



# Gases

## *Chapter 5*



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

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

**TABLE 5.1****Some Substances Found as Gases at 1 atm and 25°C****Elements****Compounds**H<sub>2</sub> (molecular hydrogen)

HF (hydrogen fluoride)

N<sub>2</sub> (molecular nitrogen)

HCl (hydrogen chloride)

O<sub>2</sub> (molecular oxygen)

HBr (hydrogen bromide)

O<sub>3</sub> (ozone)

HI (hydrogen iodide)

F<sub>2</sub> (molecular fluorine)

CO (carbon monoxide)

Cl<sub>2</sub> (molecular chlorine)CO<sub>2</sub> (carbon dioxide)

He (helium)

NH<sub>3</sub> (ammonia)

Ne (neon)

NO (nitric oxide)

Ar (argon)

NO<sub>2</sub> (nitrogen dioxide)

Kr (krypton)

N<sub>2</sub>O (nitrous oxide)

Xe (xenon)

SO<sub>2</sub> (sulfur dioxide)

Rn (radon)

H<sub>2</sub>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.

# 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.



$\text{NO}_2$  gas

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

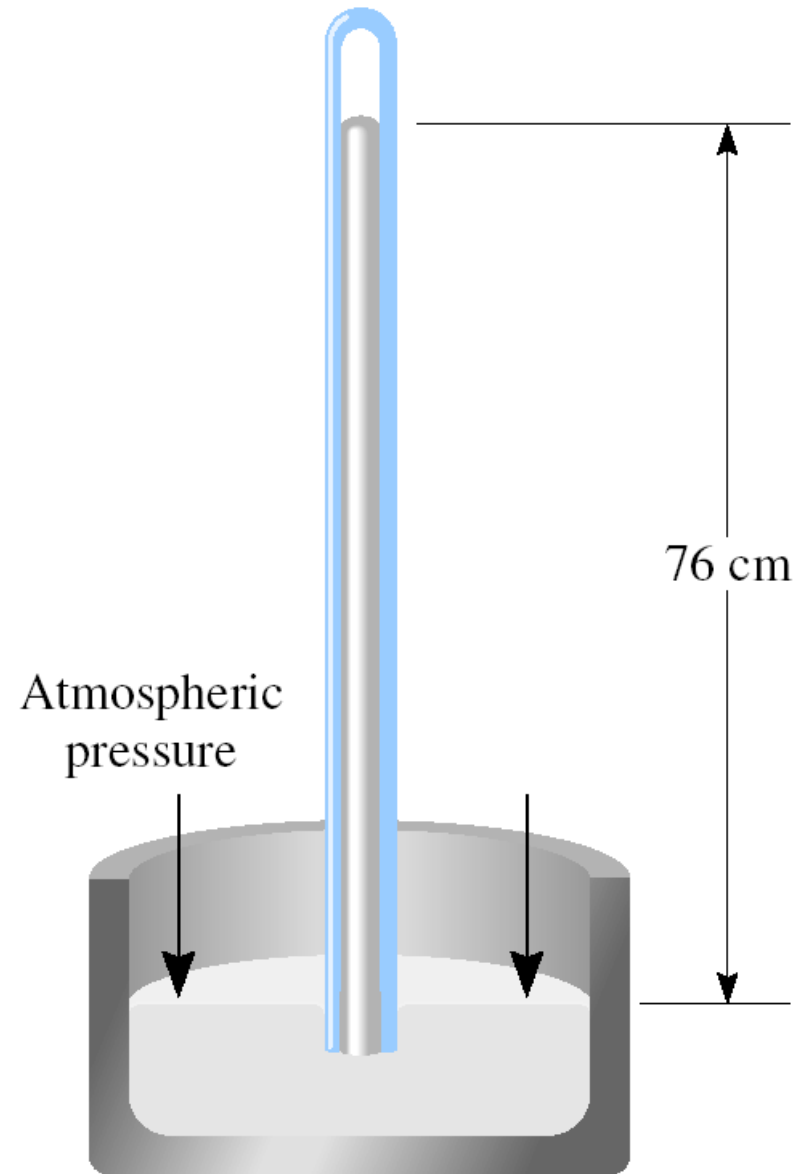
(force = mass x acceleration)

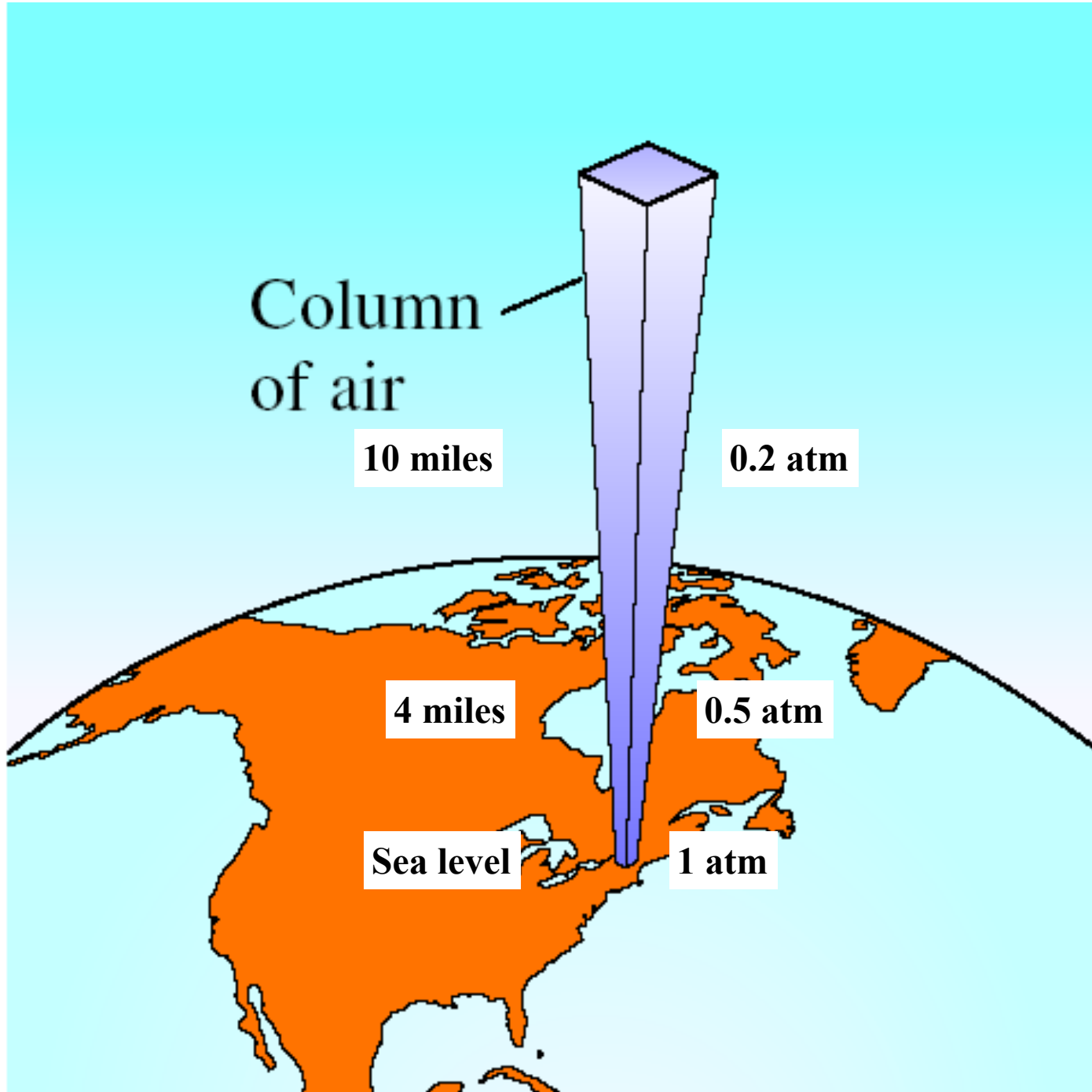
### Units of Pressure

1 pascal (Pa) = 1 N/m<sup>2</sup>

1 atm = 760 mmHg = 760 torr

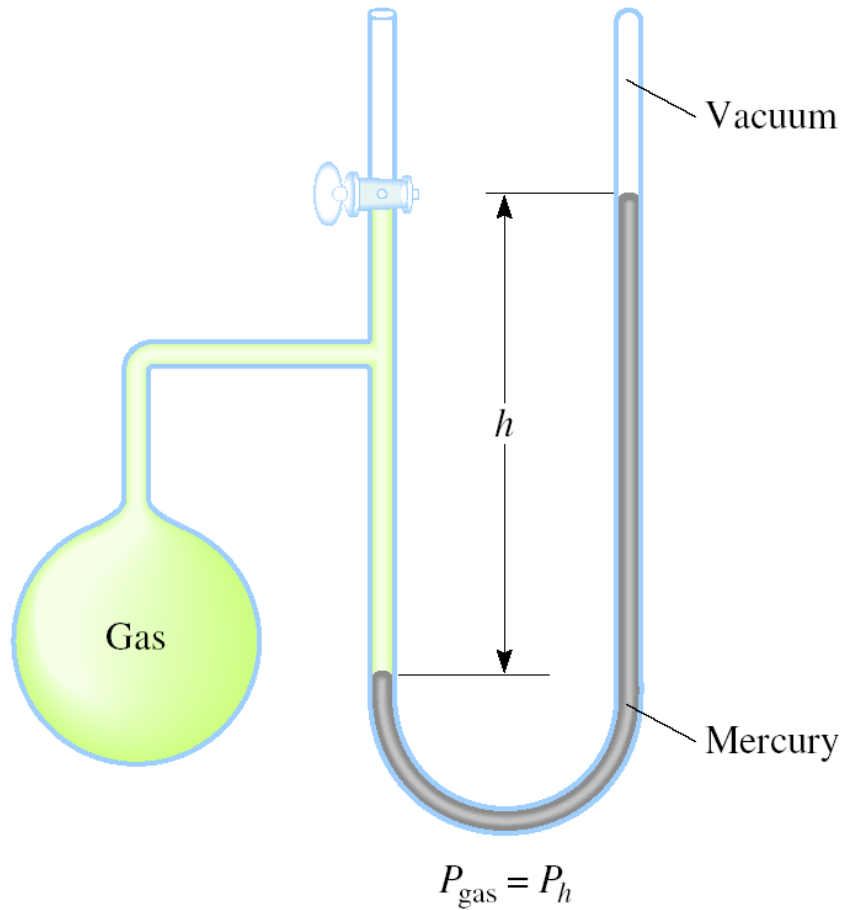
1 atm = 101,325 Pa



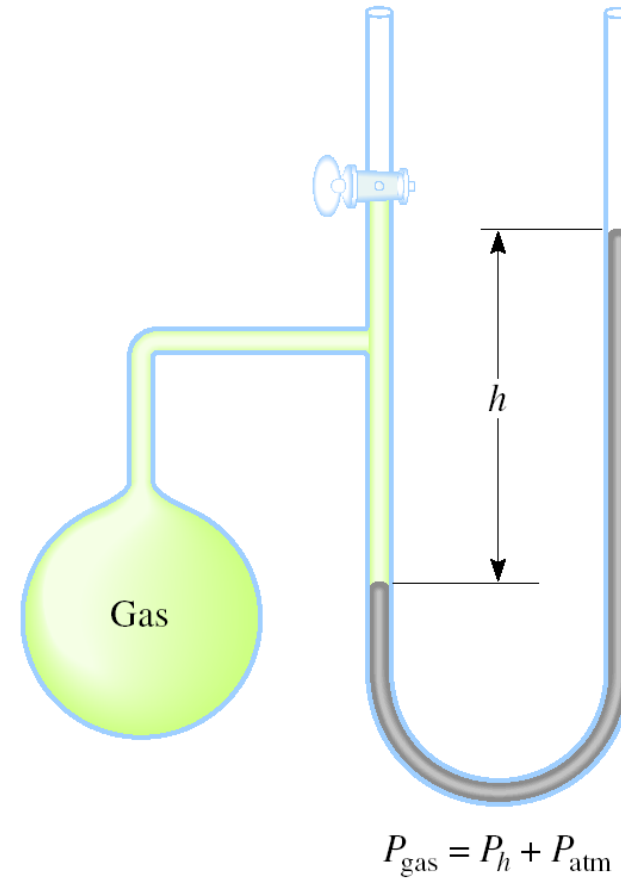


# Manometers Used to Measure Gas Pressures

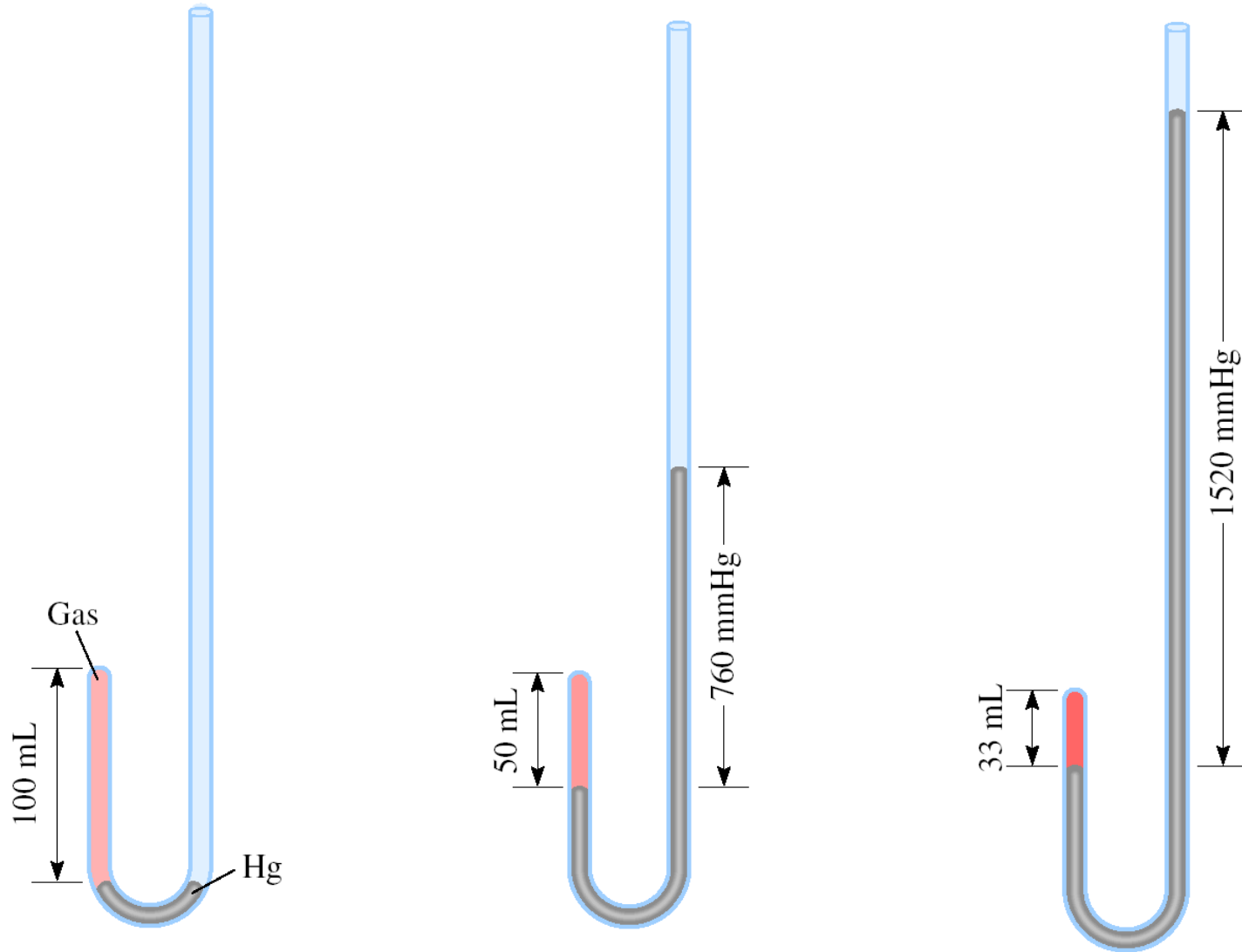
closed-tube



open-tube



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

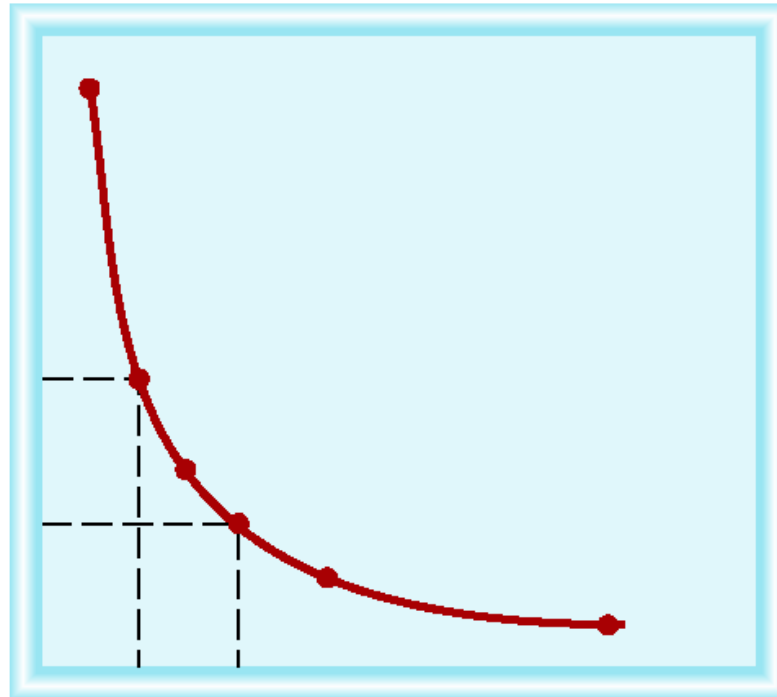


As  $P$  (h) increases

$V$  decreases



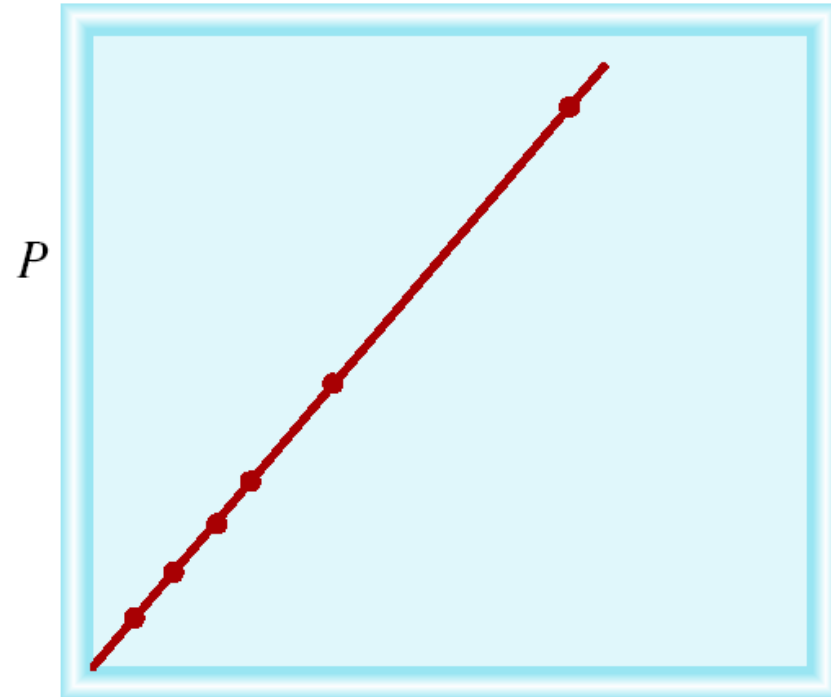
# Boyle's Law



$$P \propto 1/V$$

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

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



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?

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

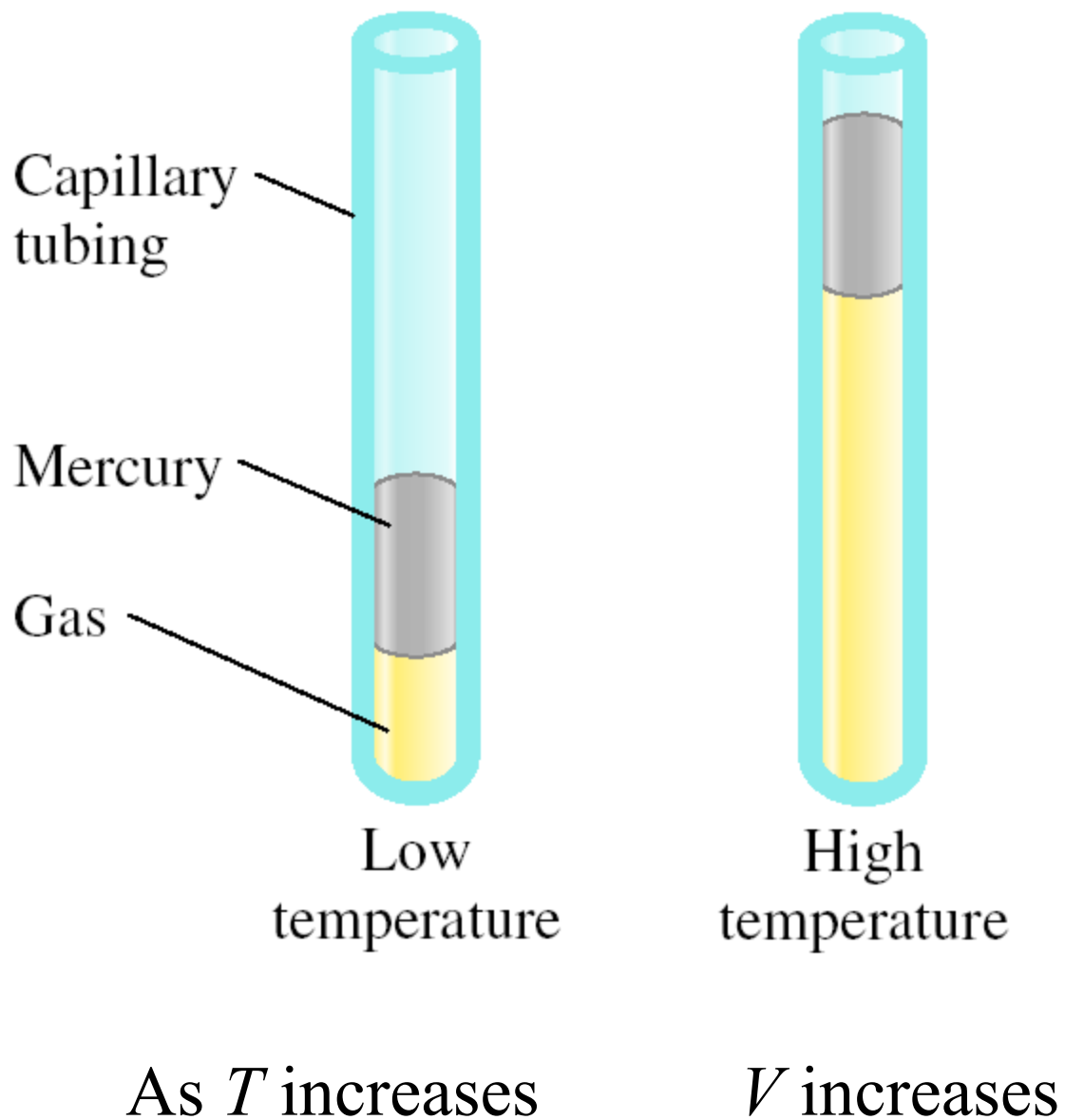
$$P_1 \times V_1 = P_2 \times V_2$$

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

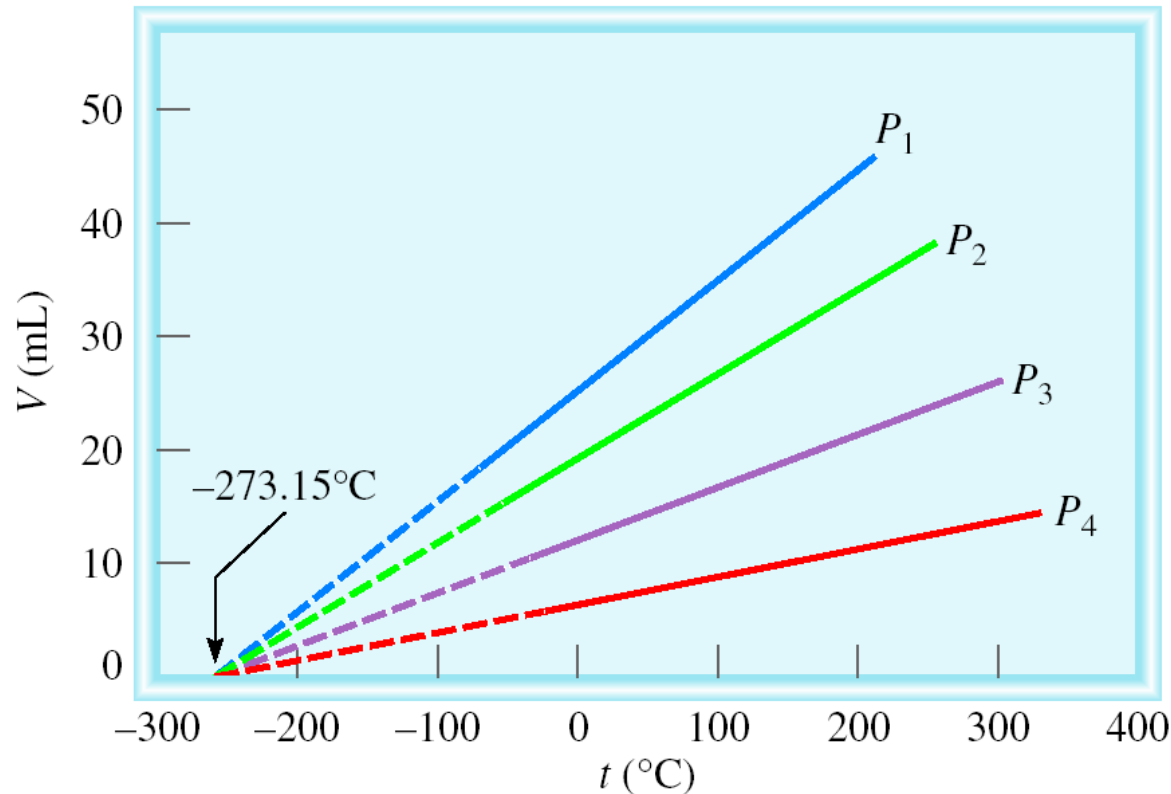
$$V_1 = 946 \text{ mL} \quad V_2 = 154 \text{ mL}$$

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

## Variation in Gas Volume with Temperature at Constant Pressure



# 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

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

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?

$$V_1/T_1 = V_2/T_2$$

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

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

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

$$T_2 = ?$$

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

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

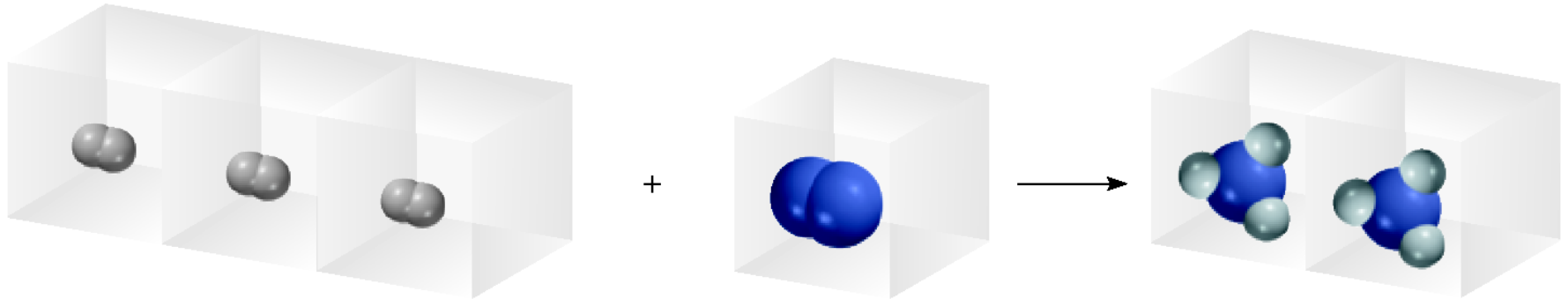
# Avogadro's Law

$V \propto$  number of moles ( $n$ )

Constant temperature  
Constant pressure

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

$$V_1 / n_1 = V_2 / n_2$$



$3\text{H}_2(\text{g})$   
3 molecules  
3 moles  
3 volumes

+  
+  
+  
+

$\text{N}_2(\text{g})$   
1 molecule  
1 mole  
1 volume

→  
→  
→  
→

$2\text{NH}_3(\text{g})$   
2 molecules  
2 moles  
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?



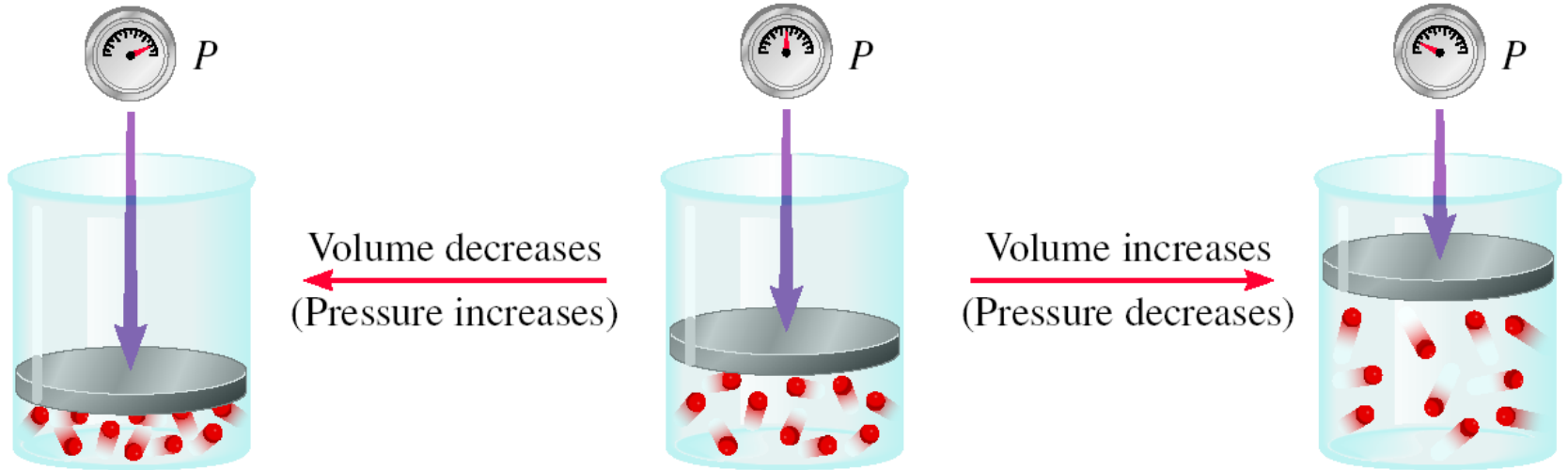
At constant  $T$  and  $P$



# Summary of Gas Laws

## Boyle's Law

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



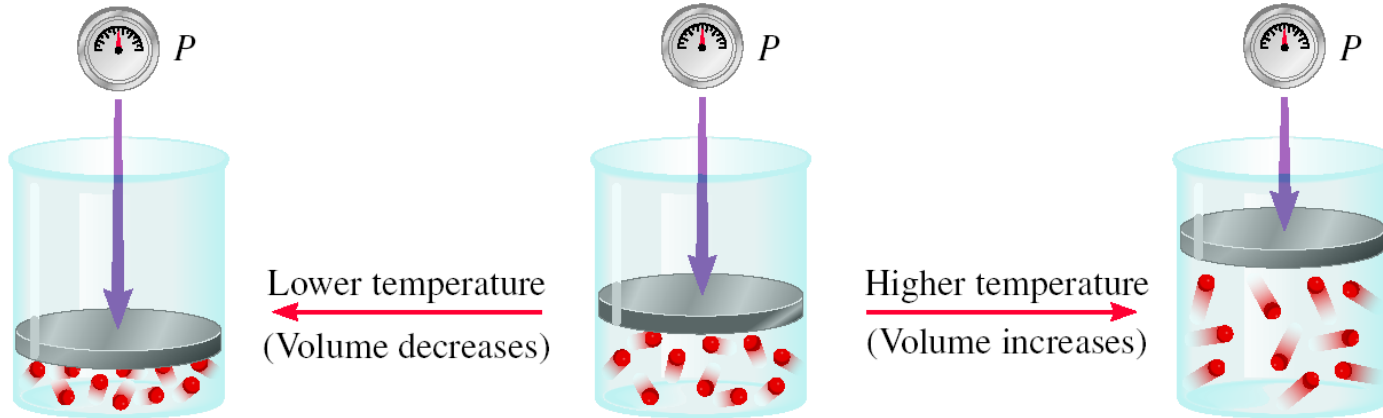
Boyle's Law

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



# Charles Law

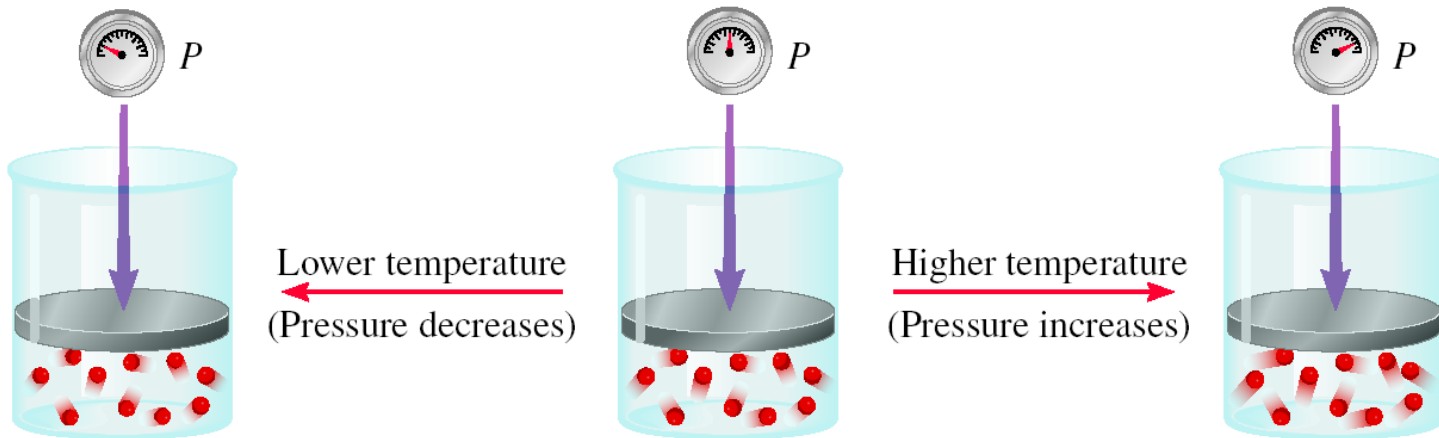
Heating or cooling a gas at constant pressure



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

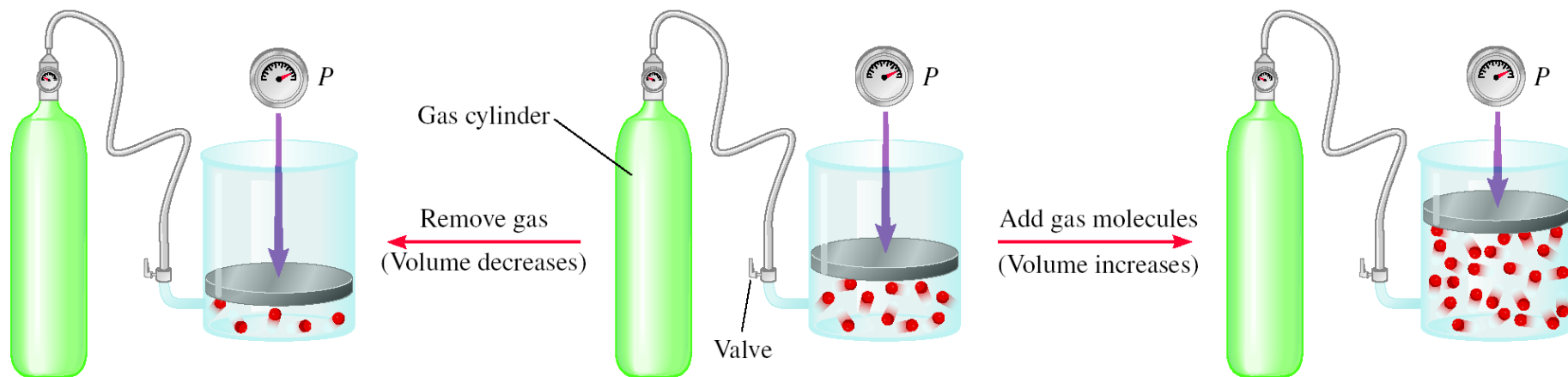


Charles's Law

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

# Avogadro's Law

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



Avogadro's Law

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

# 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  $T$ )

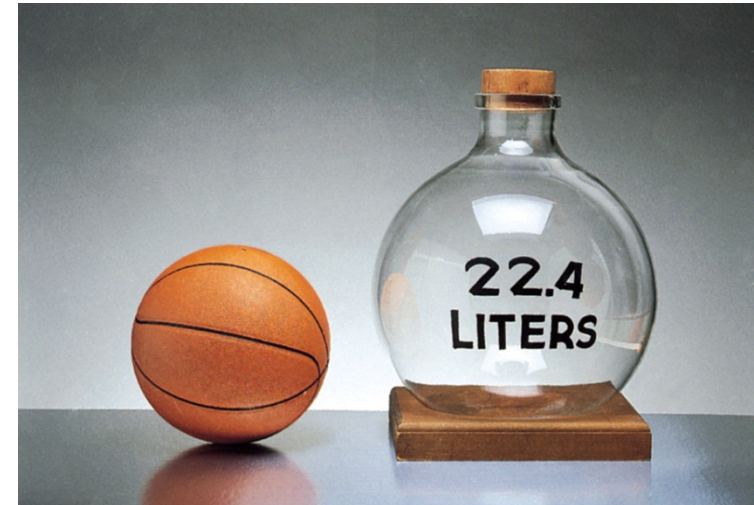
$$V \propto \frac{nT}{P}$$

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

$$PV = nRT$$

The conditions 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.



$$PV = nRT$$

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

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

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

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

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

$$PV = nRT$$

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

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

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

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

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)?

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

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

$$P_1 = 1.20 \text{ atm}$$

$$P_2 = ?$$

$$T_1 = 291 \text{ K}$$

$$T_2 = 358 \text{ K}$$

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

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



## 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 °C. What is the molar mass of the gas?

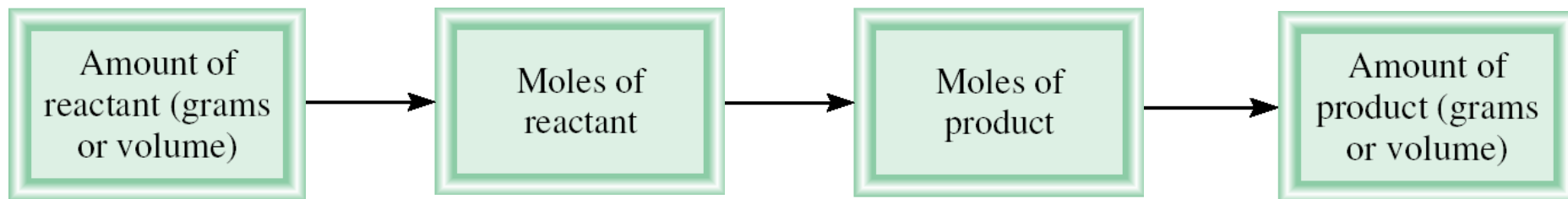
$$M = \frac{dRT}{P} \qquad d = \frac{m}{V} = \frac{4.65 \text{ g}}{2.10 \text{ L}} = 2.21 \frac{\text{g}}{\text{L}}$$

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

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



# Gas Stoichiometry



What is the volume of  $\text{CO}_2$  produced at  $37^\circ\text{C}$  and  $1.00\text{ atm}$  when  $5.60\text{ g}$  of glucose are used up in the reaction:

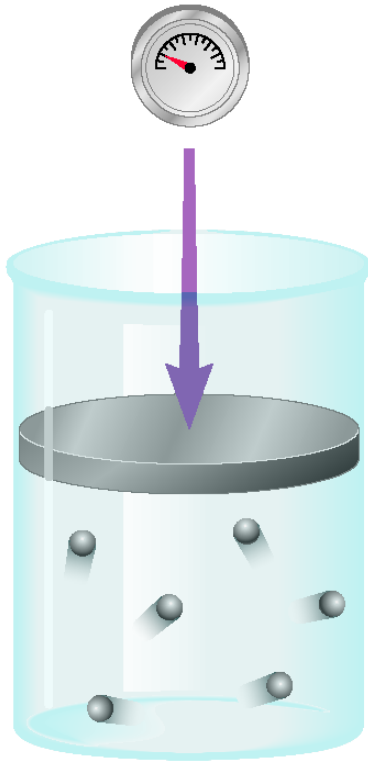


$$5.60 \text{ g } \cancel{\text{C}_6\text{H}_{12}\text{O}_6} \times \frac{1 \cancel{\text{ mol C}_6\text{H}_{12}\text{O}_6}}{180 \text{ g } \cancel{\text{C}_6\text{H}_{12}\text{O}_6}} \times \frac{6 \text{ mol CO}_2}{1 \cancel{\text{ mol C}_6\text{H}_{12}\text{O}_6}} = 0.187 \text{ mol CO}_2$$

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

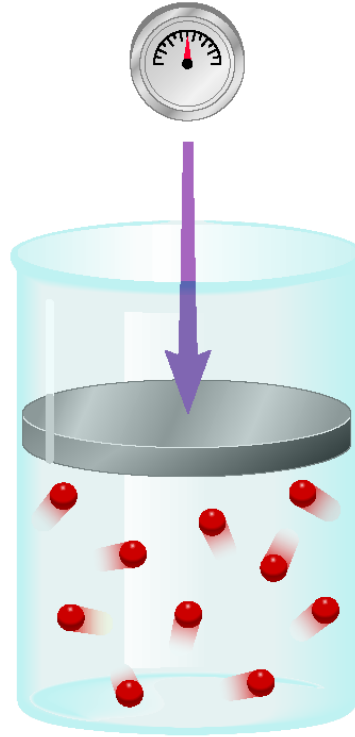
# Dalton's Law of Partial Pressures

$V$  and  $T$  are constant



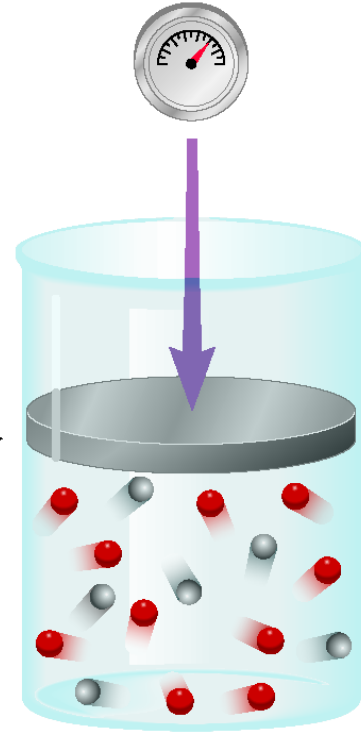
$P_1$

+



$P_2$

Combining  
the gases



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

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**

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

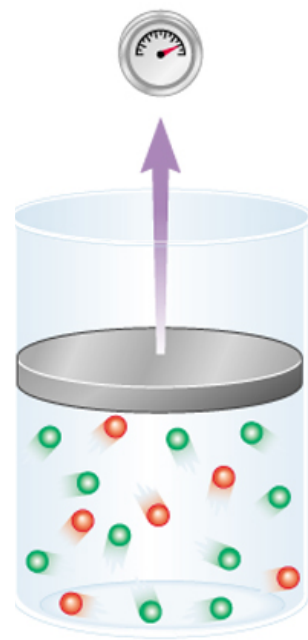
$n_B$  is the number of moles of **B**

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

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

$$P_i = X_i P_T$$

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



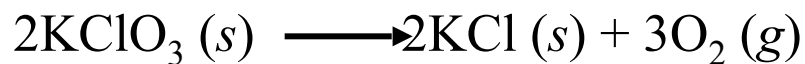
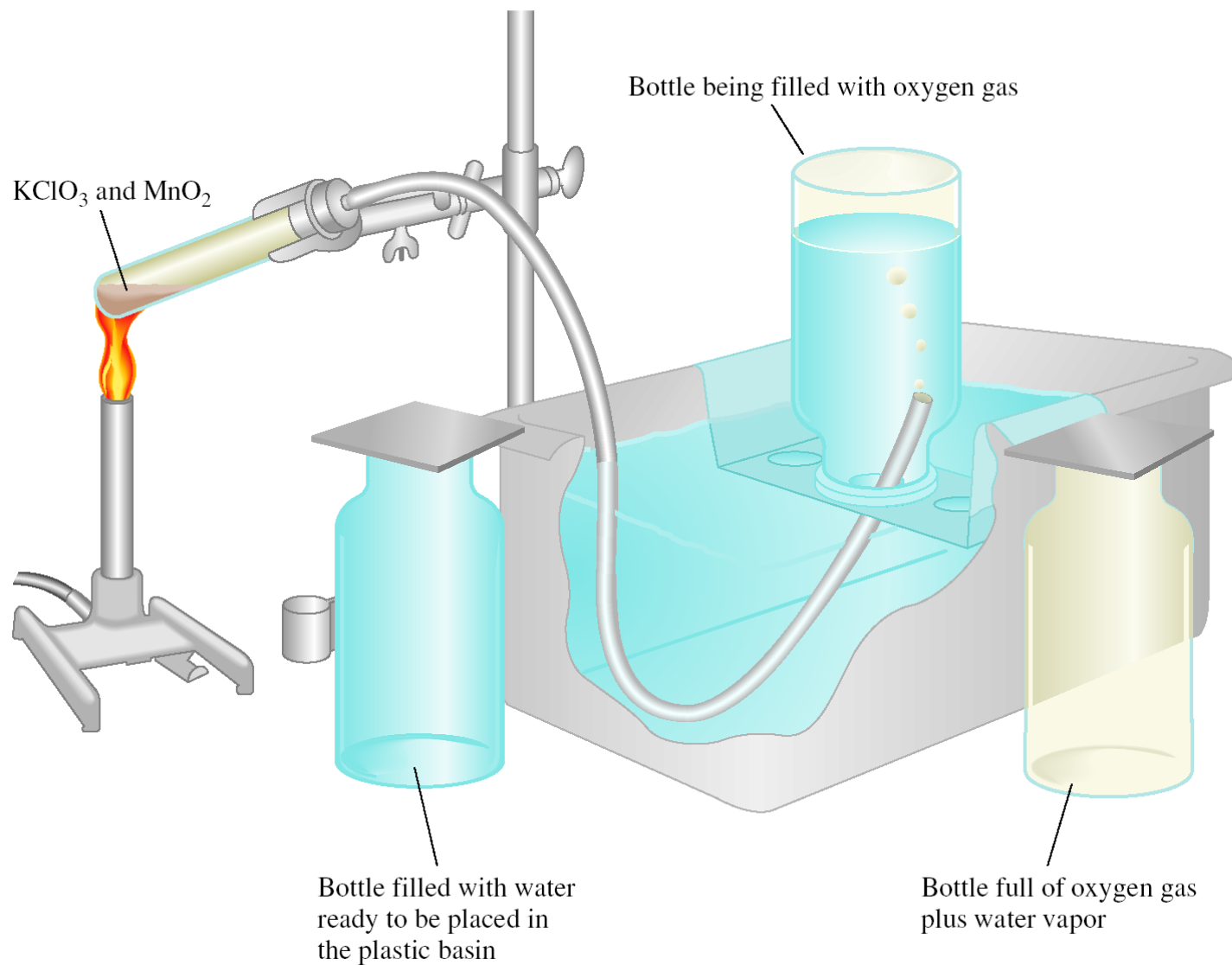
A sample of natural gas contains 8.24 moles of  $\text{CH}_4$ , 0.421 moles of  $\text{C}_2\text{H}_6$ , and 0.116 moles of  $\text{C}_3\text{H}_8$ . If the total pressure of the gases is 1.37 atm, what is the partial pressure of propane ( $\text{C}_3\text{H}_8$ )?

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

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

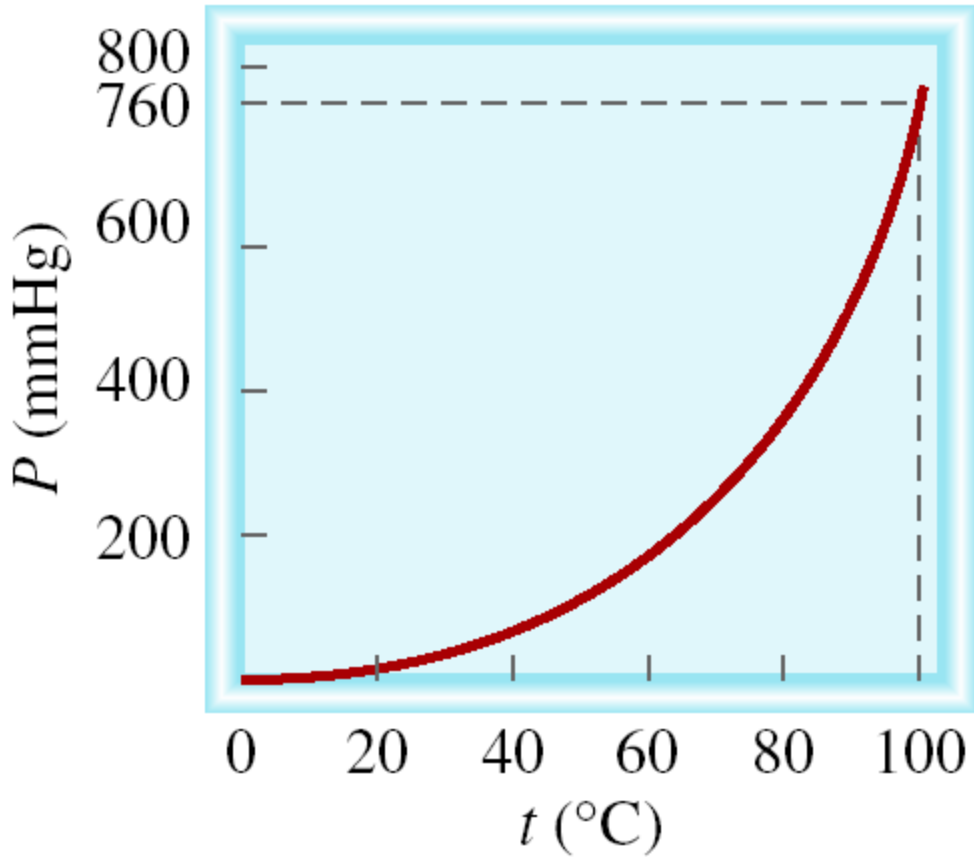
$$P_{\text{propane}} = 0.0132 \times 1.37 \text{ atm} = 0.0181 \text{ atm}$$

# Collecting a Gas over Water



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

# Vapor of Water and Temperature



**TABLE 5.3**

**Pressure of Water Vapor at Various Temperatures**

Temperature ( $^{\circ}\text{C}$ )	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

# Chemistry in Action:

## Scuba Diving and the Gas Laws

Depth (ft)	Pressure (atm)
0	1
33	2
66	3

$P \downarrow$

$V \uparrow$



# 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

$$\overline{\text{KE}} = \frac{1}{2} mu^2$$



# Kinetic theory of gases and ...

- Compressibility of Gases

- Boyle's Law

$P \propto$  collision rate with wall

Collision rate  $\propto$  number density

Number density  $\propto 1/V$

$P \propto 1/V$

- Charles' Law

$P \propto$  collision rate with wall

Collision rate  $\propto$  average kinetic energy of gas molecules

Average kinetic energy  $\propto T$

$P \propto T$

# Kinetic theory of gases and ...

- Avogadro's Law

$P \propto$  collision rate with wall

Collision rate  $\propto$  number density

Number density  $\propto n$

$P \propto n$

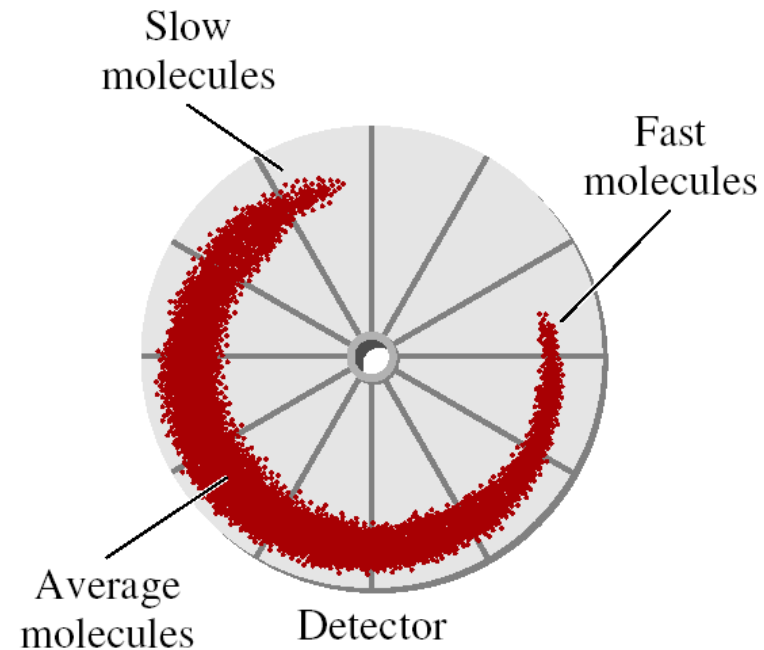
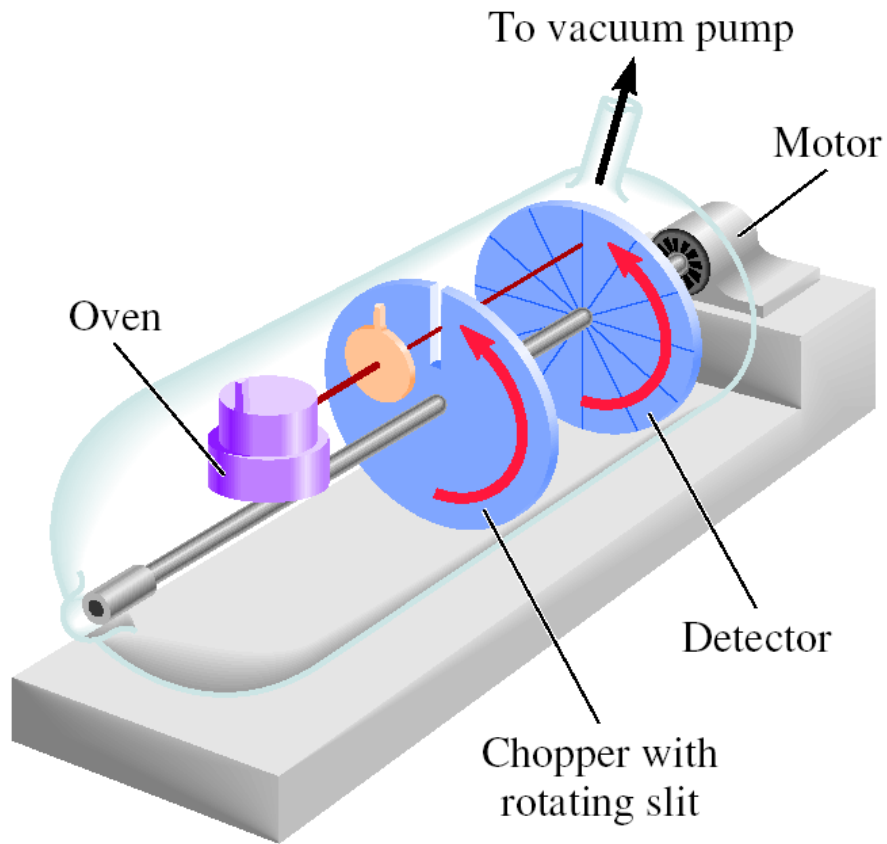
- Dalton's Law of Partial Pressures

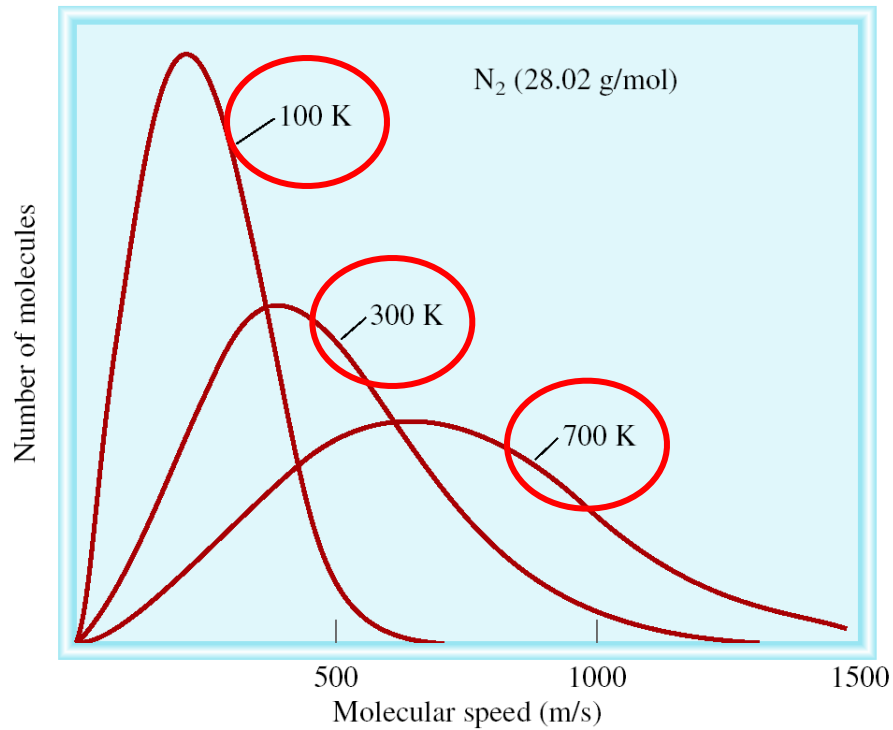
Molecules do not attract or repel one another

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

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

# Apparatus for Studying Molecular Speed Distribution

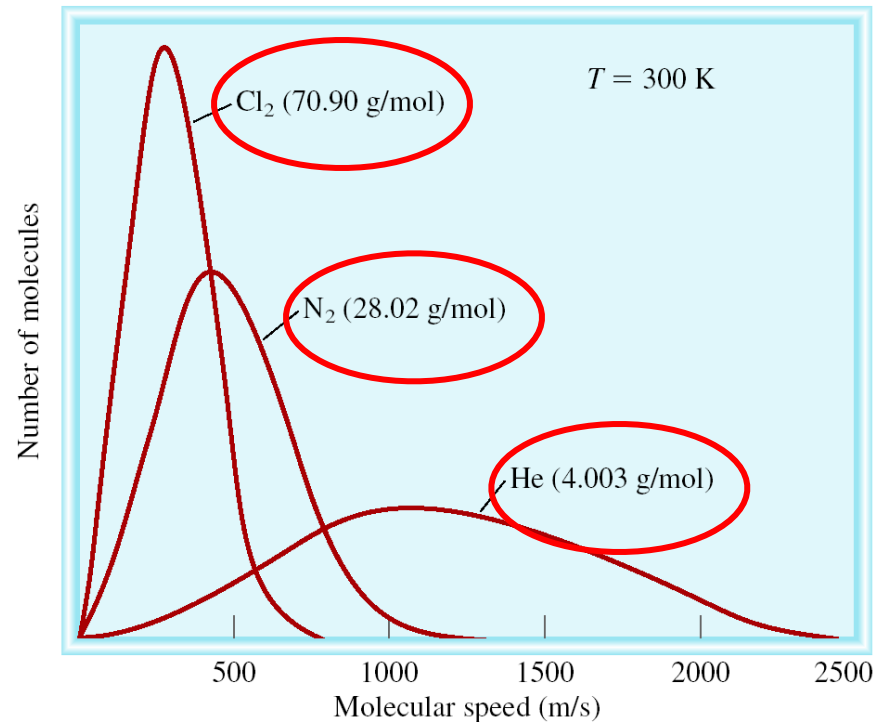




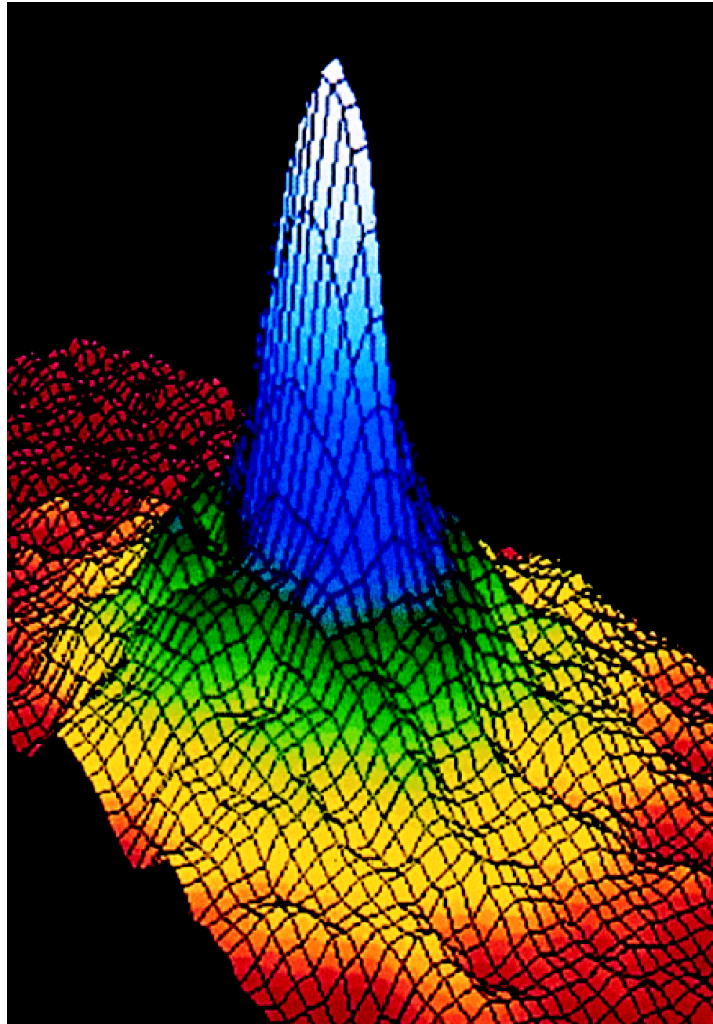
The distribution of speeds for nitrogen gas molecules at three different temperatures

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

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



# Chemistry in Action: Super Cold Atoms



*Maxwell velocity distribution of Rb atoms at about  $1.7 \times 10^{-7}$  K*

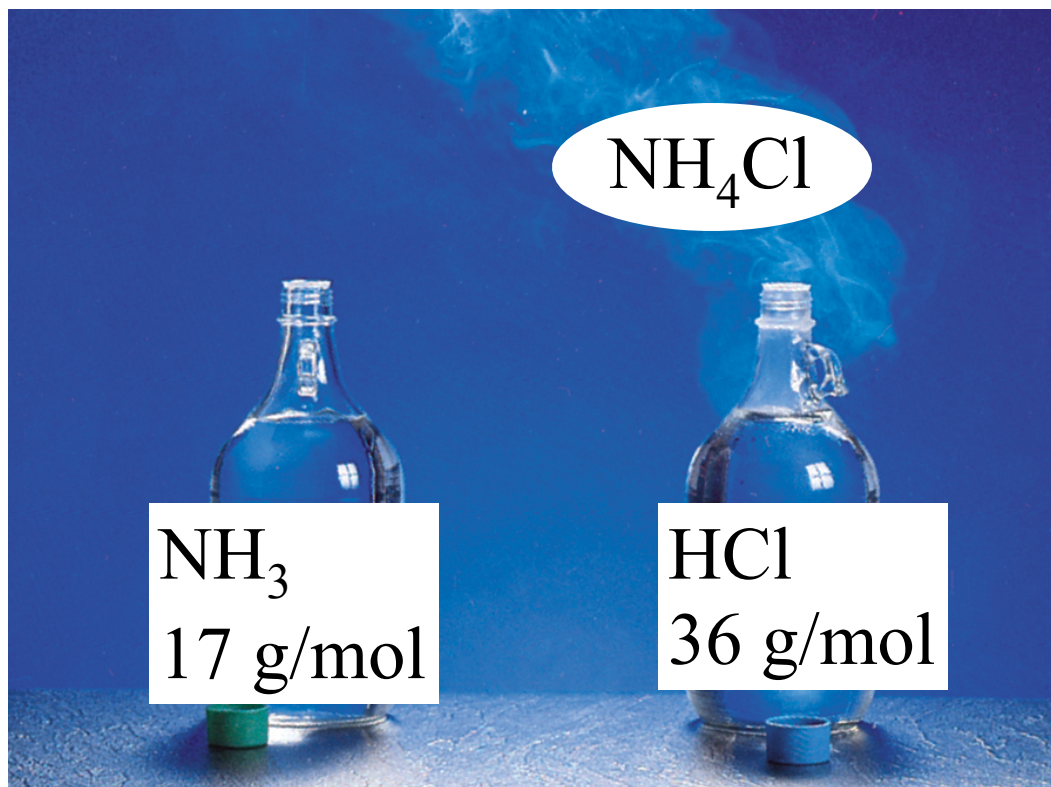
*Bose-Einstein condensate (BEC)*

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

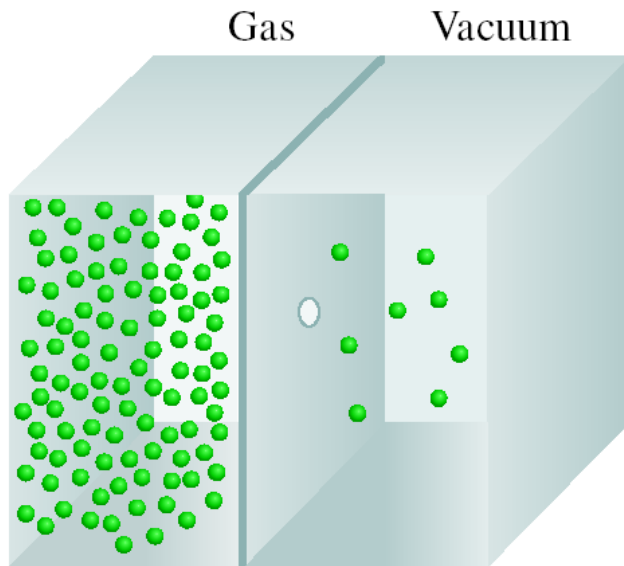


molecular path

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$



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



$$\frac{r_1}{r_2} = \frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}$$

Nickel forms a gaseous compound of the formula  $\text{Ni}(\text{CO})_x$ . What is the value of  $x$  given that under the same conditions methane ( $\text{CH}_4$ ) effuses 3.3 times faster than the compound?

$$r_1 = 3.3 \times r_2$$

$$M_2 = \left(\frac{r_1}{r_2}\right)^2 \times M_1 = (3.3)^2 \times 16 = 174.2$$

$$M_1 = 16 \text{ g/mol}$$

$$58.7 + x \cdot 28 = 174.2$$

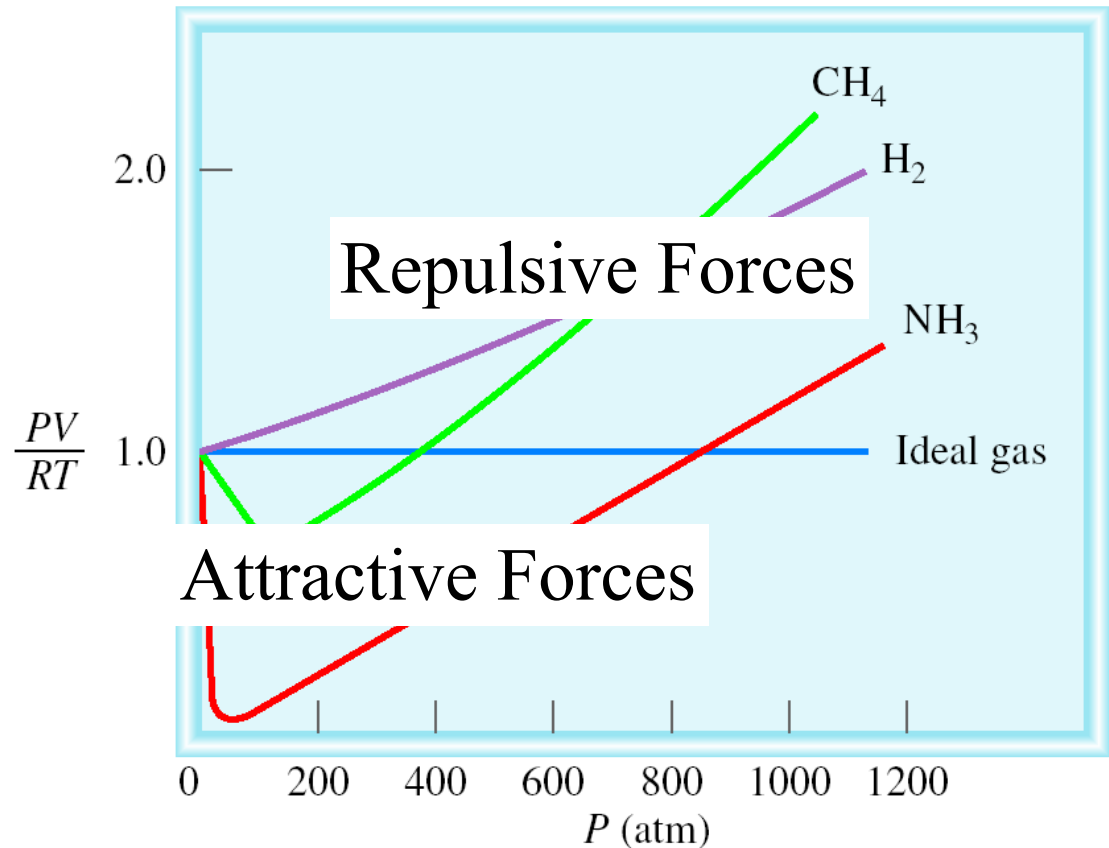
$$x = 4.1 \sim 4$$

# Deviations from Ideal Behavior

1 mole of ideal gas

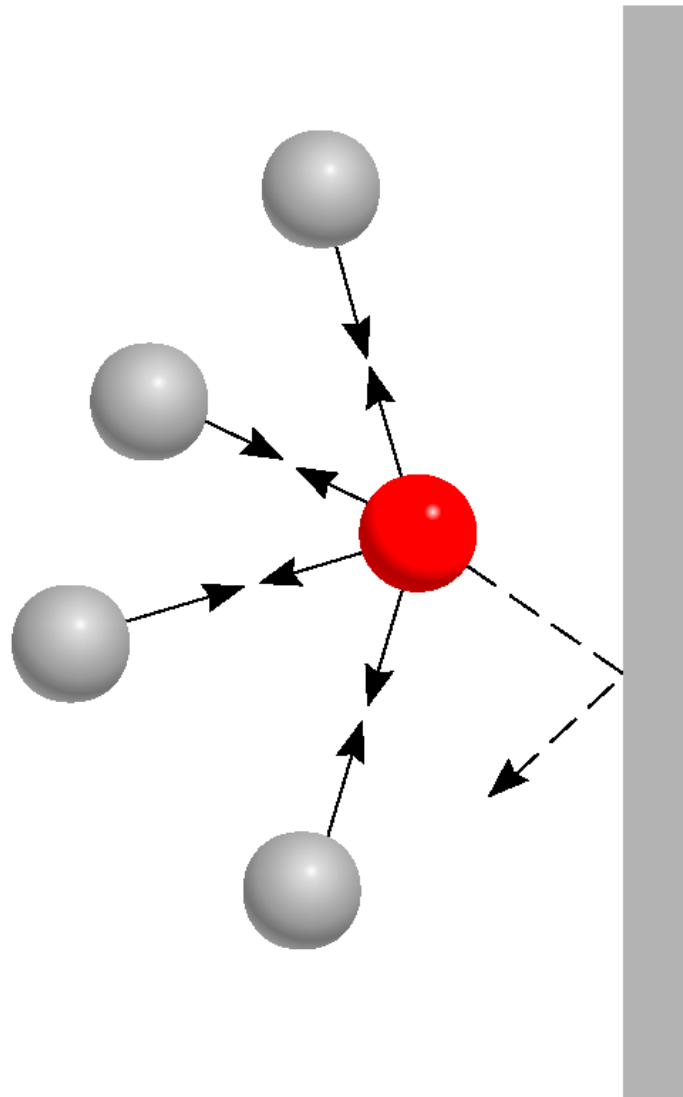
$$PV = nRT$$

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





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



*Van der Waals equation*  
nonideal gas

$$\left( P + \frac{an^2}{V^2} \right) (V - nb) = nRT$$

corrected  
pressure

corrected  
volume

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 <sub>2</sub>	0.244	0.0266
N <sub>2</sub>	1.39	0.0391
O <sub>2</sub>	1.36	0.0318
Cl <sub>2</sub>	6.49	0.0562
CO <sub>2</sub>	3.59	0.0427
CH <sub>4</sub>	2.25	0.0428
CCl <sub>4</sub>	20.4	0.138
NH <sub>3</sub>	4.17	0.0371
H <sub>2</sub> O	5.46	0.0305