Gases

Chapter 5

Elements that exist as gases at 25^{0}C and 1 atmosphere

1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
H																	**He**
Li	**Be**											**B**	**C**	**N**	**O**	**F**	**Ne**
Na	**Mg**											**Al**	**Si**	**P**	**S**	**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 can not be gases at 25 ^{0}C and 1 atm because of its strong ionic forces

Molecular compounds at 25 ^{0}C and 1 atm varies some are gases CO, HCl and others are liquid or solid $CH_3OH_{(l)}$

No simple rule to help determine if substance is g or l or s

It depends on magnitude of the intermolecular forces among molecules

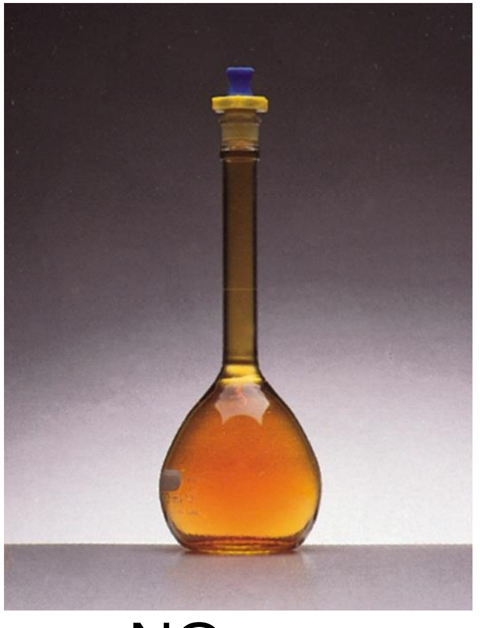
TABLE 5.1 Some Substances Found as Gases at 1 atm and 25°C

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_2O (nitrous oxide)
Xe (xenon)	SO_2 (sulfur dioxide)
Rn (radon)	H_2S (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.

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.



NO_2 gas

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

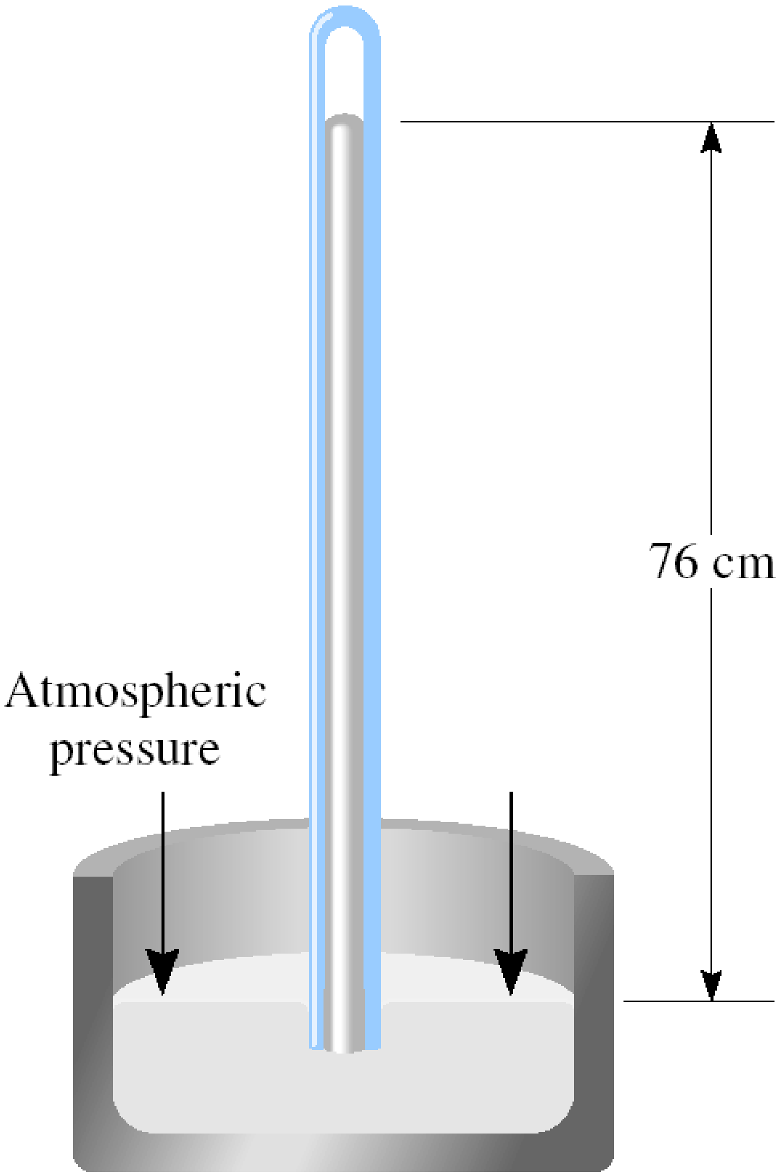
(force = mass x acceleration)

<u>Units of Pressure</u>

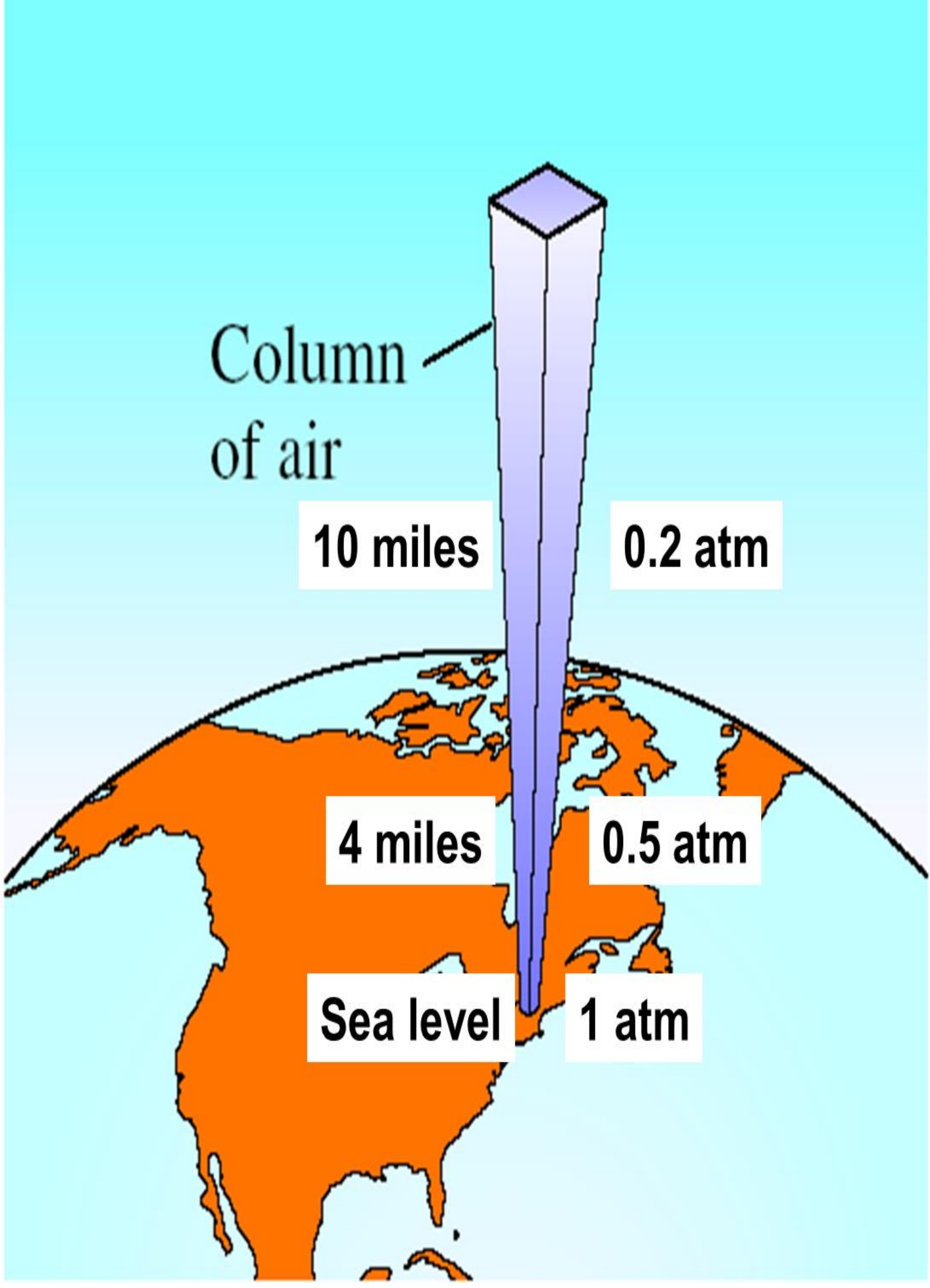
1 pascal (Pa) = 1 N/m^2

1 atm = 760 mmHg = 760 torr

1 atm = 101,325 Pa



- The result of weight of the column of air above it.
- Act on all directions (not down word only)
- Depends on location,T, Weather conditions



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?

Solution The pressure in the cabin is given by

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

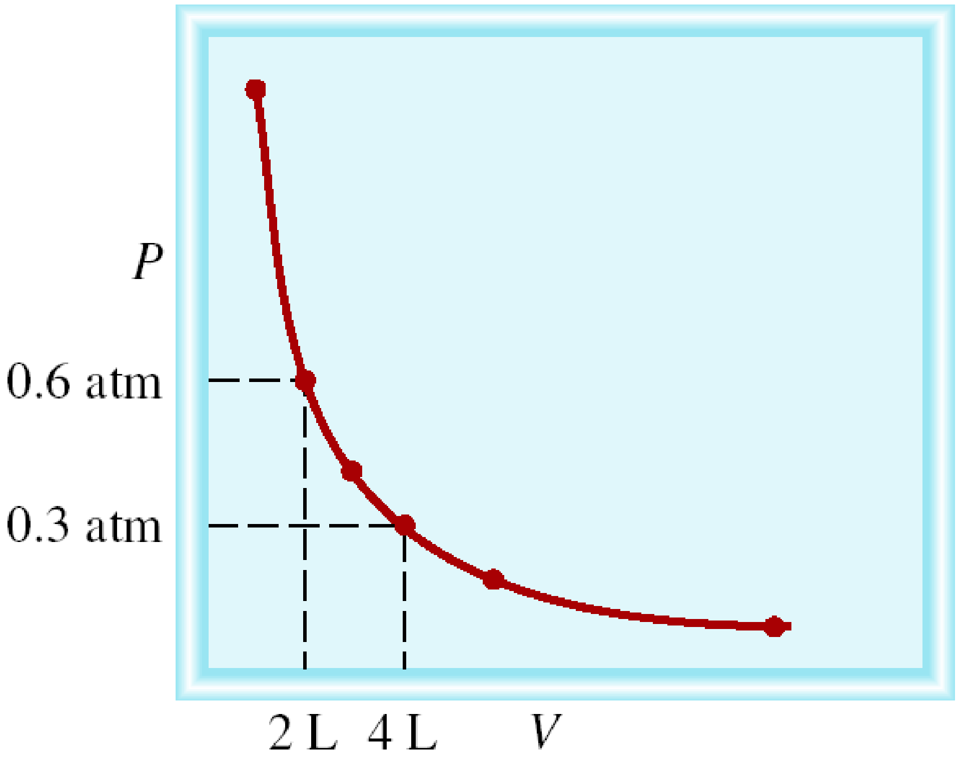
EXAMPLE 5.2

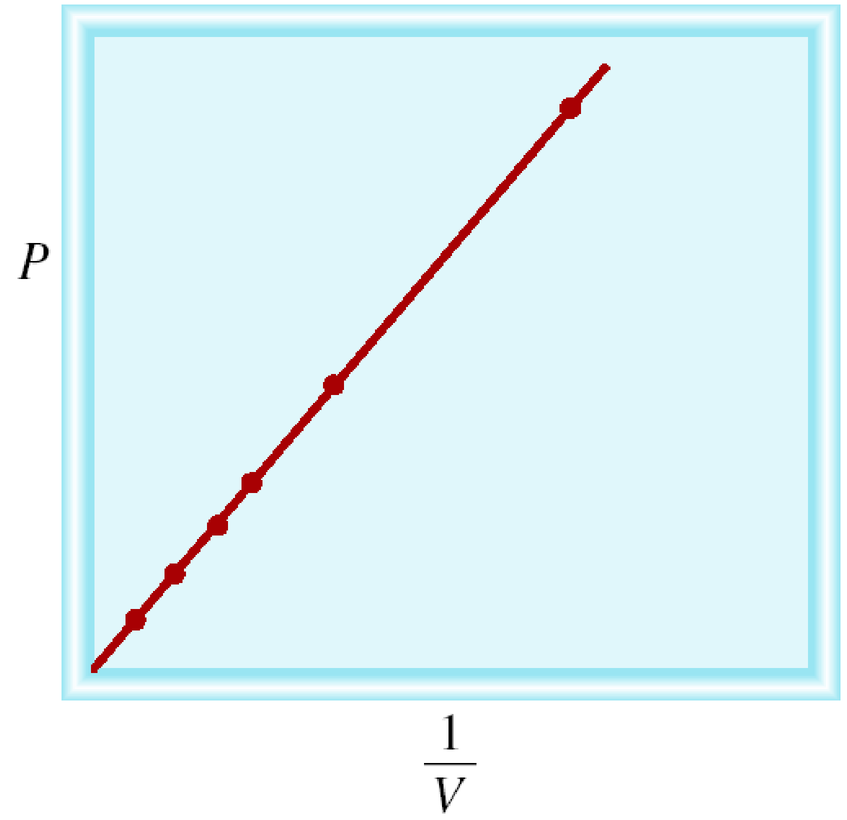
The atmospheric pressure in San Francisco on a certain day was 732 mmHg. What was the pressure in kPa?

Solution The pressure in kPa is

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

Boyle's Law





$P \alpha 1/V$

$P \times V =$ constant

$P_1 \times V_1 = P_2 \times V_2$

Constant temperature
Constant amount of gas

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?

$$P \times V = constant$$

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

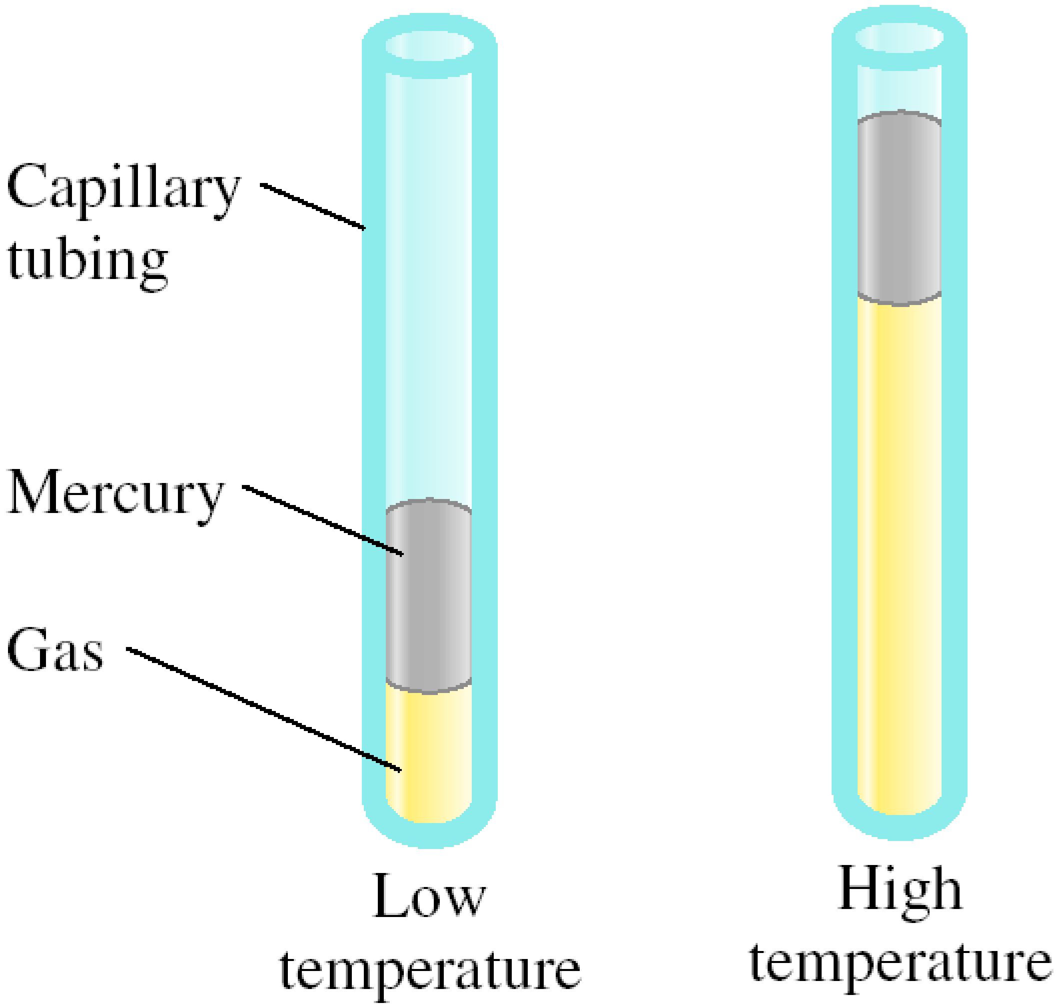
$P_1 = 726$ mmHg $\qquad P_2 = ?$

$V_1 = 946$ mL $\qquad V_2 = 154$ mL

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

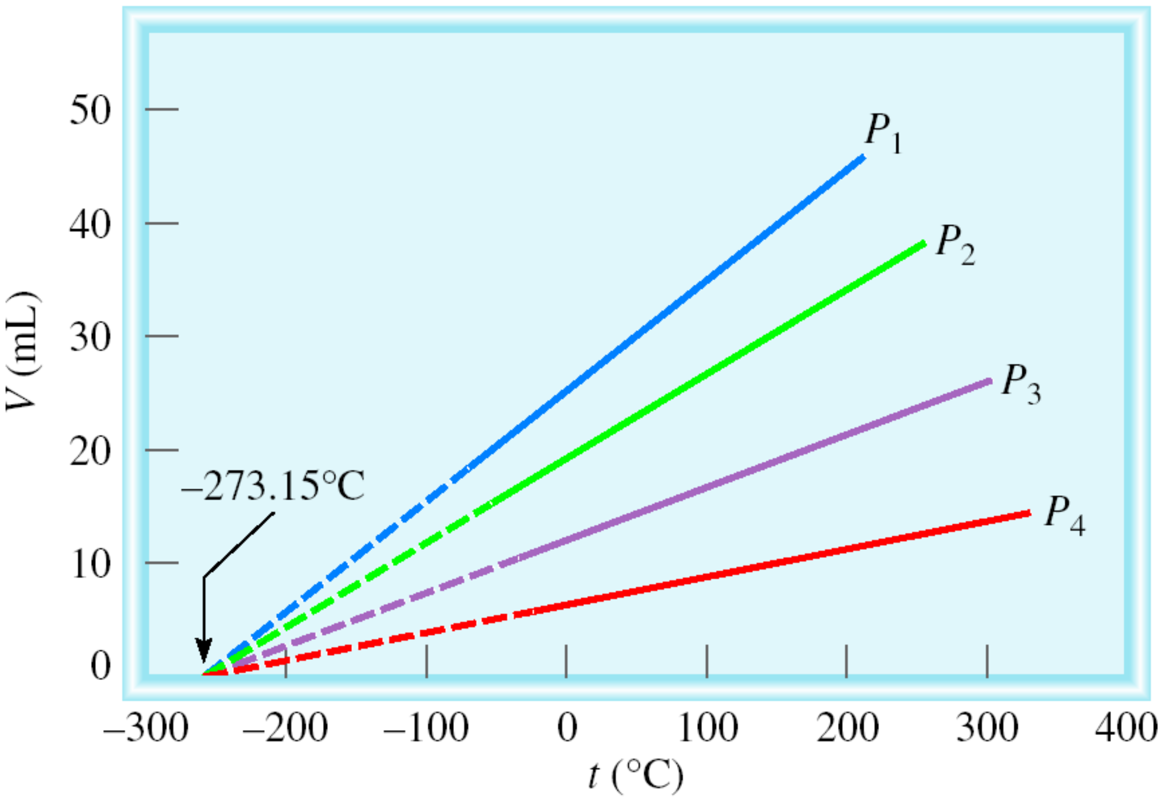
Charles' & Gay-Lussac's Law

Variation in Gas Volume with Temperature at Constant Pressure



As T increases V increases

Variation of Gas Volume with Temperature at Constant Pressure



Charles' & Gay-Lussac's Law

$V \alpha\ T$

V = constant x T

$V_1/T_1 = V_2/T_2$

Temperature **must** be in Kelvin

T (K) = t (^{0}C) + 273.15

A sample of carbon monoxide gas occupies 3.20 L at 125 ^{0}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 L $\quad\quad$ V_2 = 1.54 L

T_1 = 398.15 K $\quad\quad$ T_2 = ?

T_1 = 125 (^{0}C) + 273.15 (K) = 398.15 K

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

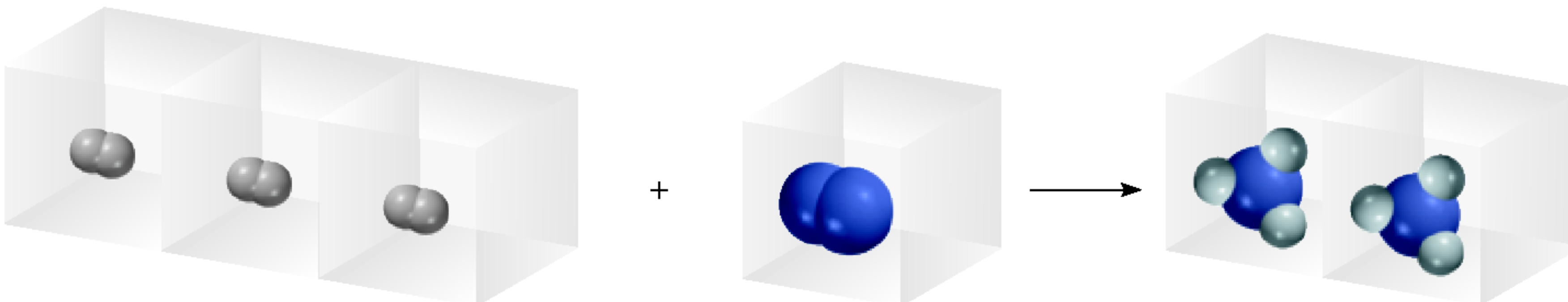
Avogadro's Law

$V \propto$ number of moles (n)

V = constant x n

Constant temperature
Constant pressure

$$V_1 / n_1 = V_2 / n_2$$



$3H_2(g)$	+	$N_2(g)$	$\longrightarrow$	$2NH_3(g)$
3 molecules	+	1 molecule	$\longrightarrow$	2 molecules
3 moles	+	1 mole	$\longrightarrow$	2 moles
3 volumes	+	1 volume	$\longrightarrow$	2 volumes

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?

$$4NH_3 + 5O_2 \longrightarrow 4NO + 6H_2O$$

$$1 \text{ mole } NH_3 \longrightarrow 1 \text{ mole } NO$$

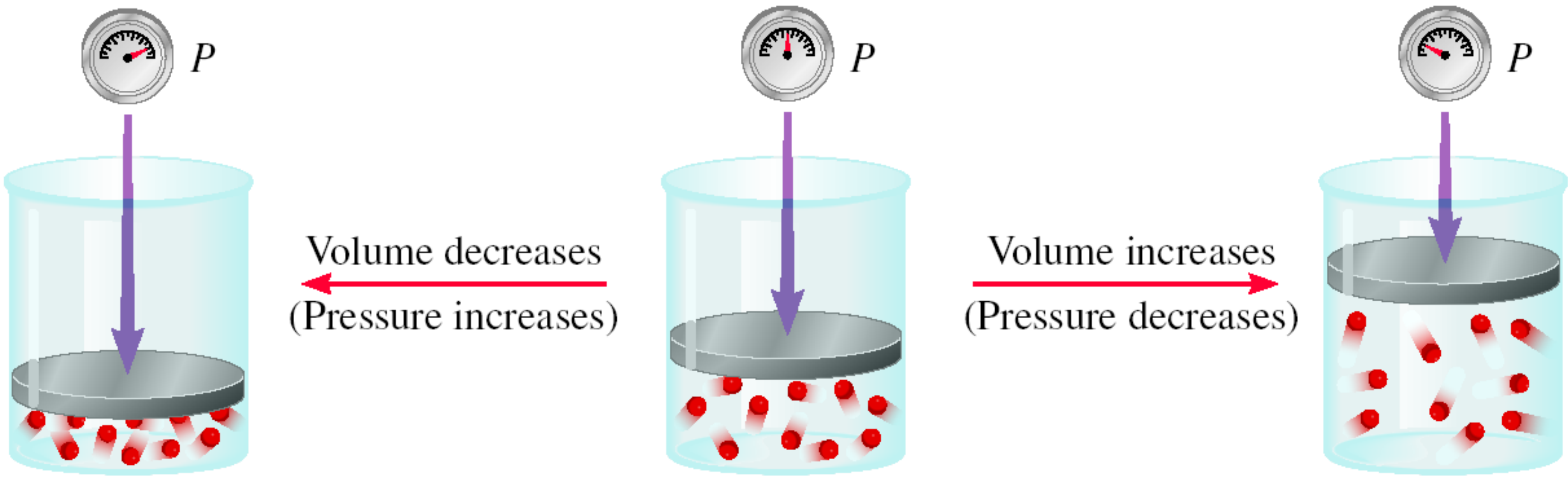
At constant *T* and *P*

$$1 \text{ volume } NH_3 \longrightarrow 1 \text{ volume } NO$$

Summary of Gas Laws

Boyle's Law

Increasing or decreasing the volume of a gas at a constant temperature

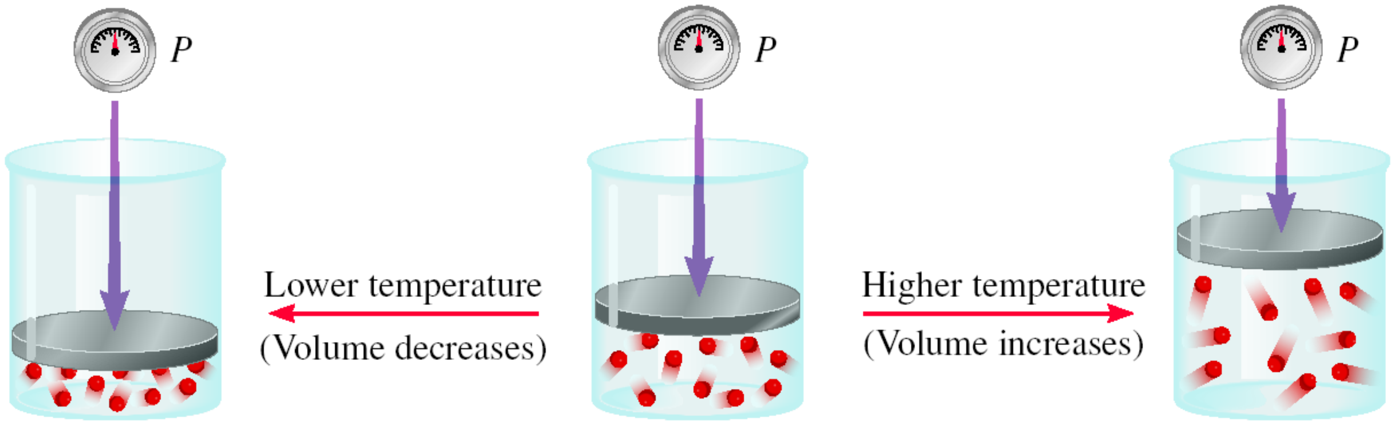


Boyle's Law

$P = (nRT)\frac{1}{V}$ nRT is constant

Charles Law

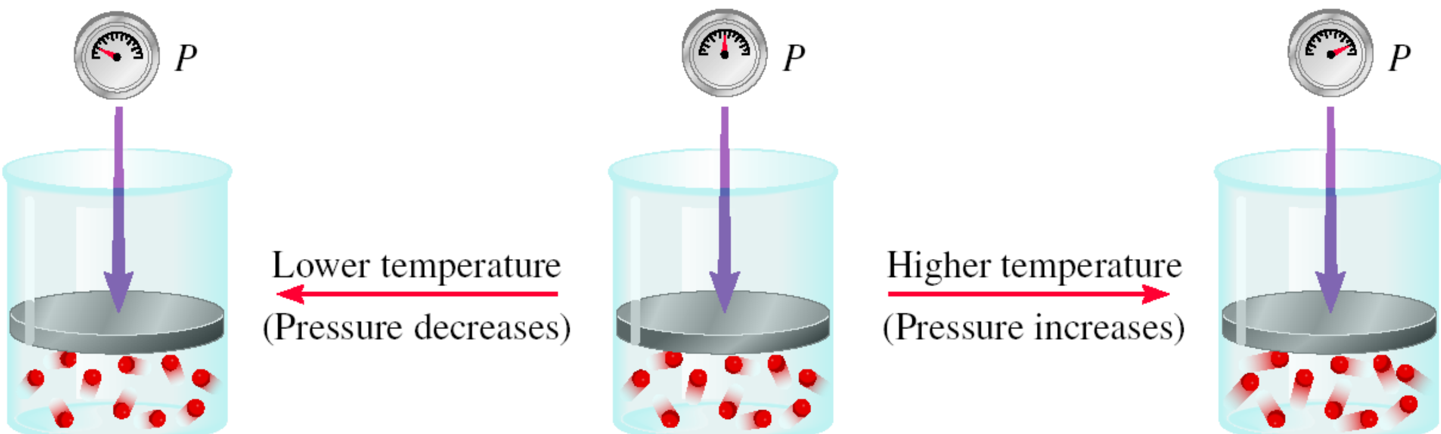
Heating or cooling a gas at constant pressure



Charles's Law

$$V = \left(\frac{nR}{P}\right) T \quad \frac{nR}{P} \text{ is constant}$$

Heating or cooling a gas at constant volume

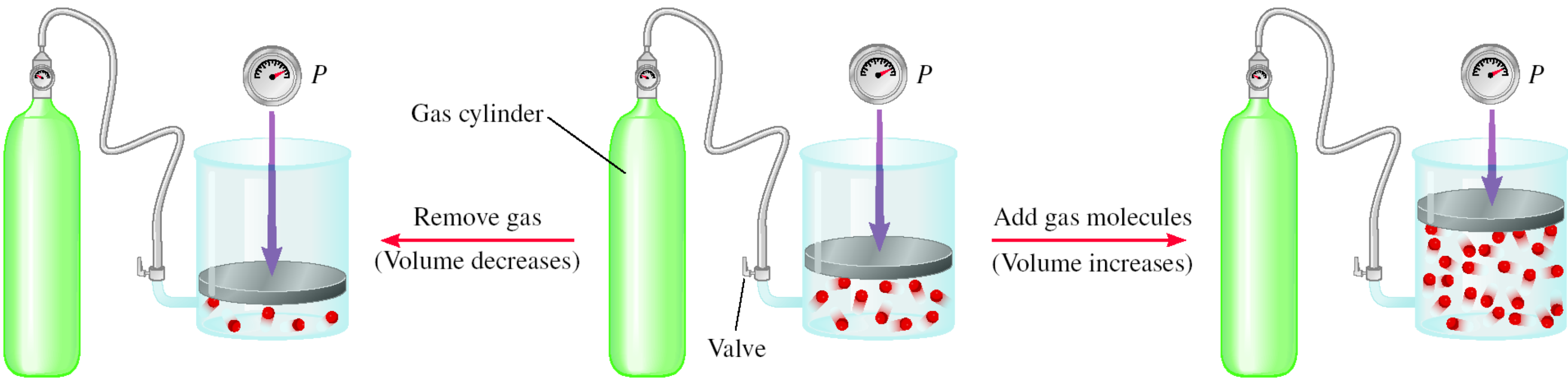


Charles's Law

$$P = \left(\frac{nR}{V}\right) T \quad \frac{nR}{V} \text{ is constant}$$

Avogadro’s Law

Dependence of volume on amount of gas at constant temperature and pressure



Avogadro’s Law

$$V = \left(\frac{RT}{P}\right) n \qquad \frac{RT}{P} \text{ is constant}$$

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*)

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

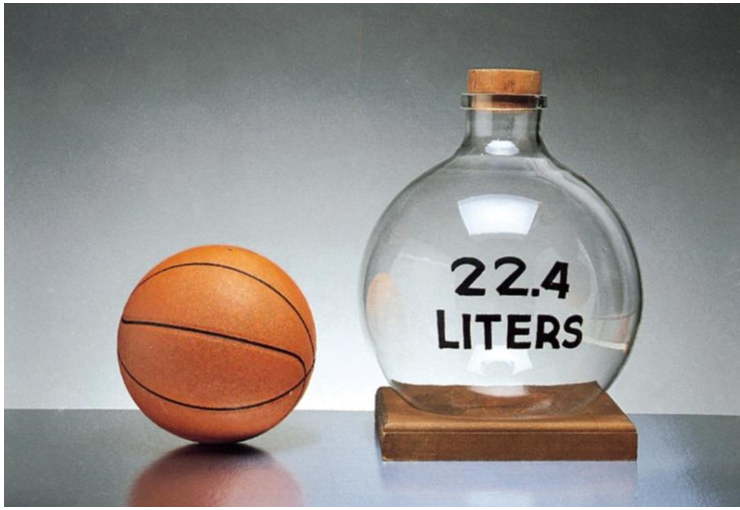
$$V = \text{constant} \times \frac{nT}{P} = R\frac{nT}{P}$$

R is the **gas constant**

$$\boxed{PV = nRT}$$

The conditions 0 ^{0}C 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.



$$PV = nRT$$

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

$R = 0.082057$ L • atm / (mol • K)

It is a hypothetical gas which follows ideal gas equation

Ideal gas:

- don't attract or repel one another
- It's volume is negligible compare to the volume of the container

If all variables changes we use:

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

In general $n_1 = n_2$ thus

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

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

$T = 0\ ^0C = 273.15\ K$

$P = 1\ atm$

$$PV = nRT$$

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

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

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

$$V = 30.7\ L$$

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 ^{0}C is heated to 85 ^{0}C at constant volume. What is the final pressure of argon in the lightbulb (in atm)?

$PV = nRT$ $\qquad$ n, V and R **are constant**

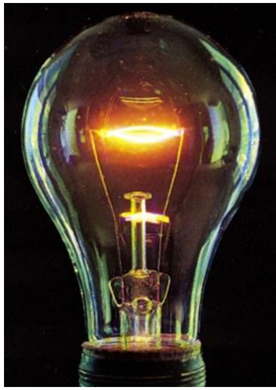
$$\frac{nR}{V} = \frac{P}{T} = \text{constant}$$

$P_1 = 1.20$ atm $\qquad$ $P_2 = ?$

$T_1 = 291$ K $\qquad$ $T_2 = 358$ K

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

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



EXAMPLE 5.3

Sulfur hexafluoride (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°C.

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

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

EXAMPLE 5.4

Calculate the volume (in liters) occupied by 7.40 g of NH_3 at STP.

Solution Recognizing that 1 mole of an ideal gas occupies 22.41 L at STP and using the molar mass of NH_3 (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 NH_3 is given by

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

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 NH_3 to number of moles of NH_3, and then applying the ideal gas equation ($V = nRT/P$). Try it.

Examples

1) 20.8 g of CH_4 gas was confined in 5.200 L vessel at 50 ^0C. Calculate the pressure exerted by the gas?

$\mathcal{M}_{CH_4}$ =16.04 g mol^{-1} 6.529 atm

2) 1.05 L balloon at 25^0C , Calculate it's volume in a summer day at 50 ^0C?

1.138 L

3) Gas volume is 2.31L at 1 atm, Calculate it's pressure in mmHg when it's volume becomes 7.32 L?

239.8mmHg

Density (d) Calculations

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

m is the mass of the gas in g

$\mathcal{M}$ is the molar mass of the gas

Molar Mass ($\mathcal{M}$) of a Gaseous Substance

$$\mathcal{M} = \frac{dRT}{P}$$

d is the density of the gas in g/L

A 2.10-L vessel contains 4.65 g of a gas at 1.00 atm and 27.0 ^{0}C. What is the molar mass of the gas?

$$\mathcal{M} = \frac{dRT}{P}$$

$$d = \frac{m}{V} = \frac{4.65 \text{ g}}{2.10 \text{ L}} = 2.21 \frac{\text{g}}{\text{L}}$$

$$\mathcal{M} = \frac{2.21 \frac{\text{g}}{\cancel{\text{L}}} \text{ x } 0.0821 \frac{\cancel{\text{L}}\bullet\cancel{\text{atm}}}{\text{mol}\bullet\cancel{\text{K}}} \text{ x } 300.15 \cancel{\text{K}}}{1 \cancel{\text{atm}}}$$

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

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.

Solution From Equation (5.12)

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

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

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

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.

Solution

$$n_{Si} = 33.0 \cancel{\text{g Si}} \times \frac{1 \text{ mol Si}}{28.09 \cancel{\text{g Si}}} = 1.17 \text{ mol Si}$$

$$n_{F} = 67.0 \cancel{\text{g F}} \times \frac{1 \text{ mol F}}{19.00 \cancel{\text{g F}}} = 3.53 \text{ mol F}$$

Therefore, the empirical formula is $Si_{1.17}F_{3.53}$, 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

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

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

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

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

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 $(SiF_3)_2$ or Si_2F_6.

Gas Stoichiometry

Amount of reactant (grams or volume) → Moles of reactant → Moles of product → Amount of product (grams or volume)

What is the volume of CO_2 produced at 37 ^{0}C and 1.00 atm when 5.60 g of glucose are used up in the reaction:

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

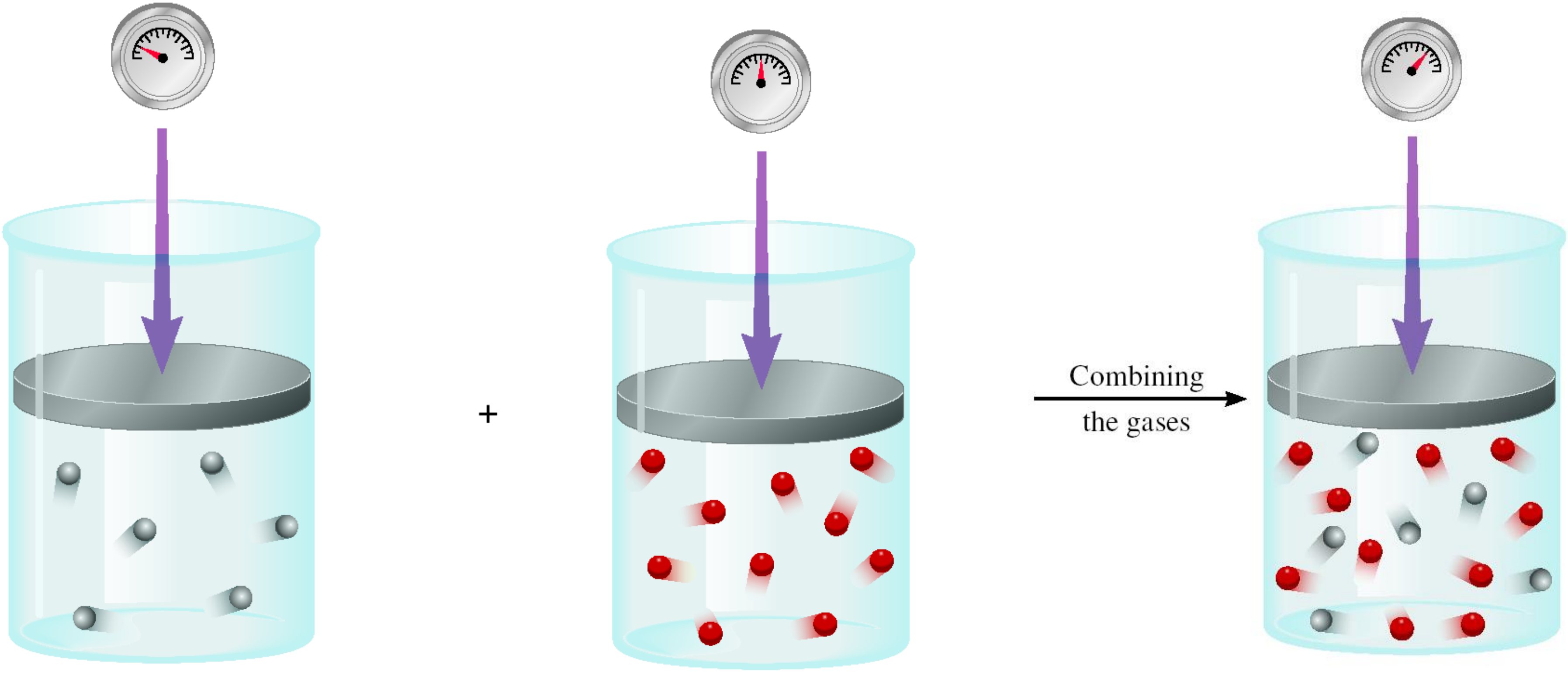
$$\text{g } C_6H_{12}O_6 \longrightarrow \text{mol } C_6H_{12}O_6 \longrightarrow \text{mol } CO_2 \longrightarrow V\ CO_2$$

$$5.60\ \cancel{\text{g } C_6H_{12}O_6} \times \frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{180\ \cancel{\text{g } C_6H_{12}O_6}} \times \frac{6\text{ mol } CO_2}{1\ \cancel{\text{mol } C_6H_{12}O_6}} = 0.187\text{ mol } CO_2$$

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

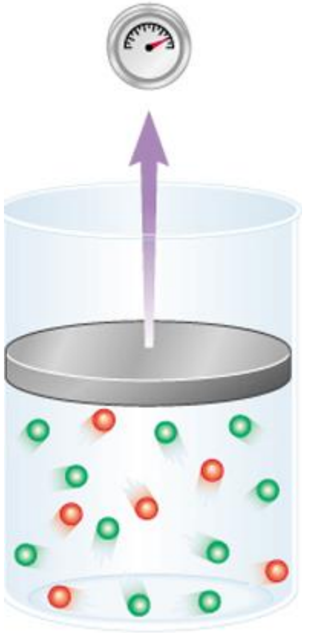
Dalton's Law of Partial Pressures

V and T are **constant**



P_1 P_2 $P_{total} = P_1 + P_2$

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



$$P_A = \frac{n_A RT}{V}$$

n_A is the number of moles of A

$$P_B = \frac{n_B RT}{V}$$

n_B is the number of moles of B

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

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

$$\boxed{P_i = X_i P_T}$$

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

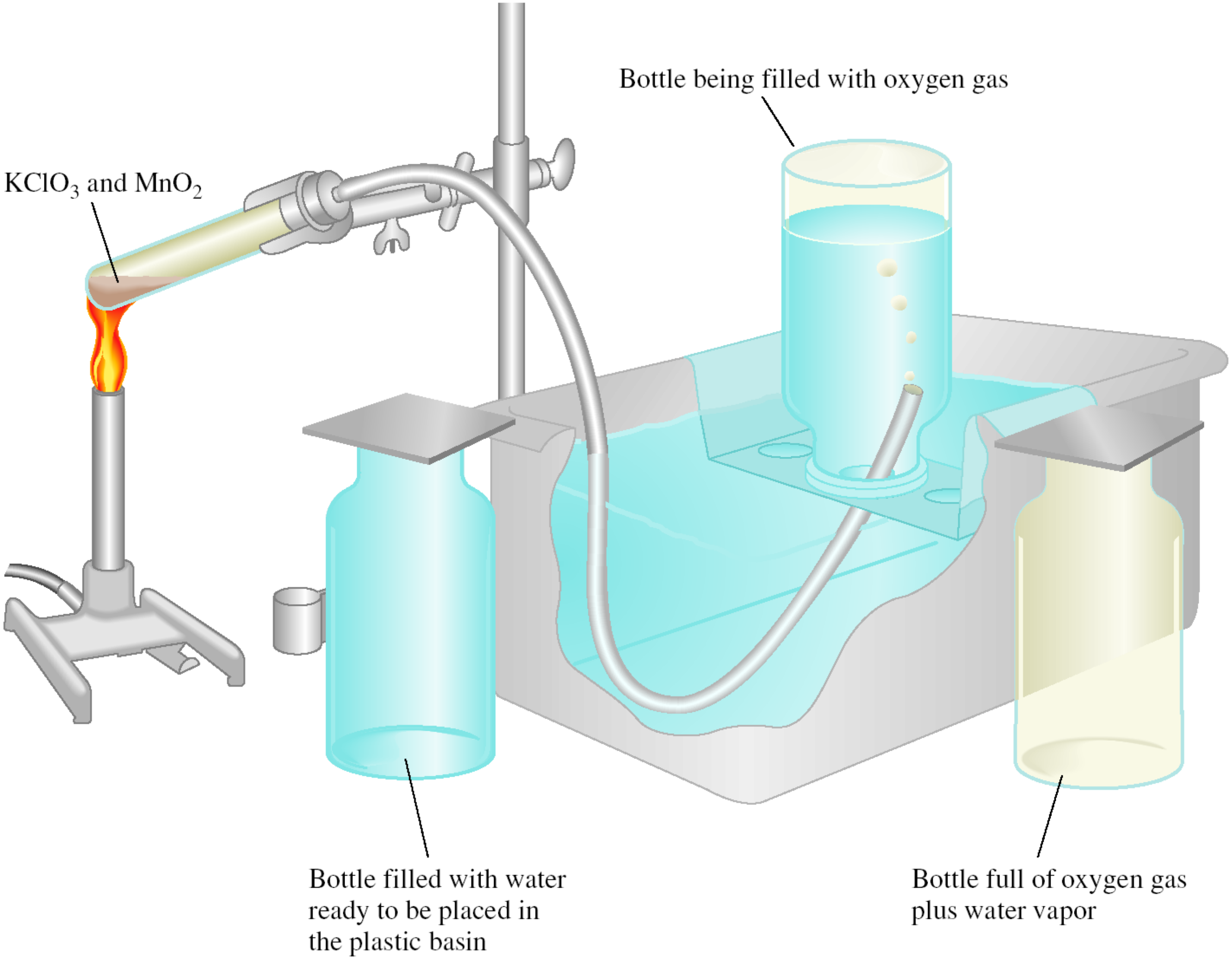
A sample of natural gas contains 8.24 moles of CH_4, 0.421 moles of C_2H_6, and 0.116 moles of C_3H_8. If the total pressure of the gases is 1.37 atm, what is the partial pressure of propane (C_3H_8)?

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

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

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

Collecting a Gas over Water



$$2KClO_3\ (s) \longrightarrow 2KCl\ (s) + 3O_2\ (g)$$

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

Vapor of Water and Temperature

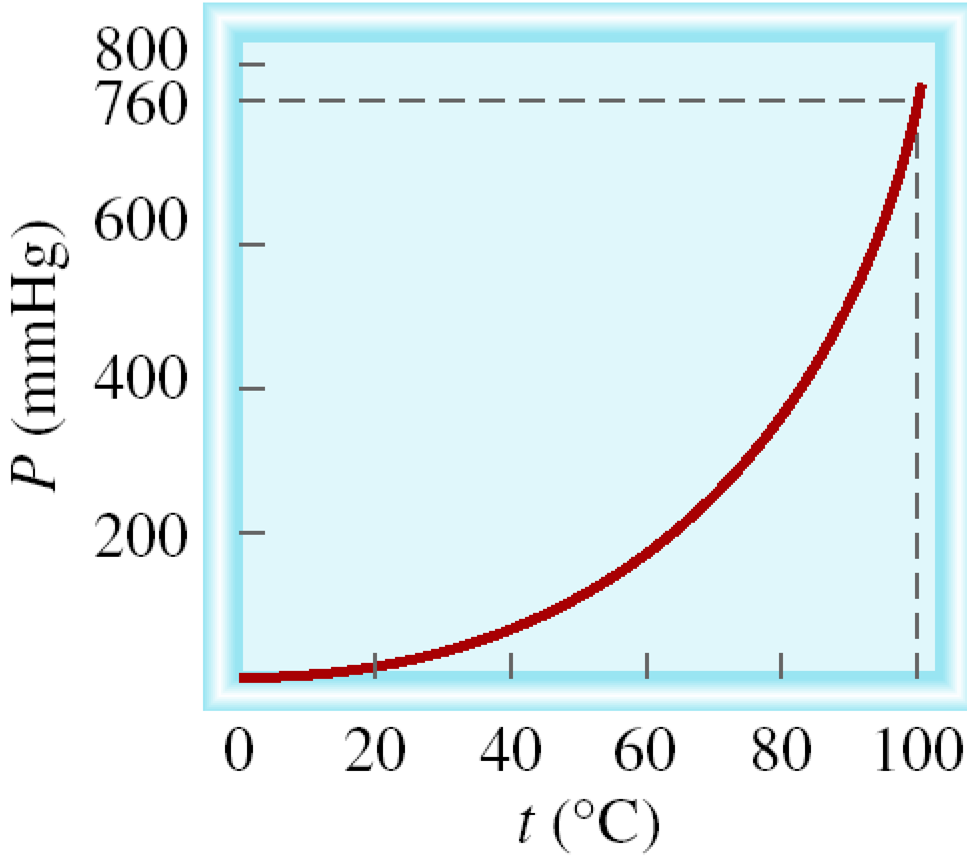


TABLE 5.3

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

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.

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 O_2 collected and the molar mass of O_2, respectively. Rearranging the equation we obtain

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

Calculate the mass of $Zn_{(s)}$ used to produce $H_{2(g)}$ over water at 25.0^{0}C in a 7.80L vessel and pressure 0.980 atm knowing that $p_{H_2O} = 23.8\ mmHg$ according to the following equation:

$$Zn_{(s)} + 2HCl_{(g)} \longrightarrow ZnCl_{2\,(aq)} + H_{2\,(g)}$$

$$m_{Zn} = 19.8\ g$$

Kinetic Molecular Theory of Gases

1. A gas is composed of molecules that are separated from each other by distances far greater than their own dimensions. The molecules can be considered to be *points*; that is, they possess mass but have negligible volume.
2. Gas molecules are in constant motion in random directions, and they frequently collide with one another. Collisions among molecules are perfectly elastic.
3. Gas molecules exert neither attractive nor repulsive forces on one another.
4. The average kinetic energy of the molecules is proportional to the temperature of the gas in kelvins. Any two gases at the same temperature will have the same average kinetic energy

$$\overline{KE} = \tfrac{1}{2} m\overline{u^2}$$

Kinetic theory of gases and …

- Compressibility of Gases
- Boyle's Law

  $P \alpha$ collision rate with wall

  Collision rate α number density

  Number density α $1/V$

  $P \alpha 1/V$

- Charles' Law

  $P \alpha$ collision rate with wall

  Collision rate α average kinetic energy of gas molecules

  Average kinetic energy α T

  $P \alpha T$

Kinetic theory of gases and …

- Avogadro's Law

  $P \alpha$ collision rate with wall

  Collision rate α number density

  Number density α n

  $P \alpha n$

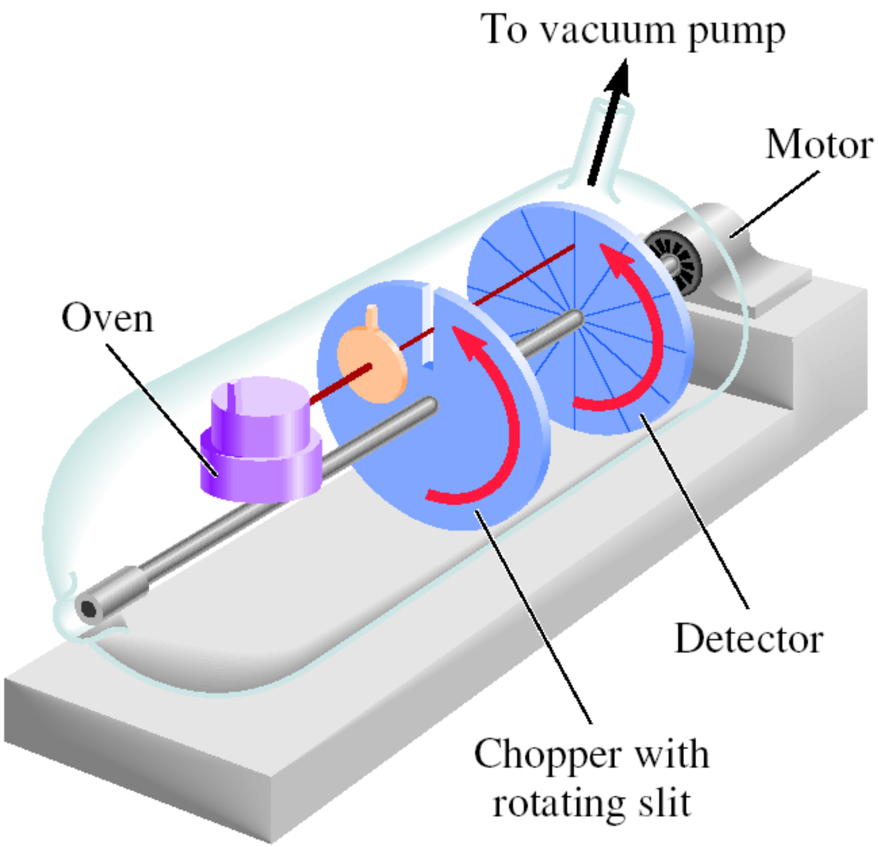
- Dalton's Law of Partial Pressures

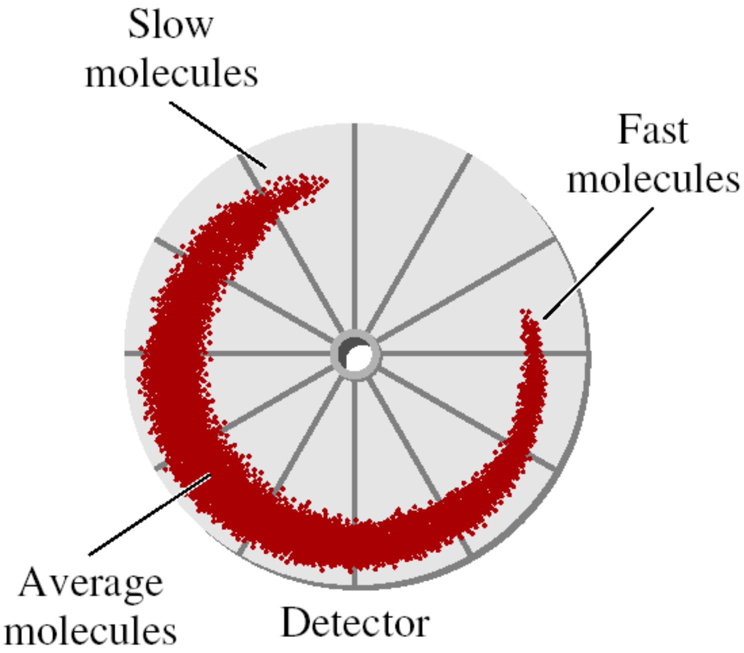
  Molecules do not attract or repel one another

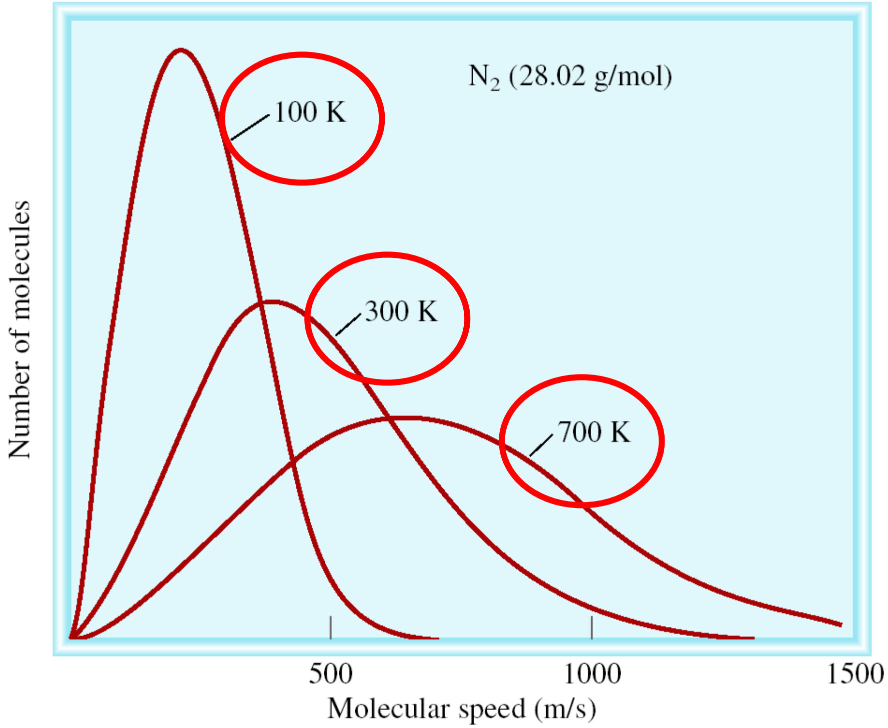
  P exerted by one type of molecule is unaffected by the presence of another gas

  $P_{\text{total}} = \Sigma P_i$

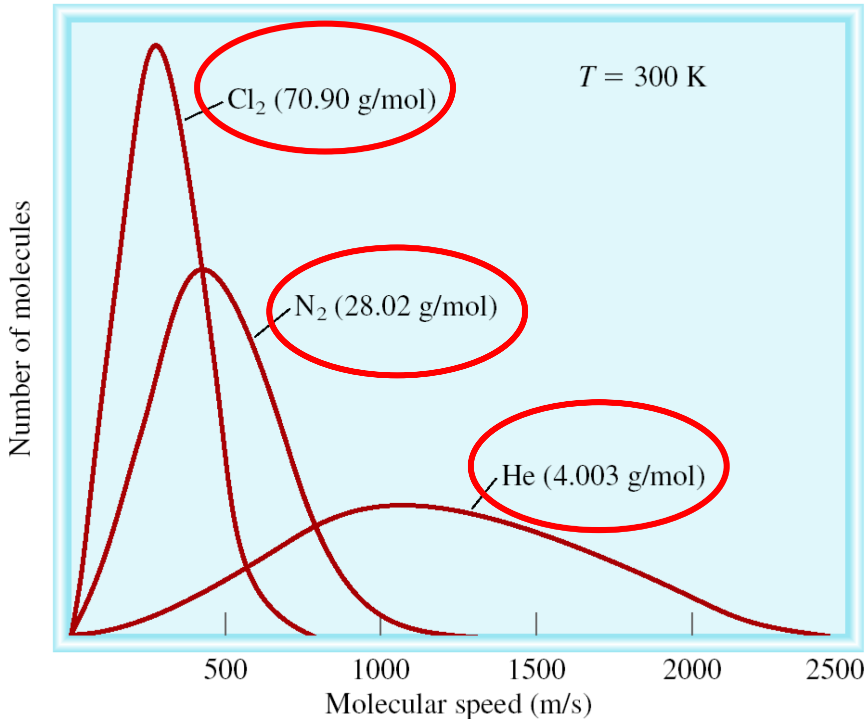
Apparatus for Studying Molecular Speed Distributiona







The distribution of speeds
of three different gases
at the same temperature



The distribution of speeds
for nitrogen gas molecules
at three different temperatures

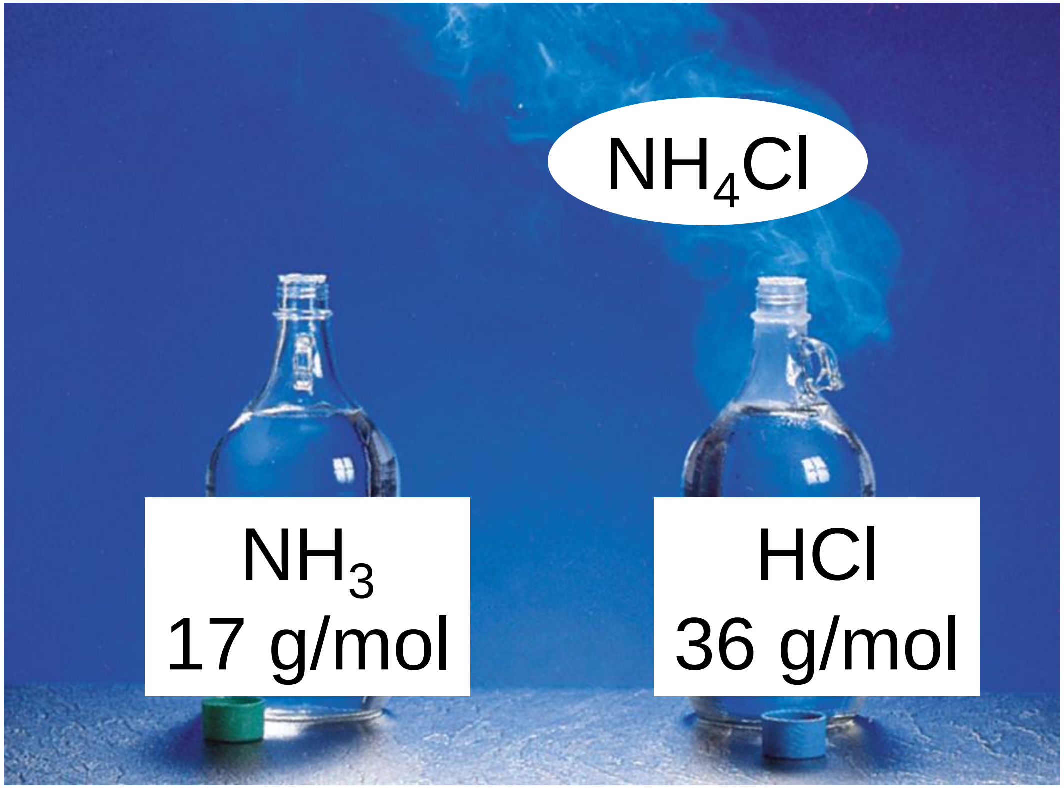
$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

The **R** must be 8.314 $J.mol^{-1}.K^{-1}$
The **M** must be in kg/mol units

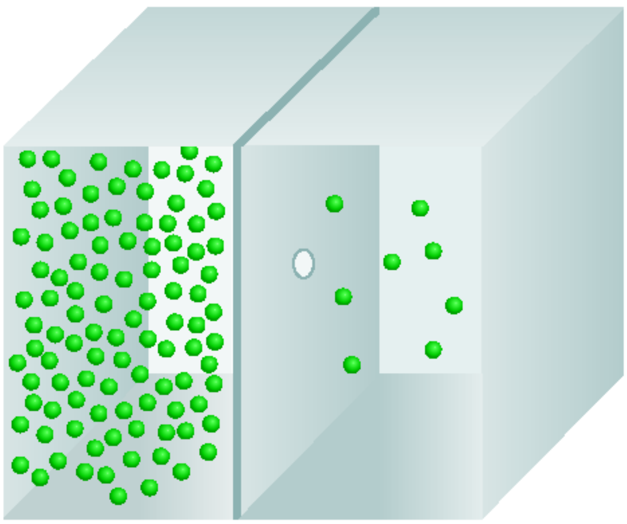
Gas diffusion is the gradual mixing of molecules of one gas with molecules of another by virtue of their kinetic properties.

molecular path

$$\frac{r_1}{r_2} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$



Gas effusion is the is the process by which gas under pressure escapes from one compartment of a container to another by passing through a small opening.



$$\frac{r_1}{r_2} = \frac{t_2}{t_1} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$

Nickel forms a gaseous compound of the formula $Ni(CO)_x$ What is the value of x given that under the same conditions methane (CH_4) effuses 3.3 times faster than the compound?

$r_1 = 3.3 \times r_2$

$$\mathcal{M}_2 = \left(\frac{r_1}{r_2}\right)^2 \times \mathcal{M}_1 = (3.3)^2 \times 16 = 174.2$$

$\mathcal{M}_1 = 16$ g/mol

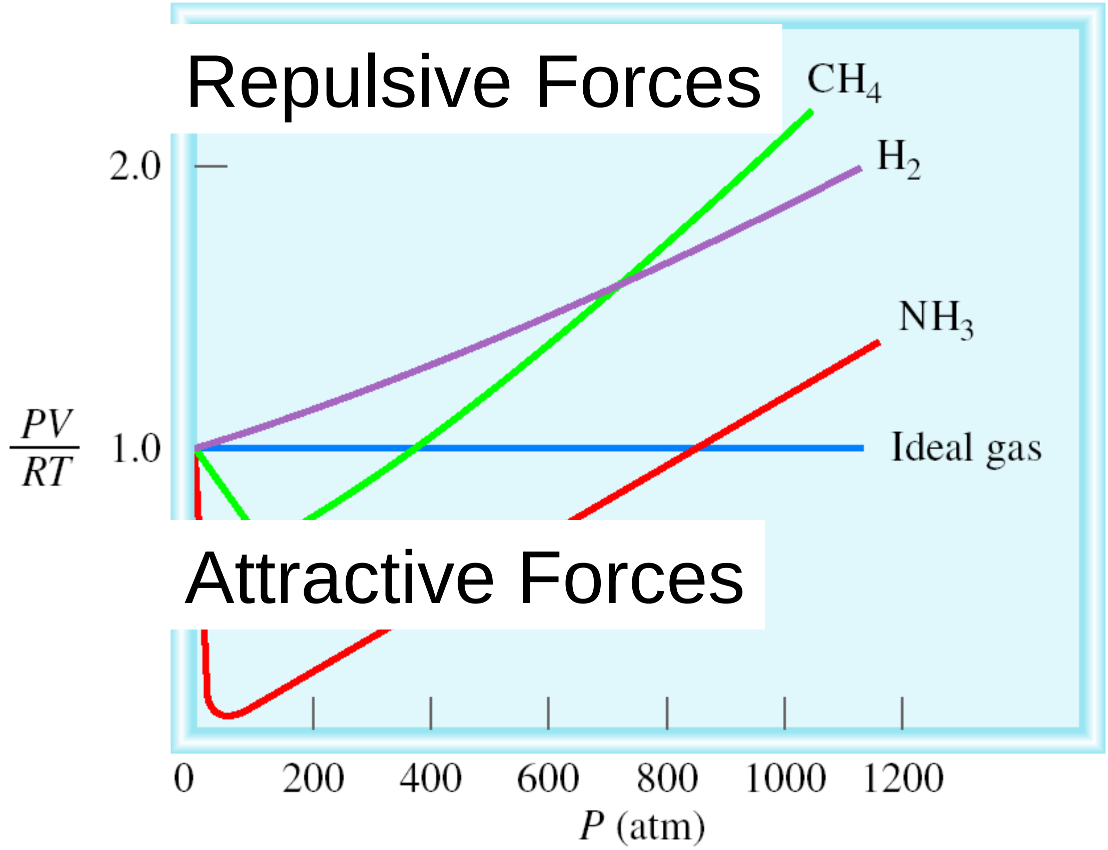
$58.7 + x \cdot 28 = 174.2$ $\quad x = 4.1 \sim 4$

Deviations from Ideal Behavior

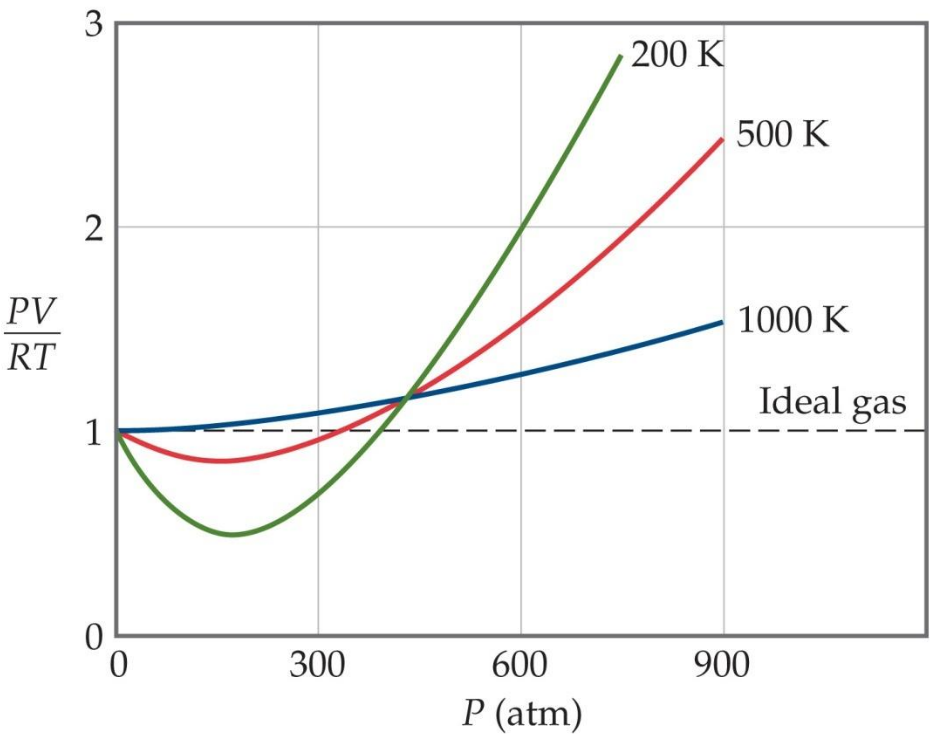
<u>1 mole of ideal gas</u>

$$PV = nRT$$

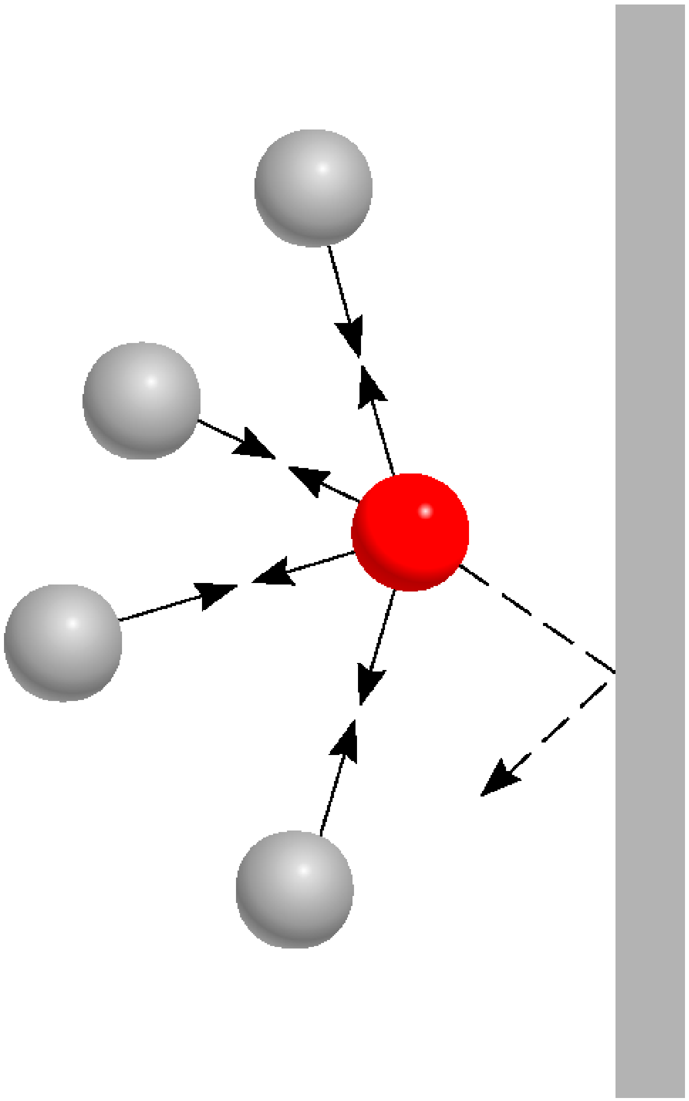
$$n = \frac{PV}{RT} = 1.0$$



Even the same gas will show wildly different behavior under high pressure at different temperatures.



Effect of intermolecular forces on the pressure exerted by a gas.



Van der Waals equation

nonideal gas

$$\underbrace{\left(P + \frac{an^2}{V^2}\right)}_{\text{corrected pressure}}\underbrace{(V - nb)}_{\text{corrected volume}} = nRT$$

TABLE 5.4

van der Waals Constants of Some Common Gases

Gas	a $\left(\frac{\text{atm} \cdot \text{L}^2}{\text{mol}^2}\right)$	b $\left(\frac{\text{L}}{\text{mol}}\right)$
He	0.034	0.0237
Ne	0.211	0.0171
Ar	1.34	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0266
H_2	0.244	0.0266
N_2	1.39	0.0391
O_2	1.36	0.0318
Cl_2	6.49	0.0562
CO_2	3.59	0.0427
CH_4	2.25	0.0428
CCl_4	20.4	0.138
NH_3	4.17	0.0371
H_2O	5.46	0.0305