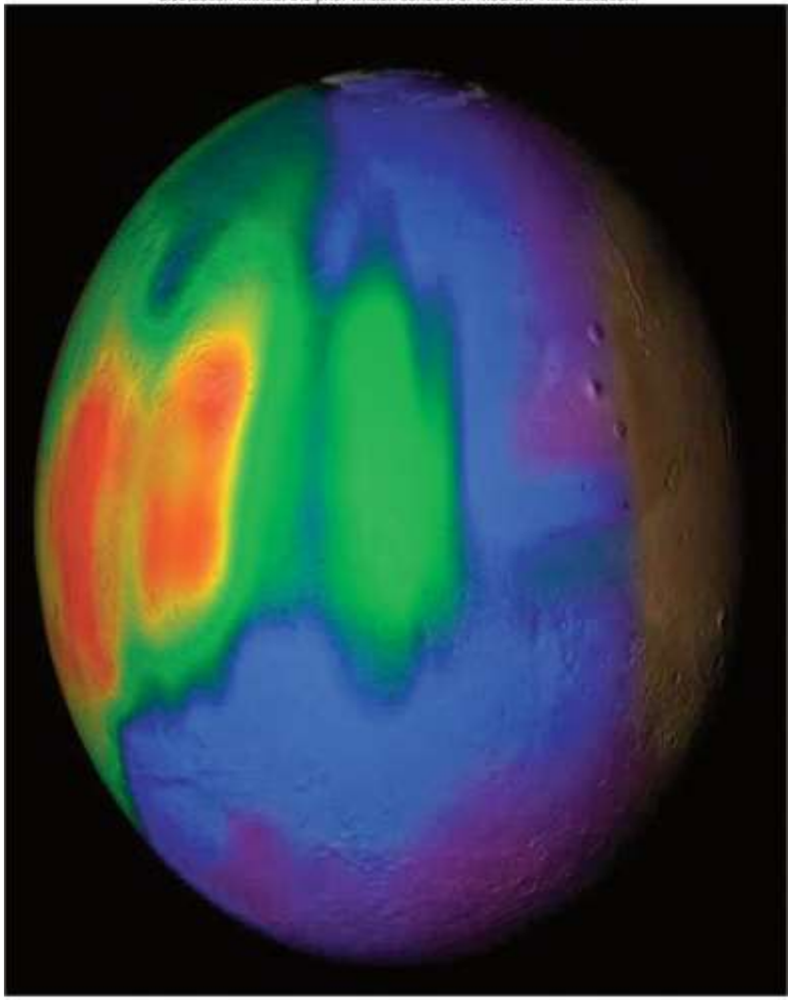
Gases

Chapter 5



NASA

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

1A	2A	3B	4B	5B	6B	7B	8B	8B	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.

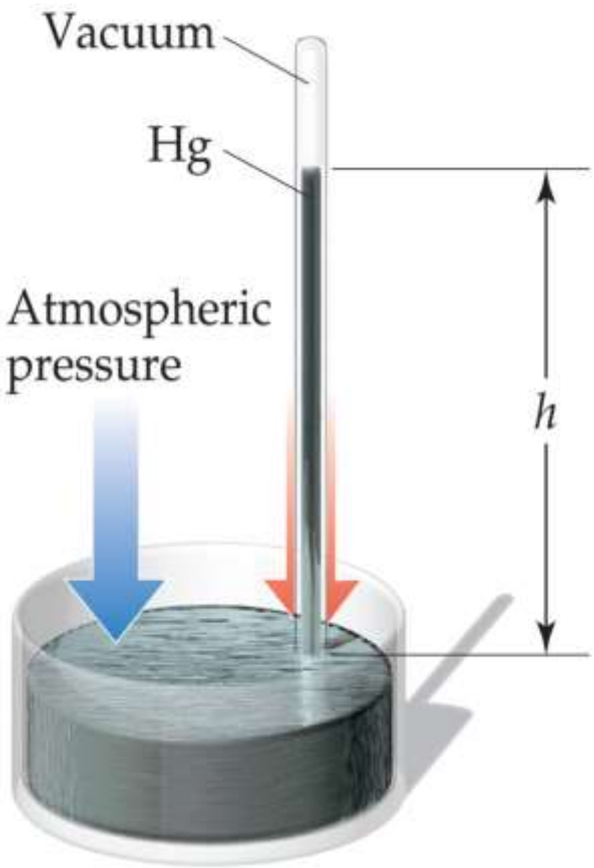
$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

(force = mass x acceleration)

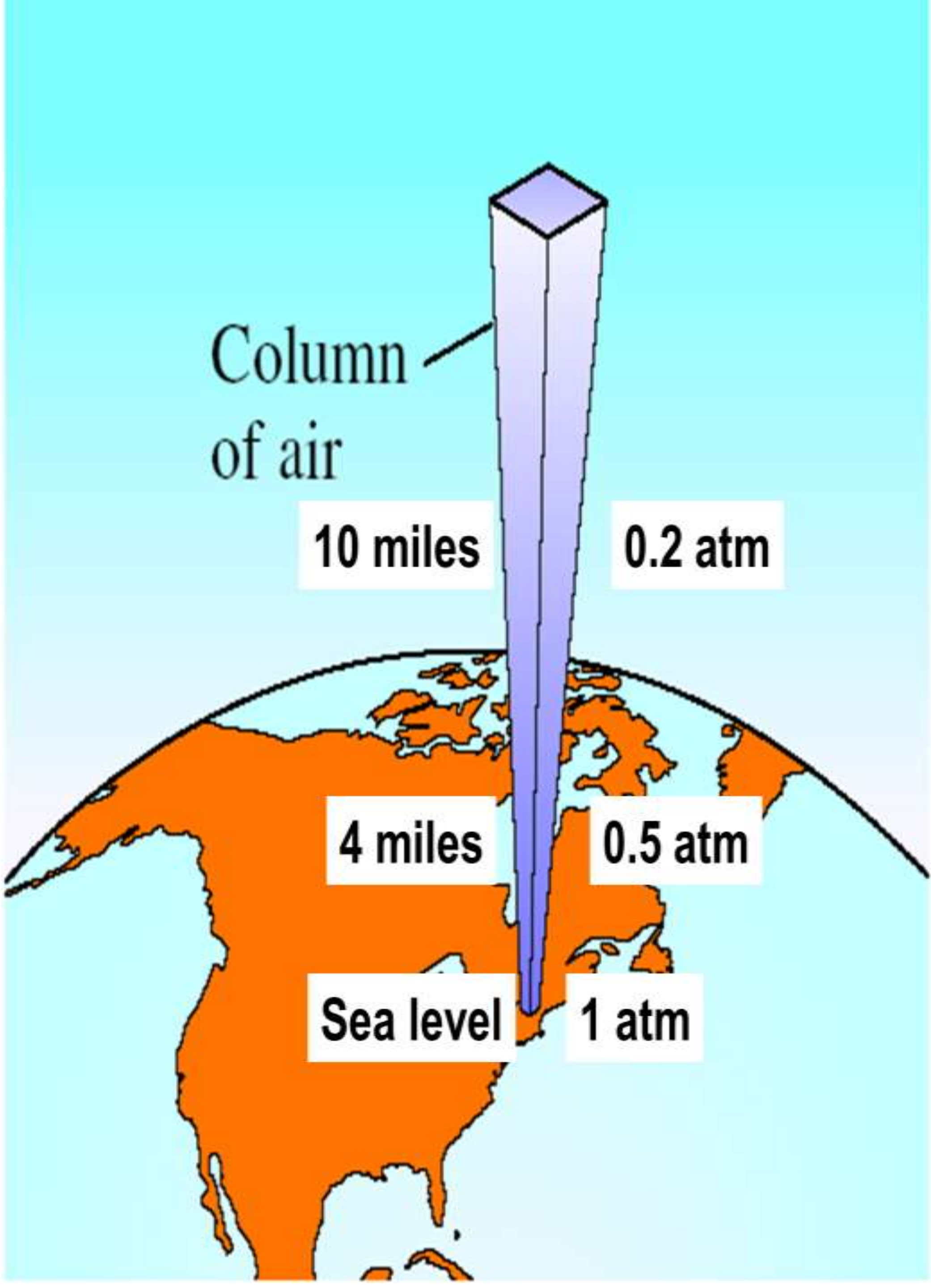
NO_2 gas

Units of Pressure

- Pascals
  - $1\ Pa = 1\ N/m^2$
- Bar
  - $1\ bar = 10^5\ Pa = 100\ kPa$
- mm Hg or torr
  - These units are literally the difference in
  - the heights measured in mm (h) of two
  - connected columns of mercury.
- Normal atmospheric pressure at sea level is
- referred to as standard pressure.
- Atmosphere It is equal to
  - 1.00 atm = 760 torr (760 mm Hg)=101.325 kPa



- 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

What is the pressure in atmospheres in the cabin if the barometer reading is 688 mmHg?

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

Solution

The pressure in the cabin is given by

$$\text{pressure} = 688\ \text{mmHg} \times \frac{1\ \text{atm}}{760\ \text{mmHg}}$$

$$= 0.905\ \text{atm}$$

Example 5.2

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

Solution

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

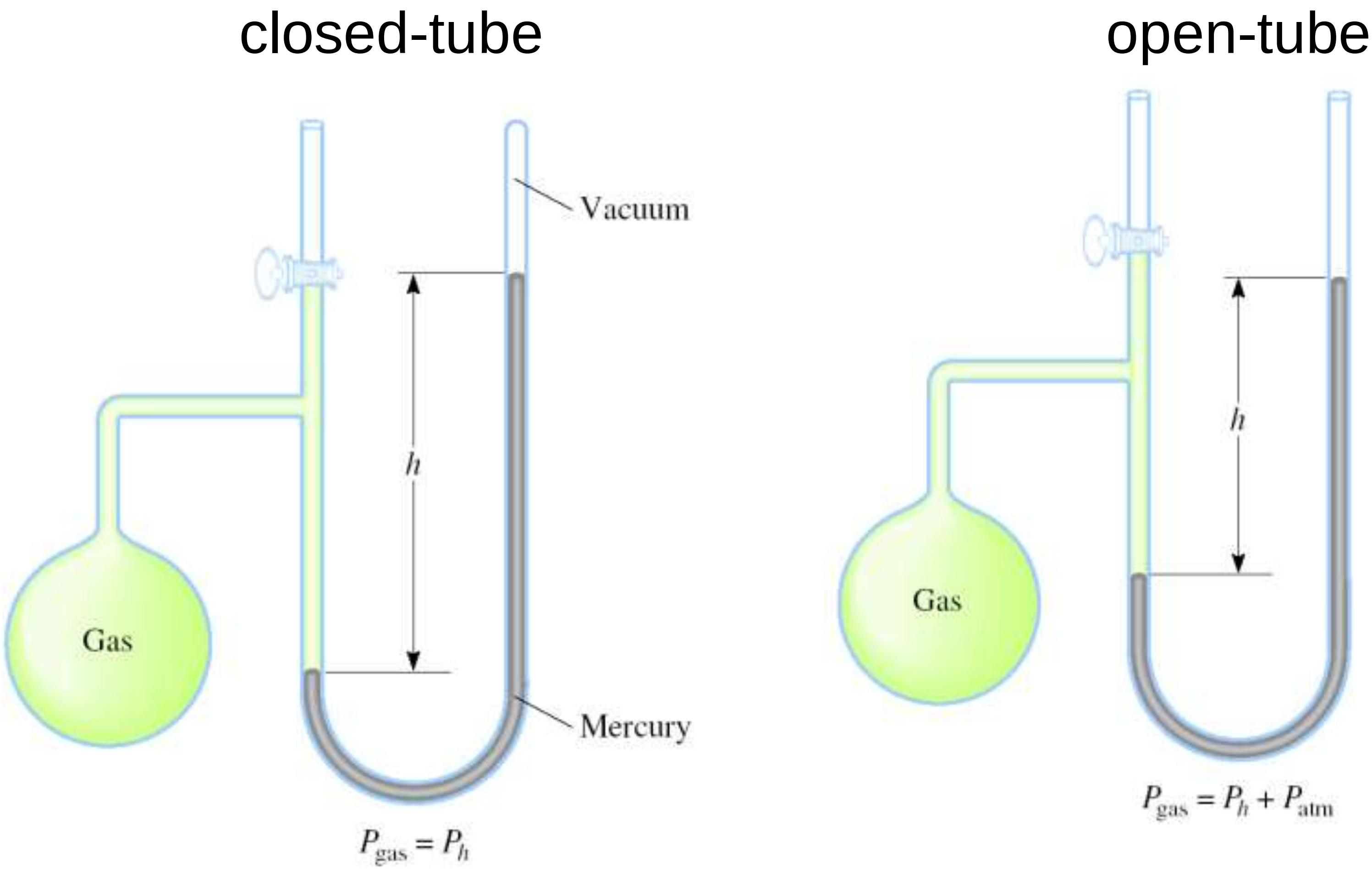
the conversion factor we needs is

The pressure in kPa is

$$\text{pressure} = 732\ \cancel{\text{mmHg}} \times \frac{1.011325 \times 10^5\,\text{Pa}}{760\ \cancel{\text{mmHg}}}$$

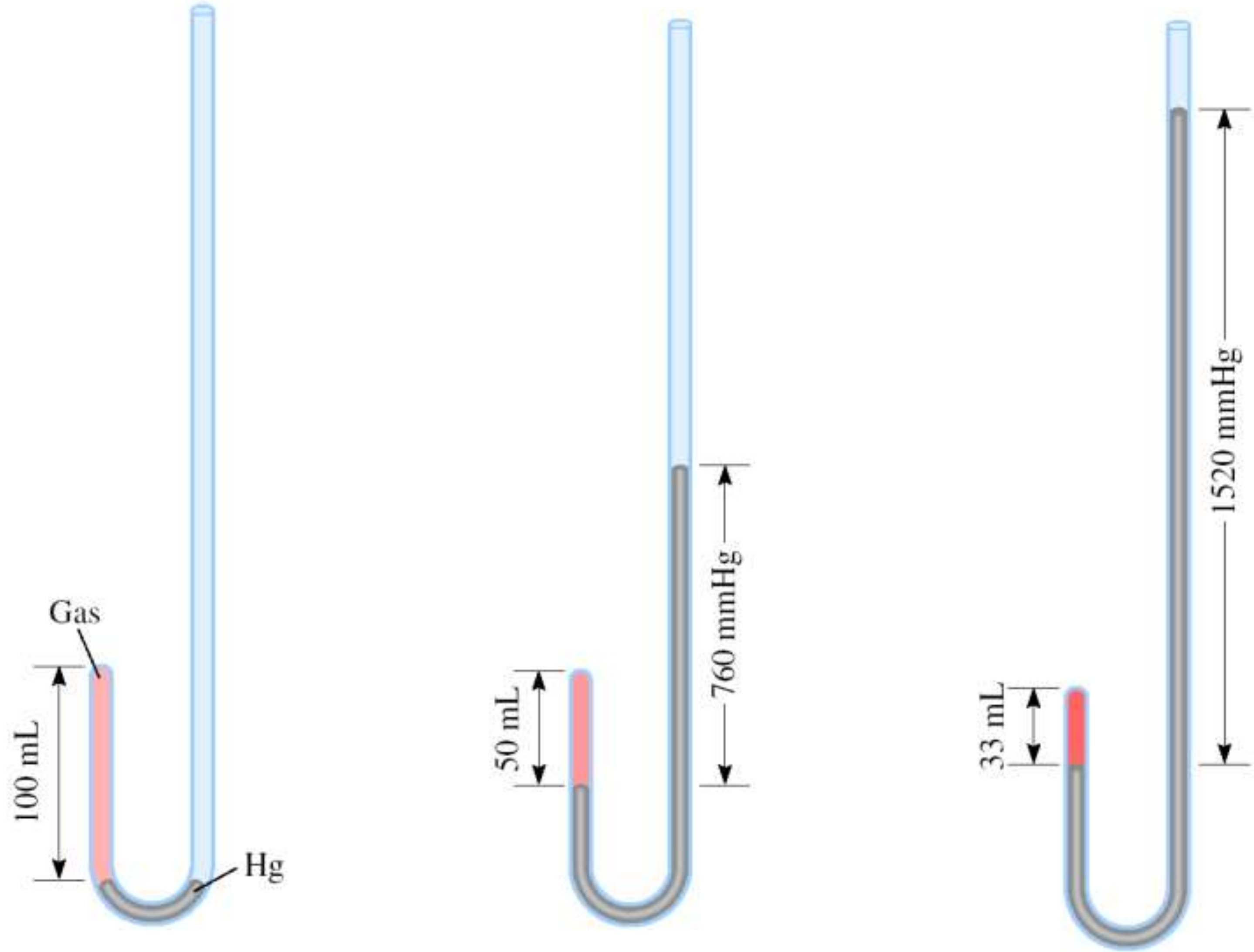
$$= 9.76 \times 10^4\ \text{Pa} \quad = 97.6\,\text{kPa}$$

Manometers Used to Measure Gas Pressures



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

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



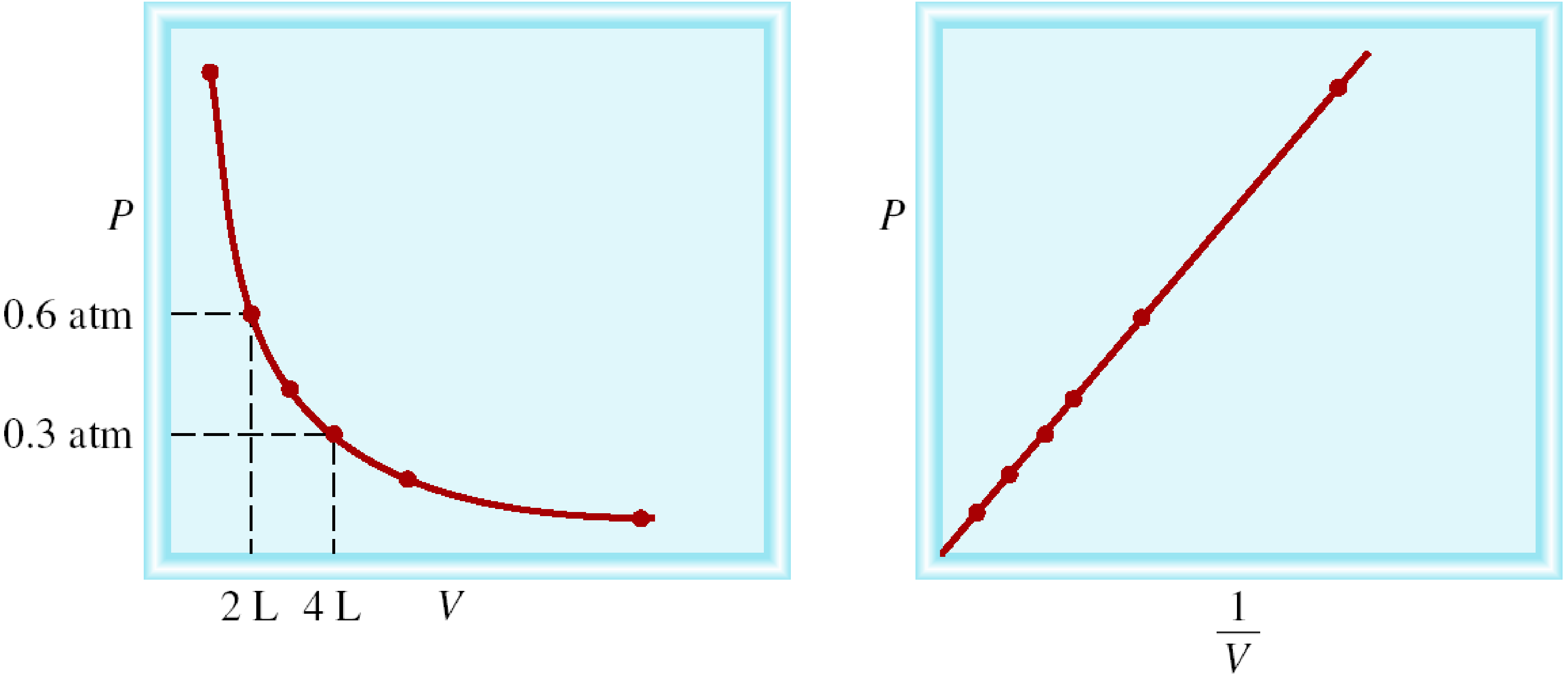
As *P* (h) increases *V* decreases

Boyle's Law

The volume of a fixed quantity of gas at constant temperature is inversely proportional to the pressure.

- $V \, \alpha \, \frac{1}{P}$ (temperature constant)
- V = volume, P = pressure, α = proportional sign, 1/ = inverse
- $V = K \times \frac{1}{P}$ (cross multiplying)
- $PV = K$
- $P_1V_1 = K = P_2V_2$
- $P_1V_1 = P_2V_2$

Boyle's Law



$P \propto 1/V$

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

$P_1 \times V_1 = P_2 \times V_2$

Constant temperature
Constant amount of gas

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?

$$P \times V = constant$$

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

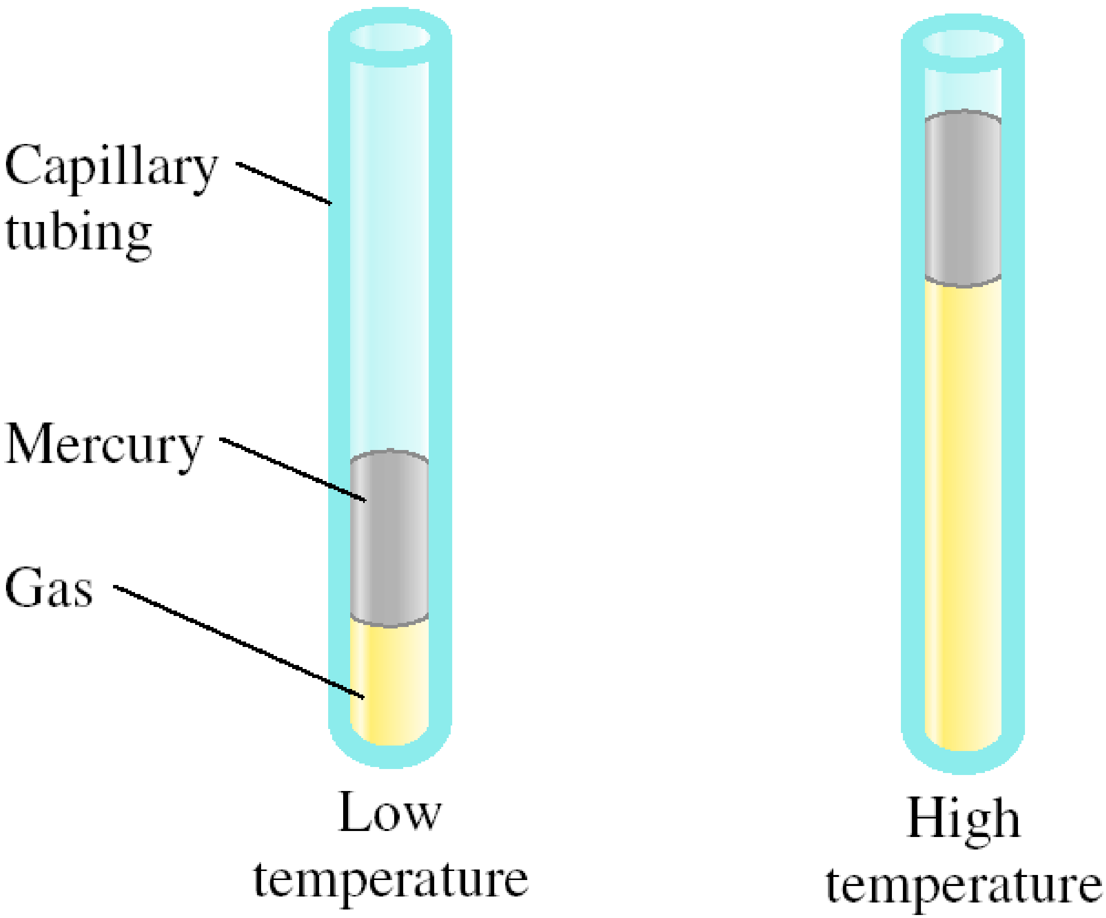
$P_1 = 726$ mmHg $\qquad P_2 = ?$

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

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

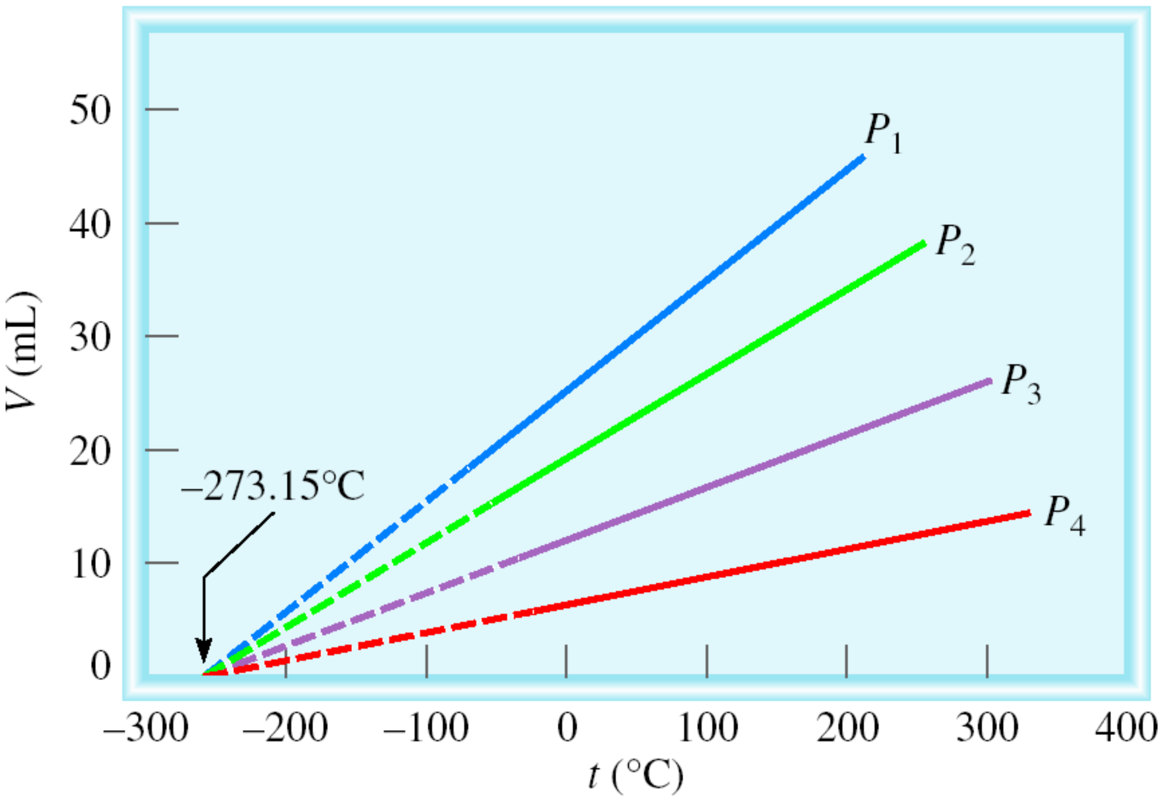
Charles' & Gay-Lussac's Law

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

$V \propto T$

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

$V_1/T_1 = V_2/T_2$

Temperature **must** be in Kelvin

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

Charles's Law

$V \ \alpha \ T$ (pressure is constant)

- $V = KT$
- $\dfrac{V}{T} = K$
- $\dfrac{V_1}{T_1} = K = \dfrac{V_2}{T_2}$
- The volume of a fixed amount of gas at constant pressure is directly proportional to its absolute temperature.

Example: In an experiment, a sample of argon gas at 225K is heated and the volume increases from 3.50L to 12.5L, calculate the final temperature.

Using, $\dfrac{V_1}{T_1} = \dfrac{V_2}{T_2}$ (cross multiplying)

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

$$T_2 = \frac{12.5L \times 225K}{3.50L} = 803.6K = 804K$$

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 $\qquad V_2 = 1.54$ L

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

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

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

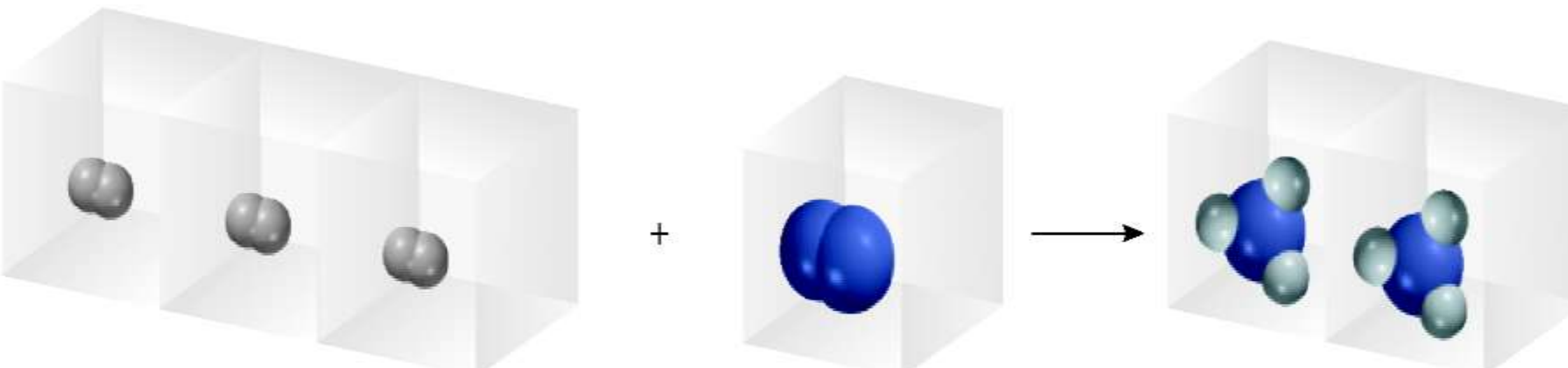
Avogadro’s Law

V a 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?

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

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

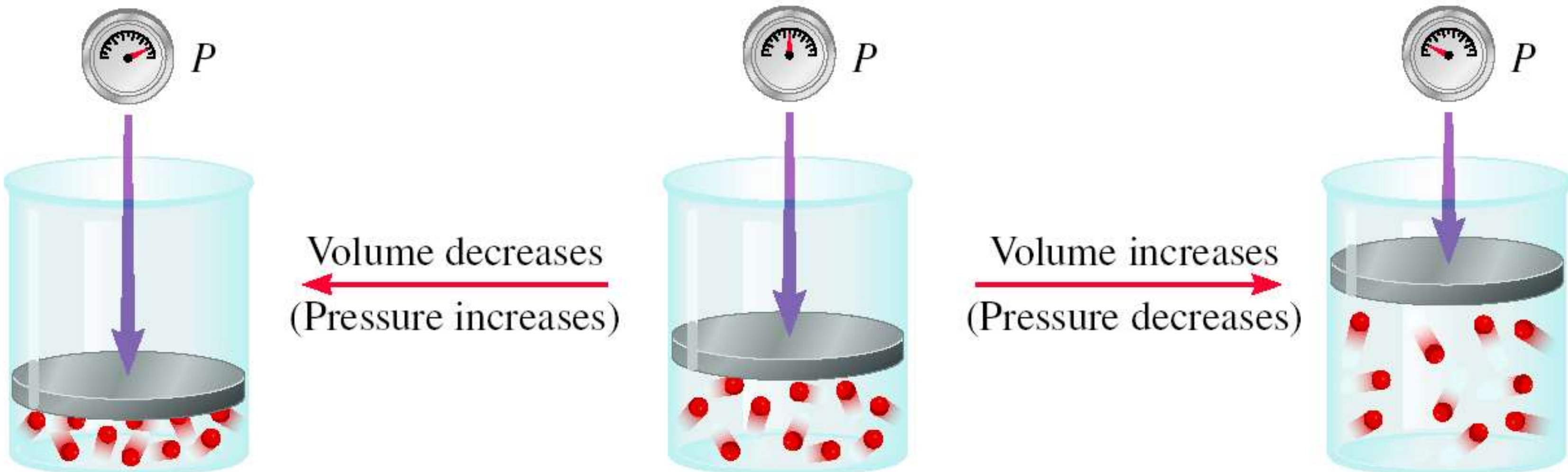
At constant *T* and *P*

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

Summary of Gas Laws

Boyle's Law

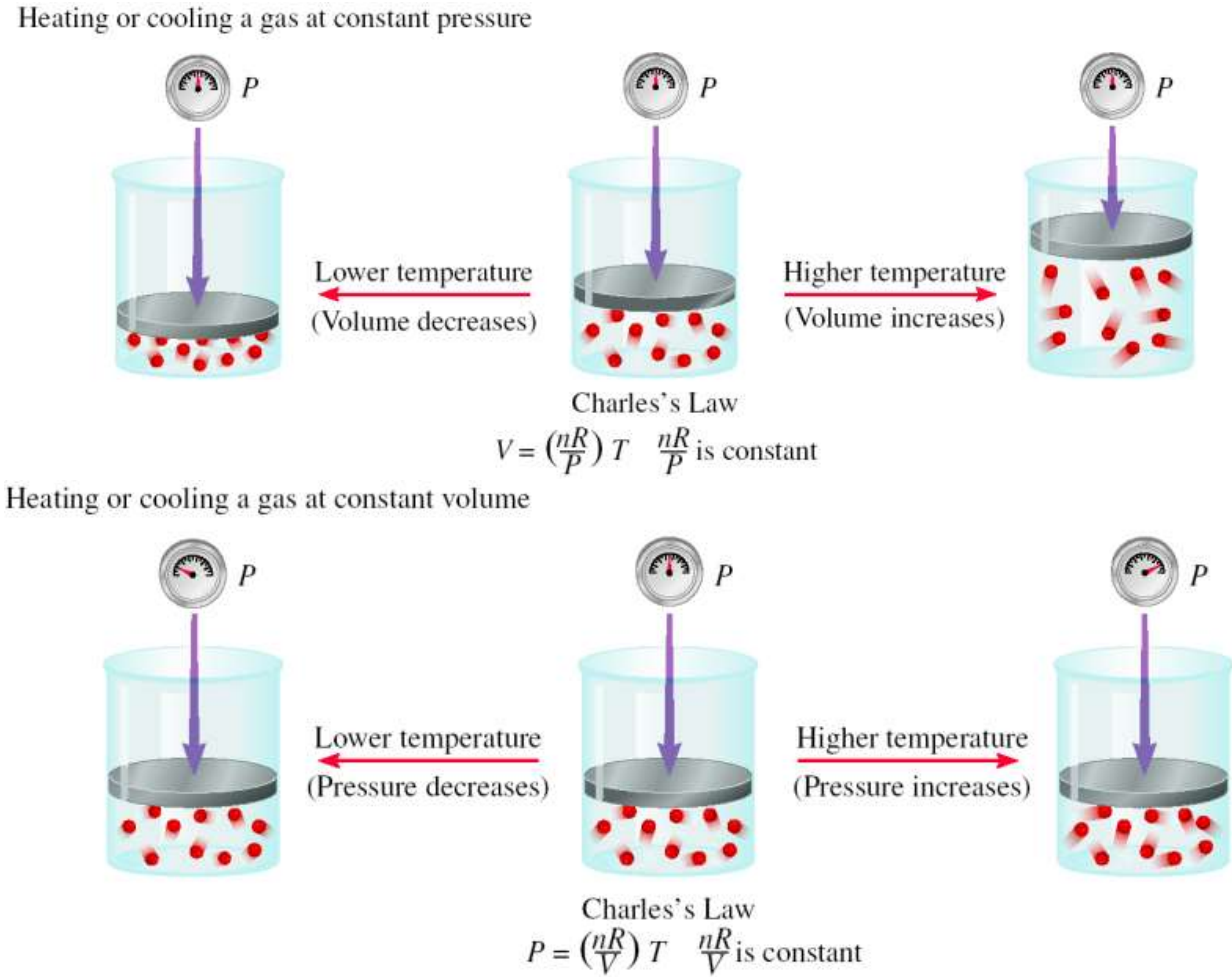
Increasing or decreasing the volume of a gas at a constant temperature



Boyle's Law

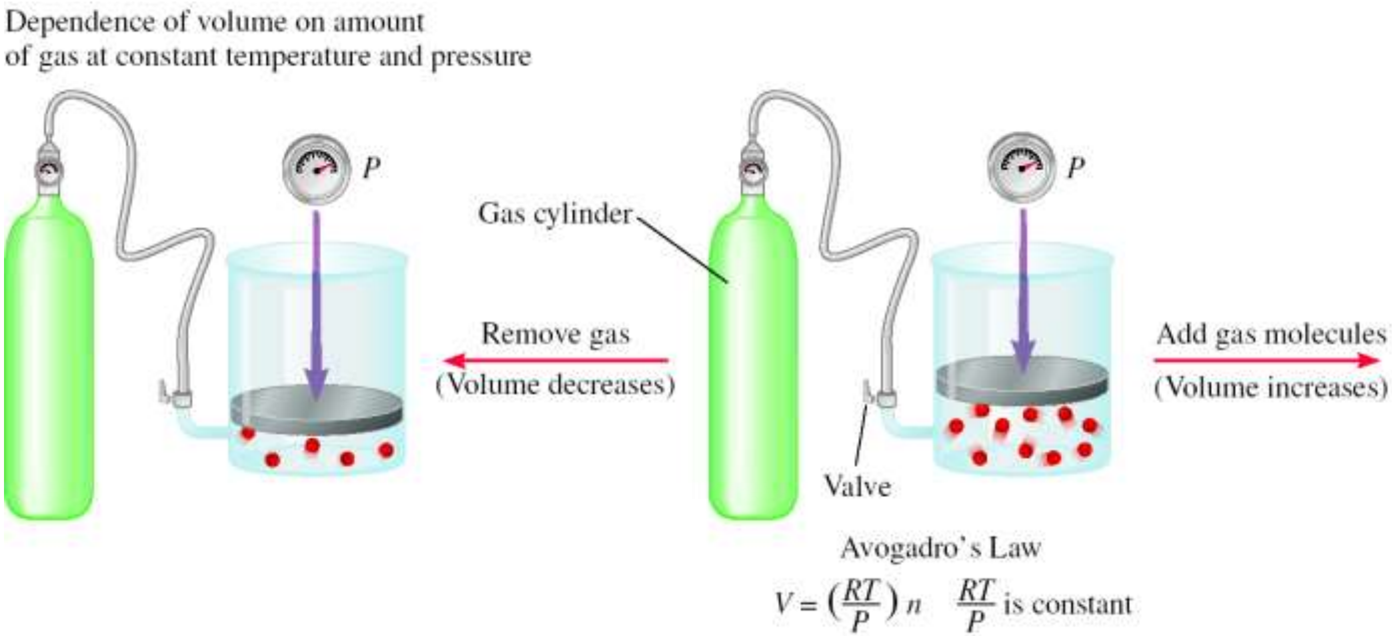
$$P = (nRT)\frac{1}{V} \quad nRT \text{ is constant}$$

Charles Law



Heating or cooling a gas at constant pressure
P
P
P
Lower temperature
(Volume decreases)
Higher temperature
(Volume increases)
Charles's Law
$V = \left(\frac{nR}{P}\right) T \quad \frac{nR}{P}$ is constant
Heating or cooling a gas at constant volume
P
P
P
Lower temperature
(Pressure decreases)
Higher temperature
(Pressure increases)
Charles's Law
$P = \left(\frac{nR}{V}\right) T \quad \frac{nR}{V}$ is constant

Avogadro’s Law



Ideal-Gas Equation

- So far we've seen that

$$V \propto 1/P \quad \text{(Boyle's law)}$$

$$V \propto T \quad \text{(Charles's law)}$$

$$V \propto n \quad \text{(Avogadro's law)}$$

- Combining these, we get

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

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

$$P = \frac{RnT}{V} \quad \text{(cross multiplying)}$$

$PV = nRT$ This is called ideal gas law.

P = Pressure; n = number of moles; V = Volume; T = Temperature; R = gas constant (0.0821 atm.L/mol)

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.

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$$PV = nRT$$

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

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

It is a hypothetical gas which follows ideal gas equation

Ideal gas:

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

If all variables change, we use:

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$ In general, $n_1 = n_2$ thus $$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

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

$$PV = nRT$$

$P = 1\ atm$ $\qquad$ $T = 0\ ^{0}\text{C} = 273.15\ \text{K}$

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

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

$$V = \frac{1.37\ \cancel{\text{mol}} \times 0.0821\ \frac{\text{L}\bullet\cancel{\text{atm}}}{\cancel{\text{mol}}\bullet\cancel{\text{K}}} \times 273.15\ \cancel{\text{K}}}{1\ \cancel{\text{atm}}} = 30.7\ \text{L}$$

A metal cylinder holds 50 L of O_2 gas at 18.5 atm and 21^{o}C, what volume will the gas occupy if the temperature is maintained at 21 ^{o}C while the pressure is reduced to 1.00 atm?

$$P_1V_1 = P_2V_2$$

$$18.5\ \text{atm} \times 50\ \text{L} = 1.00\ \text{atm} \times V_2$$

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)?

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

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

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

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

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

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

An inflated balloon has a volume of 6.0 L at sea level (1.0 atm) and is allowed to ascend in altitude until the pressure is 0.45 atm. During ascent the temperature of the gas falls from 22 ^oC to -21^oC. Calculate the volume of the balloon at its final altitude.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{1.0\ atm \times 6.0\ L}{295\ K} = \frac{0.45\ atm \times V_2}{252\ K}$$

$$V_2 = 11\ L$$

$PV = nRT$

How many liters of $N_2(g)$ at 1.00 atm and 25.0 °C are produced by the decomposition of 150. g of NaN_3?

$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$

V = ?

P = 1 atm

T = 25C + 273.15 = 298.15 K

V = nRT/P

$$n = \text{mol } N_2 = \frac{150.\text{g } NaN_3}{1} \times \frac{1 \text{ mol } NaN_3}{65.0099 \text{ g}} \times \frac{3 \text{ mol } N_2}{2 \text{ mol } NaN_3}$$

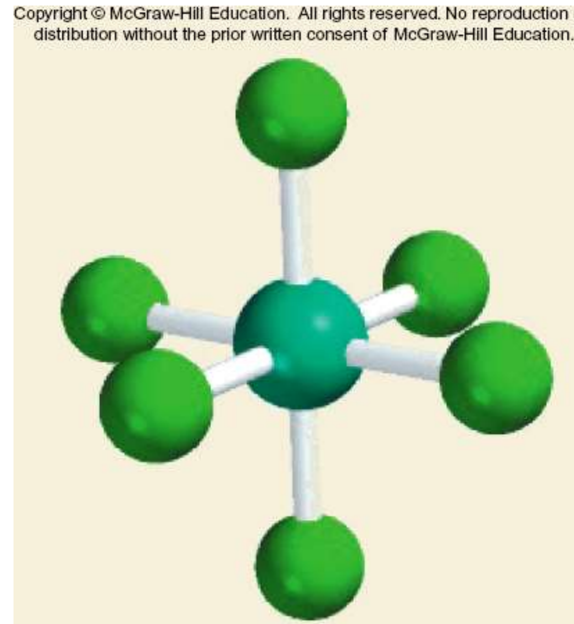
n = 3.461 mol N_2

$$V = \frac{(3.461 \text{mol } N_2)\left(0.082057 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(298.15\text{K})}{1\text{atm}}$$

V=84.62L

Example 5.3

Sulfur hexafluoride (SF_6) is a colorless and odorless gas. Due to its lack of chemical reactivity, it is used as an insulator in electronic equipment.



SF_6

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

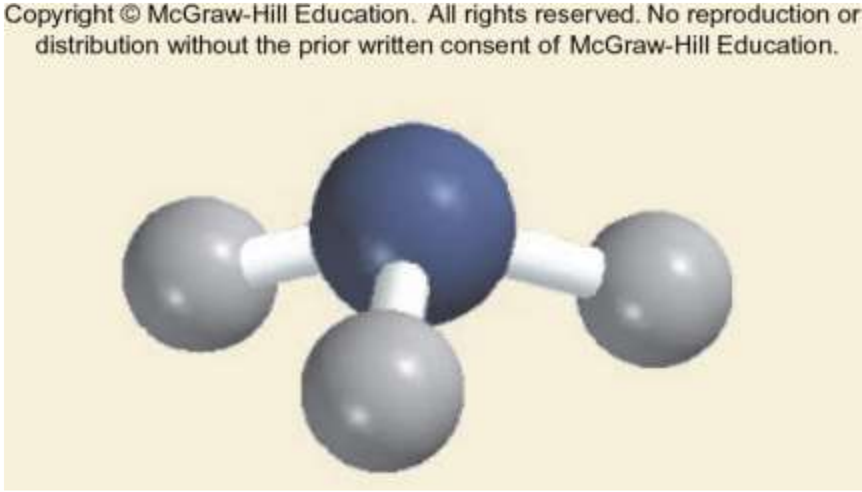
Solution

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

$$= \frac{(1.82\ \text{mol})(0.0821\ \text{L}\cdot\text{atm/K}\cdot\text{mol})(69.5 + 273)\ \text{K}}{5.43\ \text{L}}$$

Example 5.4

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



NH_3

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

$$\text{grams of } NH_3 \rightarrow \text{moles of } NH_3 \rightarrow \text{litres of } NH_3 \text{ at STP}$$

$$V = 7.40 \cancel{\text{g}}\ \cancel{NH_3} \times \frac{1 \cancel{\text{mol}}\ \cancel{NH_3}}{17.03 \text{ g} \cancel{NH_3}} \times \frac{22.41 \text{ L}}{1 \cancel{\text{mol}}\ \cancel{NH_3}} = 9.74 \text{ L}$$

Examples

1) 20.8 g of CH_4 gas was confined in 5.200 L vessel at 50 ^{0}C. 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^{0}C , Calculate its volume in a summer day at 50 ^{0}C?

1.138 L

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

239.8 mmHg

Densities of Gases

Density = mass ÷ volume

- So, $d = \frac{m}{V}$

$$\boxed{d = \frac{m}{V} = \frac{PM}{RT}}$$

What is the density of carbon tetrachloride (CCl_4) vapor at 714 torr and 125 ^oC?

$$d = \frac{PM}{RT}$$

$$d = \frac{714 \text{ torr} \times 154 \text{ g mol}^{-1}}{62.36 \text{ L torr mol}^{-1}\text{K}^{-1} \times 398 \text{ K}}$$

$$d = 4.43 \text{ g L}^{-1}$$

Molecular Mass

We can manipulate the density equation to enable us to find the molecular mass of a gas:

$$d = \frac{PM}{RT}$$

M is the molar mass of the gas

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

Calculate the average molar mass of dry air if it has a density of 1.17 g/L at 21°C and 740.0 torr.

$$M = \frac{d\,RT}{P} = 29.0 \text{ g/mol}$$

Density (d) Calculations

$$d = \frac{m}{V} = \frac{PM}{RT}$$

m is the mass of the gas in g

M is the molar mass of the gas

Molar Mass (M) of a Gaseous Substance

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

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?

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

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

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

1- A scuba diver's tank contains 0.29 kg of O_2 compressed into a volume of 2.3 L. (a) Calculate the gas pressure inside the tank at 9 ^oC. (b) What volume would this oxygen occupy at 26 ^oC and 0.95 atm?

$$PV = nRT$$

$$P \times 2.3\ L = (\ 290\ g/32\ g\ mol^{-1}) \times 0.082\ \times 282\ K$$

$$P = 91\ atm$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = 233\ L$$

2-Which gas is most dense at 1.00 atm and 298 K? CO_2, N_2O, or Cl_2.

$$d = \frac{PM}{RT} \qquad d\ CO_2 = \frac{1 \times 44}{0.082 \times 298} = 1.8\ g\ L^{-1}$$

$$d\ N_2O = 1.8\ g\ L^{-1} \qquad d\ Cl_2 = 2.91\ g\ L^{-1}$$

3- (a) Calculate the density of NO_2 gas at 0.970 atm and 35 °C. (b) Calculate the molar mass of a gas if 2.50 g occupies 0.875 L at 685 torr and 35 °C.

$$d = \frac{PM}{RT} \qquad d = \frac{0.97 \times 46}{0.082 \times 308} \qquad d = 1.77 \text{ g L}^{-1}$$

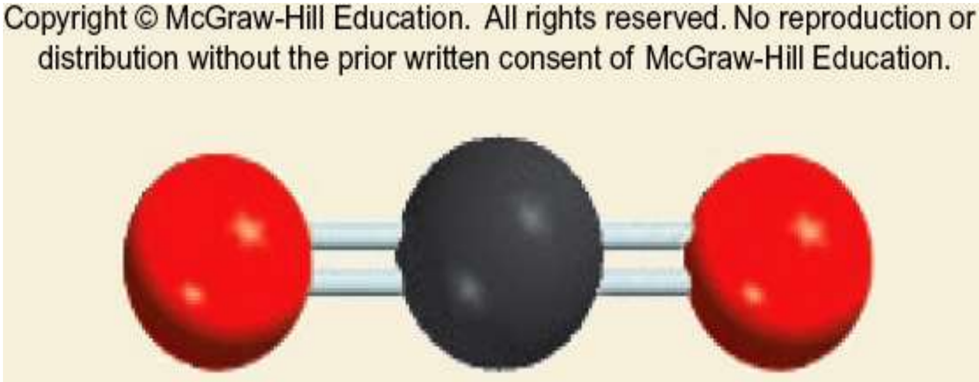
$$M = \frac{dRT}{P} \qquad d = \frac{\text{mass}}{\text{volume}}$$

$$d = \frac{2.5 \text{ g}}{0.875 \text{ L}} = 2.86 \text{ g L}^{-1}$$

$$M = \frac{2.86 \times 62.36 \times 308}{685} \qquad M = 80.2 \text{ g mol}^{-1}$$

Example 5.8

Calculate the density of carbon dioxide (CO_2) in grams per liter (g/ L) at 0.990 atm and 55°C.



Solution

To use Equation (5.11), we convert temperature to kelvins (T = 273 + 55 = 328 K) and use 44.01 g for the molar mass of CO_2

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

CO_2

$$V = \frac{nRT}{P} = \frac{(1\ \text{mol})(0.0821\ \text{L}\cdot\text{atm/K}\cdot\text{mol})(328\ \text{K})}{0.990\ \text{atm}} = 27.2\ \text{L}$$

$$d = \frac{44.01\ \text{g}}{27.2\ \text{L}} = 1.62\ \text{g/L}$$

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.

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

$$n = \frac{PV}{RT} = \frac{(2.88\ \mathrm{atm})(1.00\ \mathrm{L})}{(0.0821\ \mathrm{L \cdot atm/K \cdot mol})(309\ \mathrm{K})} = 0.1135\ \mathrm{mol}$$

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

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

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

$$n_F = 67.0\ \cancel{\text{g F}} \times \frac{1\ \text{mol F}}{28.09\ \cancel{\text{g F}}} = 3.53\ \text{mol F}$$

Therefore, the empirical formula is $Si_{1.17}F_{3.53}$, or, dividing by the smaller subscript (1.17), we obtain SiF_3

Example 5.10

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 ≈ 2). Therefore, the molecular formula of the compound must be

$$(SiF_3)_2 \text{ or } Si_2F_6.$$

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:

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

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

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

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

Example 5.11

Calculate the volume of O_2 (in liters) required for the complete combustion of 7.64 L of acetylene (C_2H_2) measured at the same temperature and pressure.

$$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$$

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Solution

According to Avogadro's law, at the same temperature and pressure, the number of moles of gases are directly related to their volumes.

$$\text{volume of } O_2 = 7.64\ \text{L}\cancel{C_2H_2} \times \frac{5\ \text{L}\,O_2}{2\ \text{L}\cancel{C_2H_2}} = 19.1\ \text{L}$$

Example 5.12

Sodium azide (NaN_3) is used in some automobile air bags. The impact of a collision triggers the decomposition of NaN_3 as follows:

$$2NaN_3(s) \rightarrow 2Na\ (s) + 3N_2\ (g)$$

The nitrogen gas produced quickly inflates the bag between the driver and the windshield and dashboard.

Calculate the volume of N_2 generated at 80°C and 823 mmHg by the decomposition of 60.0 g of NaN_3.

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An air bag can protect the driver in an automobile collision.

Example 5.12

Solution

First we calculate number of moles of N_2 produced by 60.0 g NaN_3 using the following sequence of conversions

$$\text{grams of NaN}_3 \rightarrow \text{moles of NaN}_3 \rightarrow \text{moles of N}_2$$

so that

$$\text{moles of N}_2 = 60.0\ \text{g}\ \cancel{\text{NaN}_3} \times \frac{1\ \text{mol}\ \cancel{\text{NaN}_3}}{65.02\ \text{g}\ \cancel{\text{NaN}_3}} \times \frac{3\ \text{mol N}_2}{2\ \text{mol}\ \cancel{\text{NaN}_3}}$$

$$= 1.38\ \text{mol N}_2$$

The volume of 1.38 moles of N_2 can be obtained by using the ideal gas equation:

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

$$= 36.9\ \text{L}$$

Solid $CaCO_3$ decomposes to solid CaO and CO_2 when heated. What is the pressure, in atm, of CO_2 in a 50.0 L container at 35 ^oC when 75.0 g of calcium carbonate decomposes?

A. 0.043 atm

B. 0.010 atm

B. C. 0.38 atm

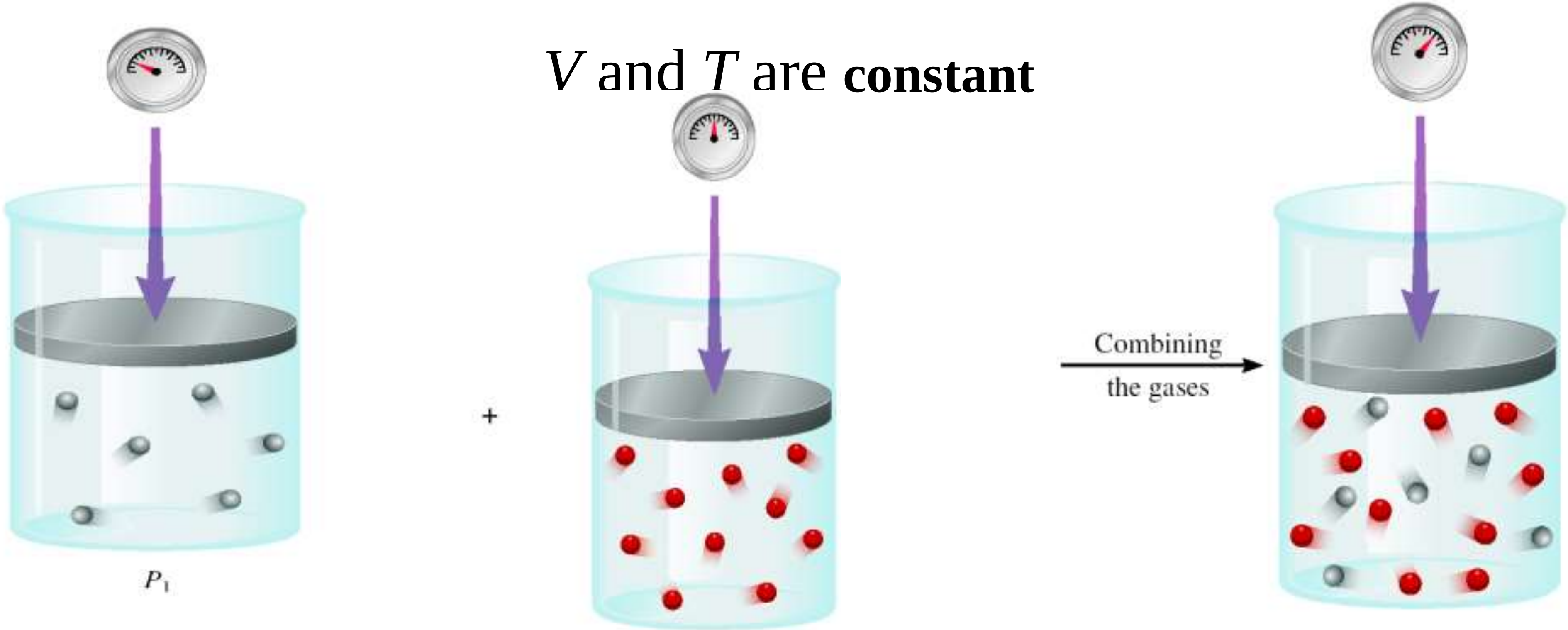
D. 0.08 atm

$$\frac{75.0 \text{ g } CaO_2 \times \dfrac{1 \text{ mol } CaCO_3}{100.1 \text{ g}} \times \dfrac{1 \text{ mol } CO_2}{1 \text{ mol } CaCO_3} \times 0.0821 \dfrac{\text{L atm}}{\text{K mol}} \times 308 \text{ K}}{50.0 \text{ L}}$$

$= 0.38$ atm

Dalton's Law of Partial Pressures

V and *T* are **constant**



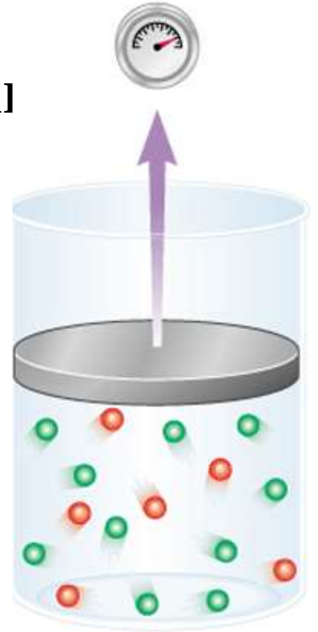
P_1 $\quad$ P_2 $\quad$ $P_{total} = P_1 + P_2$

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

- In other words,

$$P_{total} = P_1 + P_2 + P_3 + \ldots$$

Consider a case in which two gases, A and B, are in a contai[n]
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 (X_i) = $\frac{n_i}{n_T}$

Partial Pressures

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

$$\mathrm{P}_1 = n_1 \left(\frac{RT}{V}\right) \;;\; \mathrm{P}_2 = n_2 \left(\frac{RT}{V}\right) \text{ ; and so on}$$

$$\mathrm{P}_{\mathrm{t}} = (n_1 + n_2 + n_3 + ..)\frac{RT}{V} = n_{\mathrm{t}} \left(\frac{RT}{V}\right)$$

$$\frac{P_1}{P_t} = \frac{n_1 RT/V}{n_t RT/V} = \frac{n_1}{n_t}$$

$$P_1 = \frac{n_1}{n_t} P_t \quad \Rightarrow \quad \boxed{P_1 = X_1 P_t}$$

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.

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})

need to find

$$P_{\mathrm{Ne}} = X_{\mathrm{Ne}} P_{\mathrm{T}}$$

want to calculate

given

Example 5.14

we calculate the mole fraction of Ne as follows:

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

Therefore,

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

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

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

(1.21 + 0.20 + 0.586) atm = 2.00 atm.

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)?

$P_i = X_i P_T$ $\qquad$ $P_T = 1.37$ atm

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

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

250 mL of methane, CH_4, at 35 ^oC and 0.55 atm and 750 mL of propane, C_3H_8, at 35 ^oC and 1.5 atm, were introduced into a 10.0 L container. What is the final pressure, in torr, of the mixture?

A. 95.6 torr

B. 6.20×10^4 torr

C. 3.4×10^3 torr

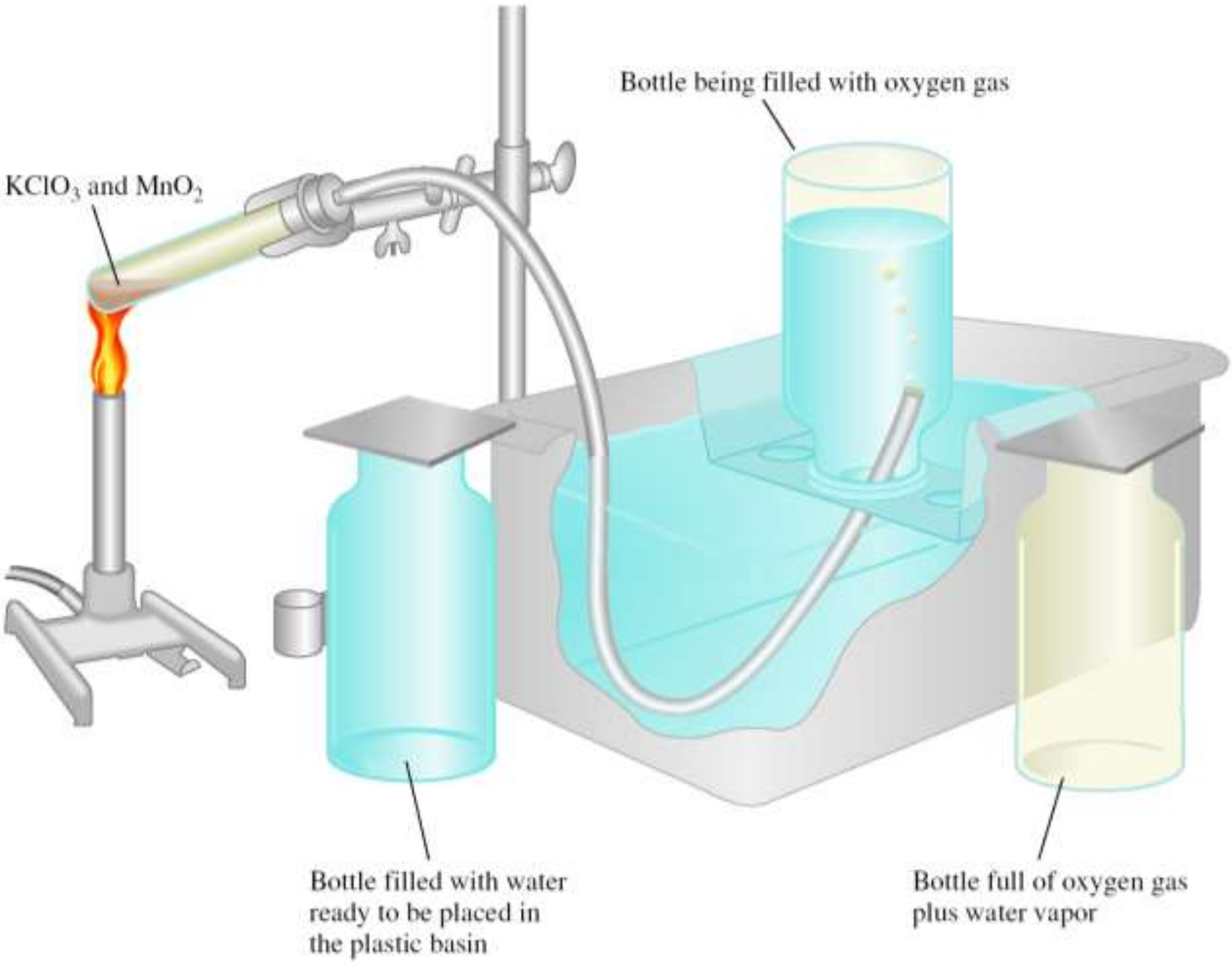
D. 760 torr

$$P(CH_4) = \frac{0.55 \text{ atm} \times 0.250 \text{ L}}{10.0 \text{ L}} = 0.0138 \text{ atm}$$

$$P(C_3H_8) = \frac{1.5 \text{ atm} \times 0.750 \text{ L}}{10.0 \text{ L}} = 0.112 \text{ atm}$$

$$P_T = (0.0138 + 0.112) \text{ atm} \times \frac{760 \text{ torr}}{\text{atm}} = 95.6 \text{ torr}$$

Collecting a Gas over Water



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

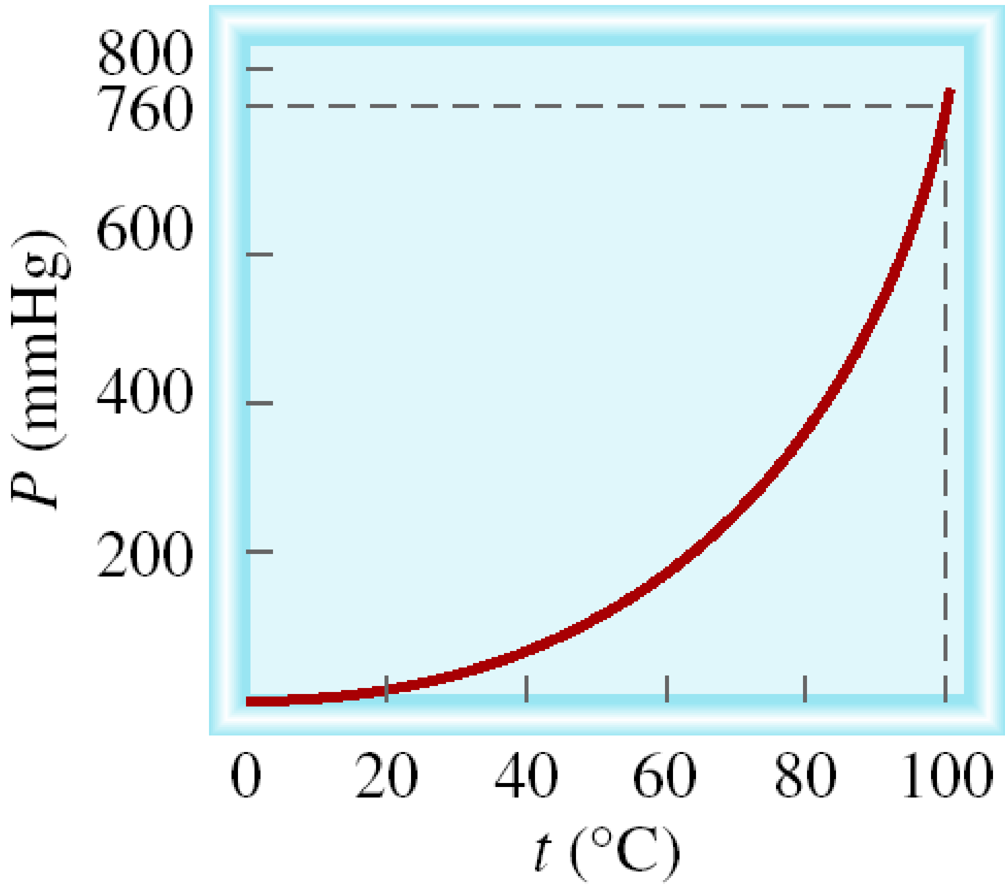
$$P_T = P_{O_2} + P_{H_2O}$$

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.

$$P_{\mathrm{O_2}} = P_{\mathrm{T}} - P_{\mathrm{H_2O}} = 762 \text{ mmHg} - 22.4 \text{ mmHg} = 740 \text{ mmHg}$$

$$PV = nRT = \frac{m}{M} RT \qquad m = \frac{PVM}{RT} = \frac{(740/760)\text{ atm}\,(0.128\text{ L})\,(32.00\text{ g/mol})}{(0.0821\text{ L}\cdot\text{atm}/K\cdot\text{mol})\,(273+24)\text{ K}}$$

$$= 0.164 \text{ g}$$

Calculate the mass of $Zn_{(s)}$ used to produce $H_{2(g)}$ over water at 25.0^0C 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

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

Kinetic theory of gases and …

- Compressibility of Gases
- Boyle's Law

  $P \ \alpha$ collision rate with wall

  Collision rate α number density

  Number density $\alpha \ 1/V$

  $P \ \alpha \ 1/V$

- Charles' Law

  $P \ \alpha$ collision rate with wall

  Collision rate α average kinetic energy of gas molecules

  Average kinetic energy $\alpha \ T$

  $P \ \alpha \ T$

Kinetic theory of gases and …

- Avogadro's Law

  $P \propto$ collision rate with wall

  Collision rate $\propto$ number density

  Number density $\propto n$

  $P \propto n$

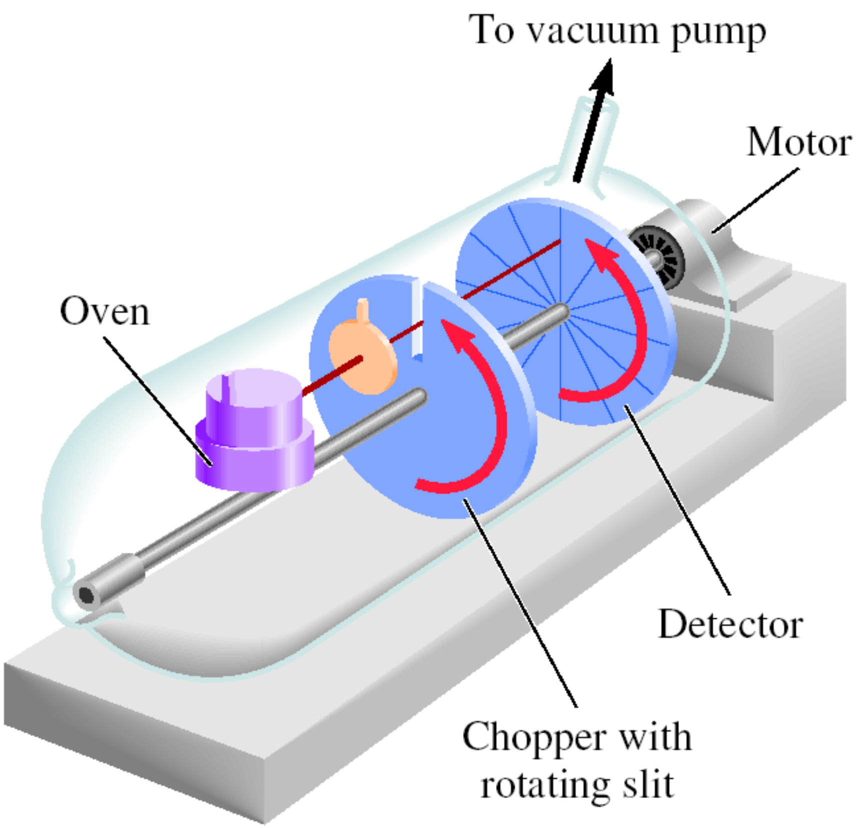
- Dalton's Law of Partial Pressures

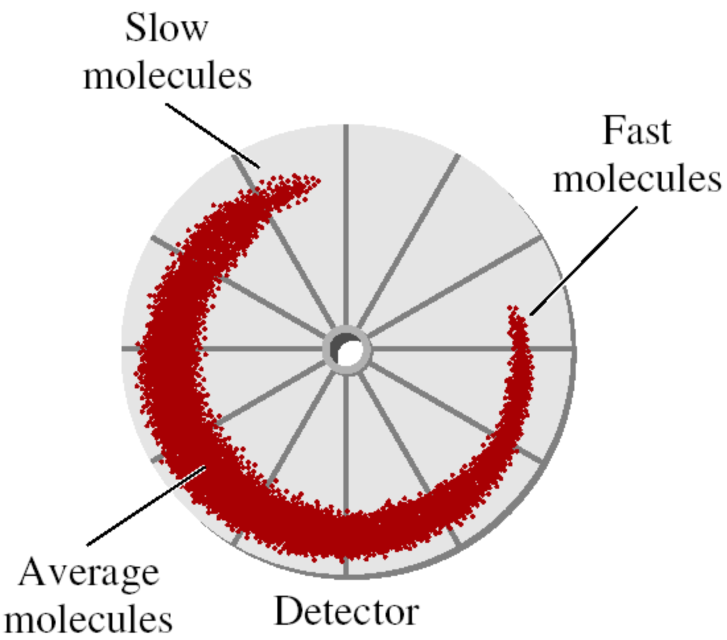
  Molecules do not attract or repel one another

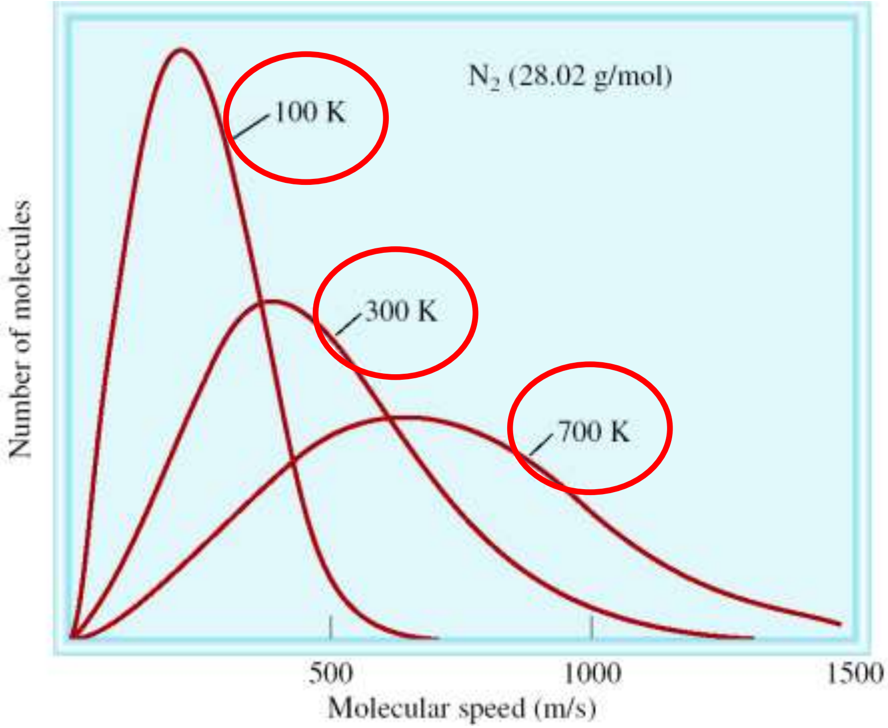
  P exerted by one type of molecule is unaffected by the presence of another gas

  $P_{\mathrm{total}} = \Sigma P_{\mathrm{i}}$

Apparatus for Studying Molecular Speed Distribution



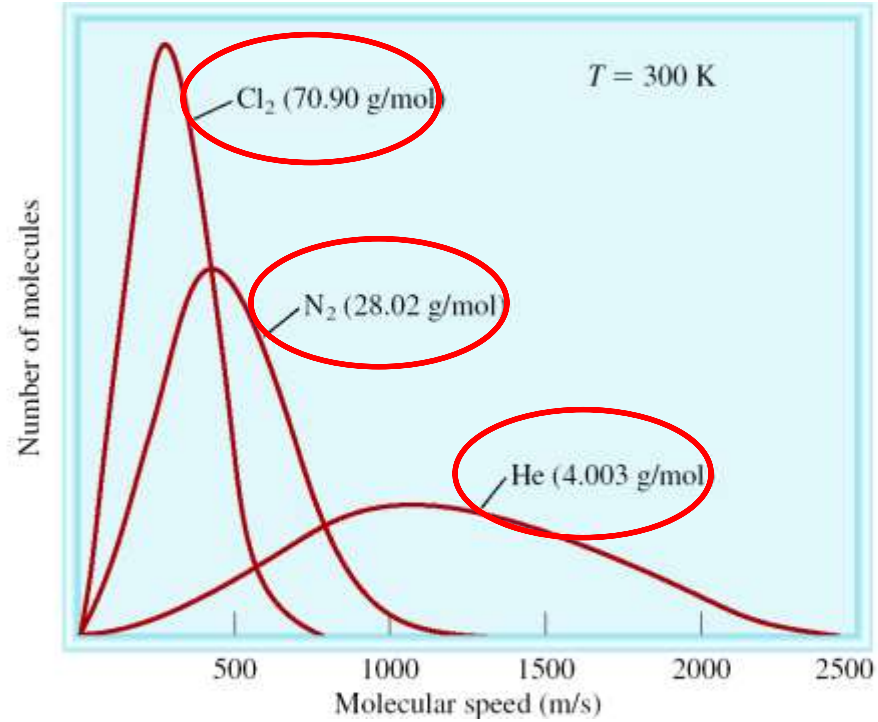




The distribution of speeds
for nitrogen gas molecules
at three **different temperatures**

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

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



Example 5.16

Calculate the root-mean-square speeds of helium atoms and nitrogen molecules in m/s at 25°C.

Solution

To calculate u_{rms}, the units of R should be 8.314 J/K. mol and, because the molar mass must be in Kg/mol.

The molar mass of He is 4.003 g/mol, or

$$1\text{J} = 1\text{Kg m}^2/\text{s}^2, \qquad 4.003\times10^{-3}\,\text{kg}/\text{mol}.$$

Example 5.16

$$u_{rms} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3(8.314 \text{ J/K}\cdot\text{mol})(298 \text{ K})}{4.003\times10^{-3} \text{ kg/mol}}} = \sqrt{1.86\times10^{6} \text{ J/kg}}$$

Using the conversion factor $1 \text{ J} = 1 \text{ kg m}^2/\text{s}^2$ we get

$$u_{rms} = \sqrt{1.86\times10^{6} \text{ kg m}^2/\text{kg}\cdot\text{s}^2} = \sqrt{1.86\times10^{6} \text{ m}^2/\text{s}^2} = 1.36\times10^{3} \text{ m/s}$$

The procedure is the same for N_2 the molar mass of which is

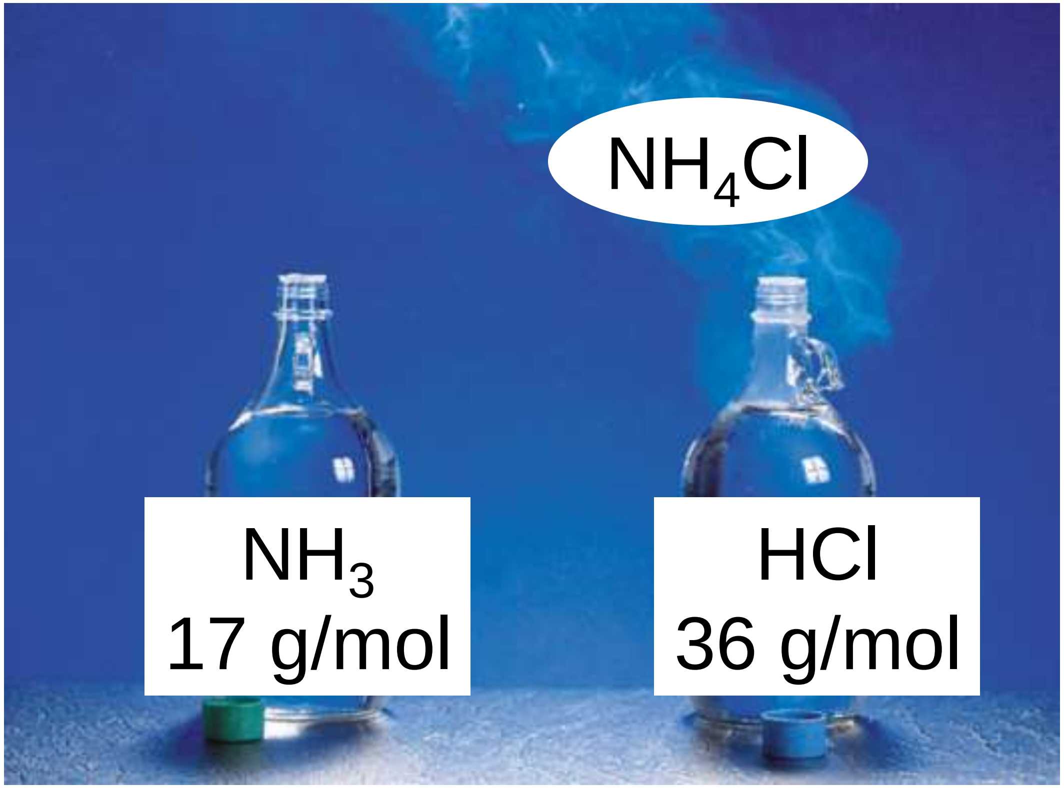
28.02 g/mol, 2.802×10^{-2} kg/mol so that we write

$$u_{rms} = \sqrt{\frac{3(8.314 \text{ J/K}\cdot\text{mol})(298 \text{ K})}{2.802\times10^{-2} \text{ kg/mol}}} = \sqrt{2.65\times10^{5} \text{ m}^2/\text{s}^2} = 515 \text{ m/s}$$

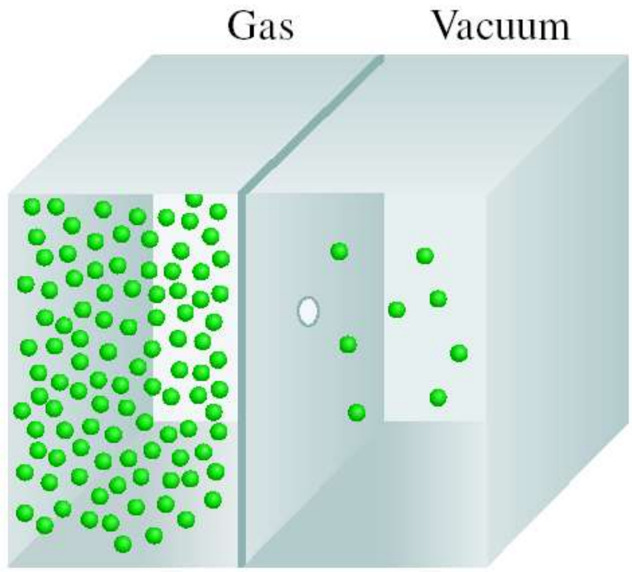
Gas diffusion is the gradual mixing of molecules of one gas with molecules of another by virtue of their kinetic properties.

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

molecular path



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

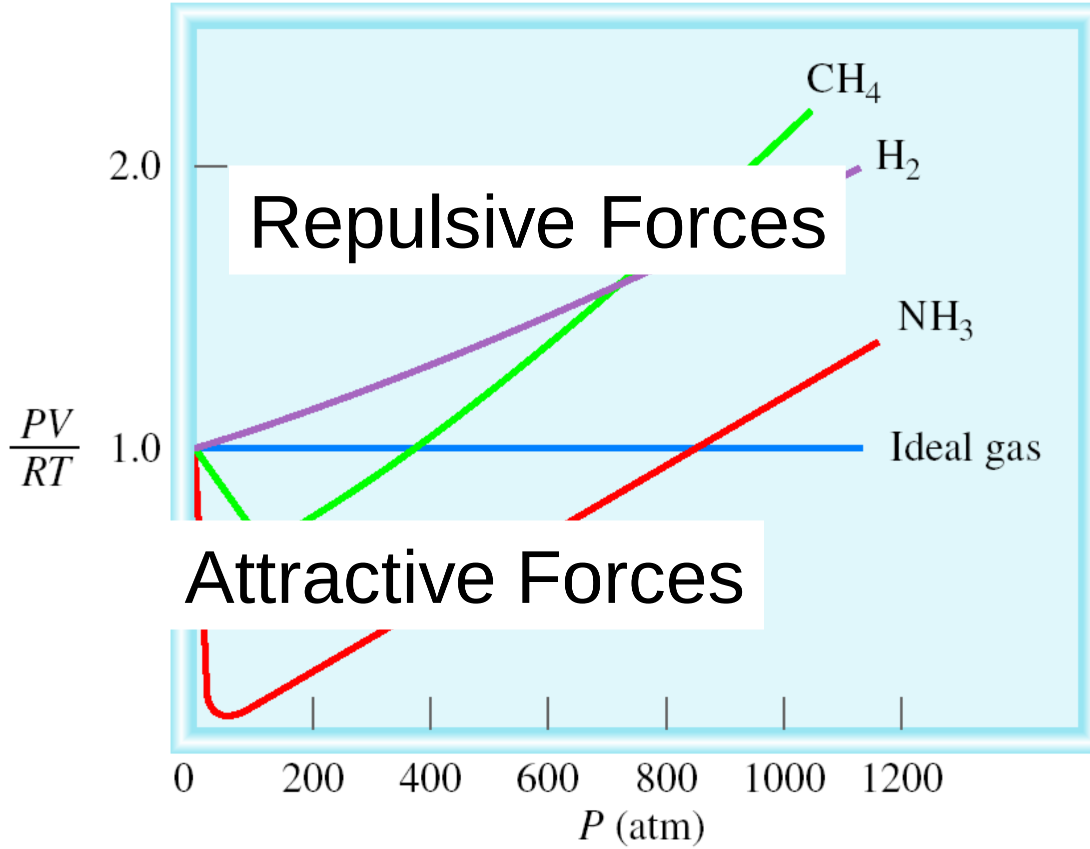
$58.7 + x \cdot 28 = 174.2 \qquad x = 4.1 \sim 4$

Deviations from Ideal Behavior

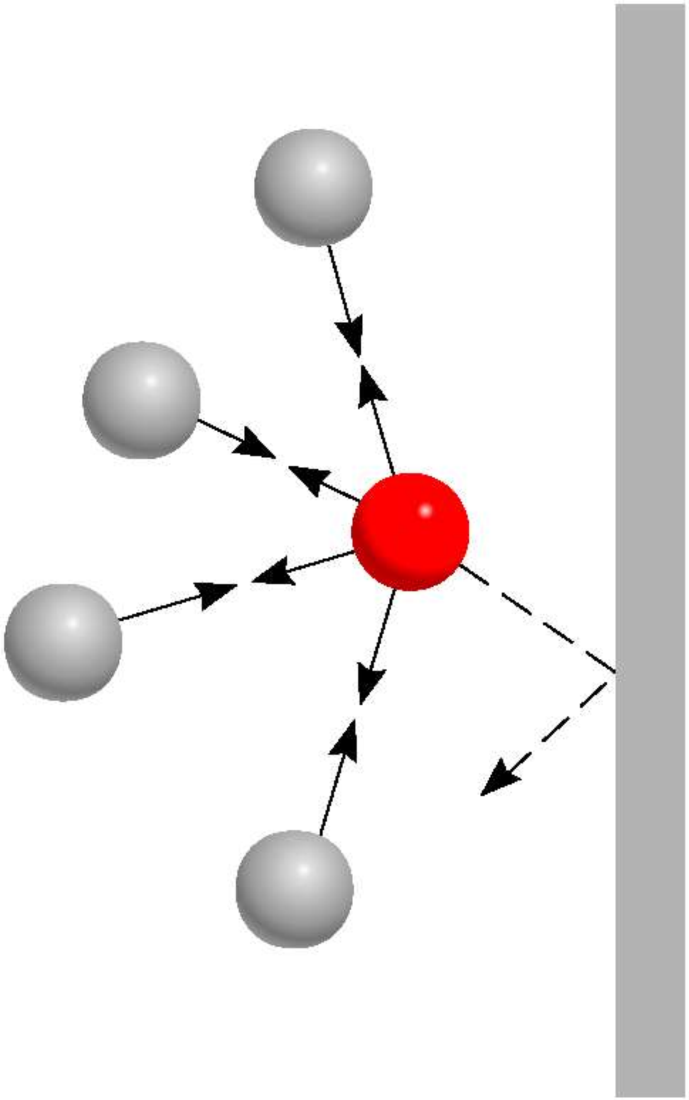
<u>1 mole of ideal gas</u>

$$PV = nRT$$

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



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

Table 5.4 Van der Waals Constants of Some Common Gases

Van der Waals equation nonideal gas

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

Gas	a $\left(\frac{\text{atm.L}^2}{\text{mol}^2}\right)$	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

Example 5.18

Given that 3.50 moles of NH_3 occupy 5.20 L at 47°C, calculate the pressure of the gas (in atm) using

a) the ideal gas equation and

b) the van der Waals equation.

Solution

a) We have the following data:

$$V = 5.20 \text{ L}$$

$$T = (47 + 273) \text{ K} = 320 \text{ K}$$

$$n = 3.50 \text{ mol}$$

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

Substituting these values in the ideal gas equation, we write

Example 5.18

$$P = \frac{nRT}{V} = \frac{(3.50\ \text{mol})(0.0821\ \text{L.atm/K.mol})(320\ \text{K})}{5.20\ \text{L}} = 17.7\ \text{atm}$$

b) We need Equation (5.18). It is convenient to first calculate the correction terms in Equation (5.18) separately. From Table 5.4, we have

$$\text{a} = 4.17\ \ \text{atm.L}^2/\text{mol}^2 \qquad \text{b} = 0.0371\ \text{L/mol}$$

so that the correction terms for pressure and volume are

$$\frac{\text{an}^2}{\text{v}^2} = \frac{(4.17\ \text{atm.L}^2/\text{mol}^2)(3.50\ \text{mol})^2}{(5.20\ \text{L})^2} = 1.89\ \text{atm}$$

$$\text{nb} = (3.50\ \text{mol})(0.0371\ \text{L/mol}) = 0.130\ \text{L}$$

$$(P + 1.89\ \text{atm})(5.20\ \text{L} - 0.130\ \text{L}) = (3.50\ \text{mol})(0.0821\ \text{L.atm/K.mol})(320\ \text{K})$$

$$P = 16.2\ \text{atm}$$