si	5A P	6A S	7A Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg							

- **Ionic compounds** do not exist as gases at 25°C and 1 atm. *because electrostatic forces between cations and anions.*
- **Molecular compounds** is more varied.
Gases; CO, CO₂, HCl, NH₃, and CH₄ (methane).
The majority are liquids or solids at room temperature.

Substances That Exist as Gases

TABLE 5.1 Some Substances Found as Gases at 1 atm and 25°C

Elements	Compounds
H ₂ (molecular hydrogen)	HF (hydrogen fluoride)
N ₂ (molecular nitrogen)	HCl (hydrogen chloride)
O ₂ (molecular oxygen)	HBr (hydrogen bromide)
O ₃ (ozone)	HI (hydrogen iodide)
F ₂ (molecular fluorine)	CO (carbon monoxide)
Cl ₂ (molecular chlorine)	CO ₂ (carbon dioxide)
He (helium)	NH ₃ (ammonia)
Ne (neon)	NO (nitric oxide)
Ar (argon)	NO ₂ (nitrogen dioxide)
Kr (krypton)	N ₂ O (nitrous oxide)
Xe (xenon)	SO ₂ (sulfur dioxide)
Rn (radon)	H ₂ S (hydrogen sulfide)
	HCN (hydrogen cyanide)*

*The boiling point of HCN is 26°C, but it is close enough to qualify as a gas at ordinary atmospheric conditions.

Substances That Exist as Gases

Physical Characteristics of Gases

- **Gases** assume the volume and shape of their containers.
- **Gases** are the most compressible state of matter.
- **Gases** will mix evenly and completely when confined to the same container.
- **Gases** have much lower densities than liquids and solids.



NO₂ gas

Pressure of a Gas

SI Units of Pressure

- **Velocity** is the change in distance with elapsed time.

$$\text{velocity} = \frac{\text{distance moved}}{\text{elapsed time}} \quad \text{m/s or cm/s}$$

- **Acceleration** is the change in velocity with time.

$$\text{acceleration} = \frac{\text{change in velocity}}{\text{elapsed time}} \quad \text{m/s}^2 \text{ (or cm/s}^2 \text{)}$$

$$\text{force} = \text{mass} \times \text{acceleration}$$

SI unit of force is the **newton (N)**,
where $1 \text{ N} = 1 \text{ kg m/s}^2$

Pressure of a Gas

SI Units of Pressure

- **Pressure** as force applied per unit area:

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

Units of Pressure

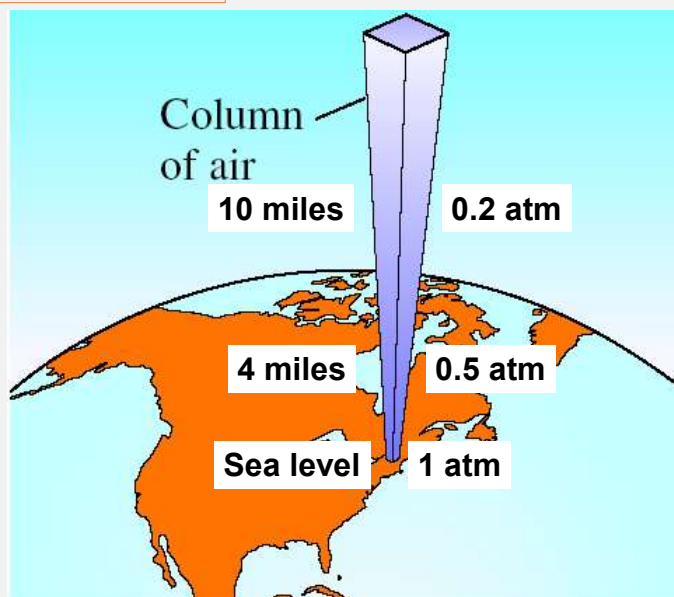
1 pascal (Pa) = 1 N/m²

1 atm = 760 mmHg = 760 torr

1 atm = 101,325 Pa

Pressure of a Gas

Atmospheric Pressure

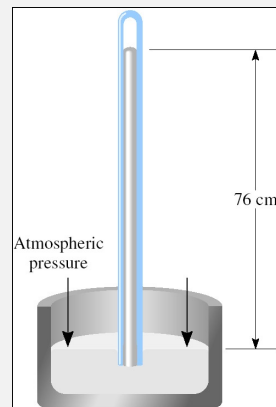


Pressure of a Gas

Atmospheric Pressure

How is atmospheric pressure measured?

- The **barometer** is probably the most familiar *instrument for measuring atmospheric pressure*.
- **Standard atmospheric pressure (1 atm)** is equal to *the pressure that supports a column of mercury exactly 760 mm (or 76 cm) high at 0°C at sea level*.



Pressure of a Gas

Atmospheric Pressure

EXAMPLE 5.1

The pressure outside a jet plane flying at high altitude falls considerably below standard atmospheric pressure. Therefore, the air inside the cabin must be pressurized to protect the passengers. What is the pressure in atmospheres in the cabin if the barometer reading is 688 mmHg?

Strategy Because 1 atm = 760 mmHg, the following conversion factor is needed to obtain the pressure in atmospheres

$$\frac{1 \text{ atm}}{760 \text{ mmHg}}$$

Solution The pressure in the cabin is given by

$$\begin{aligned} \text{pressure} &= 688 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} \\ &= 0.905 \text{ atm} \end{aligned}$$

Practice Exercise Convert 749 mmHg to atmospheres.

Pressure of a Gas

Atmospheric Pressure

EXAMPLE 5.2

The atmospheric pressure in San Francisco on a certain day was 732 mmHg. What was the pressure in kPa?

Strategy Here we are asked to convert mmHg to kPa. Because

$$1 \text{ atm} = 1.01325 \times 10^5 \text{ Pa} = 760 \text{ mmHg}$$

the conversion factor we need is

$$\frac{1.01325 \times 10^5 \text{ Pa}}{760 \text{ mmHg}}$$

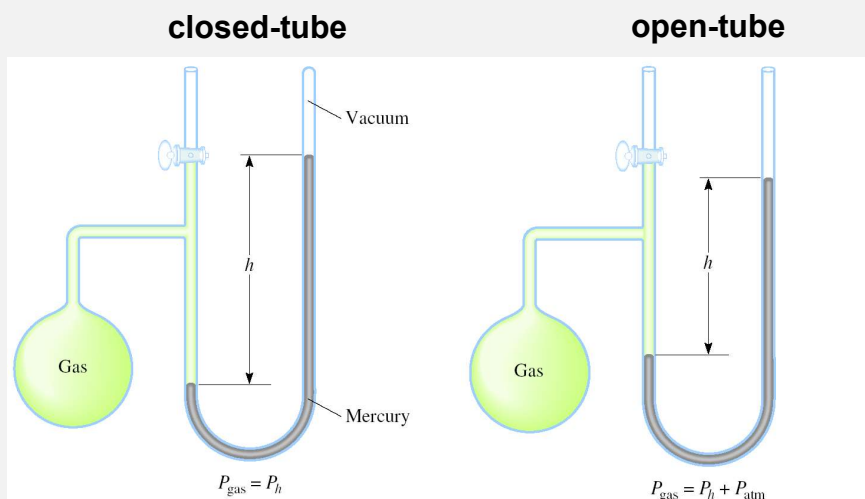
Solution The pressure in kPa is

$$\begin{aligned} \text{pressure} &= 732 \text{ mmHg} \times \frac{1.01325 \times 10^5 \text{ Pa}}{760 \text{ mmHg}} \\ &= 9.76 \times 10^4 \text{ Pa} \\ &= \mathbf{97.6 \text{ kPa}} \end{aligned}$$

Practice Exercise Convert 295 mmHg to kilopascals.

Pressure of a Gas

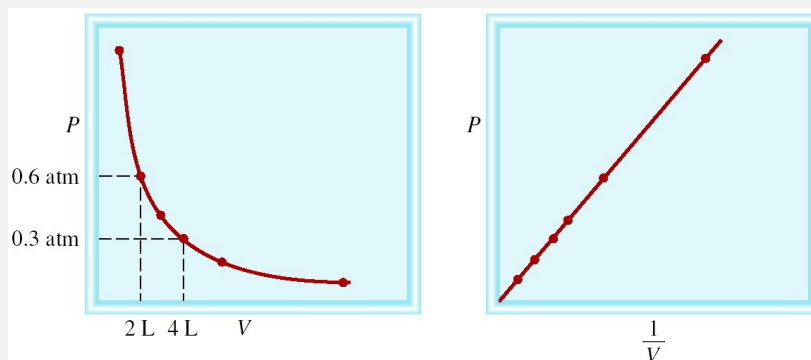
A **manometer** is a device used to measure the pressure of gases other than the atmosphere.



The Gas Laws

The Pressure-Volume Relationship: Boyle's Law

Boyle's law, states that the pressure of a fixed amount of gas at a constant temperature is inversely proportional to the volume of the gas.



$$P \propto 1/V$$

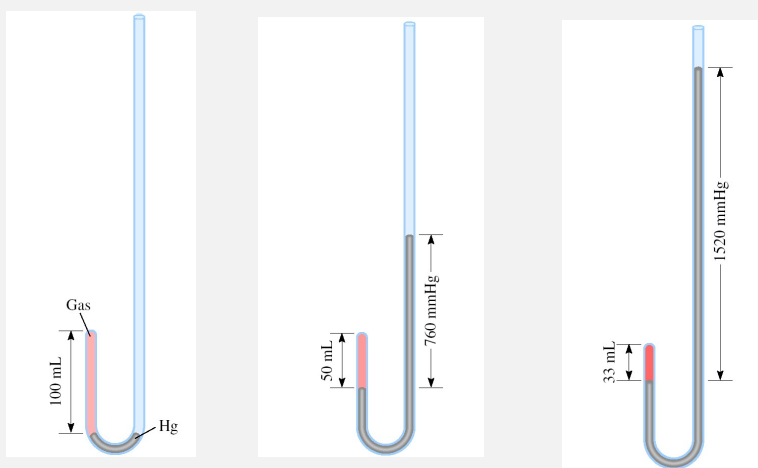
$$P \times V = \text{constant}$$

$$P_1 \times V_1 = P_2 \times V_2$$

Constant temperature
Constant amount of gas

The Gas Laws

Apparatus for Studying the Relationship Between Pressure and Volume of a Gas



As P (h) increases

V decreases

The Gas Laws

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

$$P \times V = \text{constant}$$

$$P_1 \times V_1 = P_2 \times V_2$$

$$P_1 = 726 \text{ mmHg} \quad P_2 = ?$$

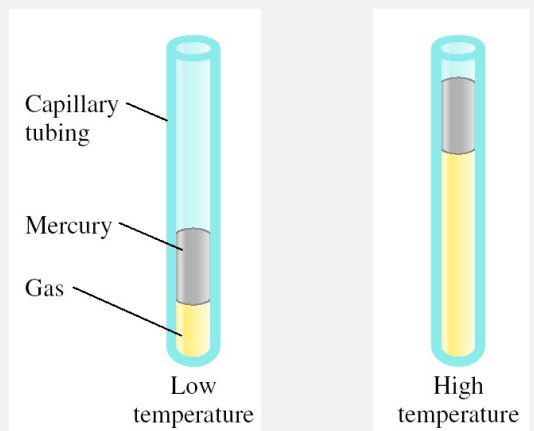
$$V_1 = 946 \text{ mL} \quad V_2 = 154 \text{ mL}$$

$$P_2 = \frac{P_1 \times V_1}{V_2} = \frac{726 \text{ mmHg} \times 946 \text{ mL}}{154 \text{ mL}} = 4460 \text{ mmHg}$$

The Gas Laws

The Temperature-Volume Relationship: Charles's and Gay-Lussac's Law

Variation in Gas Volume with Temperature at Constant Pressure



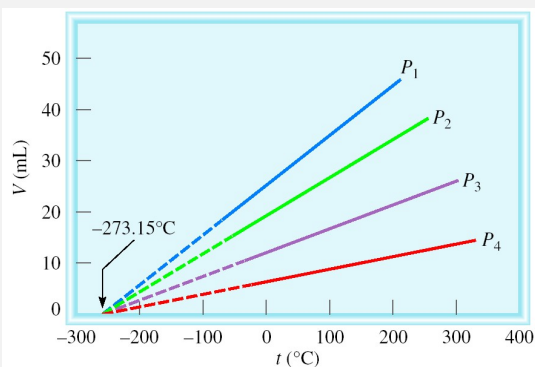
As T increases

V increases

The Gas Laws

The Temperature-Volume Relationship: Charles's and Gay-Lussac's Law

Charles's law, states that the volume of a fixed amount of gas maintained at constant pressure is directly proportional to the absolute temperature of the gas.



**Variation of Gas Volume
with Temperature at
Constant Pressure**

$$V \propto T$$

$$V = \text{constant} \times T$$

$$V_1/T_1 = V_2/T_2$$

Temperature **must** be in Kelvin

$$T (\text{K}) = t (^\circ\text{C}) + 273.15$$

The Gas Laws

A sample of carbon monoxide gas occupies 3.20 L at 125 °C. At what temperature will the gas occupy a volume of 1.54 L if the pressure remains constant?

$$V_1/T_1 = V_2/T_2$$

$$V_1 = 3.20 \text{ L}$$

$$V_2 = 1.54 \text{ L}$$

$$T_1 = 398.15 \text{ K}$$

$$T_2 = ?$$

$$T_1 = 125 (^\circ\text{C}) + 273.15 (\text{K}) = 398.15 \text{ K}$$

$$T_2 = \frac{V_2 \times T_1}{V_1} = \frac{1.54 \text{ L} \times 398.15 \text{ K}}{3.20 \text{ L}} = 192 \text{ K}$$

The Gas Laws

The Volume-Amount Relationship: Avogadro's Law

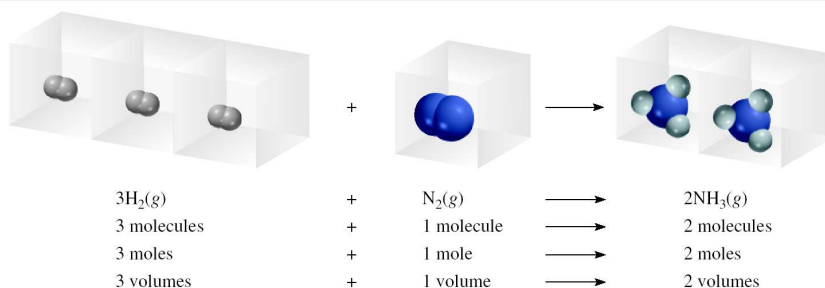
Avogadro's law, states that *at constant pressure and temperature, the volume of a gas is directly proportional to the number of moles of the gas present.*

$$V \propto \text{number of moles } (n)$$

$$V = \text{constant} \times n$$

$$V_1 / n_1 = V_2 / n_2$$

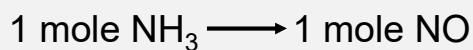
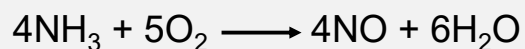
Constant temperature
Constant pressure



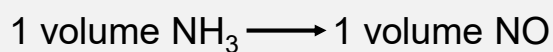
The Gas Laws

The Volume-Amount Relationship: Avogadro's Law

Ammonia burns in oxygen to form nitric oxide (NO) and water vapor. How many volumes of NO are obtained from one volume of ammonia at the same temperature and pressure?



At constant T and P



The Ideal Gas Equation

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

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

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

$$V \propto \frac{nT}{P}$$

$$V = \text{constant} \times \frac{nT}{P} = R \frac{nT}{P} \quad R \text{ is the gas constant}$$

$$PV = nRT$$

The Ideal Gas Equation

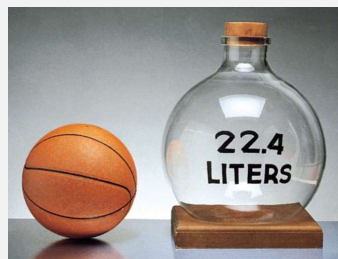
The conditions 0 °C (273.15 K) and 1 atm are called **standard temperature and pressure (STP)**.

Experiments show that at STP, 1 mole of an ideal gas occupies 22.414 L.

$$PV = nRT$$

$$R = \frac{PV}{nT} = \frac{(1 \text{ atm})(22.414 \text{ L})}{(1 \text{ mol})(273.15 \text{ K})}$$

$$R = 0.082057 \text{ L} \cdot \text{atm} / (\text{mol} \cdot \text{K})$$



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

The Ideal Gas Equation

What is the volume (in liters) occupied by 49.8 g of HCl at STP?

$$T = 0\text{ }^{\circ}\text{C} = 273.15\text{ K}$$

$$P = 1\text{ atm}$$

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

$$n = 49.8\text{ g} \times \frac{1\text{ mol HCl}}{36.45\text{ g HCl}} = 1.37\text{ mol}$$

$$V = \frac{1.37\text{ mol} \times 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \times 273.15\text{ K}}{1\text{ atm}}$$

$$V = 30.7\text{ L}$$

The Ideal Gas Equation

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

$$PV = nRT \quad n, V \text{ and } R \text{ are constant}$$

$$\frac{nR}{V} = \frac{P}{T} = \text{constant}$$

$$P_1 = 1.20\text{ atm} \quad P_2 = ?$$

$$T_1 = 291\text{ K} \quad T_2 = 358\text{ K}$$

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

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



The Ideal Gas Equation

EXAMPLE 5.3

Sulfur hexafluoride (SF₆) 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°C.

Strategy The problem gives the amount of the gas and its volume and temperature. Is the gas undergoing a change in any of its properties? What equation should we use to solve for the pressure? What temperature unit should we use?

Solution Because no changes in gas properties occur, we can use the ideal gas equation to calculate the pressure. Rearranging Equation (5.8), we write

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

Practice Exercise Calculate the volume (in liters) occupied by 2.12 moles of nitric oxide (NO) at 6.54 atm and 76°C.

The Ideal Gas Equation

EXAMPLE 5.4

Calculate the volume (in liters) occupied by 7.40 g of NH₃ at STP.

Strategy What is the volume of one mole of an ideal gas at STP? How many moles are there in 7.40 g of NH₃?

Solution Recognizing that 1 mole of an ideal gas occupies 22.41 L at STP and using the molar mass of NH₃ (17.03 g), we write the sequence of conversions as

$$\text{grams of NH}_3 \longrightarrow \text{moles of NH}_3 \longrightarrow \text{liters of NH}_3 \text{ at STP}$$

so the volume of NH₃ is given by

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

It is often true in chemistry, particularly in gas-law calculations, that a problem can be solved in more than one way. Here the problem can also be solved by first converting 7.40 g of NH₃ to number of moles of NH₃, and then applying the ideal gas equation ($V = nRT/P$). Try it.

Check Because 7.40 g of NH₃ is smaller than its molar mass, its volume at STP should be smaller than 22.41 L. Therefore, the answer is reasonable.

Practice Exercise What is the volume (in liters) occupied by 49.8 g of HCl at STP?

The Ideal Gas Equation

Density (d) Calculations

$$\frac{n}{V} = \frac{P\mathcal{M}}{RT}$$

The number of moles of the gas, n , is

$$n = \frac{m}{\mathcal{M}}$$

m is the mass of the gas in g \mathcal{M} is the molar mass of the gas

$$\frac{m}{\mathcal{M}V} = \frac{P}{RT} \qquad d = \frac{m}{V} = \frac{P\mathcal{M}}{RT}$$

Molar Mass (M) of a Gaseous Substance

$$\mathcal{M} = \frac{dRT}{P} \qquad d \text{ is the density of the gas in g/L}$$

The Ideal Gas Equation

A 2.10-L vessel contains 4.65 g of a gas at 1.00 atm and 27.0 °C.
What is the molar mass of the gas?

$$\mathcal{M} = \frac{dRT}{P} \qquad d = \frac{m}{V} = \frac{4.65 \text{ g}}{2.10 \text{ L}} = 2.21 \frac{\text{g}}{\text{L}}$$

$$\mathcal{M} = \frac{2.21 \frac{\text{g}}{\text{L}} \times 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \times 300.15 \text{ K}}{1 \text{ atm}}$$

$$\mathcal{M} = 54.5 \text{ g/mol}$$

The Ideal Gas Equation

EXAMPLE 5.9

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.

Strategy Because Equations (5.11) and (5.12) are rearrangements of each other, we can calculate the molar mass of a gas if we know its density, temperature, and pressure. The molecular formula of the compound must be consistent with its molar mass. What temperature unit should we use?

Solution From Equation (5.12)

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

Alternatively, we can solve for the molar mass by writing

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

From the given density we know there are 7.71 g of the gas in 1 L. The number of moles of the gas in this volume can be obtained from the ideal gas equation

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

Therefore, the molar mass is given by

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

We can determine the molecular formula of the compound by trial and error, using only the knowledge of the molar masses of chlorine (35.45 g) and oxygen (16.00 g). We know that a compound containing one Cl atom and one O atom would have a molar mass of 51.45 g, which is too low, while the molar mass of a compound made up of two Cl atoms and one O atom is 86.90 g, which is too high. Thus, the compound must contain one Cl atom and two O atoms and have the formula ClO_2 , which has a molar mass of 67.45 g.

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?

The Ideal Gas Equation

EXAMPLE 5.10

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.

Strategy This problem can be divided into two parts. First, it asks for the empirical formula of the compound from the percent by mass of Si and F. Second, the information provided enables us to calculate the molar mass of the compound and hence determine its molecular formula. What is the relationship between empirical molar mass and molar mass calculated from the molecular formula?

Solution We follow the procedure in Example 3.9 (p. 90) to calculate the empirical formula 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 given by

$$\begin{aligned} n_{\text{Si}} &= 33.0 \text{ g-Si} \times \frac{1 \text{ mol Si}}{28.09 \text{ g-Si}} = 1.17 \text{ mol Si} \\ n_{\text{F}} &= 67.0 \text{ g-F} \times \frac{1 \text{ mol F}}{18.998 \text{ g-F}} = 3.53 \text{ mol F} \end{aligned}$$

Therefore, the empirical formula is $\text{Si}_{1.17}\text{F}_{3.53}$, or, dividing by the smaller subscript (1.17), we obtain SiF_3 .

To calculate the molar mass of the compound, we need first to calculate the number of moles contained in 2.38 g of the compound. From the ideal gas equation

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

Because there are 2.38 g in 0.0141 mole of the compound, the mass in 1 mole, or the molar mass, is given by

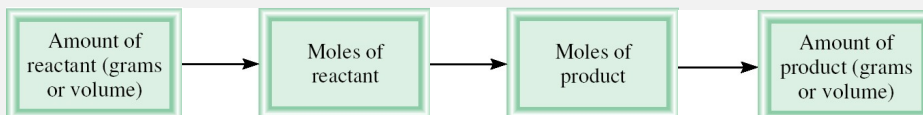
$$M = \frac{2.38 \text{ g}}{0.0141 \text{ mol}} = 169 \text{ g/mol}$$

The molar mass of the empirical formula SiF_3 is 85.09 g. Recall that the ratio (molar mass/empirical molar mass) is always an integer ($169/85.09 \approx 2$). Therefore, the molecular formula of the compound must be $(\text{SiF}_3)_2$ or Si_2F_6 .

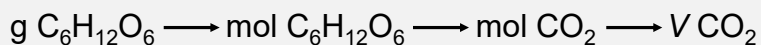
Practice Exercise A gaseous compound is 78.14 percent boron and 21.86 percent hydrogen. At 27°C, 74.3 mL of the gas exerted a pressure of 1.12 atm. If the mass of the gas was 0.0934 g, what is its molecular formula?

Gas Stoichiometry

The relationships between amounts (moles, n) and volume (V)



What is the volume of CO_2 produced at 37°C and 1.00 atm when 5.60 g of glucose are used up in the reaction:

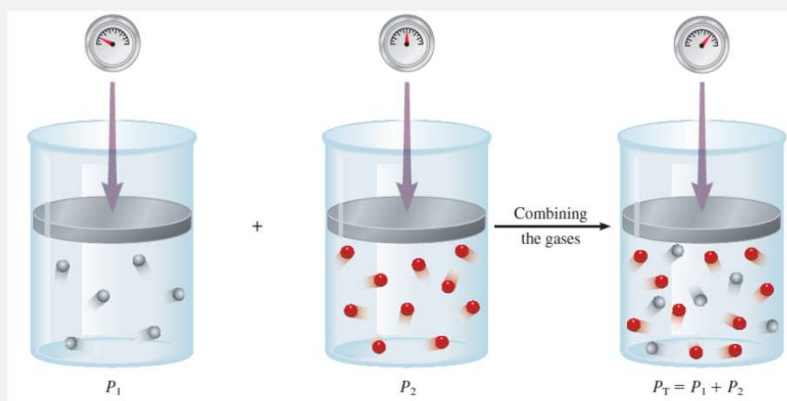


$$5.60 \text{ g C}_6\text{H}_{12}\text{O}_6 \times \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{180 \text{ g C}_6\text{H}_{12}\text{O}_6} \times \frac{6 \text{ mol CO}_2}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} = 0.187 \text{ mol CO}_2$$

$$V = \frac{nRT}{P} = \frac{0.187 \text{ mol} \times 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \times 310.15 \text{ K}}{1.00 \text{ atm}} = 4.76 \text{ L}$$

Dalton's law of Partial Pressures

Dalton's law of partial pressures, states that *the total pressure of a mixture of gases is just the sum of the pressures that each gas would exert if it were present alone.*



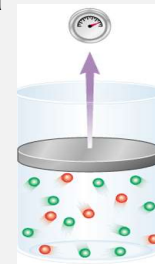
V and T are **constant**

Dalton's law of Partial Pressures

Consider a case in which two gases, A and B, are in a container of volume V.

$$P_A = \frac{n_A RT}{V} \quad n_A \text{ is the number of moles of A}$$

$$P_B = \frac{n_B RT}{V} \quad n_B \text{ is the number of moles of B}$$



$$P_T = P_A + P_B \quad X_A = \frac{n_A}{n_A + n_B} \quad X_B = \frac{n_B}{n_A + n_B}$$

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

$$P_i = X_i P_T$$

$$\text{mole fraction } (X_i) = \frac{n_i}{n_T}$$

Dalton's law of Partial Pressures

A sample of natural gas contains 8.24 moles of CH_4 , 0.421 moles of C_2H_6 , and 0.116 moles of C_3H_8 . If the total pressure of the gases is 1.37 atm, what is the partial pressure of propane (C_3H_8)?

$$P_i = X_i P_T \quad P_T = 1.37 \text{ atm}$$

$$X_{\text{propane}} = \frac{0.116}{8.24 + 0.421 + 0.116} = 0.0132$$

$$P_{\text{propane}} = 0.0132 \times 1.37 \text{ atm} = 0.0181 \text{ atm}$$

Dalton's law of Partial Pressures

EXAMPLE 5.14

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.

Strategy What is the relationship between the partial pressure of a gas and the total gas pressure? How do we calculate the mole fraction of a gas?

Solution According to Equation (5.14), the partial pressure of Ne (P_{Ne}) is equal to the product of its mole fraction (X_{Ne}) and the total pressure (P_T)

$$P_{\text{Ne}} = X_{\text{Ne}} P_T$$

need to find
want to calculate

Using Equation (5.13), we calculate the mole fraction of Ne as follows:

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

Therefore,

$$P_{\text{Ne}} = X_{\text{Ne}} P_T = 0.607 \times 2.00 \text{ atm} = 1.21 \text{ atm}$$

Similarly,

$$P_{\text{Ar}} = X_{\text{Ar}} P_T = 0.10 \times 2.00 \text{ atm} = 0.20 \text{ atm}$$

and

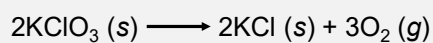
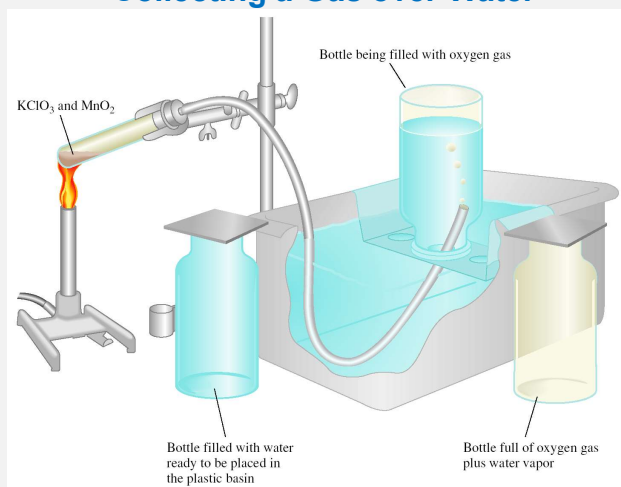
$$P_{\text{Xe}} = X_{\text{Xe}} P_T = 0.293 \times 2.00 \text{ atm} = 0.586 \text{ atm}$$

Check Make sure that the sum of the partial pressures is equal to the given total pressure; that is, $(1.21 + 0.20 + 0.586) \text{ atm} = 2.00 \text{ atm}$.

Practice Exercise A sample of natural gas contains 8.24 moles of methane (CH_4), 0.421 mole of ethane (C_2H_6), and 0.116 mole of propane (C_3H_8). If the total pressure of the gases is 1.37 atm, what are the partial pressures of the gases?

Dalton's law of Partial Pressures

Collecting a Gas over Water



Dalton's law of Partial Pressures

- Gas does not react with water and that it is not appreciably soluble in it.
- Oxygen gas is ok, but not for gases such as NH_3 .
- The oxygen gas collected in this way is not pure, however, because water vapor is also present in the bottle.
- The total gas pressure is equal to the sum of the pressures exerted by the oxygen gas and the water vapor:

$$P_T = P_{\text{O}_2} + P_{\text{H}_2\text{O}}$$

Dalton's law of Partial Pressures

Vapor of Water and Temperature

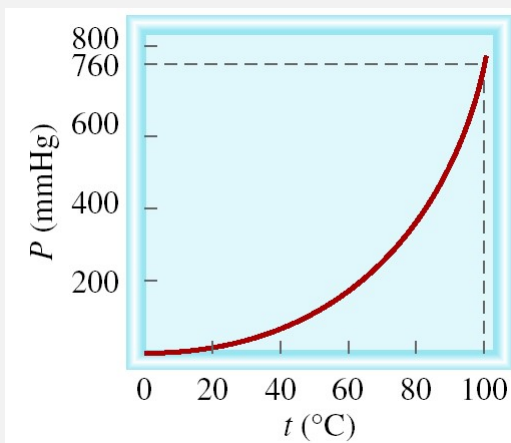


TABLE 5.3
Pressure of Water Vapor
at Various Temperatures

Temperature (°C)	Water Vapor Pressure (mmHg)
0	4.58
5	6.54
10	9.21
15	12.79
20	17.54
25	23.76
30	31.82
35	42.18
40	55.32
45	71.88
50	92.51
55	118.04
60	149.38
65	187.54
70	233.7
75	289.1
80	355.1
85	433.6
90	525.76
95	633.90
100	760.00

Dalton's law of Partial Pressures

EXAMPLE 5.15

Oxygen gas generated by the decomposition of potassium chlorate is collected as shown in Figure 5.15. 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.

Strategy To solve for the mass of O₂ generated, we must first calculate the partial pressure of O₂ in the mixture. What gas law do we need? How do we convert pressure of O₂ gas to mass of O₂ in grams?

Solution From Dalton's law of partial pressures we know that

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

Therefore,

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

From the ideal gas equation we write

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

where m and \mathcal{M} are the mass of O₂ collected and the molar mass of O₂, respectively. Rearranging the equation we obtain

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

Check The density of the oxygen gas is (0.164 g/0.128 L), or 1.28 g/L, which is a reasonable value for gases under atmospheric conditions (see Example 5.8).

Practice Exercise Hydrogen gas generated when calcium metal reacts with water is collected as shown in Figure 5.15. The volume of gas collected at 30°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°C is 31.82 mmHg.