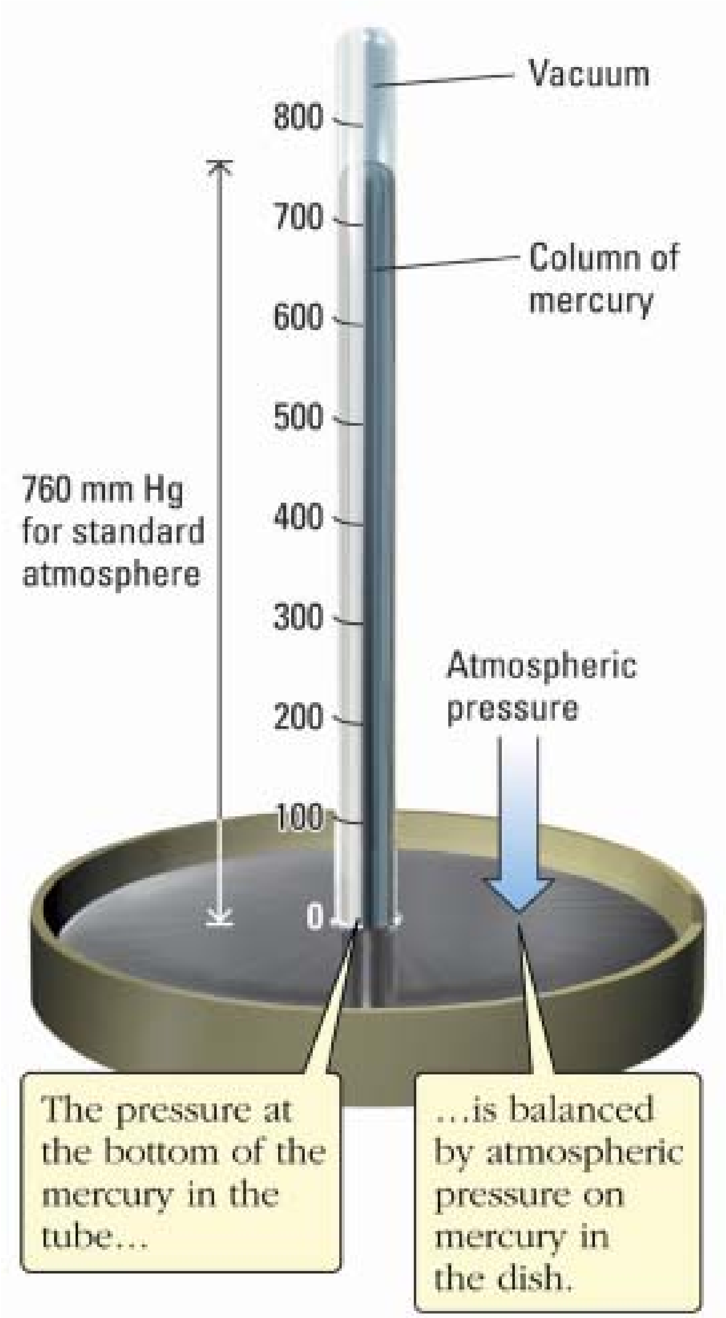
Chapter 10

Gases and the Atmosphere

Torricellian Barometer



Pressure

pressure = force/area

force = mass × acceleration

1 Standard Atmosphere = 1 atm

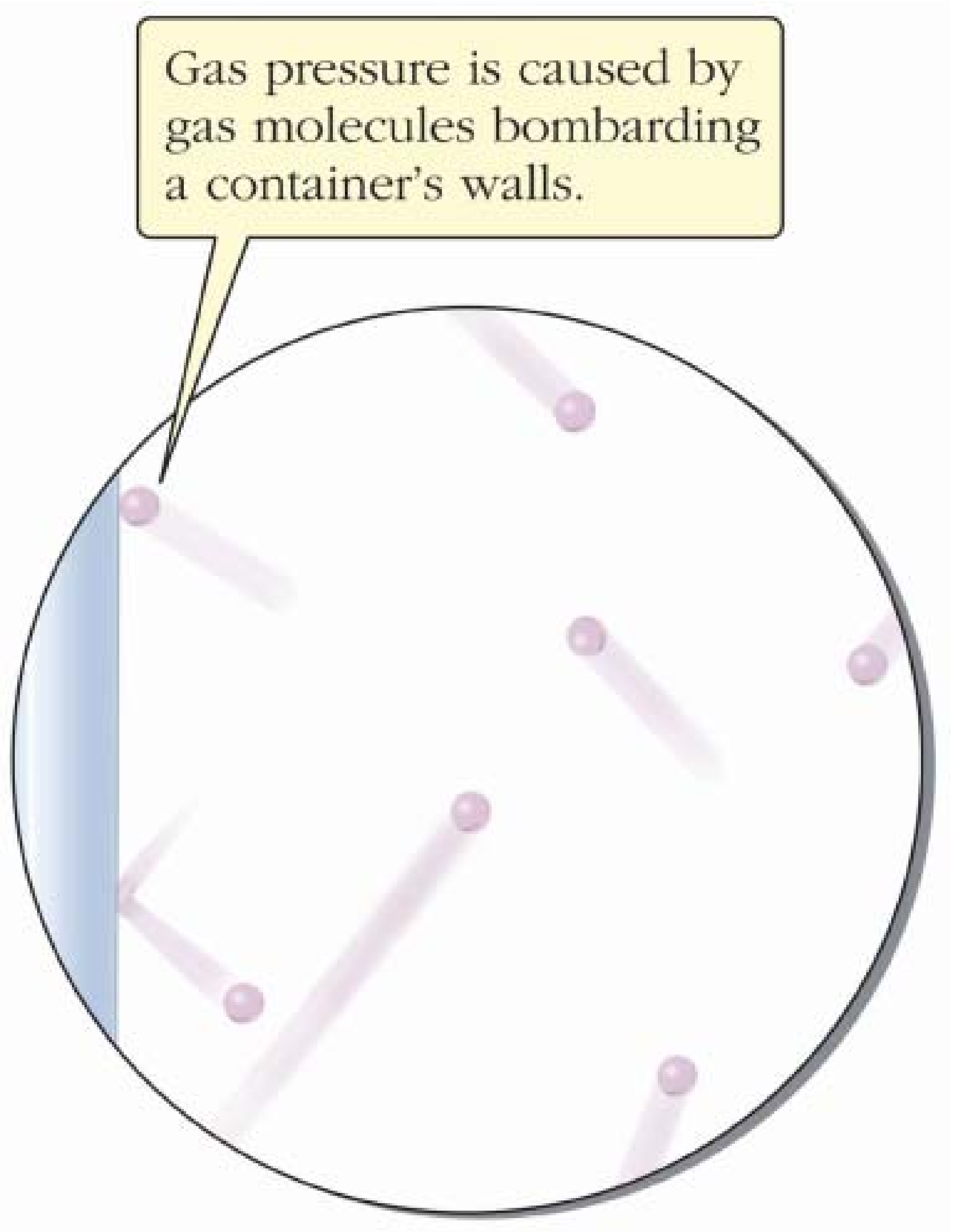
1 atm = 760 mm Hg(exactly)

760 mm Hg(exactly) = 760 torr

760 torr = 101.325 kPa

100 kPa = 1 bar

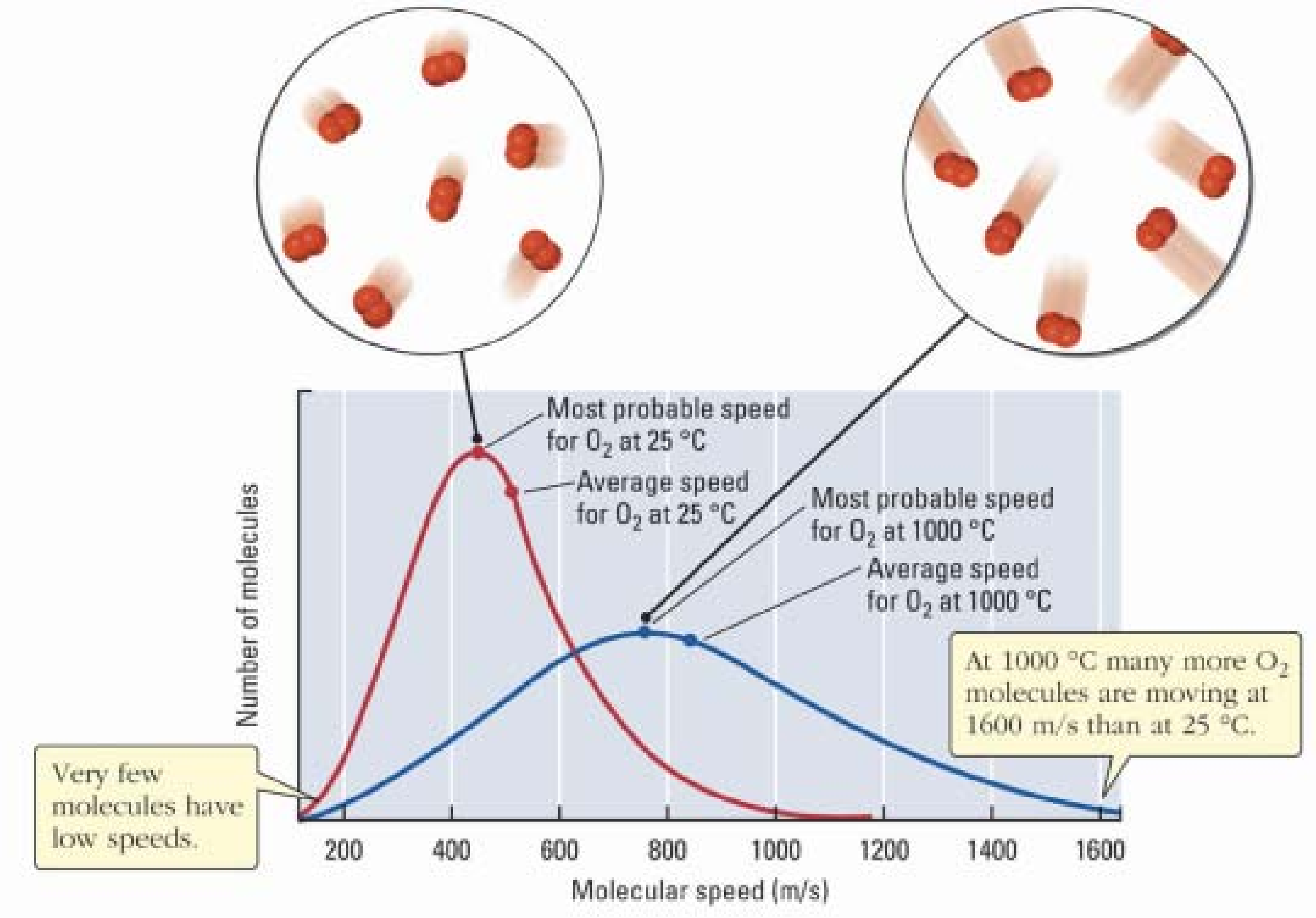
Gas Pressure



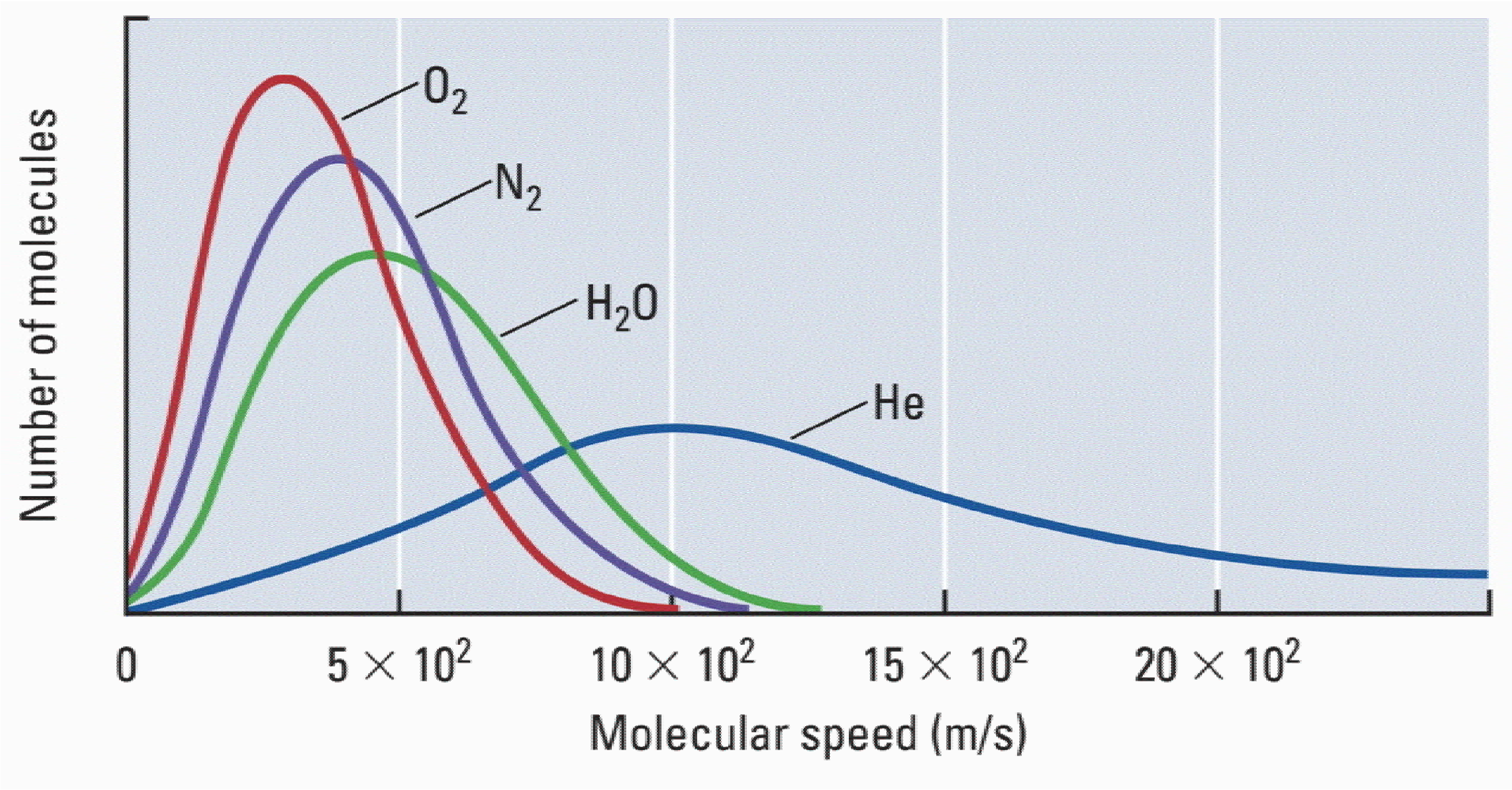
Kinetic Molecular Theory: Gases

- particles in continuous, random, rapid motion
- collisions between particles are elastic
- volume occupied by the particles is negligibly small effect on their behavior
- attractive forces between particles have a negligible effect on their behavior
- gases have no fixed volume or shape, take the volume and shape of the container

Distribution of Molecular Speeds



Molar Mass Effect on Molecular Speeds



Boyle's Law

At constant temperature and mass of gas:

$V \propto 1/P$

$V = a \times 1/P$

where a is a proportionality constant

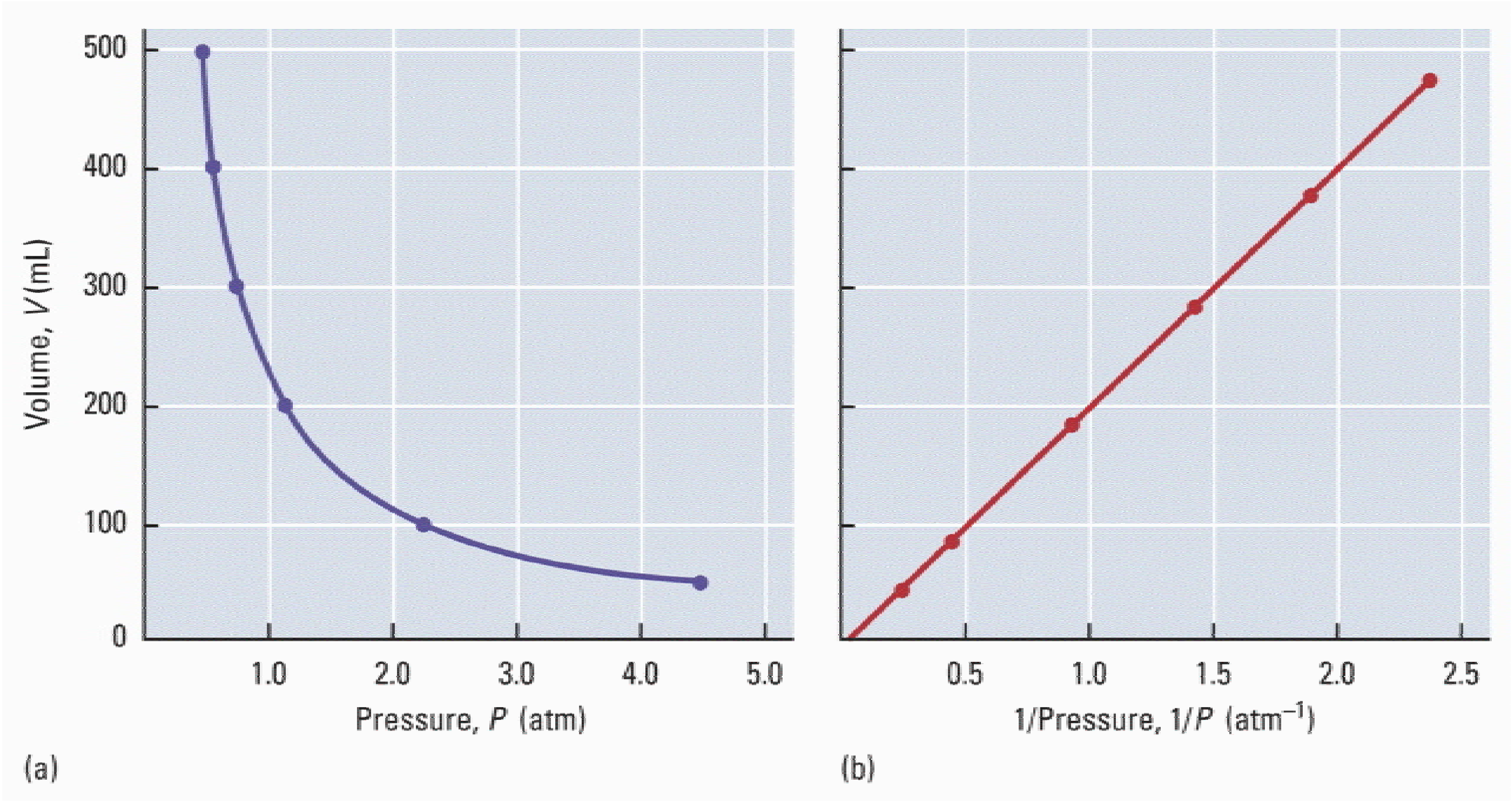
thus

$VP = a$

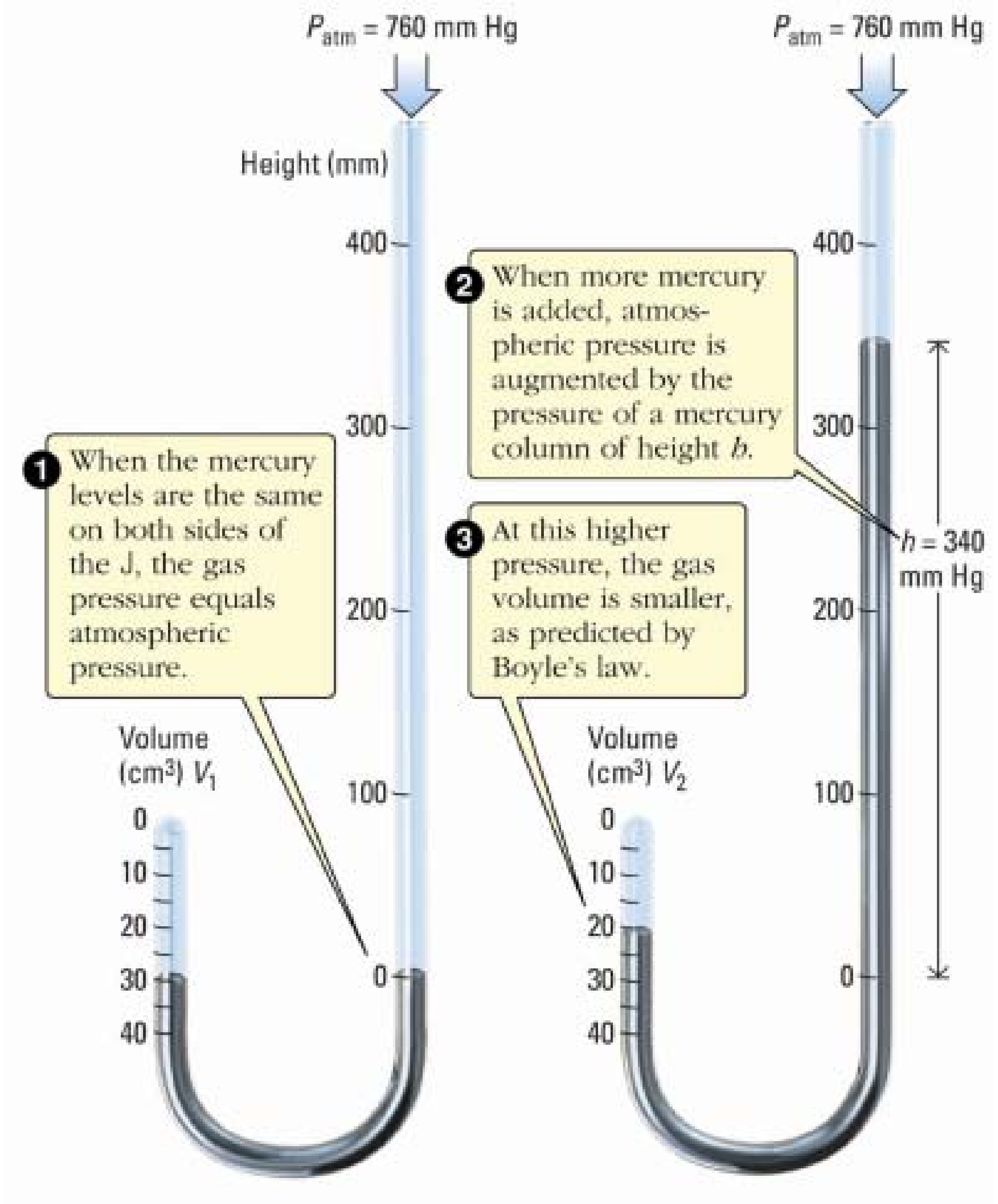
$V_1P_1 = a = V_2P_2$

$$V_1P_1 = V_2P_2$$

Volume and Pressure Plots



Boyle's Law



P_{atm} = 760 mm Hg
P_{atm} = 760 mm Hg
Height (mm)
400
300
200
100
0
400
300
200
100
0
1 When the mercury levels are the same on both sides of the J, the gas pressure equals atmospheric pressure.
2 When more mercury is added, atmospheric pressure is augmented by the pressure of a mercury column of height h.
3 At this higher pressure, the gas volume is smaller, as predicted by Boyle's law.
Volume (cm³) V_1
0
10
20
30
40
Volume (cm³) V_2
0
10
20
30
40
h = 340 mm Hg

Charles's Law

At constant pressure and mass of gas:

$$V \alpha T$$

$$V = b \times T$$

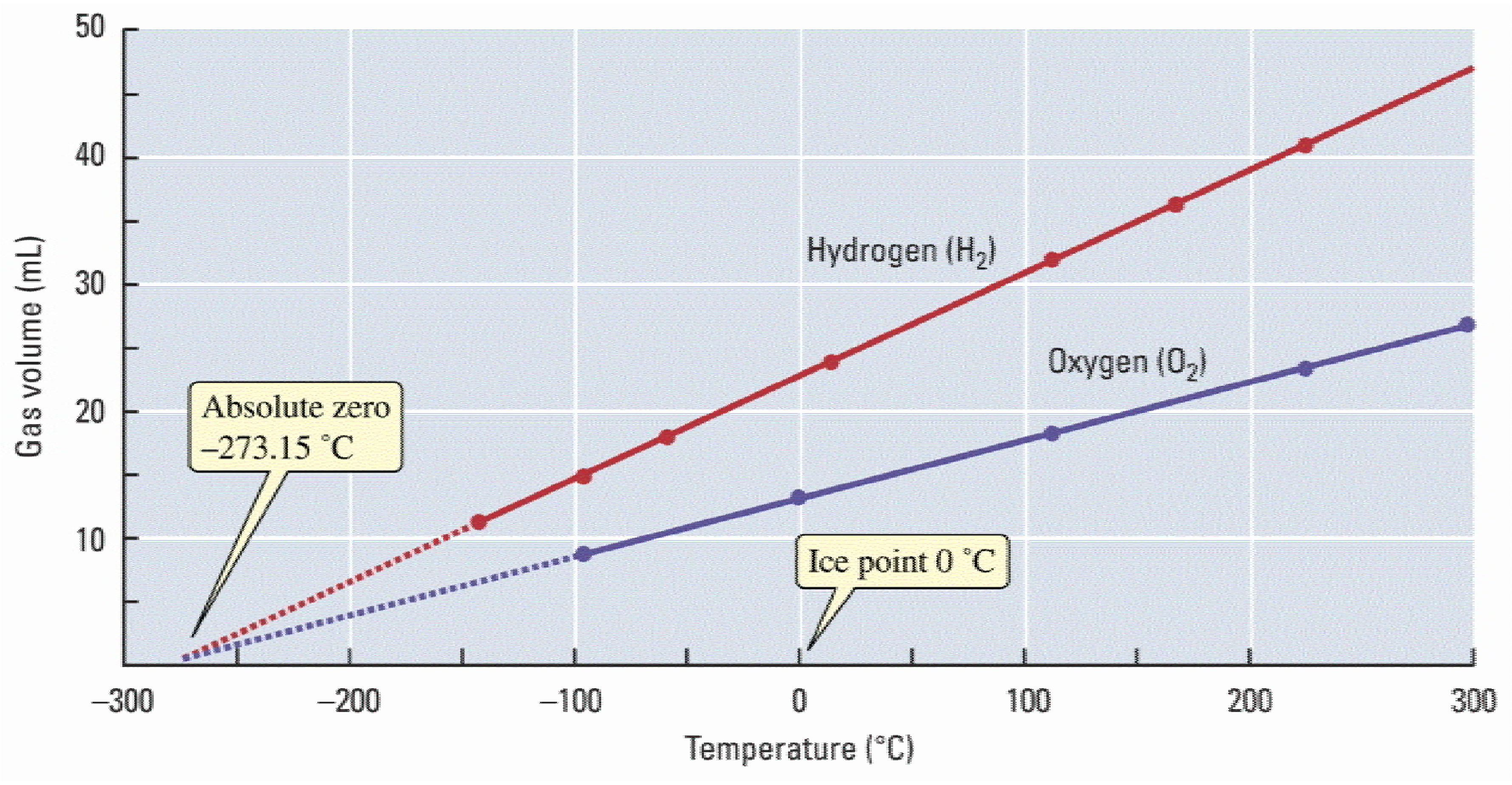
where b is a proportionality constant

$$V/T = b$$

$$V_1/T_1 = b = V_2/T_2$$

$$V_1/T_1 = V_2/T_2$$

Charles's Law



Avogadro's Law

At constant pressure and temperature

$$V \propto n$$

$$V = c \times n$$

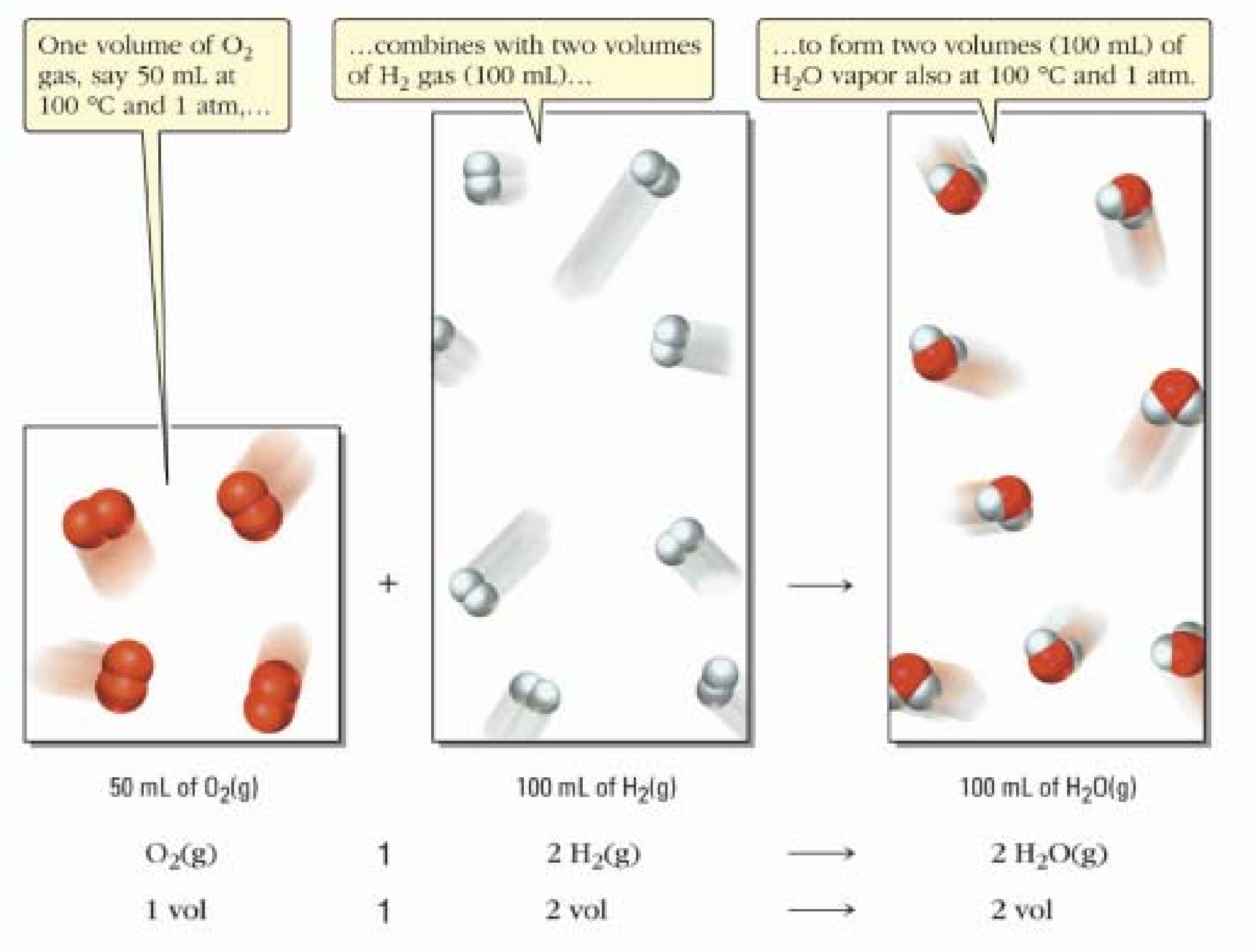
where c is a proportionality constant

$$V/n = c$$

$$V_1/n_1 = c = V_2/n_2$$

$$V_1/n_1 = V_2/n_2$$

Law of Combining Volumes



Combined Gas Law

At constant mass of gas

$V \propto T/P$

$V = d \times (T/P)$

where d is a proportionality constant

$(VP)/T = d$

$V_1P_1 / T_1 = d = V_2P_2 / T_2$

$V_1P_1 / T_1 = V_2P_2 / T_2$

Ideal Gas Law

$V \propto (n \times T)/P$

$V = R \times (n \times T)/P$

where R is proportionality constant

$P \times V = n \times R \times T$

Ideal Gas Constant

TABLE 10.4 Values of *R*, in Different Units

$$R = 0.08206 \frac{\text{L atm}}{\text{mol K}}$$

$$R = 62.36 \frac{\text{torr L}}{\text{mol K}}$$

$$R = 8.314 \frac{\text{kPa dm}^3}{\text{mol K}}$$

$$R = 8.314 \frac{\text{J}}{\text{mol K}}$$

Molecular Weight (M_w, Molar mass) from Gas Density

gas density = #g/V = d

$$PV = nRT$$

where $n = \#g/M_w$

$$PV = (\#g/M_w) \times RT$$

$$M_w = (\#g \times R \times T)/(P \times V)$$

$$M_w = (\#g/V) \times (R \times T)/P = (d \times R \times T)/P$$

Dalton's Law of Partial Pressures

The total pressure of a mixture of gases is equal to the sum of the pressures of the individual gases (partial pressures).

$$P_T = P_1 + P_2 + P_3 + P_4 + \cdot\ \cdot\ \cdot\ \cdot$$

where $P_T \rightarrow$ total pressure

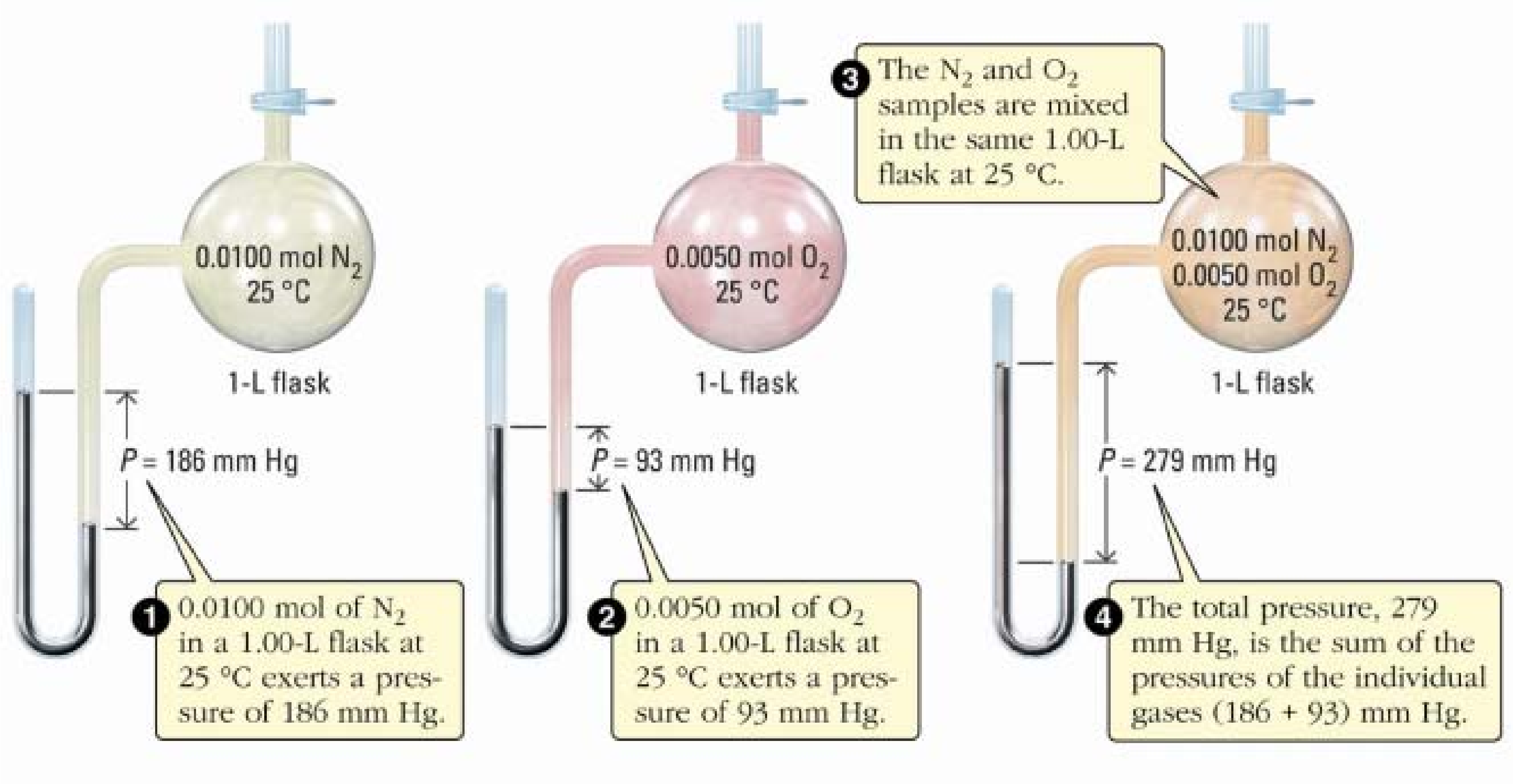
$P_1 \rightarrow$ partial pressure of gas 1

$P_2 \rightarrow$ partial pressure of gas 2

$P_3 \rightarrow$ partial pressure of gas 3

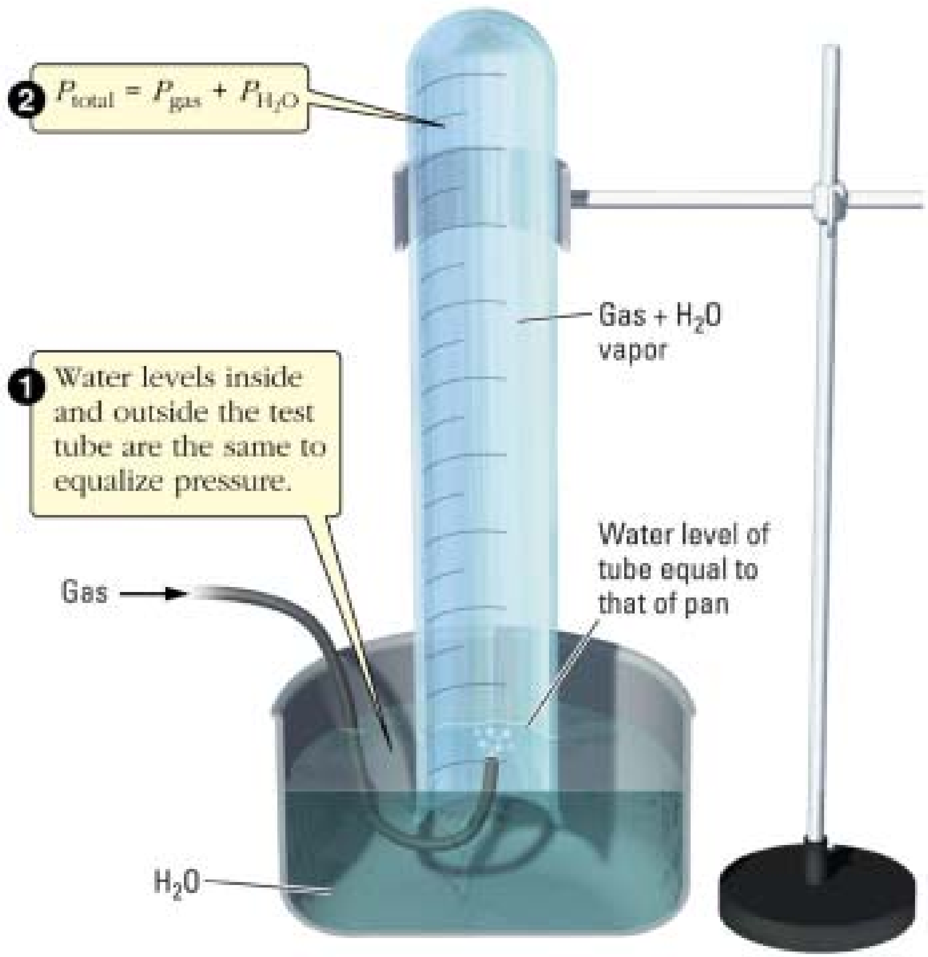
$P_4 \rightarrow$ partial pressure of gas 4

Dalton's Law



The N_2 and O_2 samples are mixed in the same 1.00-L flask at 25 °C.
0.0100 mol N_2
25 °C
1-L flask
P = 186 mm Hg
0.0050 mol O_2
25 °C
1-L flask
P = 93 mm Hg
0.0100 mol N_2
0.0050 mol O_2
25 °C
1-L flask
P = 279 mm Hg
0.0100 mol of N_2 in a 1.00-L flask at 25 °C exerts a pressure of 186 mm Hg.
0.0050 mol of O_2 in a 1.00-L flask at 25 °C exerts a pressure of 93 mm Hg.
The total pressure, 279 mm Hg, is the sum of the pressures of the individual gases (186 + 93) mm Hg.

Collecting a Gas over Water



Example: What volume will 24.0 g O_2 (M_w 32.0) occupy at 25°C and a pressure of 0.888 atm?

$$n = \frac{(24.0\ g)(1\ mol)}{(32.0\ g)} = 0.750\ mol$$

V =?; P = 0.880 atm; T = (25 + 273)K = 298K

V = nRT/P

$$= \frac{(0.750\ mol)(0.08206 L \times atm/mol \times K)(298K)}{0.888 atm}$$

= 20.7 L

Example: A student generates oxygen gas and collects it over water. If the volume of the gas is 245 mL and the barometric pressure is 756 torr at 25^{o}C, what is the volume of the "dry" oxygen gas at STP? (P_{water} = 23.8 torr at 25^{o}C)

$P_{O2} = P_{bar} - P_{water} = (756 - 23.8)$ torr = 732 torr

Example A student generates oxygen gas and collects it over water. If the volume of the gas is 245 mL and the barometric pressure is 756 torr at 25°C, what is the volume of the "dry" oxygen gas at STP?

P_{water} = 23.8 torr at 25°C; $P_{O2} = P_{bar} - P_{water}$ = (756 - 23.8) torr = 732 torr

$P_1 = P_{O2}$ = 732 torr; P_2 = SP = 760. torr

V_1 = 245mL; T_1 = 298K; T_2 = ST = 273K; V_2 = ?

$(V_1P_1/T_1) = (V_2P_2/T_2)$

$$V_2 = (V_1P_1T_2)/(T_1P_2)$$

$$= \frac{(245\text{mL})(732\text{torr})(273\text{K})}{(298\text{K})(760.\text{torr})}$$

$$= 217\text{mL}$$

Real Gases

- have a finite volume at absolute zero
- have attractive forces between gas particles

Van der Waals Equation

$$(P + n^2a/V^2)(V - nb) = nRT$$

where $a \rightarrow$ attractive forces

$b \rightarrow$ residual volume

Real versus Ideal Gases