Ch.5 Gases

The Greeks：

four fundamental element of nature：

Air, earth, water and fire

Ch.5 Gases

5-1： Measurements on gases

- 1. Volume , Amount, and Temperature

5-2： The ideal gas law

- 1. Volume is directly proportional to amount.
- 2.Volume is directly proportional to absolute temperature.
- 3. Volume is inversely proportional to pressure .

5-3： Gas law calculation

- 1. Final and initial State Problems
- 2. Calculation of P,V, n, or T
- 3. Molar Mass and Density

5-4： Stoichiometry of gaseous reactions

5-5： Gas mixtures： partial pressures and mole fractions

- 1. Wet Gases；Partial Pressure of Water
- 2. Partial Pressure and Mole Fraction

5-6： Kinetic theory of Gases

- 1. Molecular Model
- 2. Expression for Pressure P
- 3. Average Kinetic Energy of Translational Motion,Et
- 4. Averag Speed ,u
- 5.Effusion of Gases ；Graham,s Law
- 6. Distribution of Molecular Speeds

5-7： Real gases

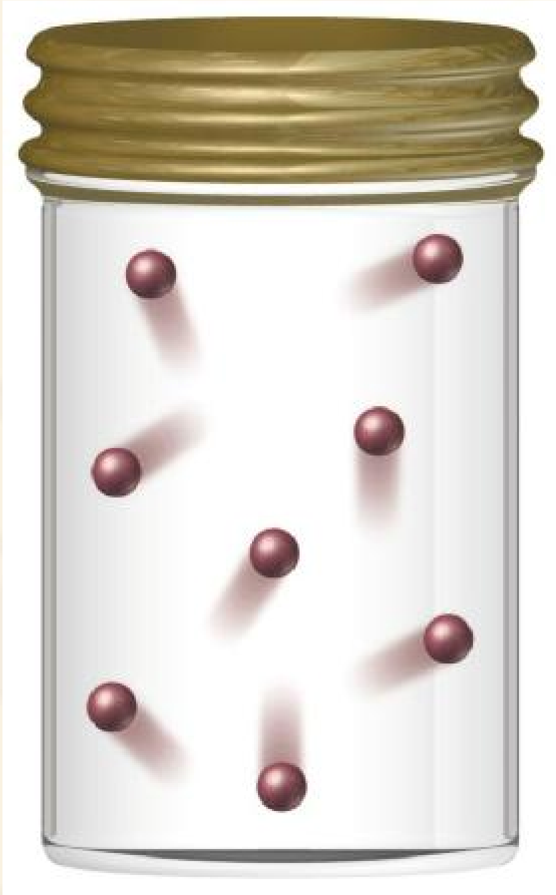
- 1. Attractive forces
- 2. Particle Volume

Gas Properties Explained

- Gases have indefinite shape and volume because the freedom of the molecules allows them to move and fill the container they're in
- Gases are compressible and have low density because of the large spaces between the molecules

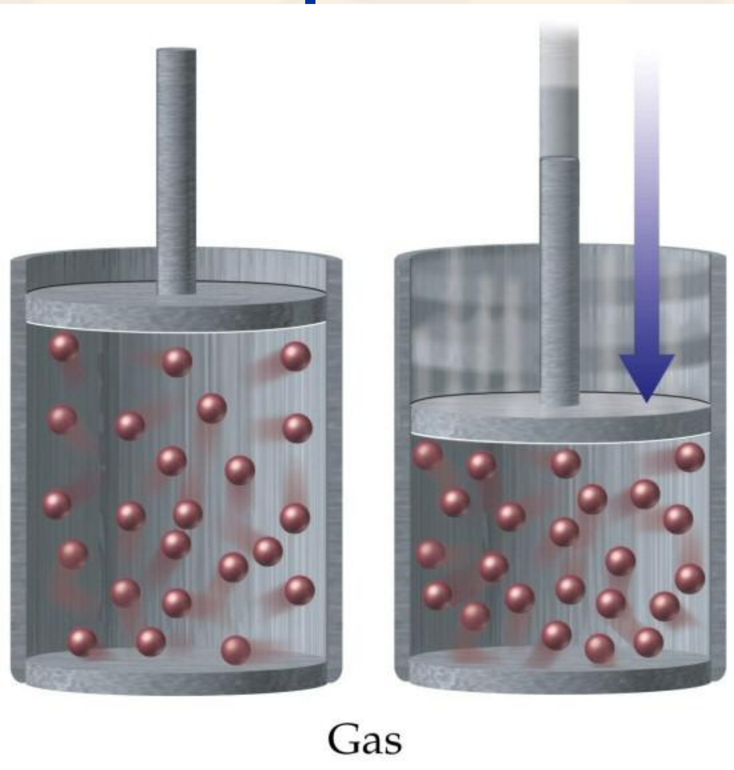
Properties – Indefinite Shape and Indefinite Volume

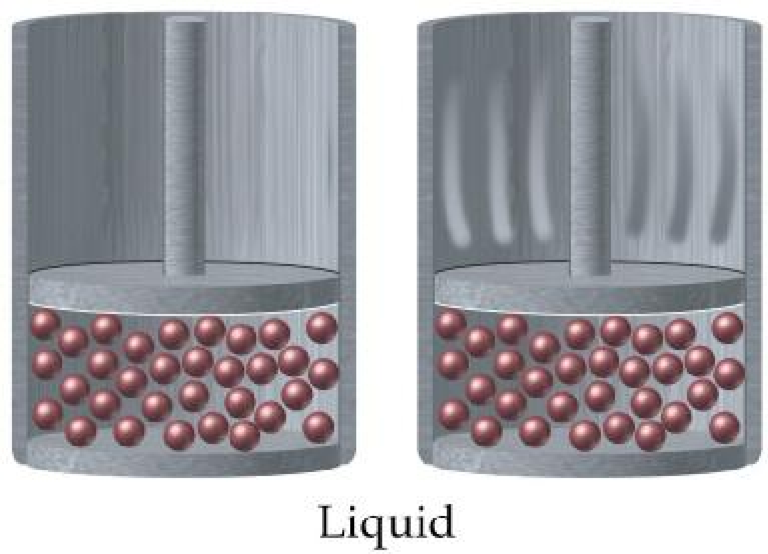
Because the gas molecules have enough kinetic energy to overcome attractions, they keep moving around and spreading out until they fill the container



As a result, gases take the shape and the volume of the container they are in.

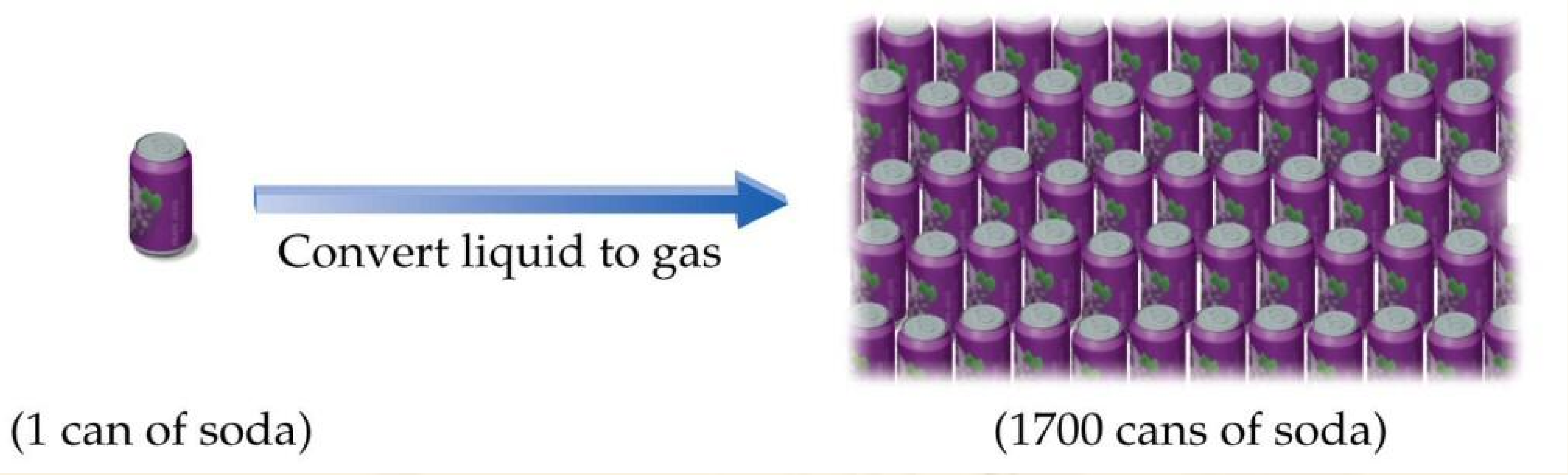
Properties - Compressibility





Because there is a lot of unoccupied space in the structure of a gas, the gas molecules can be squeezed closer together

Properties – Low Density



Because there is a lot of unoccupied space in the structure of a gas, gases have low density

The Pressure of a Gas

- result of the constant movement of the gas molecules and their collisions with the surfaces around them
- the pressure of a gas depends on several factors
  - number of gas particles in a given volume
  - volume of the container
  - average speed of the gas particles

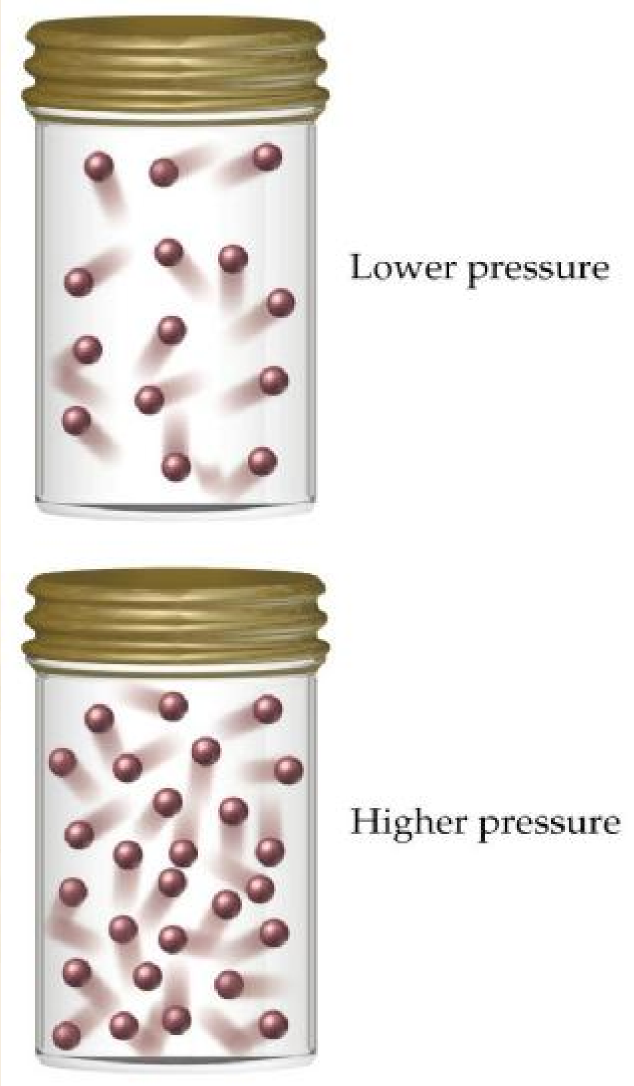


TABLE 5.1

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.

§ 5-1 Measurements on Gases (V, n, T, P)氣體的測定

Volume

™ $1L=10^3cm^3=10^{-3}m^3$

mass

™ $W=Mw\times n$

Temperature

many calculation involving the physical behavior of gases. Temperatures must be expressed on the Kelvin scale

$$T_K = T_{℃} + 273$$

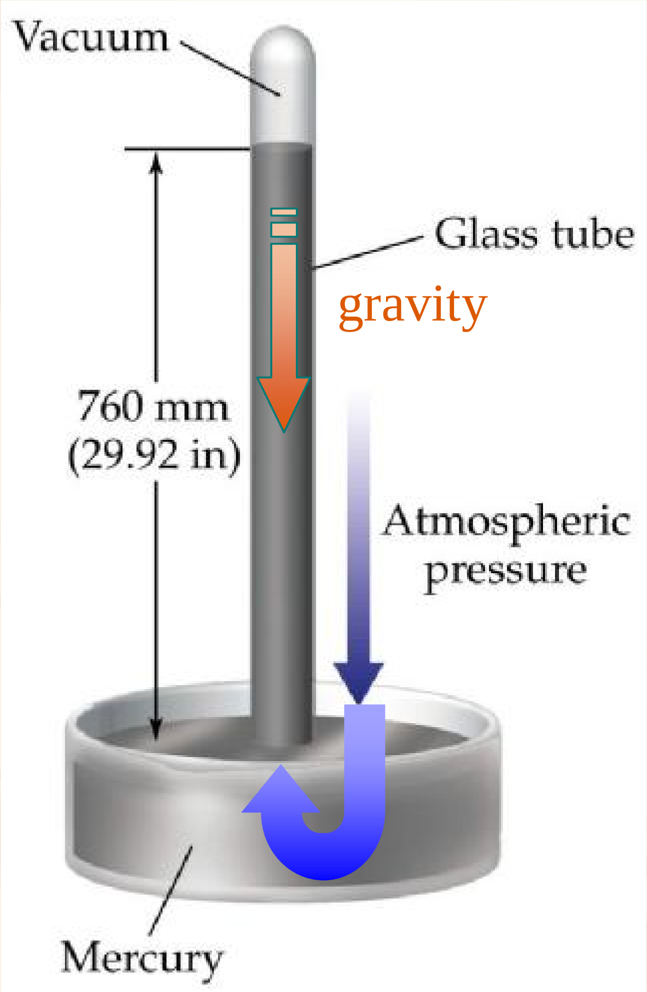
Pressure

Force per unit area

unit： psi（pound per square inch)

Measuring Air Pressure

- use a mercury **barometer (fig5.1), Torricelli**
- The pressure exerted by the mercury column exactly equals that of the atmosphere.
- force of the air on the surface of the mercury balanced by the pull of gravity on the column of mercury



Express gas pressure

1. Gas pressure is measured , it is often express in Millimeters of mercury (mmHg)
2. Gas pressure is the standard atmosphere, (atm) this is the pressure exerted by a column of mercury 760 mm high with the mercury at O℃
3. The standard unit of pressure is the pascal (Pa)

1.013bar=1atm=760mmHg

Common Units of Pressure

Unit	Average Air Pressure at Sea Level
pascal (Pa)	101,325
kilopascal (kPa)	101.325
atmosphere (atm)	1 (exactly)
millimeters of mercury (mmHg)	760 (exactly)
inches of mercury (inHg)	29.92
torr (torr)	760 (exactly)
pounds per square inch (psi, lbs./in^2)	14.7

Ex 5.1：A balloon with a volume V = 2.06 L contains 0.368 g of Helium at 22℃ and 1.08 atm. Express

(a). the volume of the balloon in cubic meters,

(b). the amount of He in mole（n）

(c). the temperature（T）in K, and

(d). the pressure （P）in both bars and millimeters of mercury

- Sol：

a). $$2.06 \times \frac{10^{-3} m^3}{1L} = 2.06 \times 10^{-3} m^3$$

b). $$n_{He} = 0.368 \times \frac{1mol}{4.003\ g} = 0.0919\ mol$$

- c). T= 22 + 273 = 295 K

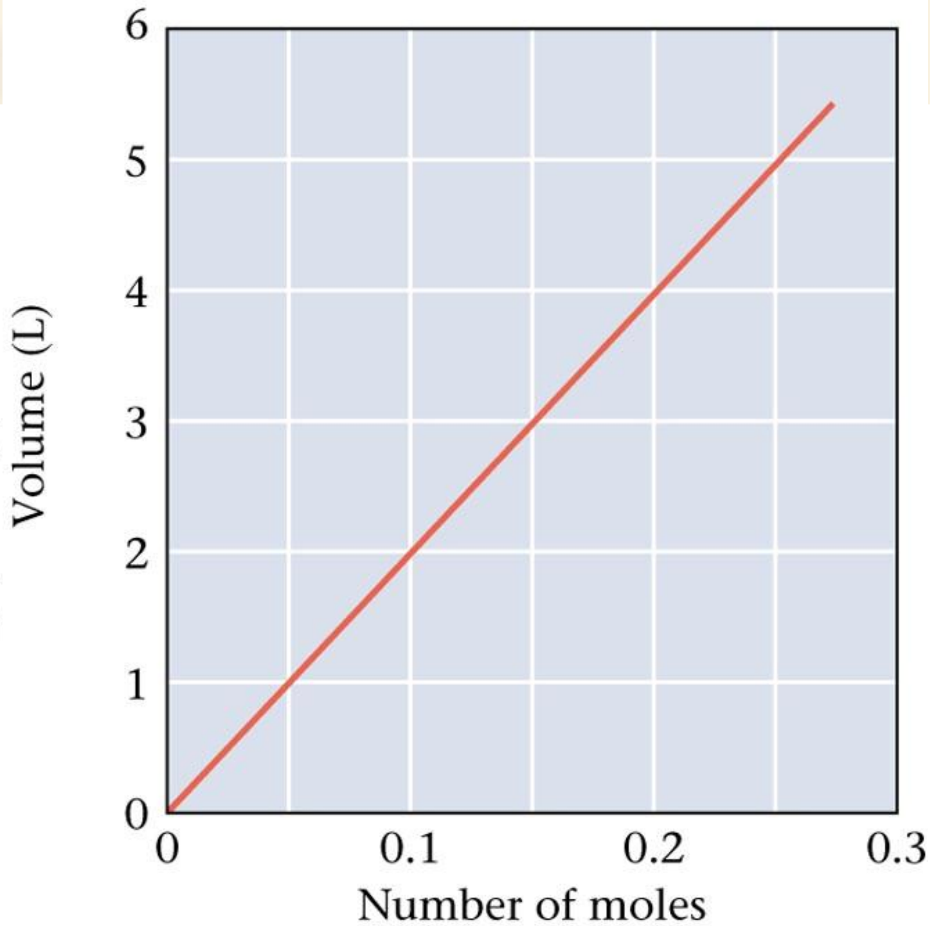
d). $$P = 1.08\, atm \times \frac{760\ mmHg}{1atm} = 821\, mmHg$$

$$P = 1.08\, atm \times \frac{1.013\, bar}{1atm} = 1.09\, bar$$

§ 5-2 The ideal gas law

1. Volume is directly proportional to amount.體積與量是成正比 (Avogadro's law)
   - $V=k_1 \times n$ (constant T，P)
2. Volume is directly proportional to absolute temperature 查理定律 (Charles's law)
   - $V=k_2 \times T$ (constant n，P)
3. Volume is inversely proportional to pressure.
   - 波以耳定律 (Boyle law)
   - $V=k_3/P$ (constant n，T)

$$V = constant \times \frac{n \times T}{P}$$



(a)

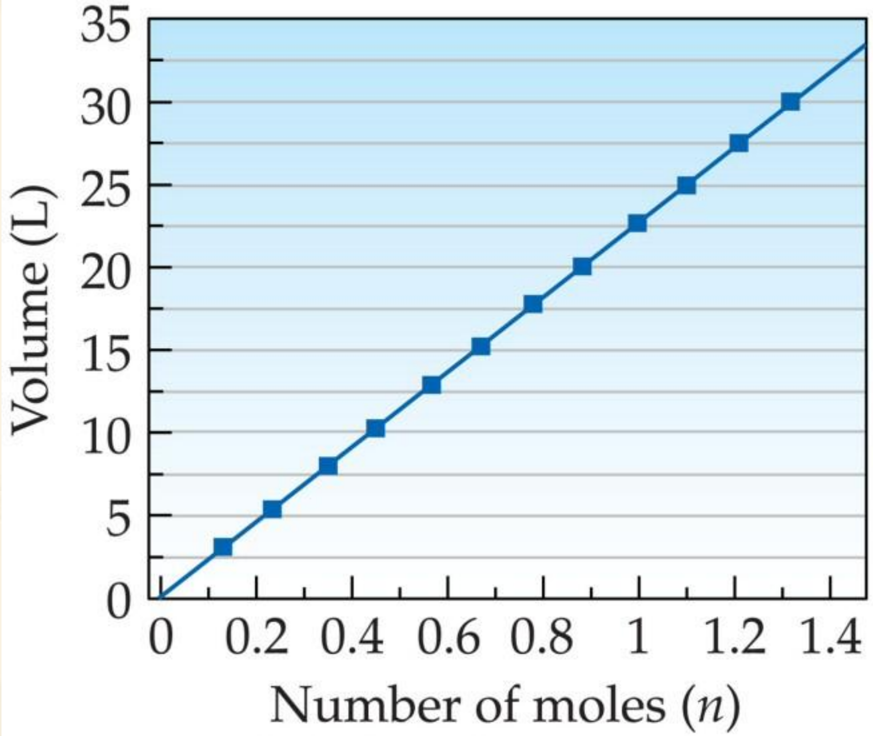
(b)

Avogadro's Law

- volume directly proportional to the number of gas molecules
  - **V** = constant x **n**
  - constant P and T
  - more gas molecules = larger volume
- count number of gas molecules by **moles**
- equal volumes of gases contain equal numbers of molecules
  - ***the gas doesn't matter***

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Avogadro’s Law



Example:	Information
A 4.8 L sample of helium gas contains 0.22 mol helium. How many additional moles of helium must be added to obtain a volume of 6.4 L?	Given: V_1 = 4.8 L, n_1 =0.22 mol V_2 = 6.4 L Find: n_2, mol and added mol Eq'n: $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

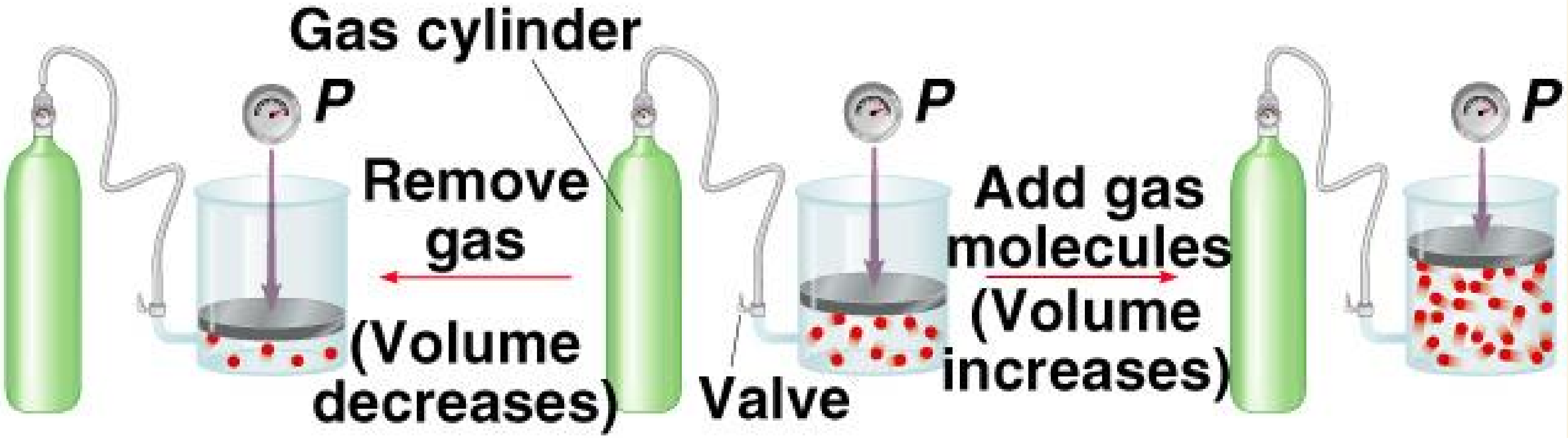
- Write a Solution Map:

$$V_1, V_2, n_1 \xrightarrow{\frac{V_1}{n_1} = \frac{V_2}{n_2}} n_2$$

when using this equation, the units of V_1 and V_2 must be the same, or you will have to convert one to the other

the units of n_1 and n_2 must be moles

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



Avogadro's Law

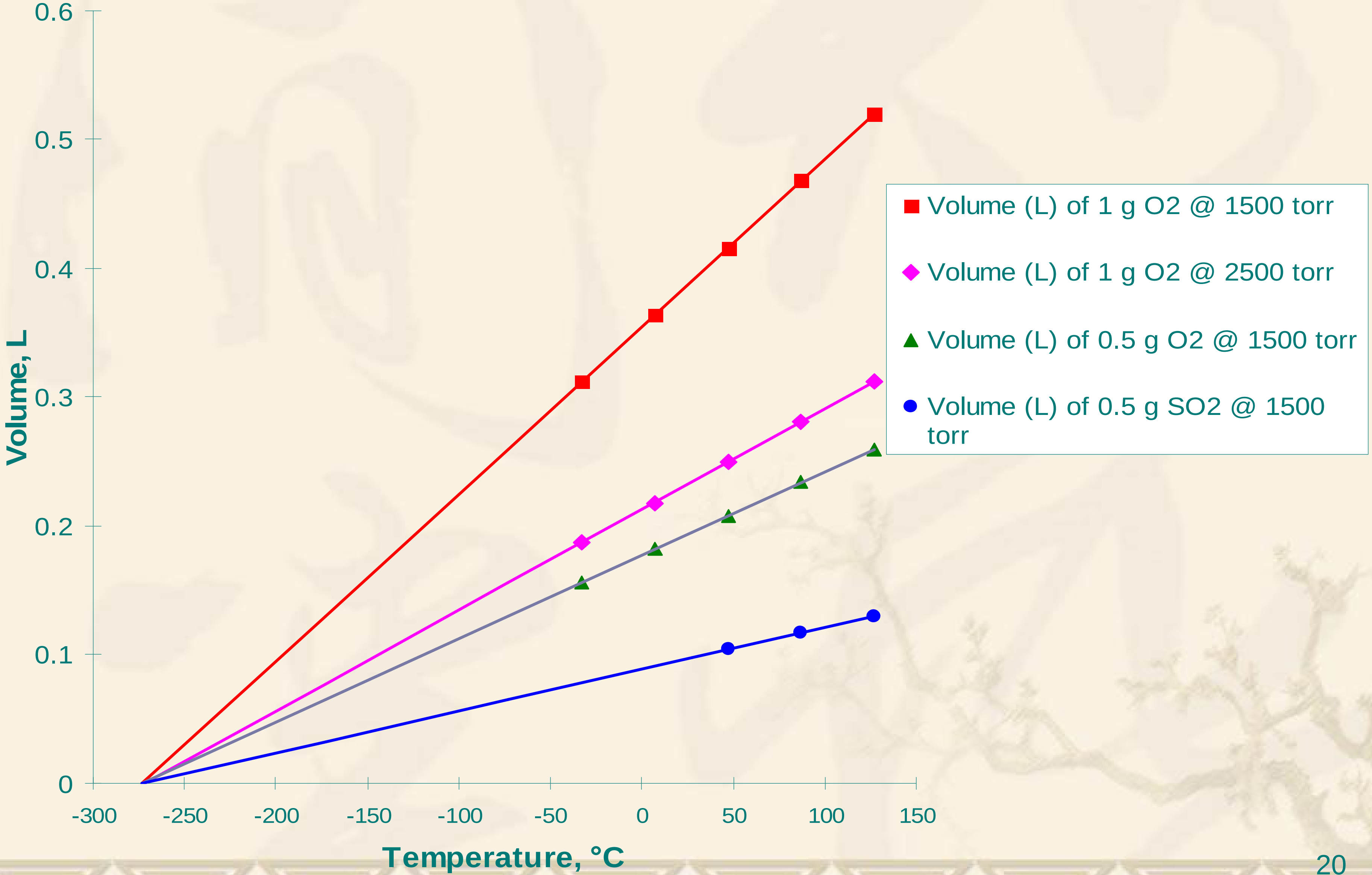
$$V = \left(\frac{RT}{P}\right) n \qquad \frac{RT}{P} \text{ is constant}$$

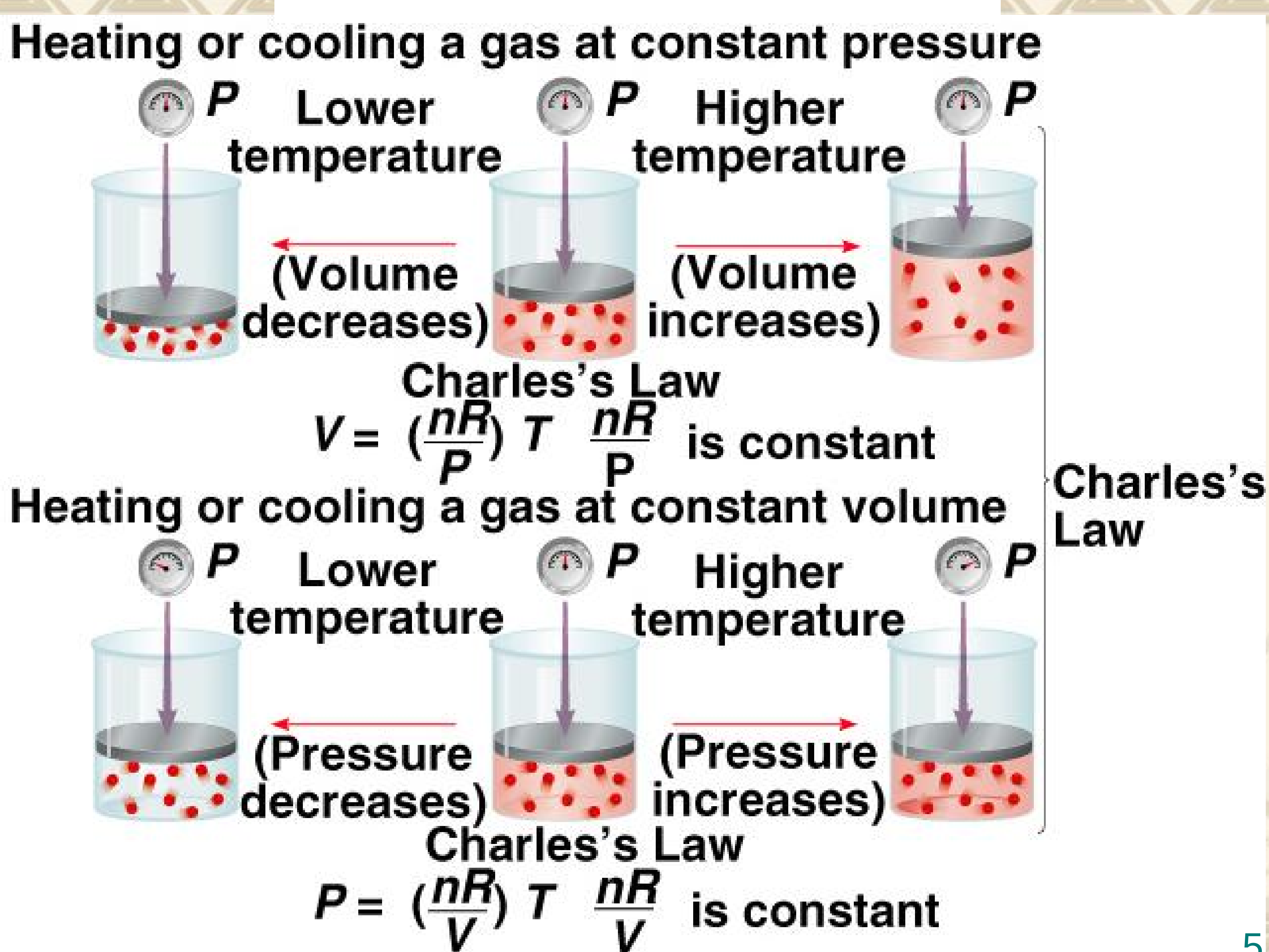
Charles' Law

- volume is directly proportional to temperature
  - constant P and amount of gas
  - graph of V vs T is straight line
- as T increases, V also increases
- Kelvin T = Celsius T + 273
- V = constant x T
  - if T measured in Kelvin

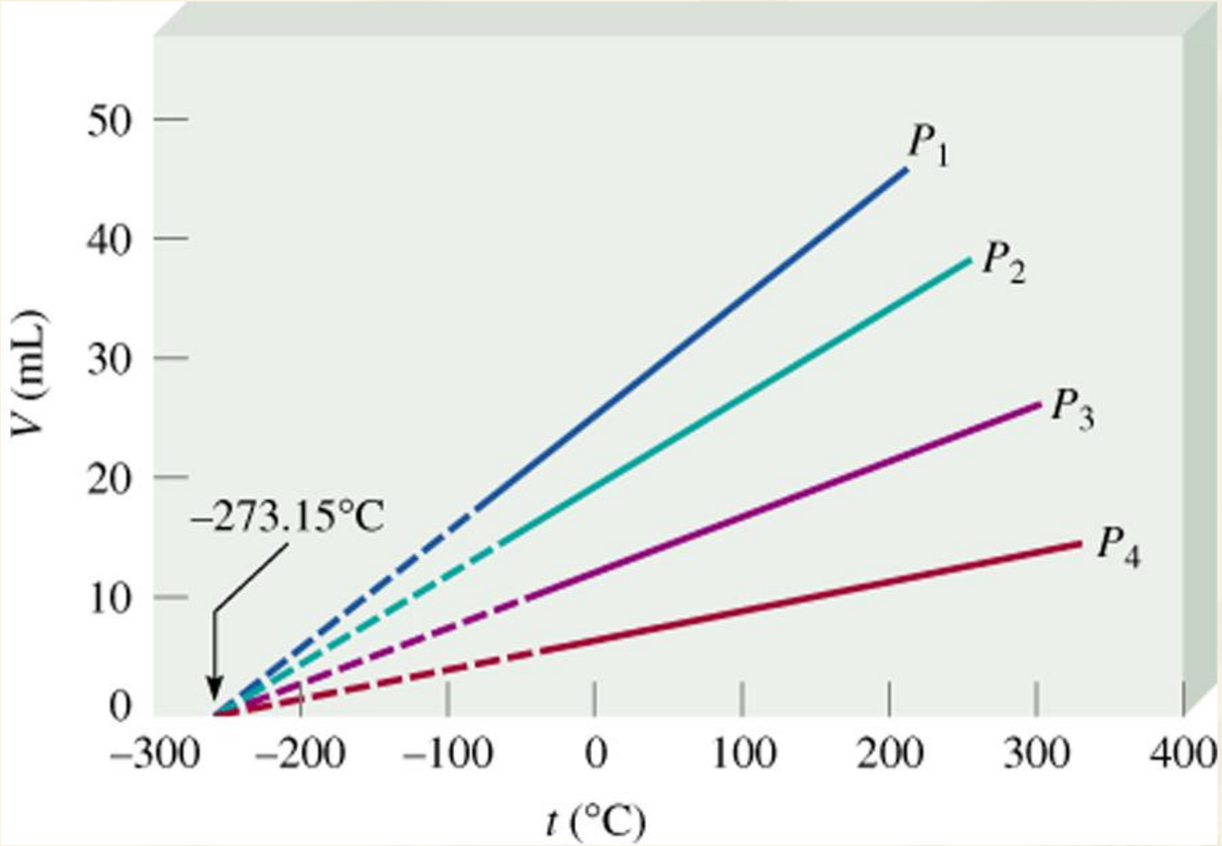
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Charle's Law & Absolute Zero




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

Variation of gas volume with temperature at constant pressure.



Charles' & Gay-Lussac's Law

$V \propto T$

V = constant x T

$V_1/T_1 = V_2/T_2$

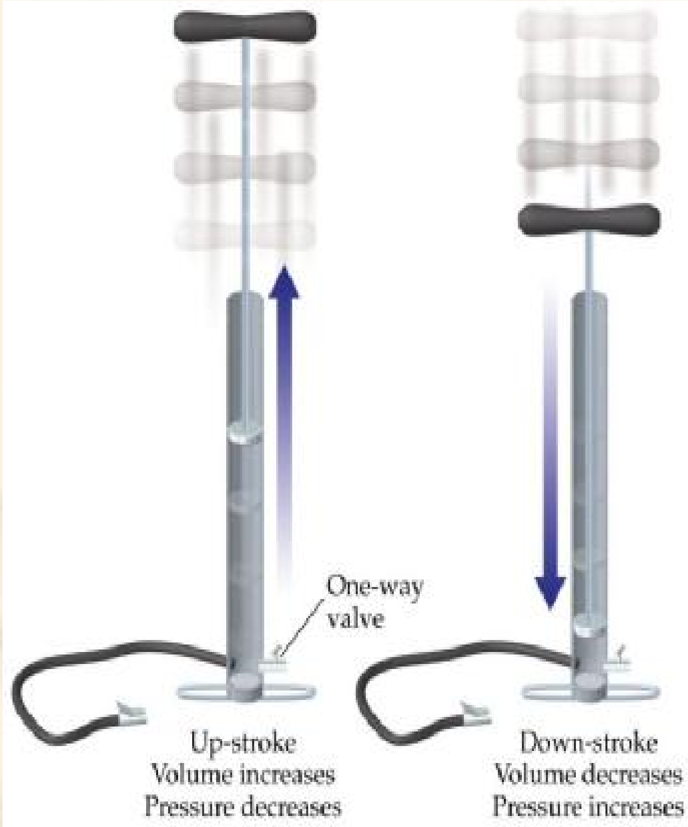
Temperature **must** be in Kelvin

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

Boyle's Law

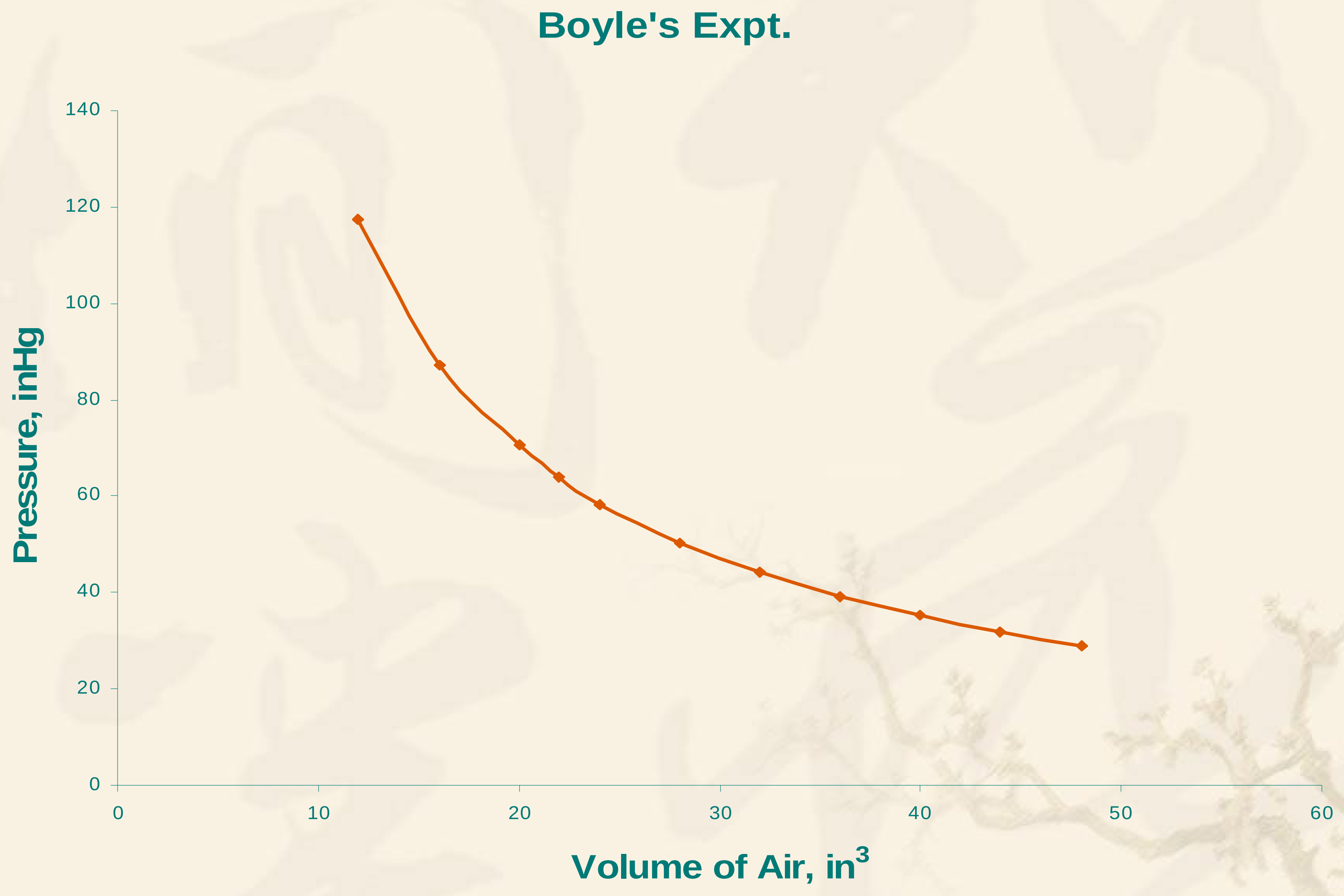
pressure of a gas is inversely proportional to its volume

- ™constant T and amount of gas
- ™graph P vs V is curve
- ™graph P vs 1/V is straight line

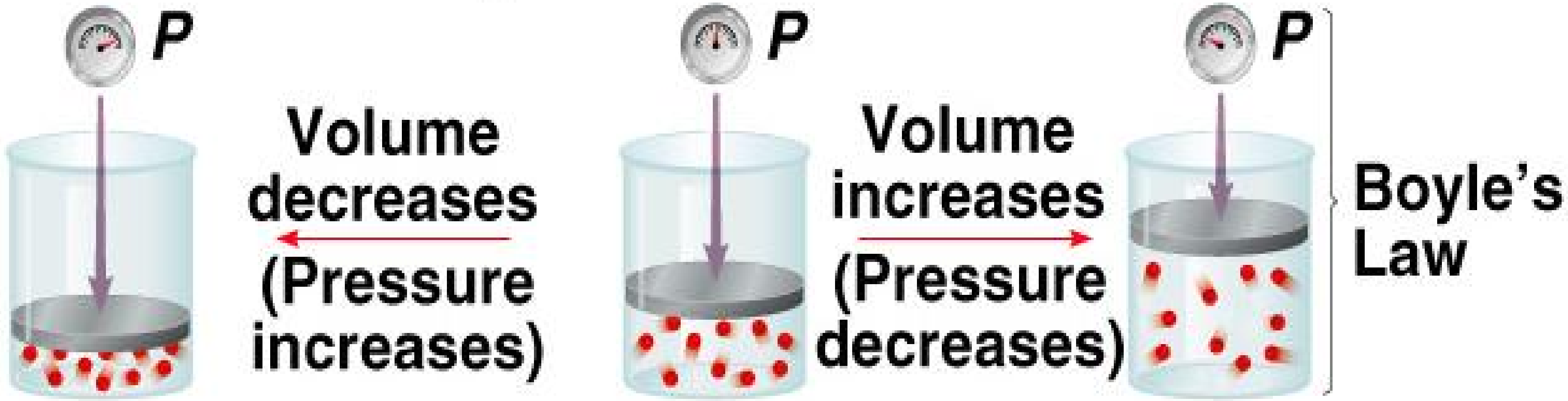


as P increases, V decreases by the same factor

- v $P \times V = \text{constant}$
- v $P_1 \times V_1 = P_2 \times V_2$


Boyle's Expt.
Pressure, inHg
Volume of Air, in3
0
20
40
60
80
100
120
140
0
10
20
30
40
50
60

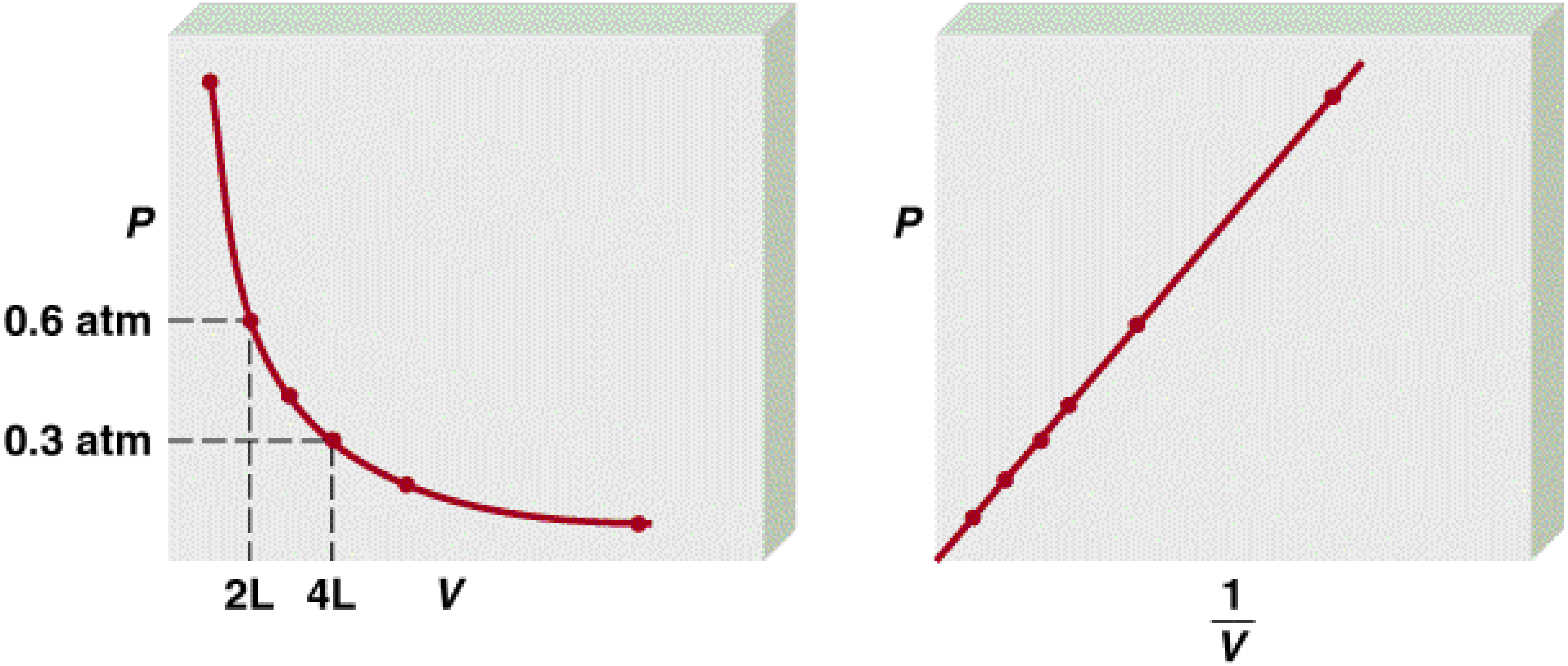
Increasing or decreasing the volume of a gas at a constant temperature



Boyle's Law

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

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

Example:
A cylinder equipped with a moveable piston has an applied pressure of 4.0 atm and a volume of 6.0 L. What is the volume if the applied pressure is decreased to 1.0 atm?

Information

Given: $P_1 = 4.0$ atm $V_1 = 6.0$ L
$P_2 = 1.0$ atm

Find: $V_2 = ?$ L

Eq'n: $P_1 \cdot V_1 = P_2 \cdot V_2$

- Write a Solution Map:

$P_1, V_1, P_2 \longrightarrow V_2$

$P_1 \bullet V_1 = P_2 \bullet V_2$

when using this equation, the units of P_1 and P_2 must be the same, or you will have to convert one to the other

for the same reason, the units of V_2 must be L to match the unit of V_1

Ideal gas law

- $V = k_1 \times n$
- $V = k_2 \times T$
- $V = k_3 / P$

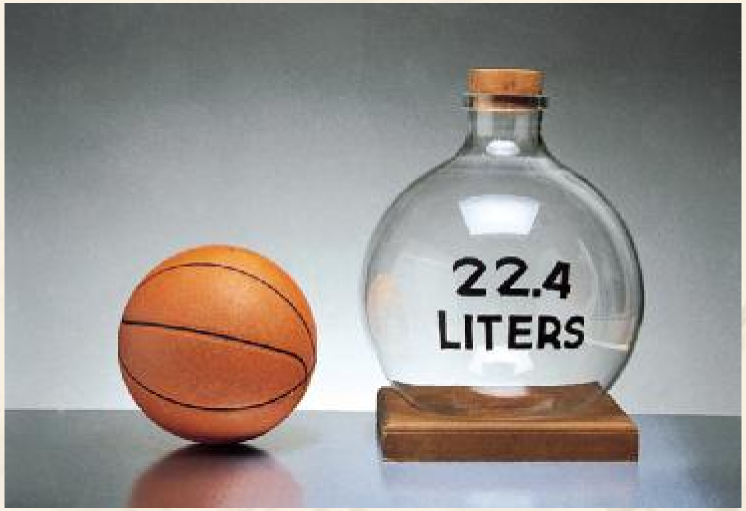
$$V = k_1 \times k_2 \times k_3 \times \frac{n \times T}{P}$$

$$= 常數 \times \frac{n \times T}{P}$$

$$PV = nRT$$

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)

Table5.1 Values of R in different Units

- Gas law $0.0821\times\frac{L\,atm}{molK}$
- Energy $8.31\frac{J}{mol\,K}$
- Molecular speed

$$8.31\times10^{3}\,\frac{g\,m^{2}}{s^{2}mol\,k} \qquad 1J=10^{3}\,\frac{g\,m^{2}}{S^{2}}$$

5.3 Gas law Calculations

The ideal gas law can be used to solve a variety of problem.

- ™The final state of gas, Knowing its initial state and the changes in P, V, n, T that occur.
- ™One of the four variables P, V, n, To Given the values of the other three.
- ™The molar mass or density of a gas

Final and Initial State Problems

- To determine the effect on V, P, n or T of a change in one or more of these variables.
  - Starting with a sample of gas at 25℃ and 1.00atm， you might be asked to calculate the pressure developed when the sample is heated to 95℃ at constant volume.
  - Initial state：$P_1V=nRT_1$
  - Final state ：$P_2V=nRT_2$

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

$$P_2 = P_1 \times \frac{T_2}{T_1} = 1.00atm \times \frac{95+273}{25+273} = 1.23atm$$

Ex:5-2 A 250mL flask, open to the atmosphere, contains 0.0110 mol of air at 0°C .on heating, part of the air escapes; how much remains in the flask at 100°C?(3位有效數字)

- Initial state：$PV=n_1RT_1$
- Final state：$PV=n_2RT_2$
- $n_1T_1=n_2T_2$

$$n_2 = n_1 \times \frac{T_1}{T_2} = 0.0110 \text{ mol} \times \frac{0 \quad + 273}{100 + 273}$$

$$= 0.00805 \text{ mol}$$

2.Calculation of P, V, n, 或T

- Values are Known for three of these quantities (perhaps V，n，and T) ；the other one (P) must be calculated.
- 例：已知V, n及T求P
- 例：已知V, n及P求T
- 例：已知P, n及T求V
- 例：已知P, V及T求n

Ex5.3: Sulfur hexafluoride is a gs that we will have a great deal more to say about in chapter 7(covalent Bonding) If 2.50g of this compound is introduced into an evacuated 500.0ml container at 83 What pressure in atmospheres is developed?

Sol:

$$V = 500.0\,mL \times \frac{1\,L}{1000\,mL} = 0.500\,L$$

$$T = 83 + 273 = 356\,K$$

$$n = 2.50\,g \times \frac{1\,mol}{146.07\,g} = 0.0171\,mol$$

$$P = \frac{nRT}{V} = \frac{0.0171 \times 0.0821 \times 356\,K}{0.5000\,L} = 1.00\,atm$$

3.Molar Mass and Density

$$PV = nRT$$

$$n = \frac{W}{Mw}$$

$$PV = \frac{W}{Mw} RT$$

$$PMw = \frac{W}{V} RT$$

$$PMw = DRT$$

$$Mw = \frac{DRT}{P}$$

$$D = \frac{PMw}{RT}$$

Ex:5-4 Acetone is widely used as a nail polish remover. A sample of liquid acetone is placed in a 3.00 L flask and vaporized by heating to 95°C at 1.02 atm. The vapor filling the flask at this temperature and pressure weighs 5.87 g，(a) What is the density of acetone vapor under these conditions b) Calculate the molar mass of acetone

$$(a)\ n = \frac{PV}{RT} = \frac{(1.02)\times(3.00\,L)}{(0.0821\,L.atm/mol.K)(368\,K)} = 0.101\,mol$$

$$(b)\ m = MM \times n$$

$$5.87 = MM \times 0.101\,mol$$

$$MM = 5.87/0.101 = 58.1\,g/mol$$

- (c)Acetone contains the three elements C,H and O. When 1.000g of acetone is burned, 2.27g of CO2 and 0.932g of H2O are formed. What is the molecular formula of acetone?

$$massC = 2.27gCO_2 \times \frac{12.01gC}{44.01gCO_2} = 0.619g\ C$$

$$massH = 0.932gH_2O \times \frac{2.016gH}{18.016gH_2O} = 0.104g\ H$$

$$massO = 1.00g - 0.619g - 0.104g = 0.277g\ O$$

$$\frac{0.619}{12.00} : \frac{0.104}{1.004} : \frac{0.277}{16.00} = \quad 0.0515 : 0.103 : 0.0173$$

$$= \quad 3 : \quad 6 : 3$$

$$D\ (Density) = \frac{P\ Mw}{RT}$$

- Pressure:
  - Compressing a gas increases its density.
  - (氣體經由壓縮其密度會增加。)
- Temperature
  - 溫度增加密度變小，所以熱空氣上升。
- Molar mass
  - Hydrogen(2.016g/mol) has the lowest molar mass and the lowest density。

5.4 Stiochiometry of Gaseous Reaction (氣體反應的化學計量)

A balanced equation can be used to relate moles or grams of substances taking part in a reaction. Where gases are involved, these relations can be extended to include volumes.

Example:

- How many liters of oxygen gas form when 294 g of $KClO_3$ completely reacts in the following reaction? Assume the oxygen gas is collected at P = 755 mmHg and T = 308 K

$$2\,KClO_3(s) \xrightarrow{\Delta} 2\,KCl(s) + 3\,O_2(g)$$

Example:

How many liters of $O_2(g)$ form when 294 g of $KClO_3$ completely reacts? Assume the $O_2(g)$ is collected at P = 755 mmHg and T = 308 K

$$2\,KClO_3(s) \xrightarrow{\Delta} 2\,KCl(s) + 3\,O_2(g)$$

Information

Given: 294 g $KClO_3$

P_{O2} = 755 mmHg, T_{O2} = 308 K

Find: V_{O2}, L

- Collect Needed Equation and Conversion Factors:

The relationship between pressure, temperature, number of moles

$$PV = nRT$$

and volume is the Ideal Gas Law

We also need the molar mass of $KClO_3$ and the mole ratio from

the chemical equation

1 mole $KClO_3$ = 122.5 g $KClO_3$

2 mol $KClO_3$ ≡ 3 mol O_2

Example:

How many liters of $O_2(g)$ form when 294 g of $KClO_3$ completely reacts? Assume the $O_2(g)$ is collected at P = 755 mmHg and T = 308 K

$$2\,KClO_3(s) \xrightarrow{\Delta} 2\,KCl(s) + 3\,O_2(g)$$

Information

Given: 294 g $KClO_3$

P_{O2} = 755 mmHg, T_{O2} = 308 K

Find: V_{O2}, L

Eq'n: PV=nRT

CF: 1 mole $KClO_3$ = 122.5 g

2 mole $KClO_3 \equiv$ 3 moles O_2

- Write a Solution Map:

$$\text{g } KClO_3 \xrightarrow{\frac{1\,\text{mol } KClO_3}{122.5\,\text{g } KClO_3}} n_{KClO3} \xrightarrow{\frac{3\,\text{mol } O_2}{2\,\text{mol } KClO_3}} n_{O2} \xrightarrow[\text{PV=nRT}]{\text{P,T,R}} V_{O2}$$

when using the Ideal Gas Equation, the units of V must be L; and the units of P must be atm, or you will have to convert the units of T must be kelvin, K

Example:

How many liters of $O_2(g)$ form when 294 g of $KClO_3$ completely reacts? Assume the $O_2(g)$ is collected at P = 755 mmHg and T = 308 K

$$2\,KClO_3(s) \xrightarrow{\Delta} 2\,KCl(s) + 3\,O_2(g)$$

Information

Given: 294 g $KClO_3$

P_{O2} = 755 mmHg, T_{O2} = 308 K

Find: V_{O2}, L

Eq'n: PV=nRT

CF: 1 mole $KClO_3$ = 122.5 g

2 mole $KClO_3 \equiv$ 3 moles O_2

SM: g $\rightarrow$ mol $KClO_3 \rightarrow$ mol $O_2 \rightarrow$ L O_2

- Apply the Solution Map:
  - find moles of O_2 made

$$294\text{ g KClO}_3 \times \frac{1\text{ mol KClO}_3}{122.5\text{ g KClO}_3} \times \frac{3\text{ mol O}_2}{2\text{ mol KClO}_3} = 3.60\text{ mol O}_2$$

Example:

How many liters of $O_2(g)$ form when 294 g of $KClO_3$ completely reacts? Assume the $O_2(g)$ is collected at P = 755 mmHg and T = 308 K

$$2\,KClO_3(s) \xrightarrow{\Delta} 2\,KCl(s) + 3\,O_2(g)$$

Information

Given: 294 g $KClO_3$

P_{O2} = 755 mmHg, T_{O2} = 308 K,

n_{O2} = 3.60 moles

Find: V_{O2}, L

Eq'n: PV=nRT

CF: 1 mole $KClO_3$ = 122.5 g

2 mole $KClO_3 \equiv$ 3 moles O_2

SM: g $\rightarrow$ mol $KClO_3 \rightarrow$ mol $O_2 \rightarrow$ L O_2

- Apply the Solution Map:
  - convert the units

$$P = 755\,\text{mmHg} \times \frac{1\,\text{atm}}{760\,\text{mmHg}} = 0.99\underline{3}42\,\text{atm}$$

Example:

How many liters of $O_2(g)$ form when 294 g of $KClO_3$ completely reacts? Assume the $O_2(g)$ is collected at P = 755 mmHg and T = 308 K

$$2\,KClO_3(s) \xrightarrow{\Delta} 2\,KCl(s) + 3\,O_2(g)$$

Information

Given: 294 g $KClO_3$

P_{O2} = 0.99$\underline{3}$42 mmHg, T_{O2} = 308 K,

n_{O2} = 3.60 moles

Find: V_{O2}, L

Eq'n: PV=nRT

CF: 1 mole $KClO_3$ = 122.5 g

2 mole $KClO_3 \equiv$ 3 moles O_2

SM: g $\rightarrow$ mol $KClO_3 \rightarrow$ mol $O_2 \rightarrow$ L O_2

- Apply the Solution Map:

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

$$\frac{n \bullet R \bullet T}{P} = V = \frac{(3.60\,\text{mol})\left(0.0821\,\frac{\text{L}\bullet\text{atm}}{\text{mol}\bullet\text{K}}\right)(298\,\text{K})}{0.99\underline{3}42\,\text{atm}}$$

$$V = 90.7\,\text{L}$$

Example:

How many liters of $O_2(g)$ form when 294 g of $KClO_3$ completely reacts? Assume the $O_2(g)$ is collected at P = 755 mmHg and T = 308 K

$$2\,KClO_3(s) \xrightarrow{\Delta} 2\,KCl(s) + 3\,O_2(g)$$

Information

Given: 294 g $KClO_3$

P_{O2} = 755 mmHg, T_{O2} = 308 K,

n_{O2} = 3.60 moles

Find: V_{O2}, L

Eq'n: PV=nRT

CF: 1 mole $KClO_3$ = 122.5 g

2 mole $KClO_3 \equiv$ 3 moles O_2

SM: g $\rightarrow$ mol $KClO_3 \rightarrow$ mol $O_2 \rightarrow$ L O_2

- Check the Solution:

$$V_{O_2} = 90.7\,L$$

The units of the answer, L, are correct.
It is hard to judge the magnitude with so many variables, but with more than 1 mole of gas at pressures near 1 atm and temperatures near 273 K – it is reasonable to have more than 22.4 L

Calculate the volume occupied by 1.00 moles of an ideal gas at STP.

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

$$(1.00 \text{ atm}) \times V = (1.00 \text{ moles})(0.0821 \frac{L \cdot \text{atm}}{\text{mol} \cdot K})(273 \text{ K})$$

$$V = 22.4 \text{ L}$$

- 1 mole of any gas at STP will occupy 22.4 L
- this volume is called the **molar volume** and can be used as a conversion factor
  - as long as you work at STP

$$1 \text{ mol} \equiv 22.4 \text{ L}$$

Ex 5-5 A nickel smelter in Sudbury, Ontario, produces 1% of the world's supply of sulfur dioxide by the reaction

$$2NiS_{(s)} + 3O_{2(g)} \rightarrow 2NiO_{(s)} + 2SO_{2(g)}$$

What volume of SO_2 at 25℃ and a pressure of one bar is produced from a metric ton of nickel (Ⅱ) sulfide

Sol:

$$n_{Nis} = 1.00 metricton \times \frac{10^6}{1 metricton} \times \frac{1mol}{90.76g} = 1.10 \times 10^4 mol$$

$$n_{so_2} = 1.10 \times 10^4 molNis \times \frac{2molSO_2}{2molNiS} = 1.10 \times 10^4 \quad SO_2$$

$$\Pr essure\ in\ aotm = 1.0bar \times \frac{1atm}{1.013bar} = 0.987\ atm$$

$$V_{SO_2} = \frac{nRT}{P} = \frac{1.10 \times 10^4 mol \times 0.0821 \times 298K}{0.987atm}$$

$$= 2.73 \times 10^5 L$$

Ex:5-6 Octane,(C_8H_{18},)is one of a the hydrocarbons in gasoline. On combustion (burning in oxygen), Octane produces carbon dioxide and water . How many liters oxygen ,measured at 0.974atm and24°C,are required to burn 1.00g of Octane

$$2C_8H_{18(l)} + 25O_{2(g)} \rightarrow 16CO_{2(g)} + 18H_2O_{(l)}$$

$$n_{O_2} = 1.00\cancel{gC_8H_{18}} \times \frac{1mol\cancel{C_8H_{18}}}{114.22\cancel{gC_8H_{18}}} \times \frac{25molO_2}{2\cancel{molC_8H_{18}}} = 0.109molO_2$$

$$V_{O_2} = \frac{nRT}{P}$$

$$= \frac{0.109molO_2 \times 0.0821L \times atm/(mol \times K) \times (24+273)K}{0.974atm}$$

$$= 2.73L$$

The law of combining volumes, proposed by Gay-Lussac in 1808:The volume ratio of any two gases in a reaction at constant temperature and pressure is the same as the reacting mole ratio.

$$2H_2O_{(l)} \rightarrow 2H_{2(g)} + O_{2(g)}$$

The volume of hydrogen produced is twice that of the other gaseous product, oxygen.

At constant temperature and pressure . The volume ratio must be the same as the mole ratio given by the coefficients of the balanced equation.

Ex5.7:Consider the reaction

$$2H_2O_{(l)} \rightarrow 2H_{2(g)} + O_{2(g)}$$

(a) What volume of H_2 at 25 ℃ and 1.00 atm is required to react with 1.00L of O_2 at the same temperature and pressure?

(b) What volume of $H_2O_{(g)}$ at 25 ℃ and 1.00 atm (d=0.997g/mL) is formed from the reaction in (a) ?

(c) What mass of H_2O is formed from the reaction in (a) ， assuming a yield of 85.2％?

Sol:

$$(a).volume\ H_2 = 1.00LO_2 \times \frac{2H_2}{1O_2} = 2.00LH_2$$

$$(b).nO_2 = \frac{Pv}{RT} = \frac{(1.0atm)(1.0L)}{(0.082)(298K)} = 0.0409mol.O_2$$

$$nH_2O = 0.0409mol.O_2 \times \frac{2molH_2O}{1molO_2} = 0.0818.molH_2O$$

$$V_{H_2O} = 0.0818mol.H_2O \times \frac{18.02g}{1mol} \times \frac{1.00mL}{0.997g} = 1.48mL$$

$$(c)yield.H_2O = 0.0818molH_2O \times \frac{18.02g}{1mol} \times 0.852 = 1.26gH_2O$$

5.5 Mixtures of Gases

The ideal gas law applies to all gases, You might expect it to apply to gas mixtures . According to Kinetic Molecular Theory, the particles in a gas behave independently.

- Air is a mixture, yet we can treat it as a single gas
- Also, we can think of each gas in the mixture independent of the other gases
  - though all gases in the mixture have the same volume and temperature
    - all gases completely occupy the container, so all gases in the mixture have the volume of the container

Gas	% in Air, by volume	Gas	% in Air, by volume
nitrogen, N_2	78	argon, Ar	78
oxygen, O_2	21	carbon dioxide, CO_2	21

Partial Pressure

- each gas in the mixture exerts a pressure independent of the other gases in the mixture
- the pressure of an component gas in a mixture is called a **partial pressure**
- the sum of the partial pressures of all the gases in a mixture equals the total pressure
  - Dalton's Law of Partial Pressures
  - $P_{total} = P_{gas\ A} + P_{gas\ B} + P_{gas\ C} + ...$

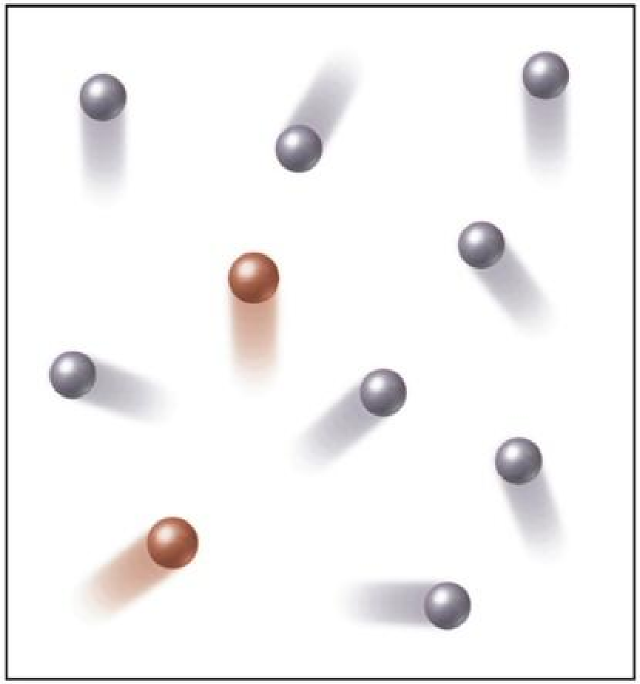
$$P_{air} = P_{N_2} + P_{O_2} + P_{Ar} = 0.78\ \text{atm} + 0.21\ \text{atm} + 0.01\ \text{atm} = 1.00\ \text{atm}$$

Finding Partial Pressure

- to find the partial pressure of a gas multiply the total pressure of the mixture by the fractional composition of the gas
- for example, in a gas mixture that is 80.0% He and 20.0% Ne that has a total pressure of 1.0 atm, the partial pressure of He would be:

P_{He} = (0.800)(1.0 atm) = 0.80 atm

  - fractional composition = percentage divided by 100



Gas mixture (80% He, 20% Ne)
P_{tot} = 1.0 atm
P_{He} = 0.80 atm
P_{Ne} = 0.20 atm

5.5 Mixtures of Gases

Gas Mixture :partial pressures and Mole fractions

$$P_{total} = n_{total} \times \frac{RT}{V} = (n_A + n_B)\frac{RT}{V} = n_A \times \frac{RT}{V} + n_B \times \frac{RT}{V}$$

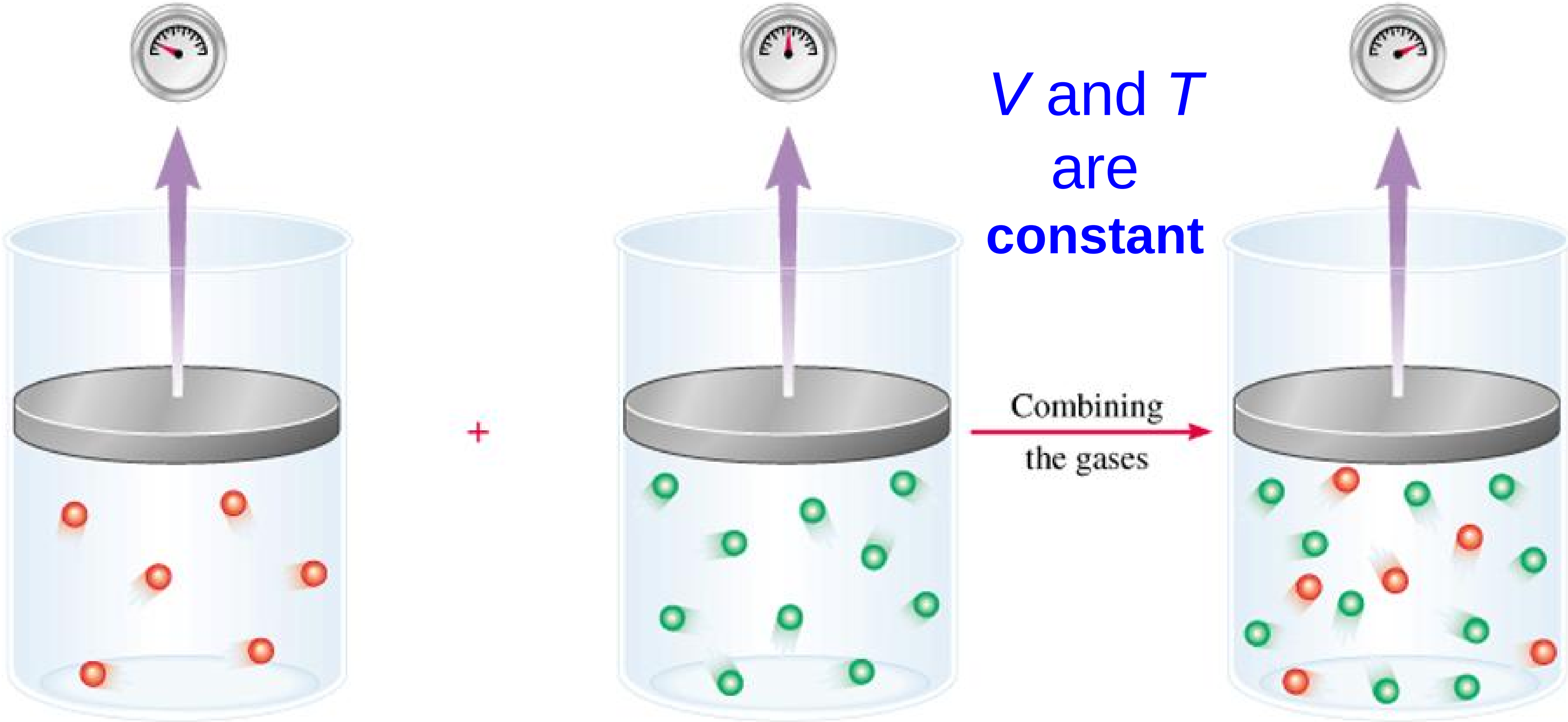
$$P_A = Partial\ pressure\ A = \frac{n_A RT}{V}$$

$$P_B = Partial\ pressure\ B = \frac{n_B RT}{V}$$

$$P_{tatal} = P_A + P_B$$

Dalton's Law of Partial Pressures

The total pressure of a gas mixture is the sum of the partial pressures of the components of the mixture.

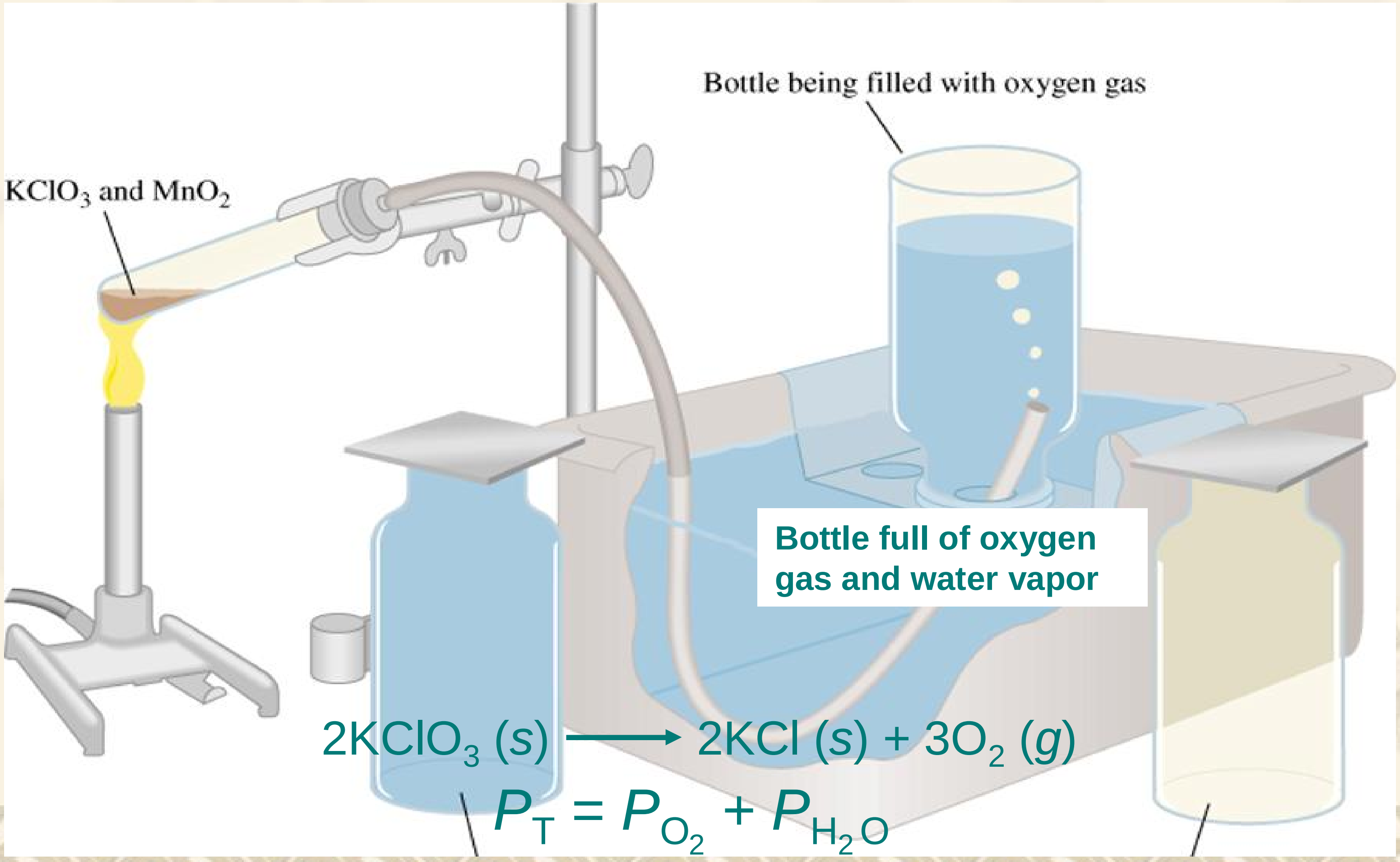


P_A $\quad$ $\boldsymbol{P_B}$ $\quad$ $P_{total} = P_A + P_B$

2.Wet Gases ;Pressure of water 水的分壓

It picks up water vapor; molecules of H_2O escape from the liquid and enter the gas phase. $P_T = P_{O_2} + P_{H_2O}$

The Partial pressure of water vapor $PH_2O$is equal to the vapor pressure of liquid water. It has a fixed value at a given temperature



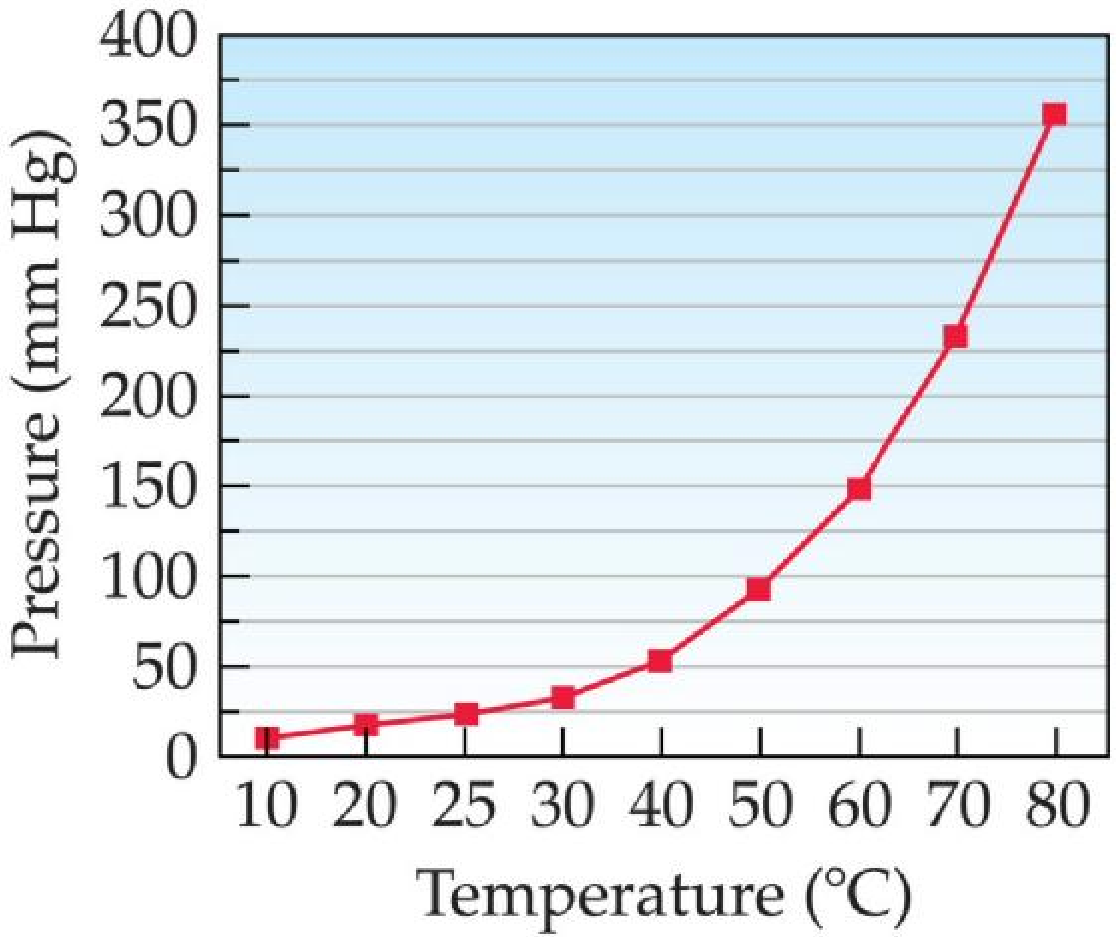
$$2KClO_3\,(s) \longrightarrow 2KCl\,(s) + 3O_2\,(g)$$

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

Collecting Gases

- gases are often collected by having them displace water from a container
- the problem is that since water evaporates, there is also water vapor in the collected gas
- the partial pressure of the water vapor, called the **vapor pressure**, depends only on the temperature
  - so you can use a table to find out the partial pressure of the water vapor in the gas you collect
- Vapor pressure，is an intensive physical property. if you collect a gas sample with a total pressure of 758 mmHg at 25°C, the partial pressure of the water vapor will be 23.8 mmHg – so the partial pressure of the dry gas will be 734 mmHg

Vapor Pressure of Water



Temp., °C	Pressure, mmHg
10	9.2
20	17.5
25	23.8
30	31.8
40	55.3
50	92.5
60	149.4
70	233.7
80	355.1

Ex5-8:A student prepares a sample of hydrogen gas by electrolyzing water at 25℃.She collects 152mL of H_2 at a total pressure of 758mmHg.Using Appendix 1 to find the vapor pressure of water, calculate

(a) The partial pressure of hydrogen.

(b). The number of moles of hydrogen collected.

$$(a)P_{H_2} = P_{tot} - P_{H_2O} = 758mmHg - 23.76mmHg = 734mm\ Hg$$

$$(b)n_{H2} = \frac{(P_{H2})V}{RT} = \frac{(734/760atm)(0.152L)}{(0.082)(298K)} = 0.00600mol\ H_2$$

3.Partial Pressure and Mole Fraction

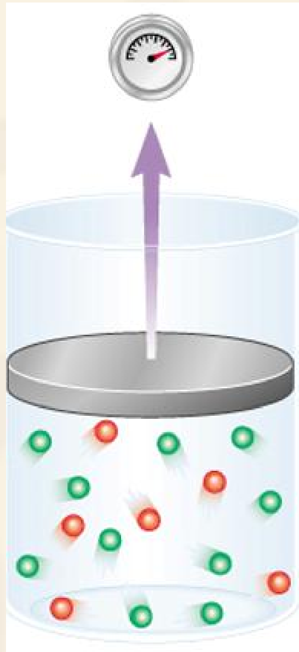
Consider a case in which two gases, A and B, are in a container of volume V.

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

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



$$P_T = P_A + P_B \qquad P_{tot} = \frac{n_{tot}RT}{V} \qquad \frac{P_A}{P_{total}} = \frac{n_A}{n_{tot}}$$

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

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

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

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

The partial pressure of a gas in a mixture is equal to its mole fraction multiplied by the total pressure.

Example 5.9 : When one mole of methane，CH_4 is heated with four moles of oxygen , the following reaction Occurs:

Assuming all of the methane is converted to $CO_2$and H_2O , What are the mole fractions of O_2, CO_2 and H_2O in the resulting mixture? If the total pressure of the mixture is 1.26 atm , What are the partial pressures?

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Sol: All the methane is consumed

$$n_{CH_4} = 0 \quad n_{CO_2} = 1.00 \quad n_{H_2O} = 2.00 \quad n_{O_2} = 4.0 - 2.0 = 2.00$$

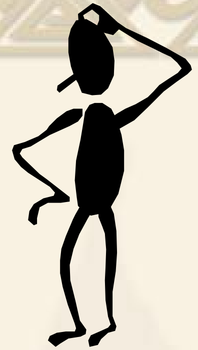
$$The\ total\ number\ of\ moles = 1.0 + 2.0 + 2.0 = 5.00$$

$$X_{co_2} = \frac{1.00}{5.00} = 0.200 \quad X_{H2O} = \frac{2.00}{5.00} = 0.400 \quad X_{CO_2} = \frac{2.00}{5.00} = 0.400$$

$$P_{H_2O} = 0.400 \times 1.26 atm = 0.504 atm$$

$$P_{CO_2} = 0.200 \times 1.26 atm = 0.252 atm$$

$$P_{O_2} = 0.400 \times 1.26 atm = 0.504 atm$$



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)?

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

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

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

5.6 kinetic theory of gases

1850-1880, James Maxwell, Rudolf Clausius , Ludwig Boltzmann , and other developed the Kinetic theory of gases.

They based it on the idea that all gases behave similarly as far as particle motion is concerned.

1.Molecular Model

Gases are mostly empty space.

- ™ 與容器相比非常小，氣體總體積可忽略。

Gas molecules are in constant, chaotic motion. 氣體分子相互及與器壁碰撞，速率恆定。

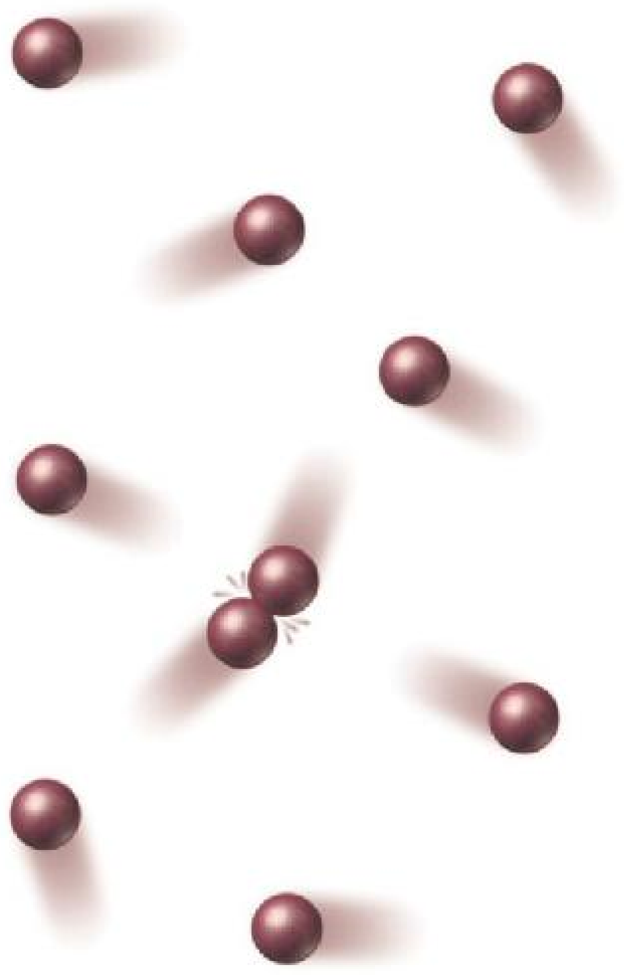
Collisions are elastic

- ™ 氣體分子間彼此無吸引力，不會相互黏在一起

Gas pressure is caused by collisions of molecules with the walls of the container

能量與碰撞頻率增加，壓力將會加大

Kinetic Molecular Theory



1 Collection of particles in constant motion

2 No attractions or repulsions between particles; collisions like billiard ball collisions

3 A lot of space between the particles compared to the size of the particles themselves

4 The speed that the particles move increases with increasing temperature

Source of Gas pressure

- Gas pressure is caused by collisions of molecules with the walls of the container.

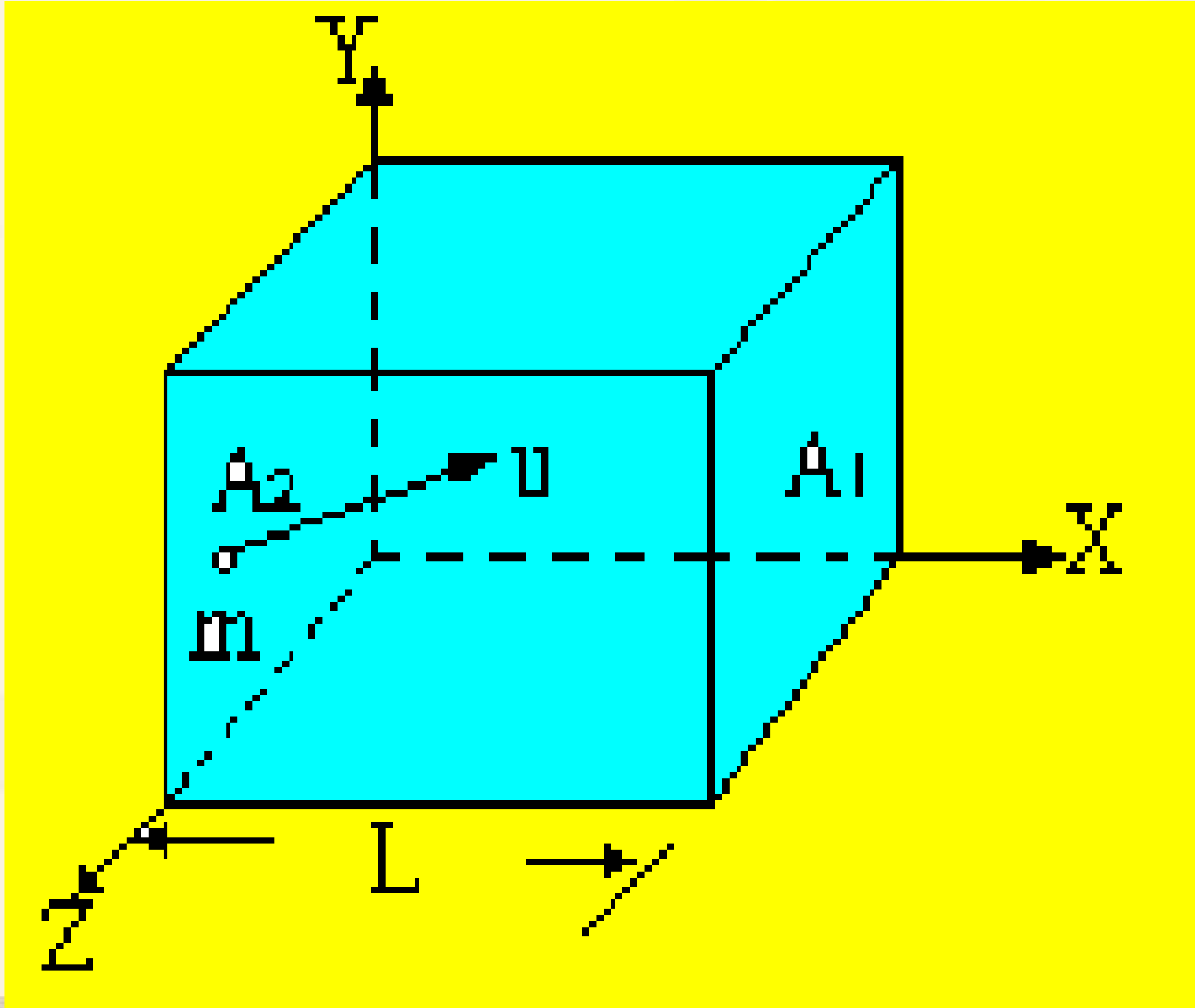
Expression for Pressure ,P

- 討論A1器壁因氣體撞擊器壁所形成的壓力：
- 氣體中一分子m以Vx撞向A1器壁，由理想氣體的假設知其必以-Vx反彈，在此碰撞中，氣體的動量改變量：

$$\Delta \mathrm{P} = \mathrm{mV_x} - m(-V_X) = \mathrm{mV_x} + mV_X = 2mV_X$$

$$P = \frac{Nmu^2}{3V}$$

- 1.N/V the concentration of gas molecules in the container. The more molecules there are in a given volume, the greater the collision frequency and so the greater the pressure.
- 2.mu^2 is a measure of the energy of collision , this equation predicts, pressure is directly related to mu^2


Y
A2
u
A1
X
m
Z
L

1. 單一分子碰撞一次動量變化：**2mu**

2. 單一分子在單位時間內之動量變化：$2mu \times \frac{u}{2l} = \frac{mu^2}{l}$

3. 容器內分子對某面碰撞之動量總變化：$F = \frac{N_0 mu^2}{3l}$

4. 某面之壓力：

$$P = \frac{F}{A} = \frac{\frac{N_0 mu^2}{3l}}{l^2} = \frac{N_0 mu^2}{3l^3} = \frac{N_0 mu^2}{3V}$$

$$PV = \frac{1}{3} N_0 mu^2$$

Average Kinetic Energy of Translational Motion E_t

$$E_t = \frac{mu^2}{2} \qquad mu^2 = \frac{3PV}{N} \qquad P = \frac{Nmu^2}{3V}$$

$$E_t = \frac{3PV}{2N} = \frac{3nRT}{2N} = \frac{3RT}{2N_A} = KT$$

1. **At a given temperature, molecules of different gases must all have the same average kinetic energy of translational motion**
   在某一溫度下，不同的氣體分子應有相同的轉移運動的平均動能。

2. **E_t is directly proportional to the Kelvin temperature**，T

氣體分子的轉移平均動能**E_t**與凱氏溫度**T**成正比 。

Average Speed ,U

$$\frac{mu^2}{2} = \frac{3RT}{2N_A} \quad u^2 = \frac{3RT}{mN_A} \quad u^2 = \frac{3RT}{MM}$$

$$averagespeed, u$$

$$u = (\frac{3RT}{MM})^2$$

$$Directly\ proportional\ to\ the\ square\ root\ of\ the\ absolute\ temperature.$$

$$\frac{u_2}{u_1} = (\frac{T_2}{T_1})^{1/2}$$

$$Inveersely\ proportional\ to\ the\ square\ root\ of\ molar\ mass(MM)$$

$$\frac{u_B}{u_A} = (\frac{MM_A}{MM_B})^{1/2}$$

Ex5.10: Calculate the average speed U of an N_2 molecule at 25 °C.

$$u = \sqrt{\frac{3RT}{Mw}}$$

$$= \sqrt{\frac{3 \times 8.31 \times 10^3 \frac{g \times m^2}{s^2 \times mol \times K} \times (25 + 273)}{28.02 \frac{g}{mol}}}$$

$$= 515 m/s$$

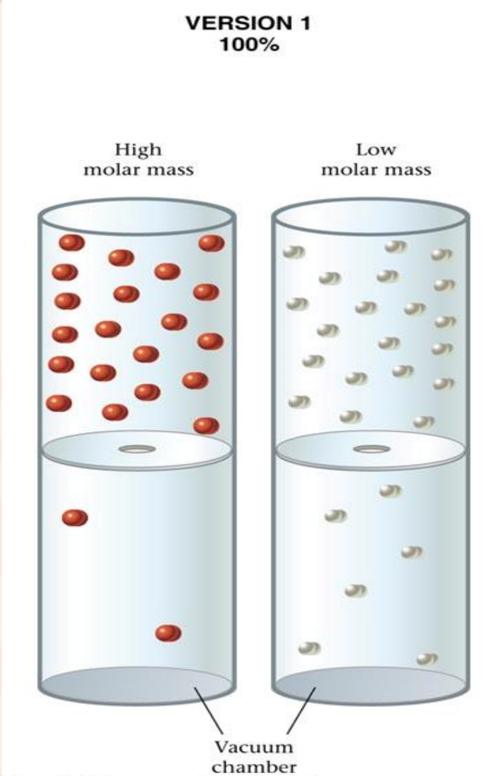
Effusion of Gases; Graham's Law

- Effusion:
- It is flow of gas molecules at low pressures through tiny pores or pinholes

$$\frac{rate\ of\ effusion\ B}{rate\ of\ effusion\ A}=\frac{u_B}{u_A} \qquad \frac{u_B}{u_A}=(\frac{MM_A}{MM_B})^{1/2}$$

$$\frac{rate\ of\ effusion\ B}{rate\ of\ effusion\ A}=(\frac{MM_A}{MM_B})^{1/2}$$

$$\frac{rate\ of\ effusion\ {}^{235}_{92}UF_6}{rate\ of\ effusion\ {}^{238}_{92}UF_6}=(\frac{352.0}{349})^{1/2}=1.004$$



- 同溫同壓下,氣體逸散的速率(mol/time)與氣體莫耳質量之均方根成反比
- The equation was discover by the Thomas Graham in 1829 .Among them the separation of the components of air. Graham's law
- At a given temperature and pressure , the rate of effusion of a gas , in moles per unit time, is inversely proportional to the square root of its molar mass.

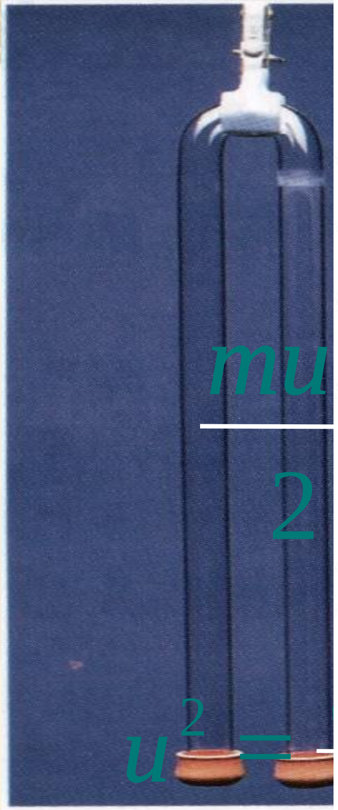
Ex:5.11 In an effusion experiment , argon gas is allowed to expand through a tiny opening into an evacuated flask of volume 120mL for 32.0s. At which point the pressure in the flask is found to be 12.5mmHg. This experiment is repeated with a gas X of unknown molar mass at the same T and P . It is found that the pressure in the flask builds up to 12.5mmHg after 48 s. C alculate the molar mass of X 。

$$\frac{rate\ Ar}{rate\ X} = \frac{n/32.0s}{n/48.0s} = \frac{48.0}{32.0} = 1.50$$

$$Applying \quad Graham's\ law$$

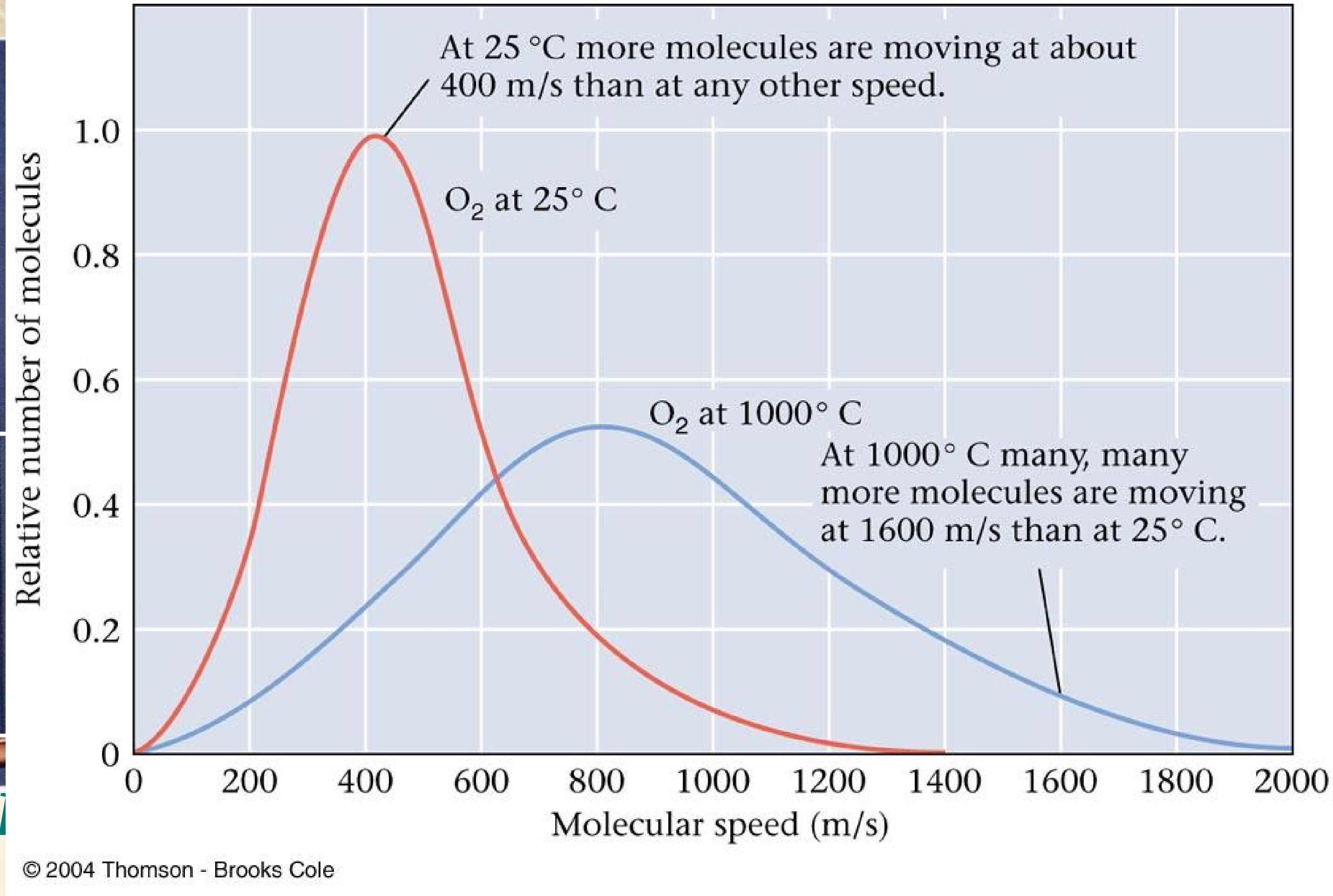
$$1.50 = \left(\frac{MM_x}{MM_{Ar}}\right)^{1/2}$$

$$MM_x = (39.95g/mol) \times (1.50)^2 = 89.9g/mol$$



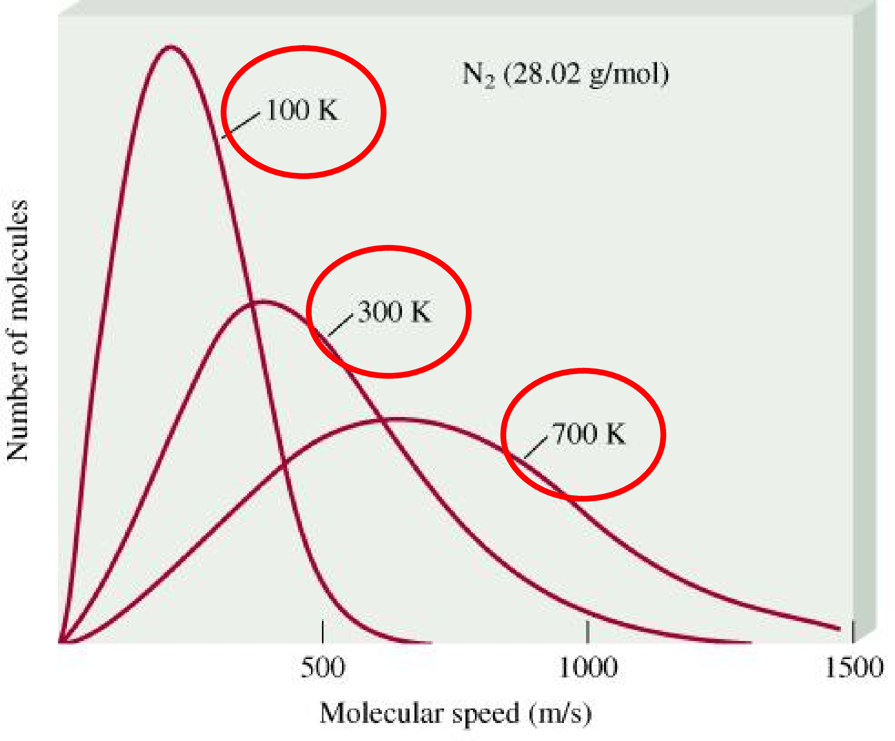
$\frac{mu}{2}$

$u^2 =$



1. 平均速度與絕對溫度的平方根成正比。

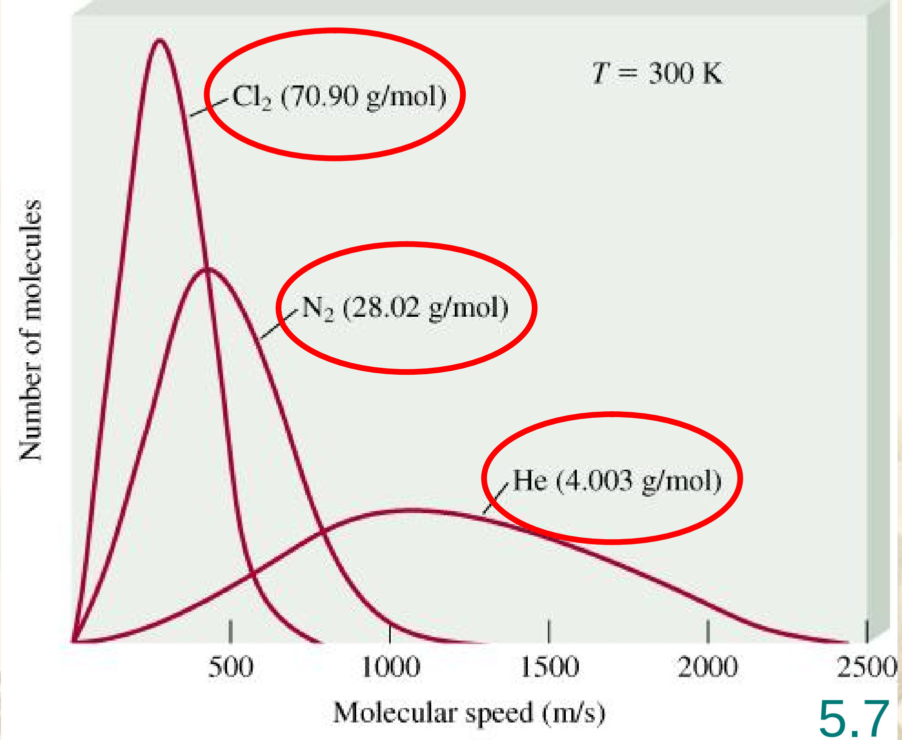
2. 平均速度與莫耳質量的平方根成反比。



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

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

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

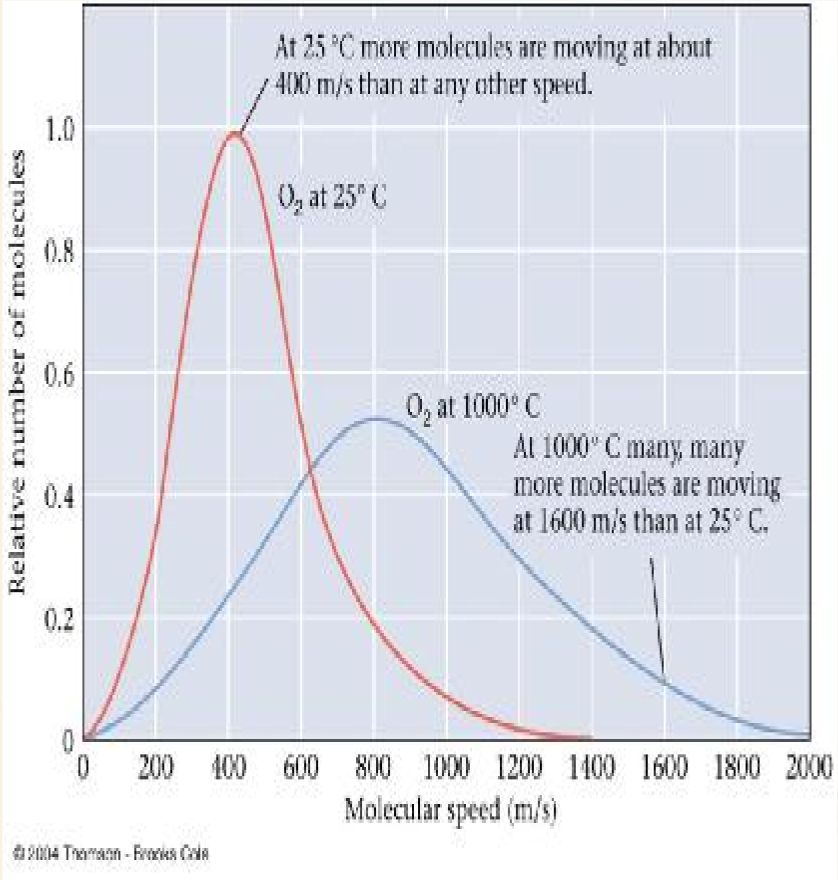


Distribution of Molecular Speeds

$$\frac{mu^2}{2} = \frac{3RT}{2N_A}$$

$$u^2 = \frac{3RT}{mN_A} = \frac{3RT}{Mw}$$

$$u = \sqrt{\frac{3RT}{Mw}}$$



1.Temperature increases, the speed of the molecules increases.

2. Mass o f moles increases , the speed of the molecules decreases.

5.7 Real Gases

Molar volume $=V_m = V/n$; ideal gas law $V^0_m=RT/P$

Table 5.2 the deviations from ideality become larger at high pressures and low temperatures; ideality become smaller at Low pressures and hight temperatures.

1.In general ，the closer a gas is to the liquid state, the more it will deviate from the ideal gas law.愈是氣體狀態愈接近理想氣體方程式。

- Height temperature、Low pressure(非極性)
- 愈接近液體狀體，則與理想氣體方程式誤差愈大。

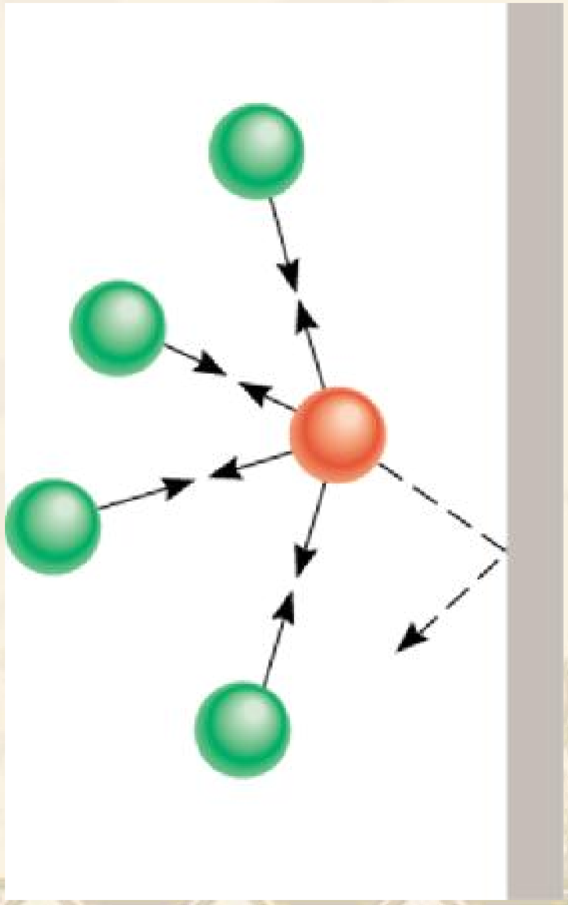
2.From a molecular standpoint , deviations from the ideal gas law arise because it neglects two factors:

- (1). Attractive forces between gas particles.
- (2) . The finite volume of gas particles.

Attractive forces

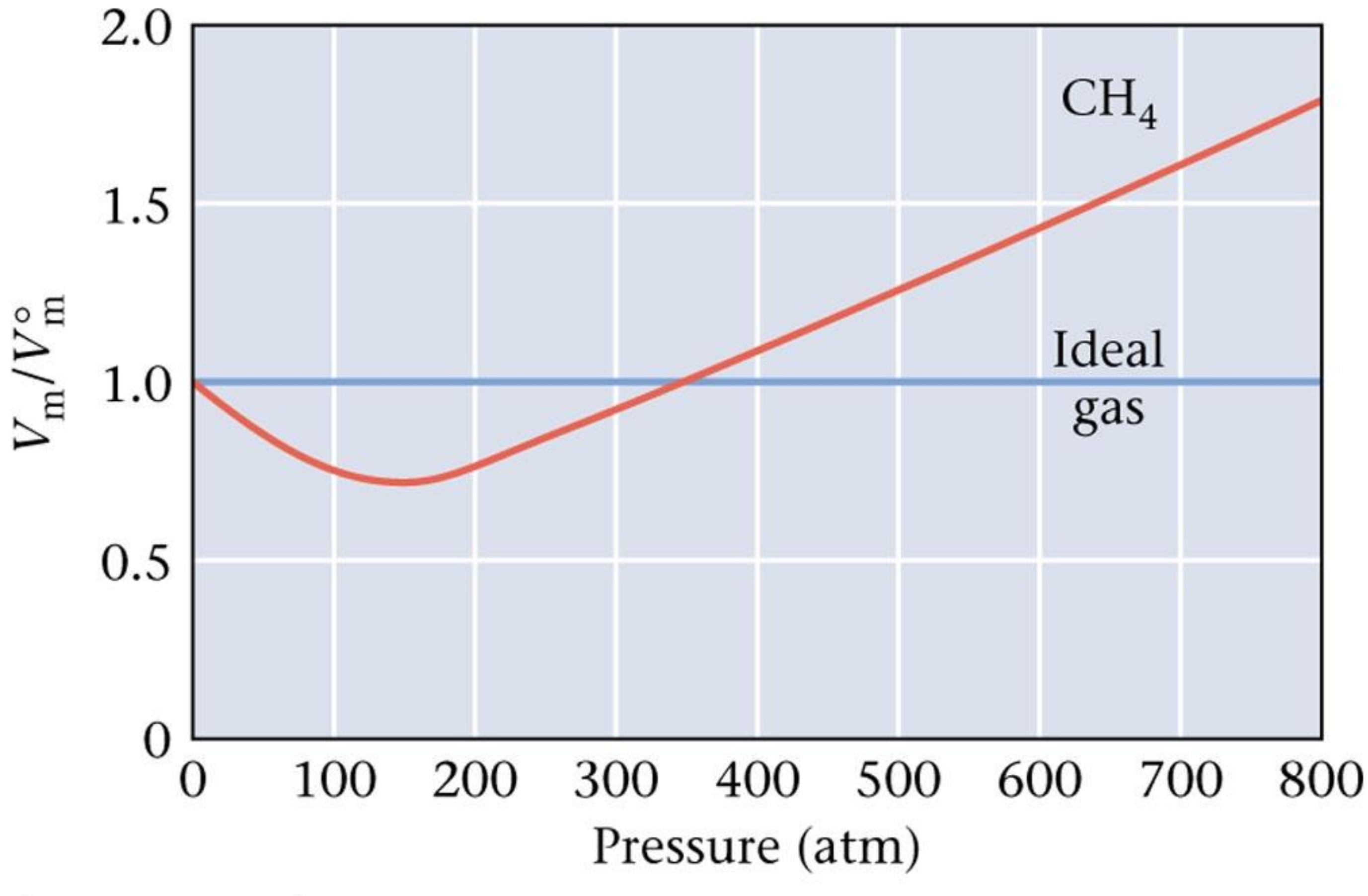
- 氣體間有吸引力，將氣體分子拉近而降低彼此間的距離，氣體分子緊縮而縮小體積

$$\frac{V_m - V_m^{\circ}}{V_m^{\circ}} < 0$$



Practical Volume

$$\frac{V_m - V_m^{\circ}}{V_m^{\circ}} > 0$$

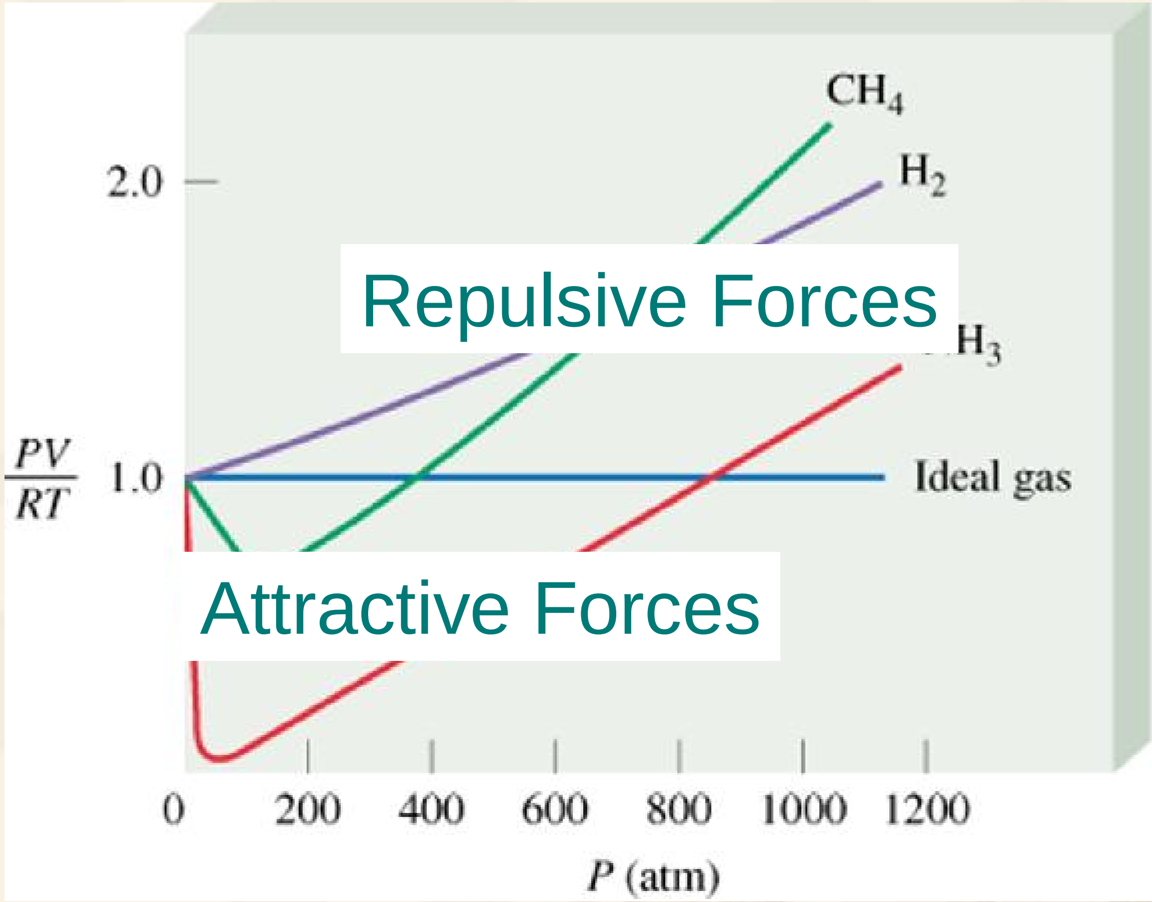


Deviations from Ideal Behavior

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

$$PV = nRT$$

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



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 $\left(\frac{atm \cdot L^2}{mol^2}\right)$	b $\left(\frac{L}{mol}\right)$
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

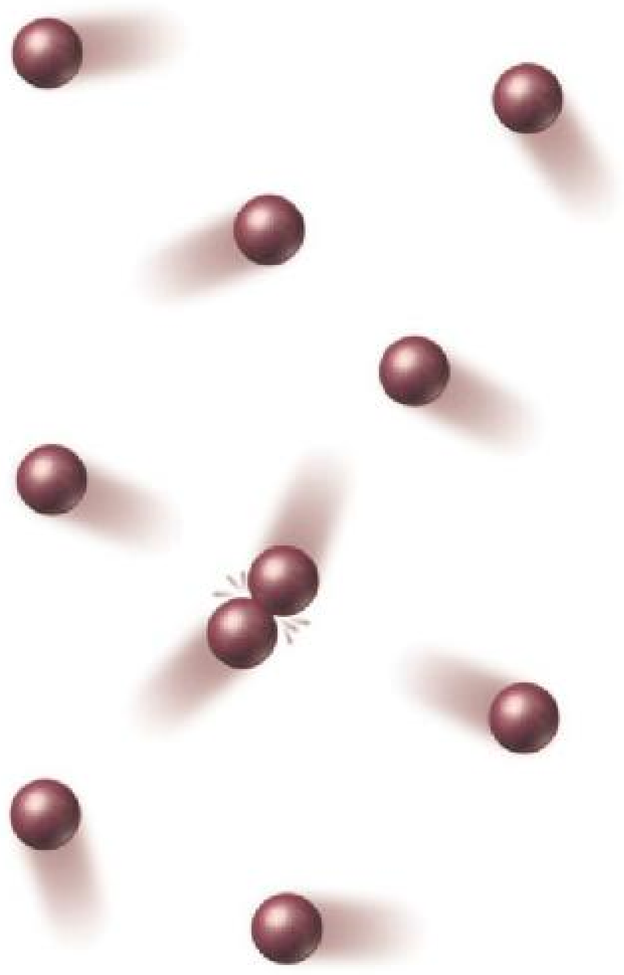
Kinetic Molecular Theory

- the particles of the gas, (either atoms or molecules), are constantly moving
- the attraction between particles is negligible
- when the moving particles hit another particle or the container, they do not stick; but they bounce off and continue moving in another direction
  - like billiard balls

Kinetic Molecular Theory

- there is a lot of empty space between the particles
  - compared to the size of the particles
- the average kinetic energy of the particles is directly proportional to the Kelvin temperature
  - as you raise the temperature of the gas, the average speed of the particles increases
    - but don't be fooled into thinking all the particles are moving at the same speed!!

Kinetic Molecular Theory



1 Collection of particles in constant motion

2 No attractions or repulsions between particles; collisions like billiard ball collisions

3 A lot of space between the particles compared to the size of the particles themselves

4 The speed that the particles move increases with increasing temperature

Gas Properties Explained

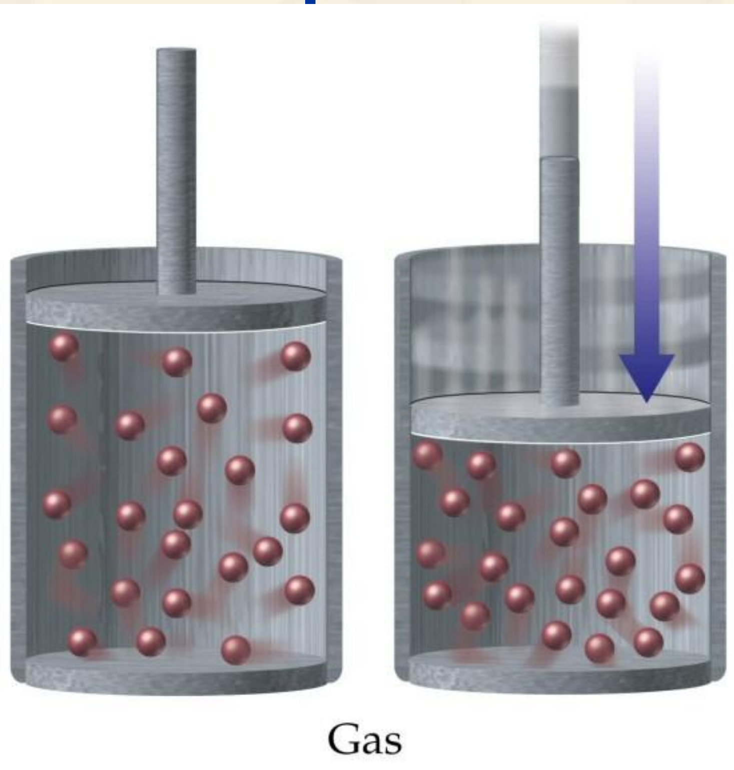
- Gases have indefinite shape and volume because the freedom of the molecules allows them to move and fill the container they're in
- Gases are compressible and have low density because of the large spaces between the molecules

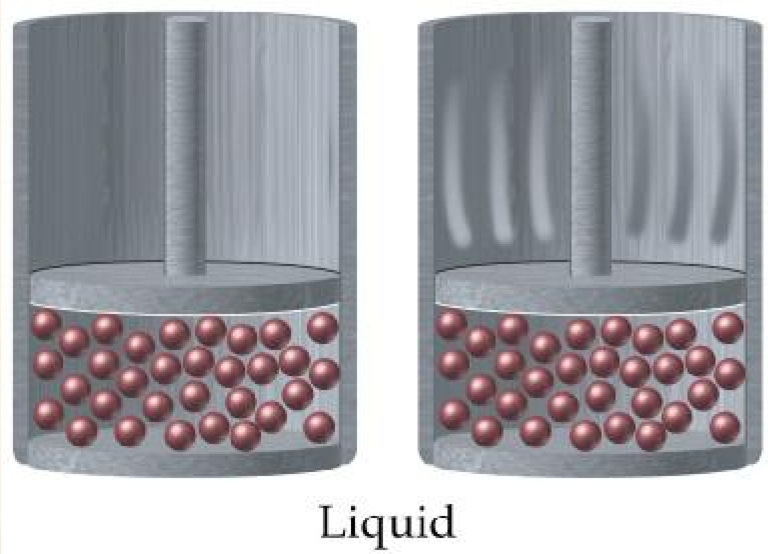
Properties – Indefinite Shape and Indefinite Volume

Because the gas molecules have enough kinetic energy to overcome attractions, they keep moving around and spreading out until they fill the container

As a result, gases take the shape and the volume of the container they are in.

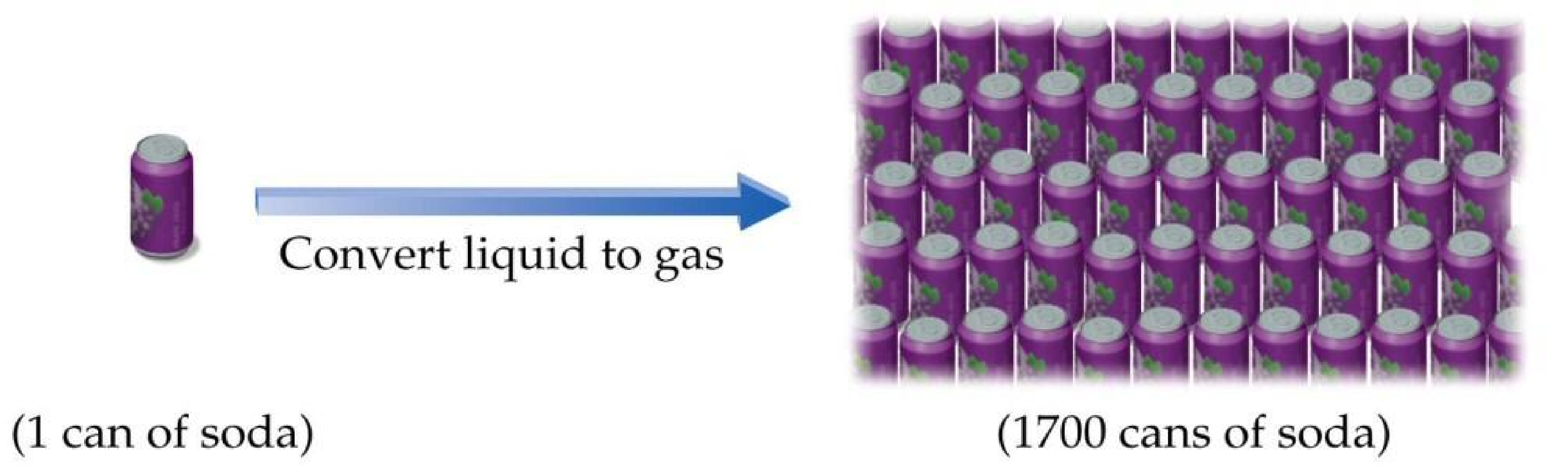
Properties - Compressibility





Because there is a lot of unoccupied space in the structure of a gas, the gas molecules can be squeezed closer together

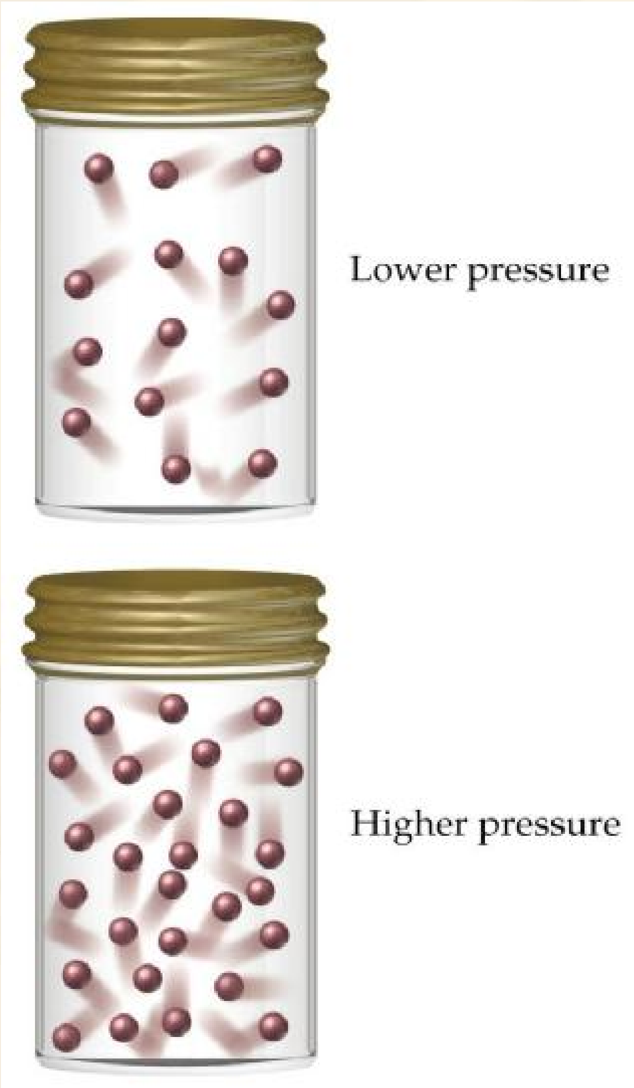
Properties – Low Density



Because there is a lot of unoccupied space in the structure of a gas, gases have low density

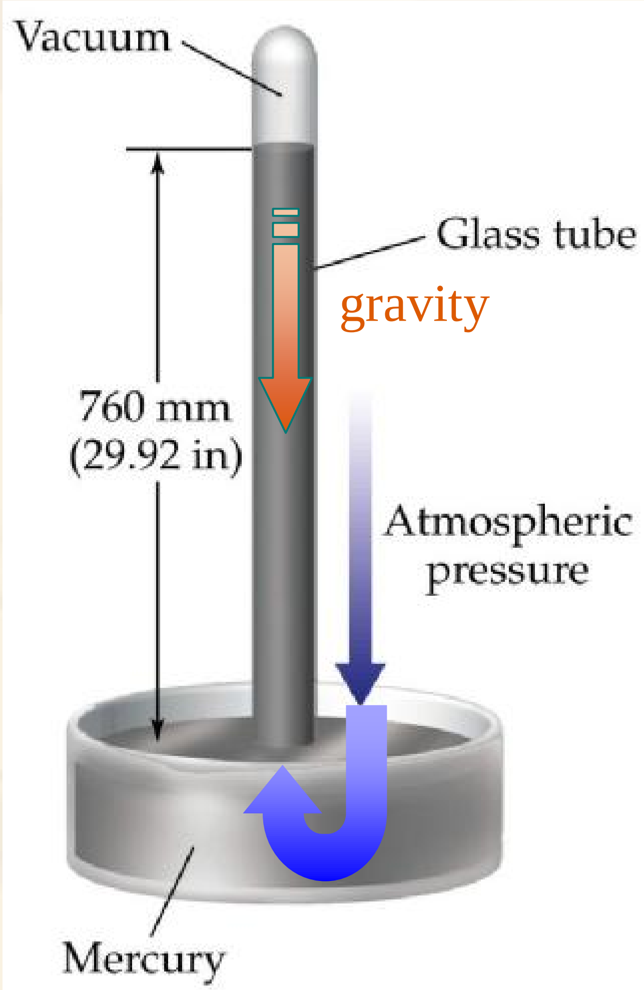
The Pressure of a Gas

- result of the constant movement of the gas molecules and their collisions with the surfaces around them
- the pressure of a gas depends on several factors
  - number of gas particles in a given volume
  - volume of the container
  - average speed of the gas particles



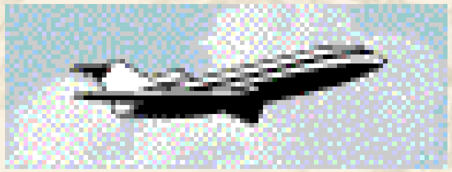
Measuring Air Pressure

- use a **barometer**
- column of mercury supported by air pressure
- force of the air on the surface of the mercury balanced by the pull of gravity on the column of mercury



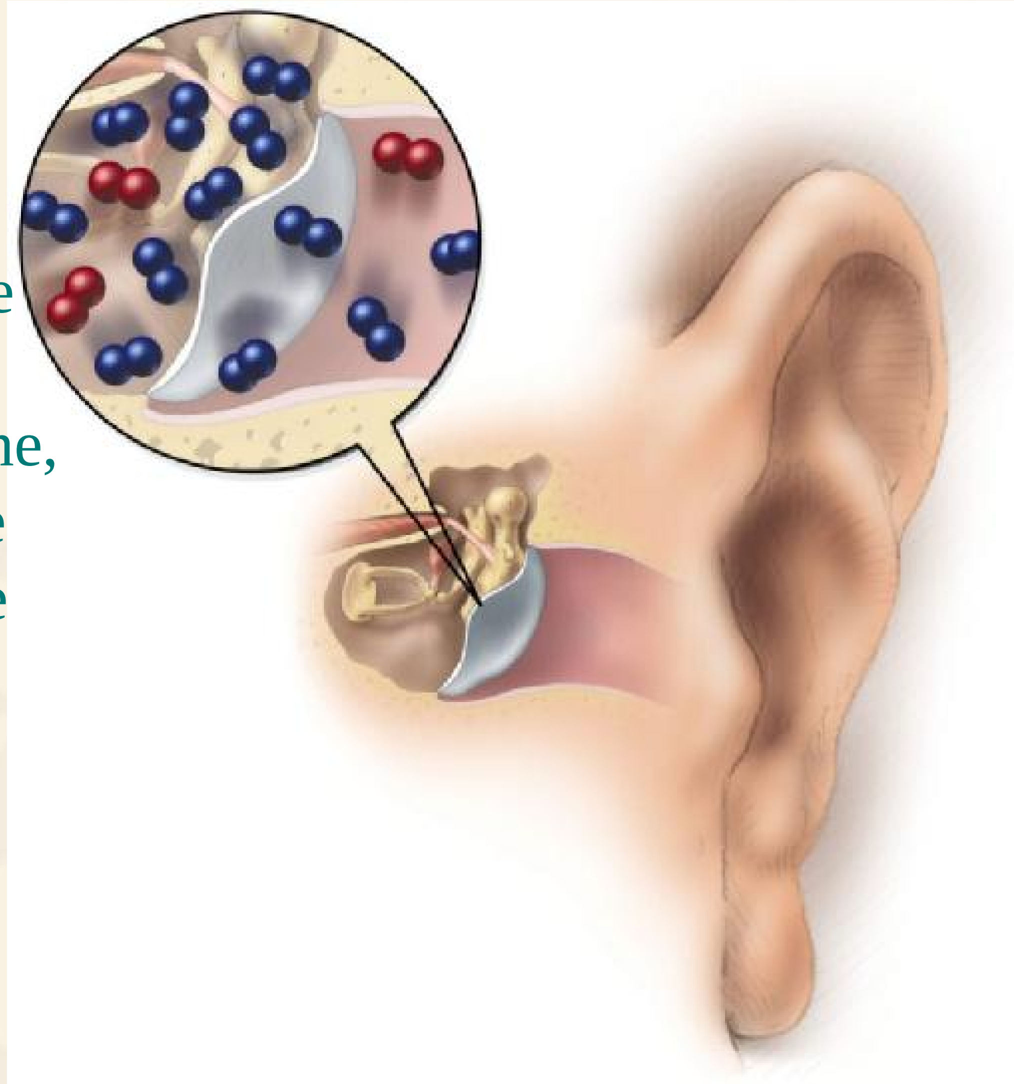
Atmospheric Pressure & Altitude

- the higher up in the atmosphere you go, the lower the atmospheric pressure is around you
  - at the surface the atmospheric pressure is 14.7 psi, but at 10,000 ft is is only 10.0 psi
- rapid changes in atmospheric pressure may cause your ears to "pop" due to an imbalance in pressure on either side of your ear drum



Pressure Imbalance in Ear

If there is a difference in pressure across the eardrum membrane, the membrane will be pushed out – what we commonly call a "popped eardrum."



Example 11.1: Converting Between Pressure Units

Example:

- A high-performance road bicycle is inflated to a total pressure of 125 psi. What is the pressure in millimeters of mercury?

Example:
A high-performance road bicycle is inflated to a total pressure of 125 psi. What is the pressure in millimeters of mercury?

- Write down the given quantity and its units.

Given: 125 psi

Example: A high-performance road bicycle is inflated to a total pressure of 125 psi. What is the pressure in millimeters of mercury?	Information Given: 125 psi

- Write down the quantity to find and/or its units.

  Find: ? mmHg

Example: A high-performance road bicycle is inflated to a total pressure of 125 psi. What is the pressure in millimeters of mercury?	Information Given: 125 psi Find: ? mmHg

- Collect Needed Conversion Factors:

$$14.7 \text{ psi} = 760 \text{ mmHg}$$

Example: A high-performance road bicycle is inflated to a total pressure of 125 psi. What is the pressure in millimeters of mercury?	Information Given: 125 psi Find: ? mmHg CF: 14.7 psi = 760 mmHg

- Write a Solution Map for converting the units :

$$\text{psi} \longrightarrow \text{mmHg}$$

$$\frac{760\ \text{mmHg}}{14.7\ \text{psi}}$$

Example: A high-performance road bicycle is inflated to a total pressure of 125 psi. What is the pressure in millimeters of mercury?	Information Given: 125 psi Find: ? mmHg CF: 14.7 psi = 760 mmHg SM: psi → mmHg

- Apply the Solution Map:

$$125\ \cancel{\text{psi}} \times \frac{760\ \text{mmHg}}{14.7\ \cancel{\text{psi}}} = \text{mmHg}$$

$= 6.46259 \times 10^3$ mmHg

- Sig. Figs. & Round: $= 6.46 \times 10^3$ mmHg

the 760 is an exact number and does not effect the significant figures

Example: A high-performance road bicycle is inflated to a total pressure of 125 psi. What is the pressure in millimeters of mercury?	Information Given: 125 psi Find: ? mmHg CF: 14.7 psi = 760 mmHg SM: psi → mmHg

- Check the Solution:

$$125 \text{ psi} = 6.46 \times 10^3 \text{ mmHg}$$

The units of the answer, mmHg, are correct.
The magnitude of the answer makes sense since mmHg are smaller than psi.

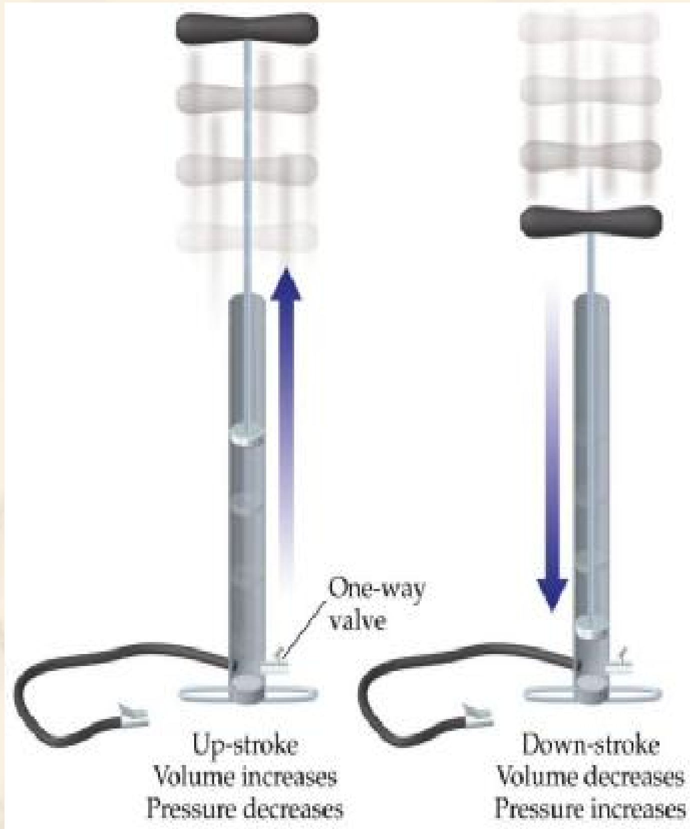
Boyle's Law

pressure of a gas is inversely proportional to its volume

- constant T and amount of gas
- graph P vs V is curve
- graph P vs 1/V is straight line

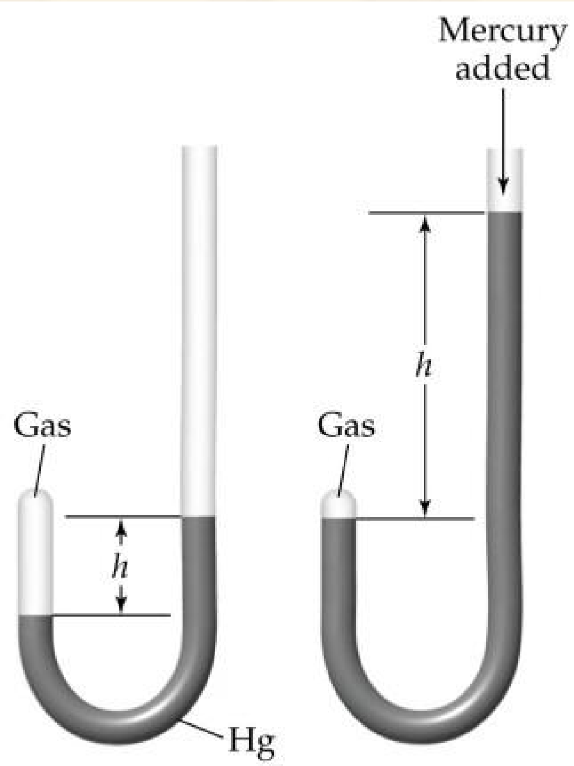
as P increases, V decreases by the same factor

- $P \times V = \text{constant}$
- $P_1 \times V_1 = P_2 \times V_2$

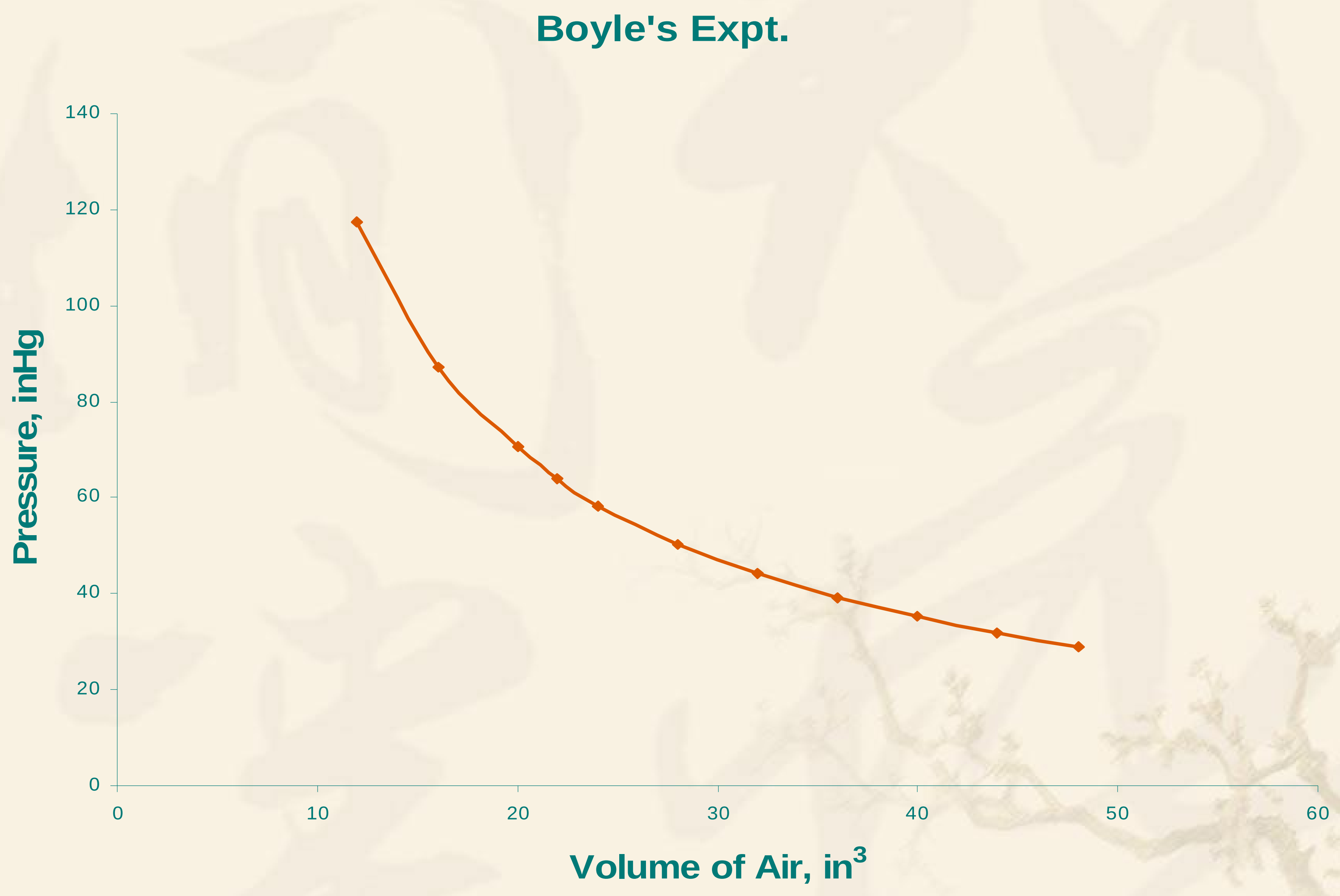


Boyle's Experiment

- added Hg to a J-tube with air trapped inside
- used length of air column as a measure of

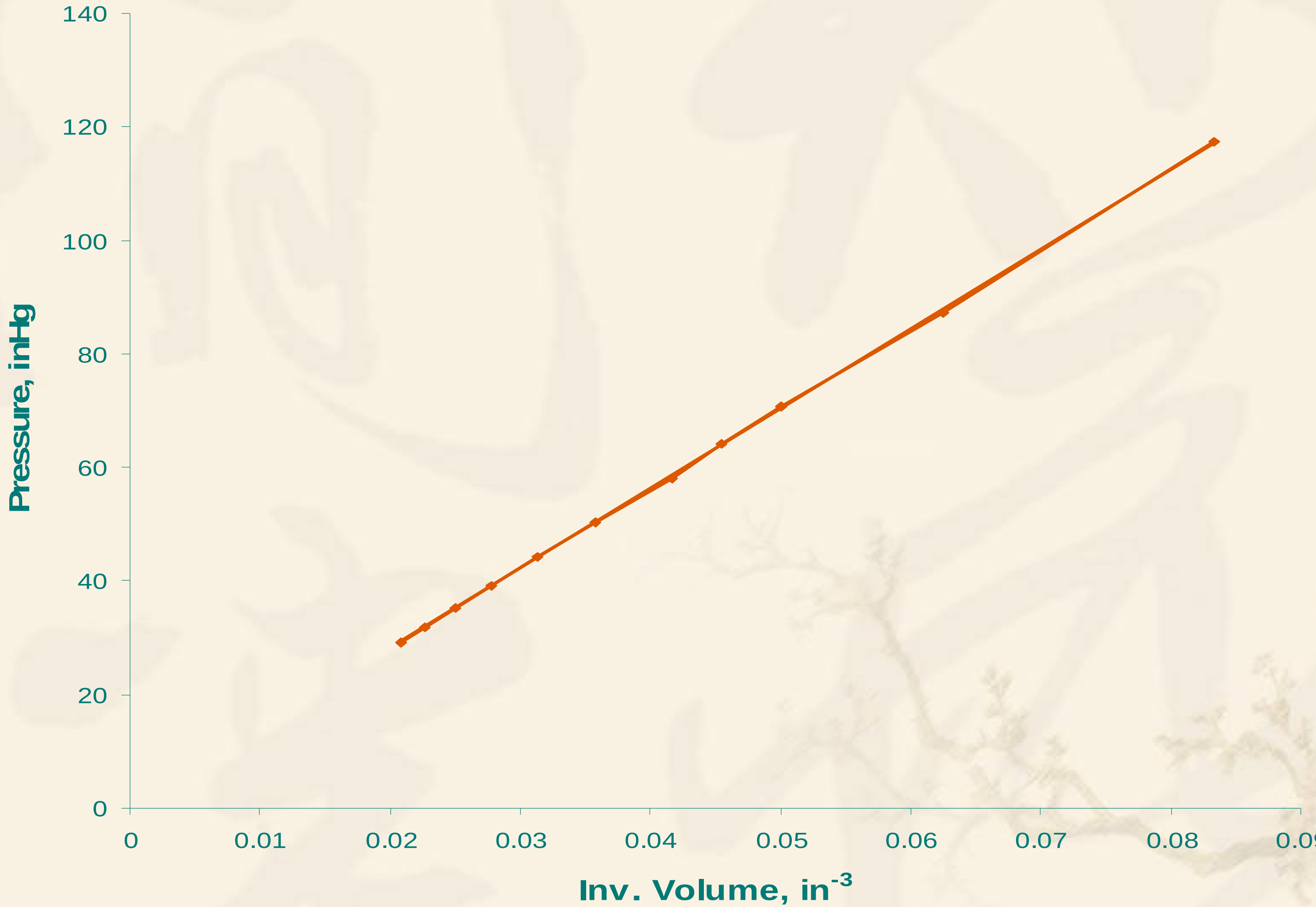


Length of Air in Column (in)	*Difference in Hg Levels (in)*
48	0.0
44	2.8
40	6.2
36	10.1
32	15.1
28	21.2
24	29.7
22	35.0



Boyle's Expt.
Pressure, inHg
140
120
100
80
60
40
20
0
0
10
20
30
40
50
60
Volume of Air, in3

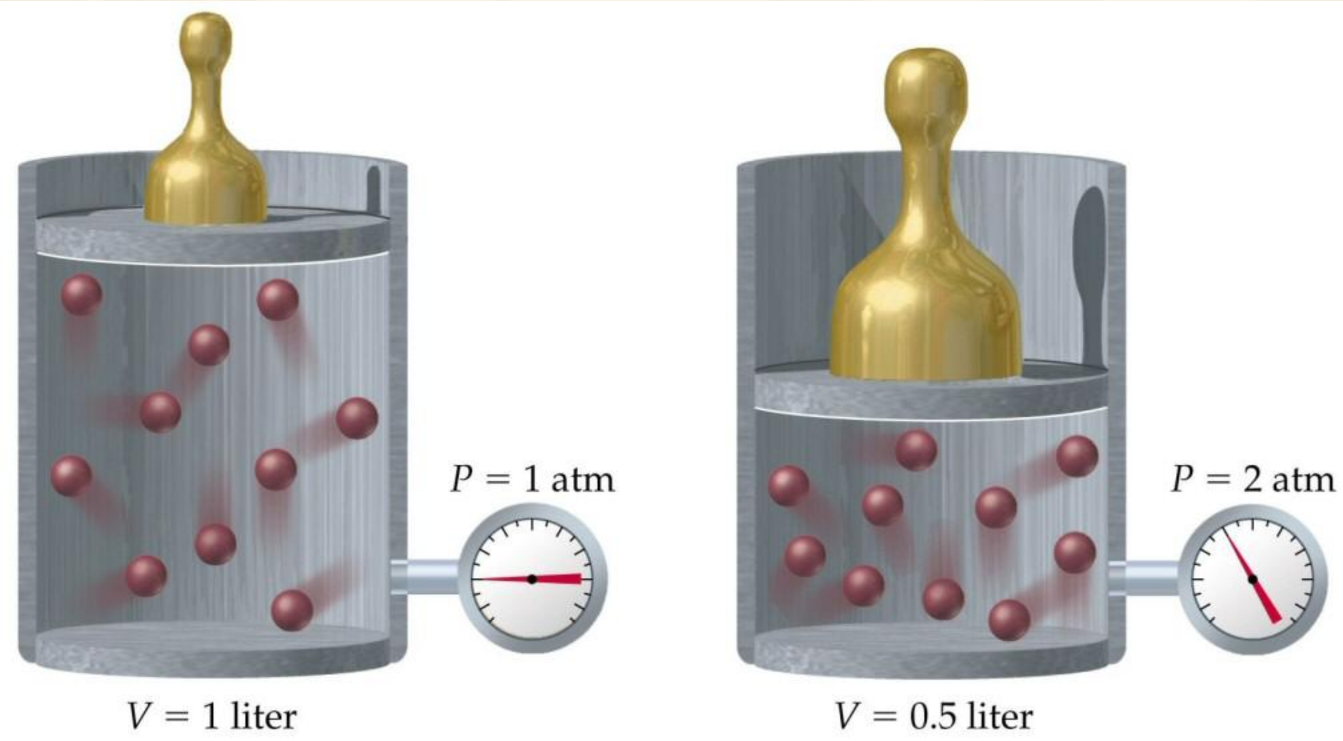
Inverse Volume vs Pressure of Air, Boyle's Expt.



Boyle's Experiment, P x V

Pressure	Volume	P x V
29.13	48	1400
33.50	42	1400
41.63	34	1400
50.31	28	1400
61.31	23	1400
74.13	19	1400
87.88	16	1400
115.56	12	1400

When you double the pressure on a gas, the volume is cut in half, (as long as the temperature and amount of gas do not change)

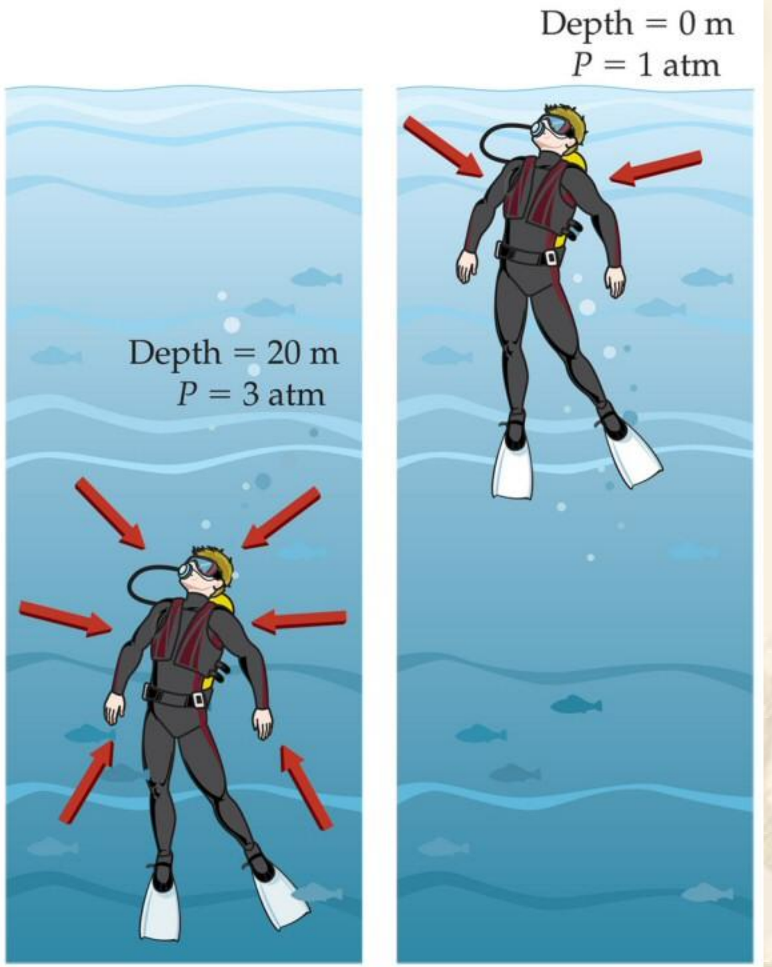


Boyle's Law & Breathing

- inhale
  - diaphragm & rib muscles contract
  - chest cavity expands - volume increase
  - pressure inside lungs drops below air pressure
  - air flows into lung to equilibrate pressure
    - gases move from hi pressure to low
- exhale
  - diaphragm & rib muscles relax
  - chest cavity volume decreases
  - pressure inside lungs rises above air pressure
  - air flows out of lung to equilibrate pressure
- normal healthy person can generate a lung pressure of 1.06 atm

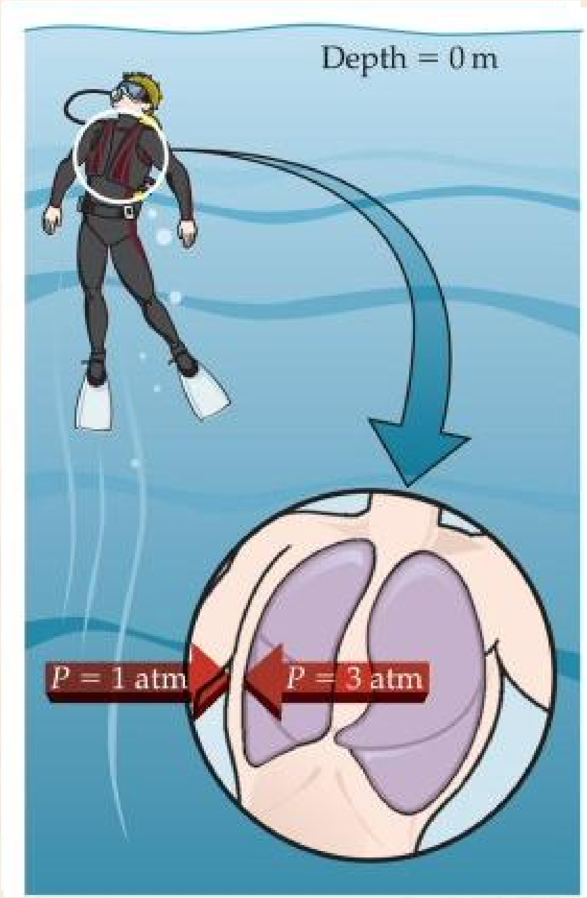
Boyle's Law and Diving

- since water is denser than air, for each 10 m you dive below the surface the pressure on your lungs increases 1 atm
  - at 20 m the total pressure is 3 atm
- if your tank contained air at 1 atm pressure you would not be able to inhale it into your lungs

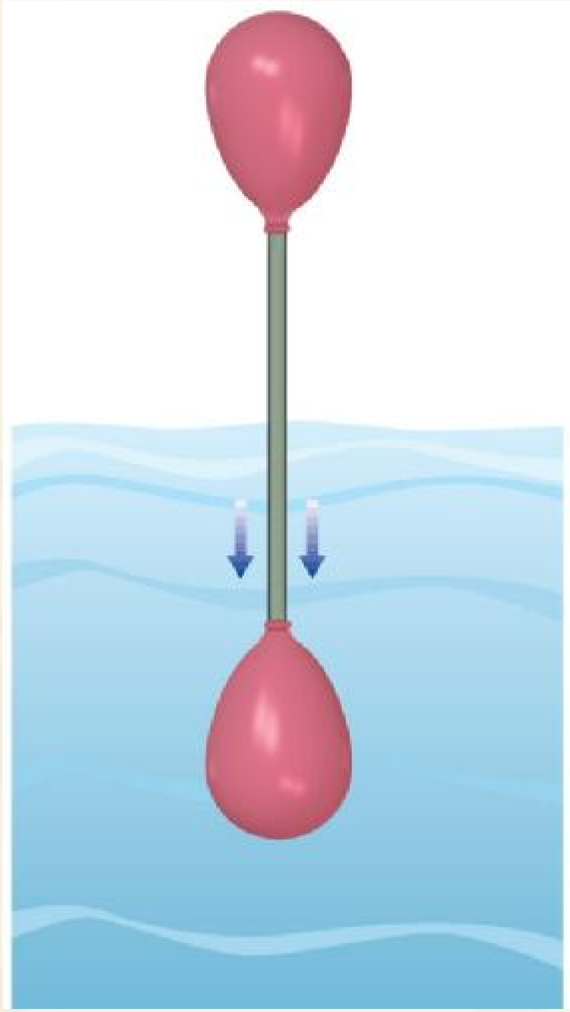


Boyle's Law and Diving

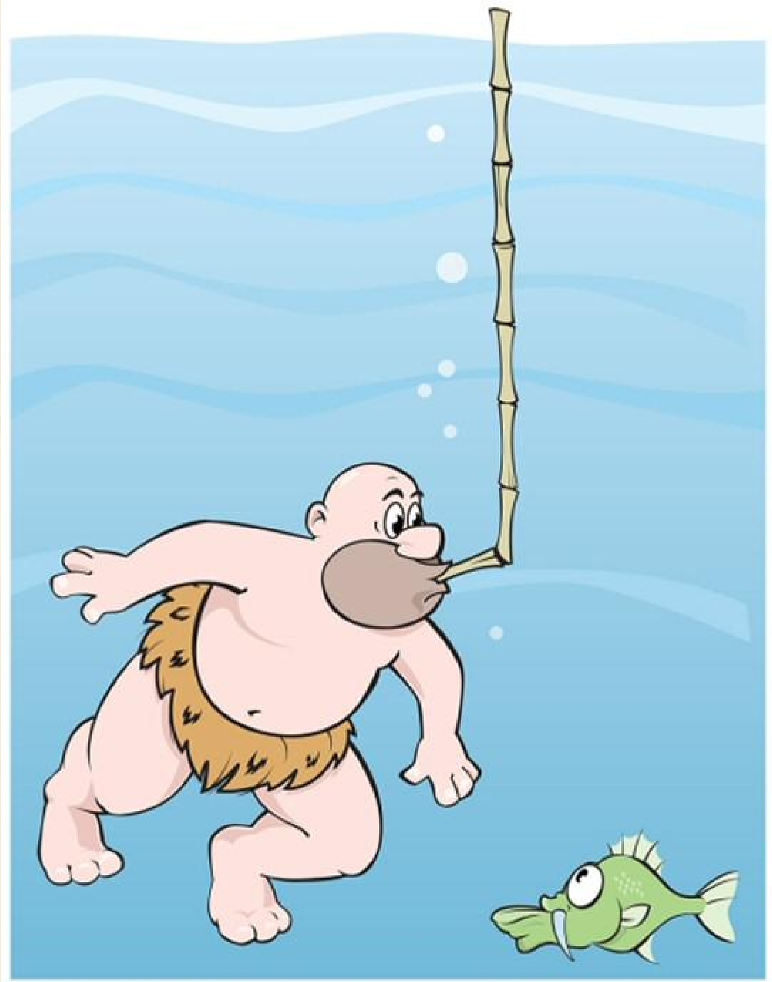
- scuba tanks have a regulator so that the air in the tank is delivered at the same pressure as the water surrounding you
- if a diver holds her breath and rises quickly, so that the outside pressure drops to 1 atm; according to Boyle's Law, what should happen to the volume of air in the lungs?



Which Way Would Air Flow?



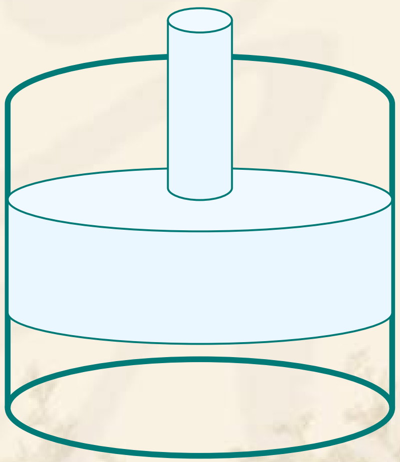
Is this possible at a depth of 20 m?



Example 11.2: Boyle's Law

Example:

- A cylinder equipped with a moveable piston has an applied pressure of 4.0 atm and a volume of 6.0 L. What is the volume if the applied pressure is decreased to 1.0 atm?



Example:
A cylinder equipped with a moveable piston has an applied pressure of 4.0 atm and a volume of 6.0 L. What is the volume if the applied pressure is decreased to 1.0 atm?

- Write down the given quantity and its units.

Given: $P_1 = 4.0$ atm $V_1 = 6.0$ L

$P_2 = 1.0$ atm

Example:
A cylinder equipped with a moveable piston has an applied pressure of 4.0 atm and a volume of 6.0 L. What is the volume if the applied pressure is decreased to 1.0 atm?

Information

Given: P_1 = 4.0 atm V_1 = 6.0 L
P_2 = 1.0 atm

- Write down the quantity to find and/or its units.

Find: V_2, L

Example:

A cylinder equipped with a moveable piston has an applied pressure of 4.0 atm and a volume of 6.0 L. What is the volume if the applied pressure is decreased to 1.0 atm?

Information

Given: P_1 = 4.0 atm V_1 = 6.0 L

P_2 = 1.0 atm

Find: V_2 = ? L

- Collect Needed Equation:

  The relationship between pressure and volume is Boyle's Law

$$P_1 \bullet V_1 = P_2 \bullet V_2$$

Example:

A cylinder equipped with a moveable piston has an applied pressure of 4.0 atm and a volume of 6.0 L. What is the volume if the applied pressure is decreased to 1.0 atm?

Information

Given: $P_1 = 4.0$ atm $V_1 = 6.0$ L

$P_2 = 1.0$ atm

Find: $V_2 = ?$ L

Eq'n: $P_1 \cdot V_1 = P_2 \cdot V_2$

- Write a Solution Map:

$$P_1, V_1, P_2 \longrightarrow V_2$$

$$P_1 \bullet V_1 = P_2 \bullet V_2$$

when using this equation, the units of P_1 and P_2 must be the same, or you will have to convert one to the other

for the same reason, the units of V_2 must be L to match the unit of V_1

Example:
A cylinder equipped with a moveable piston has an applied pressure of 4.0 atm and a volume of 6.0 L. What is the volume if the applied pressure is decreased to 1.0 atm?

Information

Given: P_1 = 4.0 atm V_1 = 6.0 L
P_2 = 1.0 atm

Find: V_2 = ? L

Eq'n: $P_1 \cdot V_1 = P_2 \cdot V_2$

SM: $P_1, V_1, P_2 \rightarrow V_2$

- Apply the Solution Map:

$$P_1 \bullet V_1 = P_2 \bullet V_2$$

$$\frac{P_1 \bullet V_1}{P_2} = V_2$$

$$\frac{(4.0\ \text{atm}) \bullet (6.0\ \text{L})}{(1.0\ \text{atm})} = V_2 = 24\ \text{L}$$

- Sig. Figs. & Round:

Example:
A cylinder equipped with a moveable piston has an applied pressure of 4.0 atm and a volume of 6.0 L. What is the volume if the applied pressure is decreased to 1.0 atm?

Information

Given: $P_1 = 4.0$ atm $V_1 = 6.0$ L

$P_2 = 1.0$ atm

Find: $V_2 = ?$ L

Eq'n: $P_1 \cdot V_1 = P_2 \cdot V_2$

SM: $P_1, V_1, P_2 \rightarrow V_2$

- Check the Solution:

$$V_2 = 24 \text{ L}$$

The units of the answer, L, are correct.
The magnitude of the answer makes sense - since the pressure is decreasing the volume should be increasing.

Temperature Scales

Celsius	Kelvin	Fahrenheit	Rankine	
100°C	373 K	212°F	671 R	BP Water
0°C	273 K	32°F	459 R	MP Ice
-38.9°C	234.1 K	-38°F	421 R	BP Mercury
-183°C	90 K	-297°F	162 R	BP Oxygen
-269°C	4 K	-452°F	7 R	BP Helium
-273°C	0 K	-459 °F	0 R	Absolute Zero

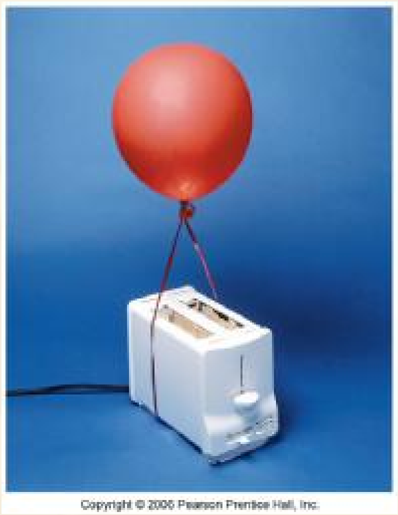
Standard Conditions

- Common reference points for comparing
- standard pressure = 1.00 atm
- standard temperature = 0°C
  - 273 K
- STP

Volume and Temperature

- In a rigid container, raising the temperature increases the pressure
- For a cylinder with a piston, the pressure outside and inside stay the same
- To keep the pressure from rising, the piston moves out increasing the volume of the cylinder
  - as volume increases, pressure decreases

Volume and Temperature



As a gas is heated, it expands. This causes the density of the gas to decrease.

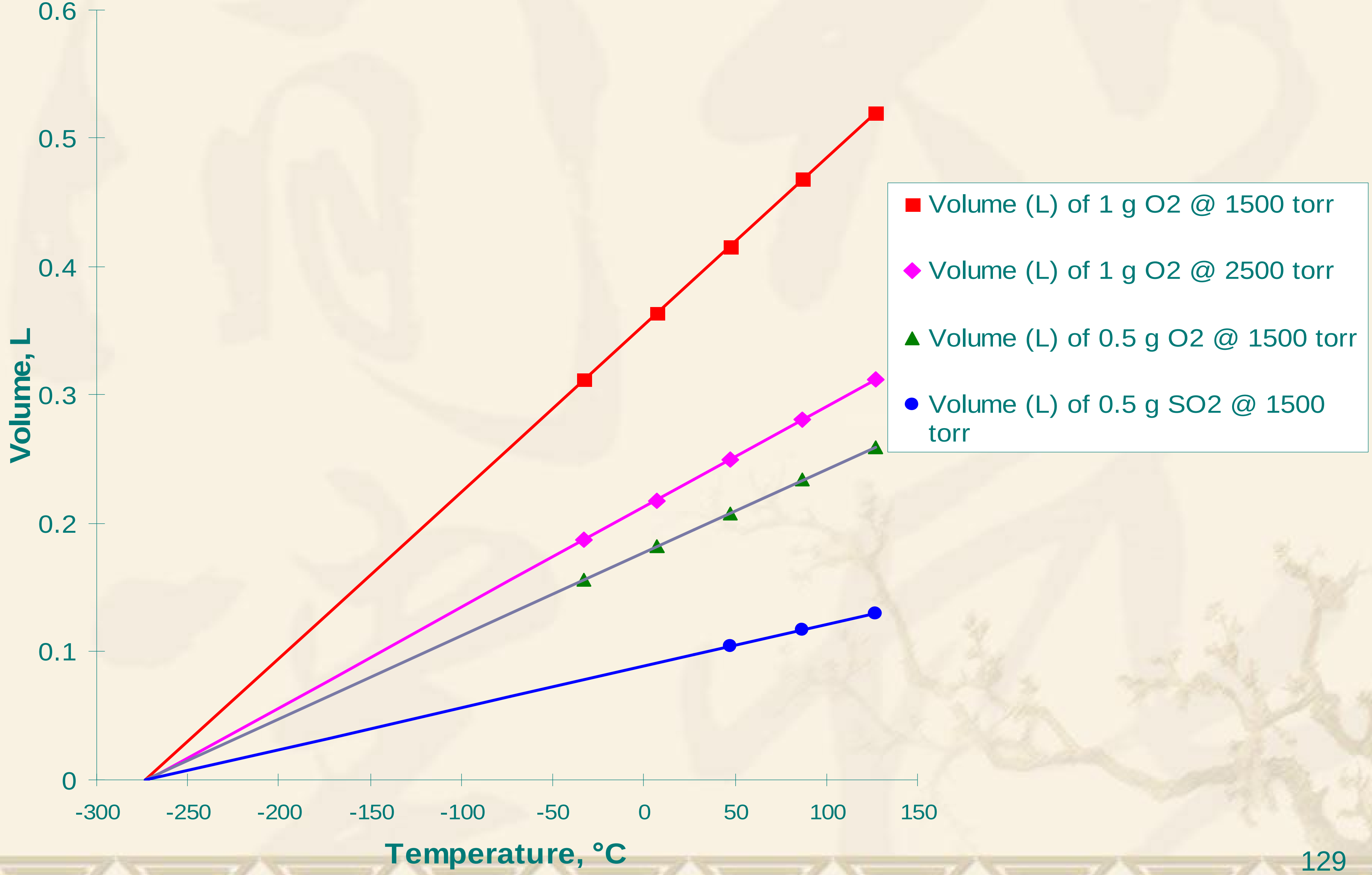
Because the hot air in the balloon is less dense than the surrounding air, it rises.

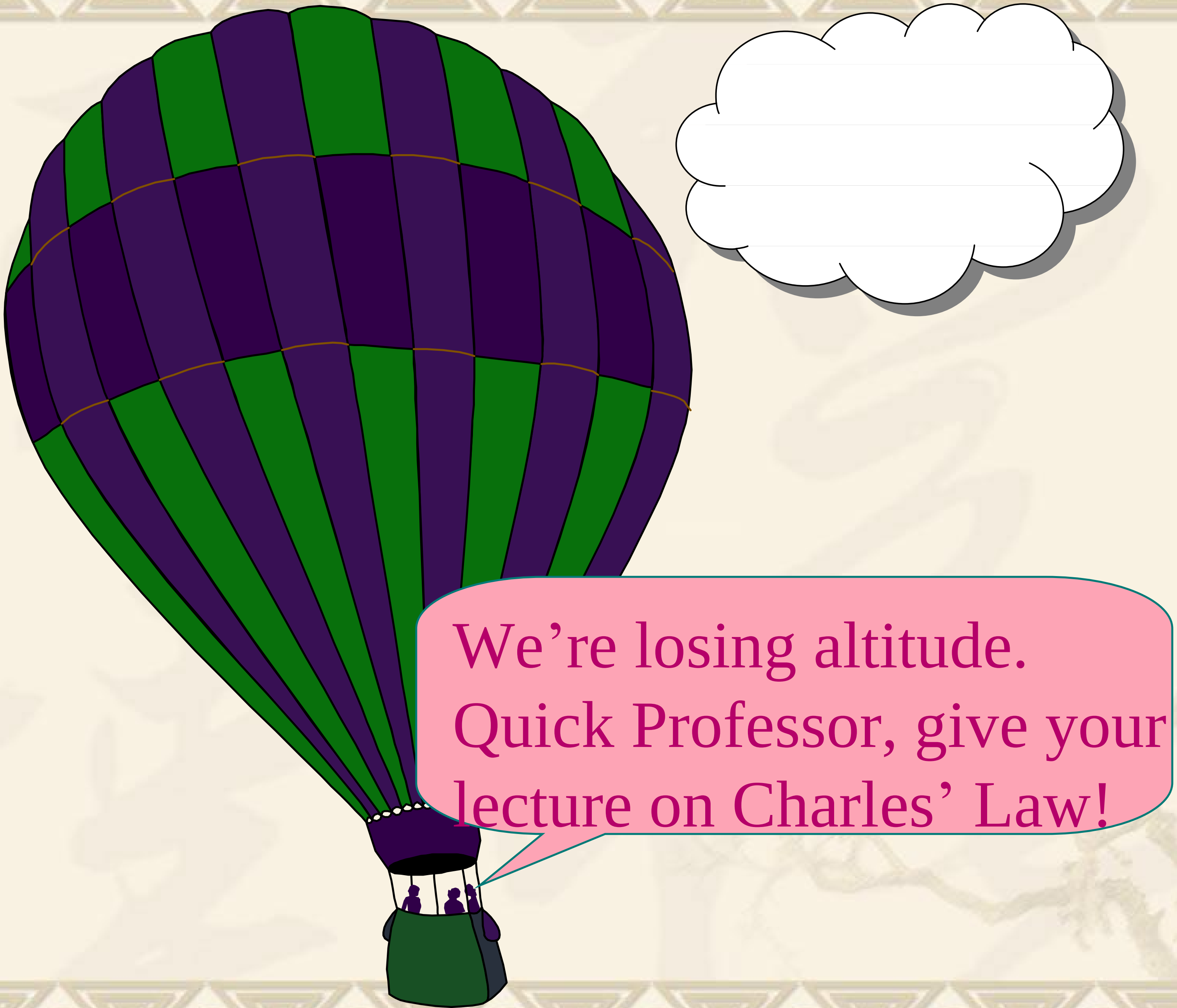
Charles' Law

- volume is directly proportional to temperature
  - constant P and amount of gas
  - graph of V vs T is straight line
- as T increases, V also increases
- Kelvin T = Celsius T + 273
- V = constant x T
  - if T measured in Kelvin

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

Charle's Law & Absolute Zero

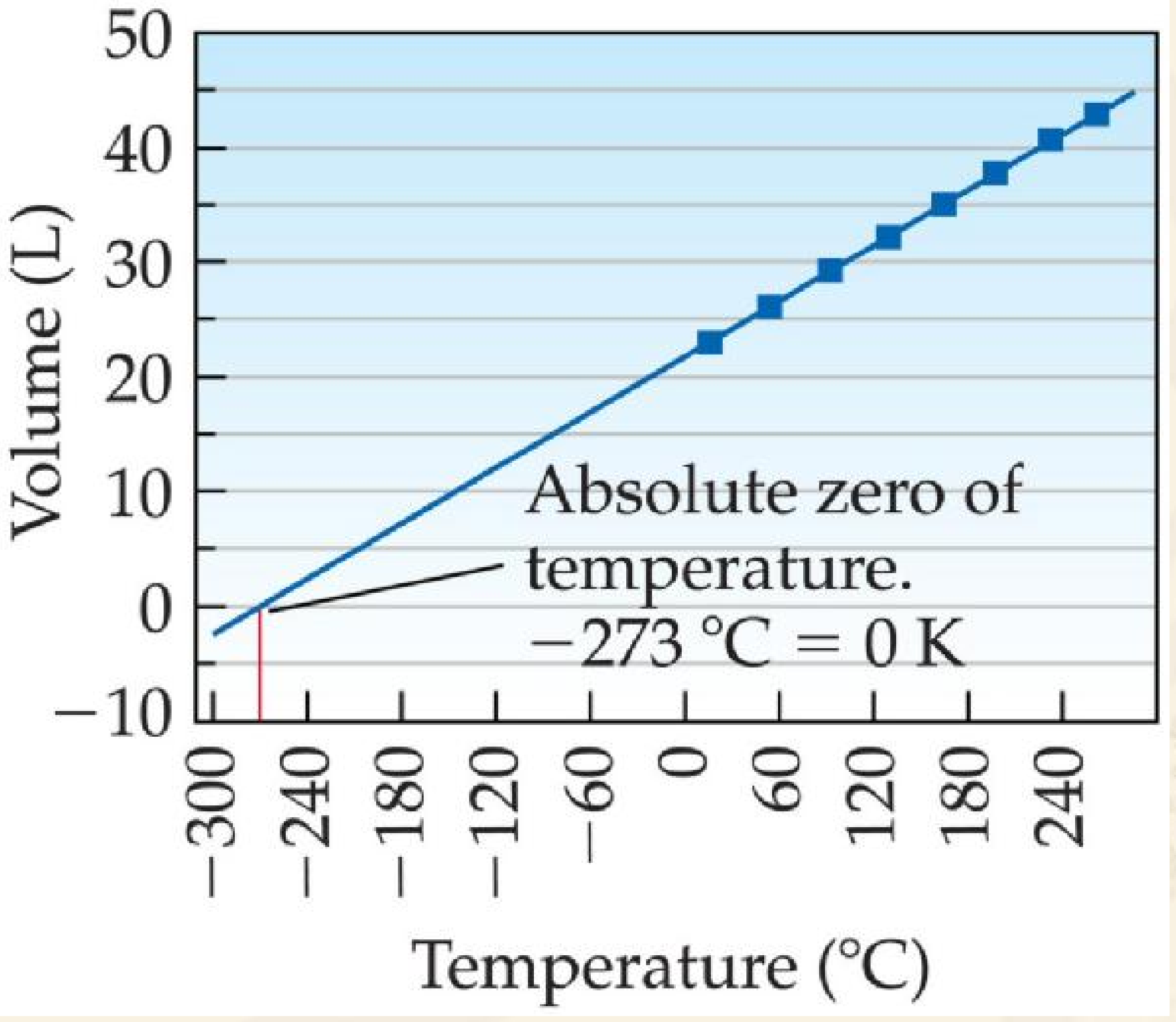



We’re losing altitude.
Quick Professor, give your
lecture on Charles’ Law!

Absolute Zero

- theoretical ***temperature*** at which a gas would have zero volume and no pressure
  - Kelvin calculated by extrapolation
- 0 K = -273.15 °C = -459 °F = 0 R
- never attainable
  - though we've gotten real close!
- all gas law problems use the Kelvin temperature scale!

Determining Absolute Zero

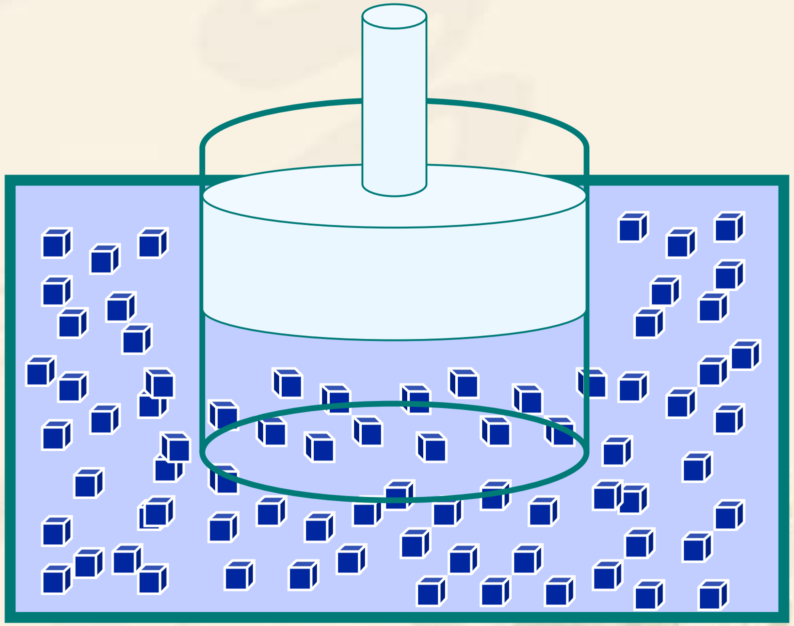


William Thomson, the Lord of Kelvin, extrapolated the line graphs of Volume vs. Temp. to determine the theoretical temp. a gas would have a volume of 0.

Example 11.3: Charles' Law

Example:

- A sample of gas has a volume of 2.80 L at an unknown temperature. When the sample is submerged in ice water at 0°C, its volume decreases to 2.57 L. What was the initial temperature in kelvin and in celsius? (assume constant pressure)



Example:
A gas has a volume of 2.80 L at an unknown temperature. When the sample is at 0°C, its volume decreases to 2.57 L. What was the initial temperature in kelvin and in celsius?

- Write down the given quantity and its units.

Given: $V_1 = 2.80\ L$

$V_2 = 2.57\ L$ $t_2 = 0^{\circ}C$

Example:

A gas has a volume of 2.80 L at an unknown temperature. When the sample is at 0°C, its volume decreases to 2.57 L. What was the initial temperature in kelvin and in celsius?

Information

Given: $V_1 = 2.80$ L

$V_2 = 2.57$ L $t_2 = 0$°C

- Write down the quantity to find and/or its units.

Find: $temp_1$, in K and °C

Example:
A gas has a volume of 2.80 L at an unknown temperature. When the sample is at 0°C, its volume decreases to 2.57 L. What was the initial temperature in kelvin and in celsius?

Information

Given: V_1 = 2.80 L

V_2 = 2.57 L t_2 = 0°C

Find: $temp_1$ in K and °C

- Collect Needed Equation:

  The relationship between temperature and volume is Charles' Law

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

Example:

A gas has a volume of 2.80 L at an unknown temperature. When the sample is at 0°C, its volume decreases to 2.57 L. What was the initial temperature in kelvin and in celsius?

Information

Given: $V_1 = 2.80$ L
$V_2 = 2.57$ L $t_2 = 0°C$

Find: temp$_1$ in K and °C

Eq'n: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

- Write a Solution Map:

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

when using this equation, the units of V_1 and V_2 must be the same, or you will have to convert one to the other

the units of T_1 and T_2 must be kelvin, K

Example:
A gas has a volume of 2.80 L at an unknown temperature. When the sample is at 0°C, its volume decreases to 2.57 L. What was the initial temperature in kelvin and in celsius?

Information

Given: $V_1 = 2.80$ L

$V_2 = 2.57$ L $\quad t_2 = 0°C$

Find: $temp_1$ in K and °C

Eq'n: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

SM: $V_1, V_2\ T_2 \rightarrow T_1$

- Apply the Solution Map:

$$T_2(K) = t_2(°C) + 273$$

$$T_2 = 0 + 273$$

$$T_2 = 273\ K$$

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

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

$$\frac{(2.80\ L)\bullet(273\ K)}{(2.57\ L)} = T_1$$

$$297\ K = T_1$$

Example:
A gas has a volume of 2.80 L at an unknown temperature. When the sample is at 0°C, its volume decreases to 2.57 L. What was the initial temperature in kelvin and in celsius?

Information
Given: V_1 = 2.80 L T_1 = 297 K
V_2 = 2.57 L t_2 = 0°C, T_2 = 273 K
Find: temp_1 in K and °C
Eq'n: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
SM: $V_1, V_2\ T_2 \rightarrow T_1$

- Apply the Solution Map:
  - convert to celsius

$$t_1(°C) = T_1(K) - 273$$

$$t_1 = 297 - 273$$

$$t_1 = 24\ °C$$

Example:
A gas has a volume of 2.80 L at an unknown temperature. When the sample is at 0°C, its volume decreases to 2.57 L. What was the initial temperature in kelvin and in celsius?

Information

Given: V_1 = 2.80 L T_1 = 297 K
V_2 = 2.57 L t_2 = 0°C, T_2 = 273 K

Find: temp in K and °C

Eq'n: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

SM: $V_1, V_2, T_2 \rightarrow T_1$

- Check the Solution:

$$T_1 = 297\ \text{K or } t_1 = 24\ ^\circ\text{C}$$

The units of the answer, K and °C, are correct.
The magnitude of the answer makes sense -
since the volume is decreasing
the temperature should be decreasing.

The Combined Gas Law

- Boyle's Law shows the relationship between pressure and volume
  - at constant temperature
- Charles' Law shows the relationship between volume and absolute temperature
  - at constant pressure
- the two laws can be combined together to give a law that predicts what happens to the volume of a sample of gas when both the pressure and temperature change

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

Example 11.4:
The Combined Gas Law

範例5.6 過氧化氫(H_2O_2)是商業上用以製造漂髮劑的作用成份。在25°C，1atm的條件下，欲生成1.00L的氧氣需要多少克的過氧化氫?其反應式為 $2H_2O_{2(aq)} \rightarrow O_{2(g)} + 2H_2O_{(l)}$

$$n_{O_2} \rightarrow n_{H_2O_2} \rightarrow m_{H_2O_2}$$

$$n_{O_2} = \frac{PV}{RT} = \frac{1.00atm \times 1.00L}{0.0821L \times atm/(mol \times K) \times (25+273)K} = 0.0409molO_2$$

$$n_{H_2O_2} = 0.0409molO_2 \times \frac{2molH_2O_2}{1molO_2} = 0.0818molH_2O_2$$

$$m_{H_2O_2} = 0.0818molH_2O_2 \times \frac{34.02gH_2O_2}{1molH_2O_2} = 2.78gH_2O_2$$

Example:

- A sample of gas has an initial volume of 158 mL at a pressure of 735 mmHg and a temperature of 34°C. If the gas is compressed to a volume of 108 mL and heated to 85°C, what is the final pressure in mmHg?

Example:
A sample of gas has a volume of 158 mL at a pressure of 735 mmHg and a temperature of 34°C. The gas is compressed to a volume of 108 mL and heated to 85°C, what is the final pressure in mmHg?

- Write down the given quantity and its units.

Given: $V_1 = 158$ mL, $P_1 = 735$ mmHg, $t_1 = 34^\circ C$

$V_2 = 108$ mL, $t_2 = 85^\circ C$

Example:
A sample of gas has a volume of 158 mL at a pressure of 735 mmHg and a temperature of 34°C. The gas is compressed to a volume of 108 mL and heated to 85°C, what is the final pressure in mmHg?

Information

Given: V_1 = 158 mL, P_1 = 755 mmHg,
t_1 = 34°C
V_2 = 108 mL, t_2 = 85°C

- Write down the quantity to find and/or its units.

Find: P_2, mmHg

Example: A sample of gas has a volume of 158 mL at a pressure of 735 mmHg and a temperature of 34°C. The gas is compressed to a volume of 108 mL and heated to 85°C, what is the final pressure in mmHg?	Information Given: V_1 = 158 mL, P_1 = 755 mmHg, t_1 = 34°C V_2 = 108 mL, t_2 = 85°C Find: P_2, mmHg

- Collect Needed Equation:

The relationship between pressure, temperature and volume is the Combined Gas Law

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

Example:

A sample of gas has a volume of 158 mL at a pressure of 735 mmHg and a temperature of 34°C. The gas is compressed to a volume of 108 mL and heated to 85°C, what is the final pressure in mmHg?

Information

Given: V_1 = 158 mL, P_1 = 755 mmHg,

t_1 = 34°C

V_2 = 108 mL, t_2 = 85°C

Find: P_2, mmHg

Eq'n: $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$

- Write a Solution Map:

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

when using this equation, the units of V_1 and V_2, and the units of P_1 and P_2, must be the same, or you will have to convert one to the other

the units of T_1 and T_2 must be kelvin, K

Example:

A sample of gas has a volume of 158 mL at a pressure of 735 mmHg and a temperature of 34°C. The gas is compressed to a volume of 108 mL and heated to 85°C, what is the final pressure in mmHg?

Information

Given: V_1 = 158 mL, P_1 = 755 mmHg,

t_1 = 34°C

V_2 = 108 mL, t_2 = 85°C

Find: P_2, mmHg

Eq'n: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

SM: $P_1, V_1, V_2, T_1, T_2 \rightarrow P_2$

- Apply the Solution Map:

$$T_1(\text{K}) = t_1(°\text{C}) + 273$$

$$T_1 = 34 + 273$$

$$T_1 = 307\ \text{K}$$

$$T_2(\text{K}) = t_2(°\text{C}) + 273$$

$$T_2 = 85 + 273$$

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

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

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

$$\frac{(755\ \text{mmHg}) \bullet (158\ \text{mL}) \bullet (358\ \text{K})}{(307\ \text{K}) \bullet (108\ \text{mL})} = P_2$$

$$1.25 \times 10^3\ \text{mmHg} = P_2$$

Example:
A sample of gas has a volume of 158 mL at a pressure of 735 mmHg and a temperature of 34°C. The gas is compressed to a volume of 108 mL and heated to 85°C, what is the final pressure in mmHg?

Information

Given: V_1 = 158 mL, P_1 = 755 mmHg,
t_1 = 34°C → T_1 = 307 K
V_2 = 108 mL, t_2 = 85°C → T_2 = 358 K

Find: P_2, mmHg

Eq'n: $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$

SM: P_1, V_1, V_2, T_1, T_2 → P_2

- Check the Solution:

$$P_2 = 1.25 \times 10^3 \text{ mmHg}$$

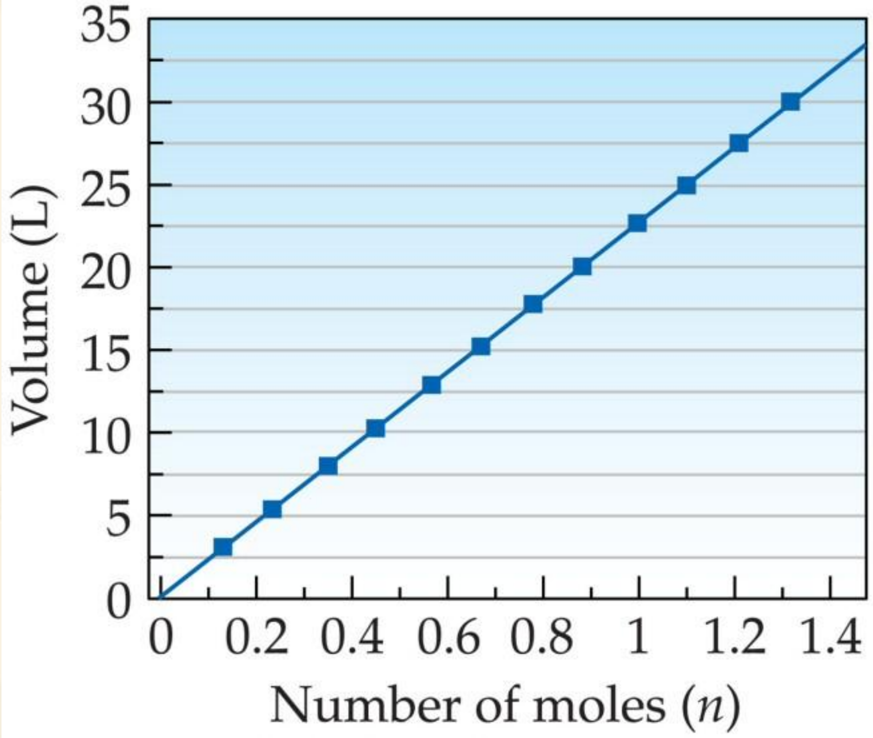
The units of the answer, mmHg, are correct.
The magnitude of the answer makes sense -
since the volume is decreasing and temperature is increasing
the pressure should be increasing.

Avogadro’s Law

- volume directly proportional to the number of gas molecules
  - **V** = constant x **n**
  - constant P and T
  - more gas molecules = larger volume
- count number of gas molecules by **moles**
- equal volumes of gases contain equal numbers of molecules
  - ***the gas doesn’t matter***

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Avogadro's Law



Example 11.5: Avogadro's Law

Example:

- A 4.8 L sample of helium gas contains 0.22 mol helium. How many additional moles of helium must be added to obtain a volume of 6.4 L? (assume constant pressure and temperature)

Example:

A 4.8 L sample of helium gas contains 0.22 mol helium. How many additional moles of helium must be added to obtain a volume of 6.4 L?

- Write down the given quantity and its units.

Given: $V_1 = 4.8\ L$ $n_1 = 0.22\ mol$

$V_2 = 6.4\ L$

Example: A 4.8 L sample of helium gas contains 0.22 mol helium. How many additional moles of helium must be added to obtain a volume of 6.4 L?	Information Given: V_1 = 4.8 L, n_1 =0.22 mol V_2 = 6.4 L

- Write down the quantity to find and/or its units.

Find: n_2, mol; and moles added

Example:	Information
A 4.8 L sample of helium gas contains 0.22 mol helium. How many additional moles of helium must be added to obtain a volume of 6.4 L?	Given: V_1 = 4.8 L, n_1 =0.22 mol V_2 = 6.4 L Find: n_2, mol and added mol

- Collect Needed Equation:

The relationship between temperature and volume is Avogadro's Law

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Example:	Information
A 4.8 L sample of helium gas contains 0.22 mol helium. How many additional moles of helium must be added to obtain a volume of 6.4 L?	Given: V_1 = 4.8 L, n_1 =0.22 mol V_2 = 6.4 L Find: n_2, mol and added mol Eq'n: $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

- Write a Solution Map:

$$V_1, V_2, n_1 \xrightarrow{\quad} n_2$$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

when using this equation, the units of V_1 and V_2 must be the same, or you will have to convert one to the other

the units of n_1 and n_2 must be moles

Example:

A 4.8 L sample of helium gas contains 0.22 mol helium. How many additional moles of helium must be added to obtain a volume of 6.4 L?

Information

Given: V_1 = 4.8 L, n_1 =0.22 mol

V_2 = 6.4 L

Find: n_2, mol and added mol

Eq'n: $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

SM: $V_1, V_2, n_1 \rightarrow n_2$

- Apply the Solution Map:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

$$n_2 = \frac{V_2 \bullet n_1}{V_1} \qquad n_2 = \frac{(6.4\,\mathrm{L}) \bullet (0.22\,\mathrm{mol})}{(4.8\,\mathrm{L})}$$

$$n_2 = 0.29\,\mathrm{mol}$$

Example:

A 4.8 L sample of helium gas contains 0.22 mol helium. How many additional moles of helium must be added to obtain a volume of 6.4 L?

Information

Given: V_1 = 4.8 L, n_1 =0.22 mol

V_2 = 6.4 L

Find: n_2, mol and added mol

Eq'n: $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

SM: $V_1, V_2, n_1 \rightarrow n_2$

- Apply the Solution Map:
  - to get the added moles, subtract n_1 from n_2

$$\text{added moles} = n_2 - n_1$$

$$\text{added moles} = 0.29 - 0.22$$

$$\text{added moles} = 0.07 \text{ moles}$$

Example:
A 4.8 L sample of helium gas contains 0.22 mol helium. How many additional moles of helium must be added to obtain a volume of 6.4 L?

Information

Given: $V_1 = 4.8$ L, $n_1 = 0.22$ mol

$V_2 = 6.4$ L

Find: n_2, mol and added mol

Eq'n: $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

SM: $V_1, V_2, n_1 \rightarrow n_2$

- Check the Solution:

$$n_2 = 0.29 \text{ moles; added } 0.07 \text{ moles}$$

The units of the answer, moles, are correct.
The magnitude of the answer makes sense - since the volume is increasing the total number of moles should be increasing.

Ideal Gas Law

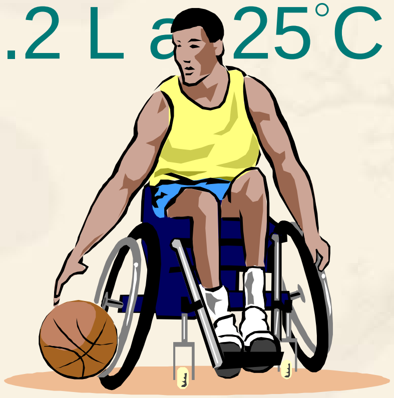
- By combing the gas laws we can write a general equation
- **R** is called the **Gas Constant**
- the value of **R** depends on the units of P and V
  - we will use 0.0821 $\frac{\text{atm} \bullet \text{L}}{\text{mol} \bullet \text{K}}$ and convert P to atm and V to L
- use the Ideal Gas law when have a gas at one condition, use the Combined Gas Law when you have gas whose condition is changing

$$\frac{(P)\bullet(V)}{(n)\bullet(T)} = R \quad \text{or} \quad PV = nRT$$

Example 11.7:
The Ideal Gas Law Requiring Unit Conversion

Example:

- Calculate the number of moles of gas in a basketball inflated to a total pressure of 24.2 psi with a volume of 3.2 L at 25°C



Example:
Calculate the number of moles of gas in a basketball inflated to a total pressure of 24.2 psi with a volume of 3.2 L at 25°C

- Write down the given quantity and its units.

  Given: $V = 3.2$ L, $P = 24.2$ psi, $t = 25°C$

Example:
Calculate the number of moles of gas in a basketball inflated to a total pressure of 24.2 psi with a volume of 3.2 L at 25°C

Information
Given: V = 3.2 L, P = 24.2 psi,
t = 25°C

- Write down the quantity to find and/or its units.

Find: n, mol

Example:	Information
Calculate the number of moles of gas in a basketball inflated to a total pressure of 24.2 psi with a volume of 3.2 L at 25°C	Given: V = 3.2 L, P = 24.2 psi, t = 25°C Find: n, mol

- Collect Needed Equation:

The relationship between pressure, temperature, number of moles and volume is the Ideal Gas Law

$$\frac{(P)\bullet(V)}{(n)\bullet(T)} = R \quad \text{or} \quad PV = nRT$$

Example:
Calculate the number of moles of gas in a basketball inflated to a total pressure of 24.2 psi with a volume of 3.2 L at 25°C

Information
Given: V = 3.2 L, P = 24.2 psi,
t = 25°C
Find: n, mol
Eq'n: PV = nRT

- Write a Solution Map:

$$\boxed{P,V,T,R} \xrightarrow{\quad PV = nRT \quad} \boxed{n}$$

when using the Ideal Gas Equation, the units of V must be L;
and the units of P must be atm, or you will have to convert

the units of T must be kelvin, K

Example:
Calculate the number of moles of gas in a basketball inflated to a total pressure of 24.2 psi with a volume of 3.2 L at 25°C

Information
Given: V = 3.2 L, P = 24.2 psi,
t = 25°C
Find: n, mol
Eq'n: PV = nRT
SM: P,V,T,R → n

- Apply the Solution Map:
  - convert the units

$$T(K) = t(°C) + 273$$
$$T = 25 + 273$$
$$T = 298\ K$$

$$P = 24.2\ \text{psi} \times \frac{1\ \text{atm}}{14.7\ \text{psi}} = 1.6\underline{4}62\ \text{atm}$$

Example:
Calculate the number of moles of gas in a basketball inflated to a total pressure of 24.2 psi with a volume of 3.2 L at 25°C

Information

Given: V = 3.2 L, P = 1.6$\underline{4}$62 atm,
T = 298 K

Find: n, mol

Eq'n: PV = nRT

SM: P,V,T,R → n

- Apply the Solution Map:

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

$$\frac{P \bullet V}{R \bullet T} = n = \frac{(1.6\underline{4}62\ \text{atm})(3.2\ \text{L})}{\left(0.0821\ \frac{\text{L}\bullet\text{atm}}{\text{mol}\bullet\text{K}}\right)(298\ \text{K})}$$

$$n = 0.22\ \text{mol}$$

Example:
Calculate the number of moles of gas in a basketball inflated to a total pressure of 24.2 psi with a volume of 3.2 L at 25°C

Information
Given: V = 3.2 L, P = 1.6$\underline{4}$62 atm,
T = 298 K
Find: n, mol
Eq'n: PV = nRT
SM: P,V,T,R → n

- Check the Solution:

$$n = 0.22 \text{ moles}$$

The units of the answer, moles, are correct.
It is hard to judge the magnitude with so many variables, but with a volume less than 22.4 L at pressures near 1 atm and temperatures near 273 K – it is reasonable to have less than 1 mole of gas

Molar Mass of a Gas

- one of the methods chemists use to determine the molar mass of an unknown substance is to heat a weighed sample until it becomes a gas, measure the temperature, pressure and volume, and use the Ideal Gas Law

$$\text{Molar Mass} = \frac{\text{mass in grams}}{\text{moles}}$$

Example 11.8:
Molar Mass Using The Ideal Gas Law and a Mass Measurement

Example:

- A sample of a gas has a mass of 0.311 g. Its volume is 0.225 L at a temperature of 55°C and a pressure of 886 mmHg. Find its molar mass.

Example:

A sample of a gas has a mass of 0.311 g. Its volume is 0.225 L at a temperature of 55°C and a pressure of 886 mmHg. Find its molar mass.

- Write down the given quantity and its units.

Given: V = 0.225 L, P = 886 mmHg, t = 55°C

m = 0.311 g

Example: A sample of a gas has a mass of 0.311 g. Its volume is 0.225 L at a temperature of 55°C and a pressure of 886 mmHg. Find its molar mass.	Information Given: V = 0.225 L, P = 886 mmHg, t = 55°C, m = 0.311 g

- Write down the quantity to find and/or its units.

Find: molar mass (g/mol)

Example:

A sample of a gas has a mass of 0.311 g. Its volume is 0.225 L at a temperature of 55°C and a pressure of 886 mmHg. Find its molar mass.

Information

Given: V = 0.225 L, P = 886 mmHg,
t = 55°C, m = 0.311 g

Find: molar mass, (g/mol)

- Collect Needed Equations:

The relationship between pressure, temperature, number of moles and volume is the Ideal Gas Law

$$PV = nRT$$

The relationship between mass and moles is the molar mass

$$\text{Molar Mass} = \frac{\text{mass in grams}}{\text{moles}}$$

Example:
A sample of a gas has a mass of 0.311 g. Its volume is 0.225 L at a temperature of 55°C and a pressure of 886 mmHg. Find its molar mass.

Information
Given: V = 0.225 L, P = 886 mmHg, t = 55°C, m = 0.311 g
Find: molar mass, (g/mol)
Eq'n: PV = nRT; MM = mass/moles

- Write a Solution Map:

$$\text{P,V,T,R} \xrightarrow[\text{PV = nRT}]{} \text{n} \xrightarrow{\text{mass}} \text{Molar Mass}$$

when using the Ideal Gas Equation, the units of V must be L; and the units of P must be atm, or you will have to convert

the units of T must be kelvin, K

Example:	Information
A sample of a gas has a mass of 0.311 g. Its volume is 0.225 L at a temperature of 55°C and a pressure of 886 mmHg. Find its molar mass.	Given: V = 0.225 L, P = 886 mmHg, t = 55°C, m = 0.311 g Find: molar mass, (g/mol) Eq'n: PV = nRT; MM = mass/moles SM: P,V,T,R → n & mass → mol. mass

- Apply the Solution Map:
  - convert the units

$$T(K) = t(°C) + 273$$

$$T = 55 + 273$$

$$T = 328\ K$$

$$P = 886\ mmHg \times \frac{1\ atm}{760\ mmHg} = 1.1\underline{6}58\ mmHg$$

Example: A sample of a gas has a mass of 0.311 g. Its volume is 0.225 L at a temperature of 55°C and a pressure of 886 mmHg. Find its molar mass.	Information Given: V = 0.225 L, P = 1.1658 atm, t = 328 K, m = 0.311 g Find: molar mass, (g/mol) Eq'n: PV = nRT; MM = mass/moles SM: P,V,T,R → n & mass → mol. mass

- Apply the Solution Map:

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

$$\frac{P \bullet V}{R \bullet T} = n = \frac{(1.16\underline{5}8\ \text{atm})(0.225\ \text{L})}{\left(0.0821\ \frac{\text{L}\bullet\text{atm}}{\text{mol}\bullet\text{K}}\right)(328\ \text{K})}$$

$$n = 9.7\underline{4}06 \times 10^{-3}\ \text{mol}$$

$$\text{Molar Mass} = \frac{\text{mass in grams}}{\text{moles}}$$

$$\text{Molar Mass} = \frac{0.311\text{g}}{9.7406 \times 10^{-3}\ \text{moles}}$$

$$\text{Molar Mass} = 31.9\ \text{g/mol}$$

Example: A sample of a gas has a mass of 0.311 g. Its volume is 0.225 L at a temperature of 55°C and a pressure of 886 mmHg. Find its molar mass.	Information Given: V = 0.225 L, P = 1.1658 atm, t = 328 K, m = 0.311 g Find: molar mass, (g/mol) Eq'n: PV = nRT; MM = mass/moles SM: P,V,T,R → n & mass → mol. mass

- Check the Solution:

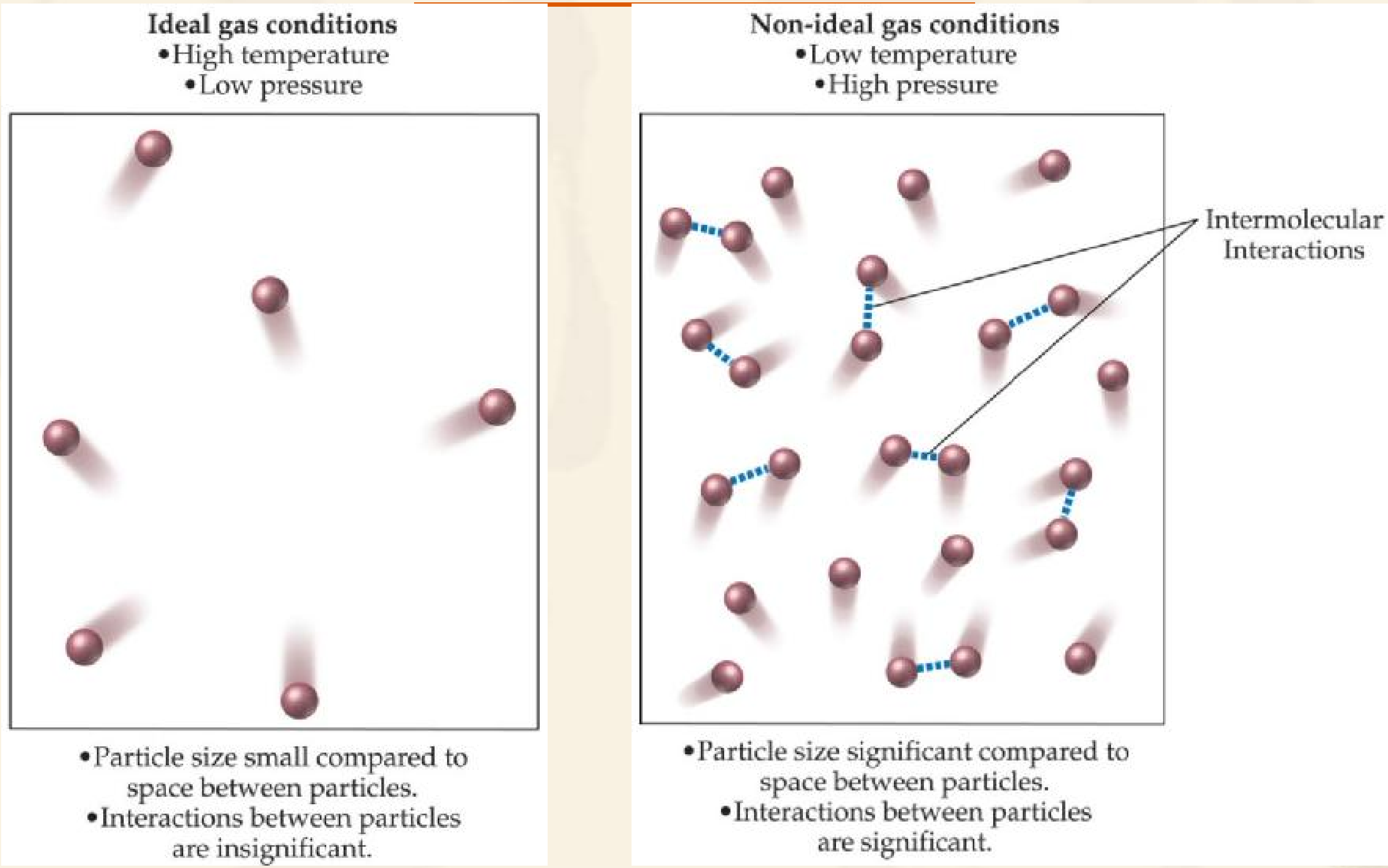
molar mass = 31.9 g/mol

The units of the answer, g/mol, are correct.
It is hard to judge the magnitude with so many variables.

Ideal vs. Real Gases

- Real gases often do not behave like ideal gases at high pressure or low temperature
- Ideal gas laws assume
  1) no attractions between gas molecules
  2) gas molecules do not take up space
  - ™ based on the Kinetic-Molecular Theory
- at low temperatures and high pressures these assumptions are not valid

Ideal vs. Real



Mixtures of Gases

- According to Kinetic Molecular Theory, the particles in a gas behave independently
- Air is a mixture, yet we can treat it as a single gas
- Also, we can think of each gas in the mixture independent of the other gases
  - though all gases in the mixture have the same volume and temperature
    - all gases completely occupy the container, so all gases in the mixture have the volume of the container

Gas	% in Air, by volume	Gas	% in Air, by volume
nitrogen, N_2	78	argon, Ar	78
oxygen, O_2	21	carbon dioxide, CO_2	21

Partial Pressure

- each gas in the mixture exerts a pressure independent of the other gases in the mixture
- the pressure of an component gas in a mixture is called a **partial pressure**
- the sum of the partial pressures of all the gases in a mixture equals the total pressure
  - Dalton's Law of Partial Pressures
  - $P_{total} = P_{gas\ A} + P_{gas\ B} + P_{gas\ C} + ...$

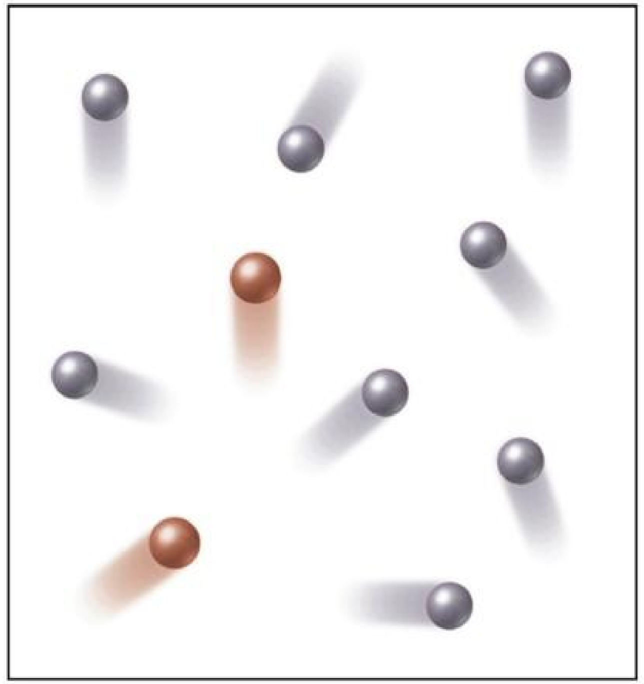
$$P_{air} = P_{N_2} + P_{O_2} + P_{Ar} = 0.78\,\text{atm} + 0.21\,\text{atm} + 0.01\,\text{atm} = 1.00\,\text{atm}$$

Finding Partial Pressure

- to find the partial pressure of a gas, multiply the total pressure of the mixture by the fractional composition of the gas
- for example, in a gas mixture that is 80.0% He and 20.0% Ne that has a total pressure of 1.0 atm, the partial pressure of He would be:

P_{He} = (0.800)(1.0 atm) = 0.80 atm

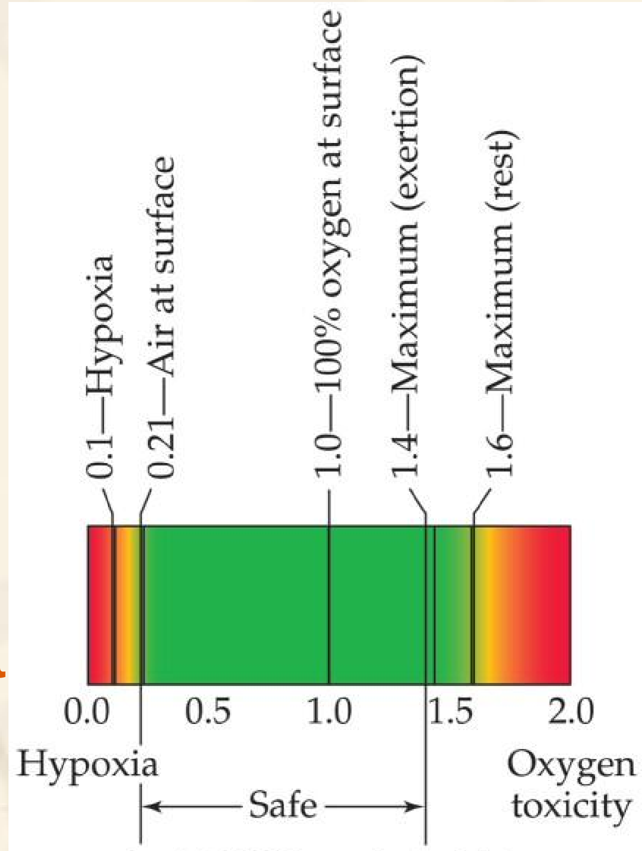
  - fractional composition = percentage divided by 100



Gas mixture (80% He, 20% Ne)
P_{tot} = 1.0 atm
P_{He} = 0.80 atm
P_{Ne} = 0.20 atm

Mountain Climbing & Partial Pressure

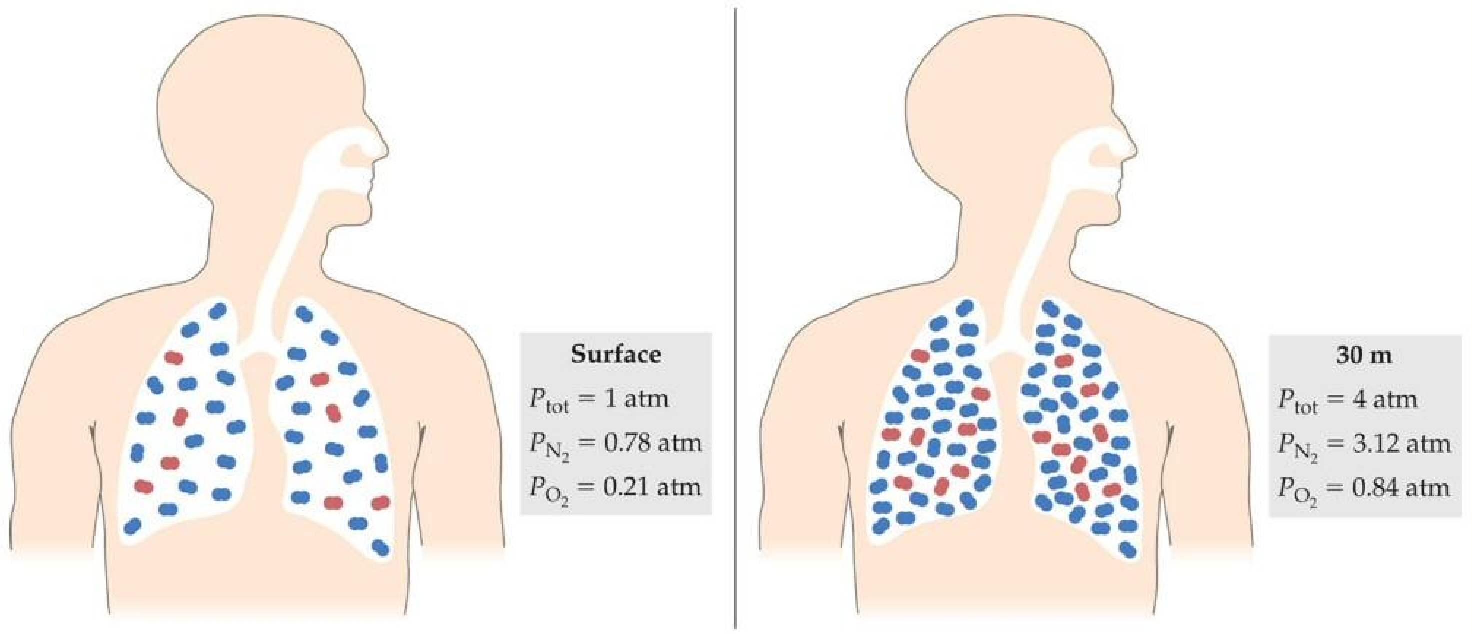
- our bodies are adapted to breathe O_2 at a partial pressure of 0.21 atm
  - Sherpa, people native to the Himalaya mountains, are adapted to the much lower partial pressure of oxygen in their air
- partial pressures of O_2 lower than 0.1 atm will lead to **hypoxia**
  - unconsciousness or death
- climbers of Mt Everest must carry O_2 in cylinders to prevent hypoxia
  - on top of Mt Everest, P_{air} = 0.311 atm, so P_{O2} = 0.065 atm



Deep Sea Divers & Partial Pressure

- its also possible to have too much O_2, a condition called **oxygen toxicity**
  - P_{O2} > 1.4 atm
  - oxygen toxicity can lead to muscle spasms, tunnel vision and convulsions
- its also possible to have too much N_2, a condition called **nitrogen narcosis**
  - also known as Rapture of the Deep
- when diving deep, the pressure of the air divers breathe increases – so the partial pressure of the oxygen increases
  - at a depth of 55 m the partial pressure of O_2 is 1.4 atm
  - divers that go below 50 m use a mixture of He and O_2 called heliox that contains a lower percentage of O_2 than air

Partial Pressure vs. Total Pressure

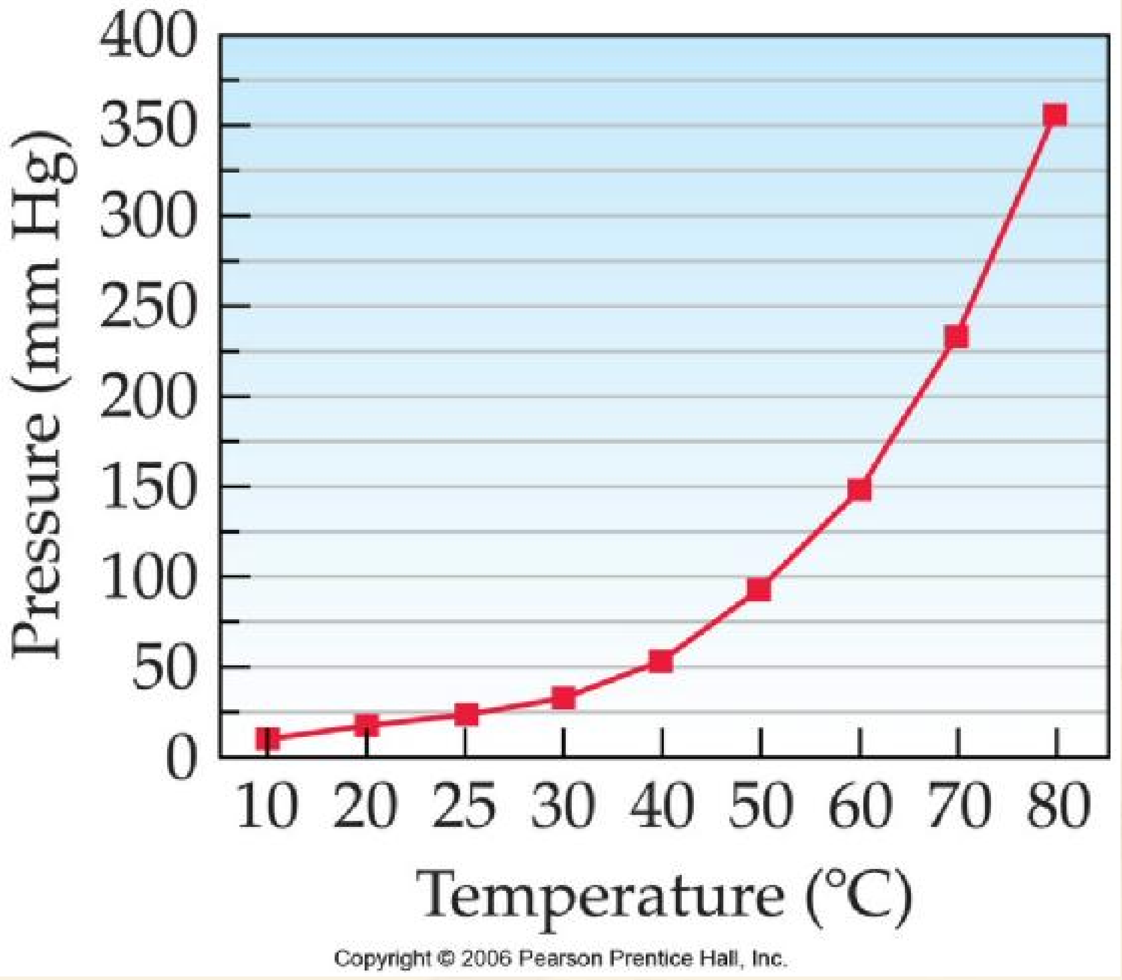


At a depth of 30 m, the total pressure of air in the divers lungs, and the partial pressure of all the gases in the air, are quadrupled!

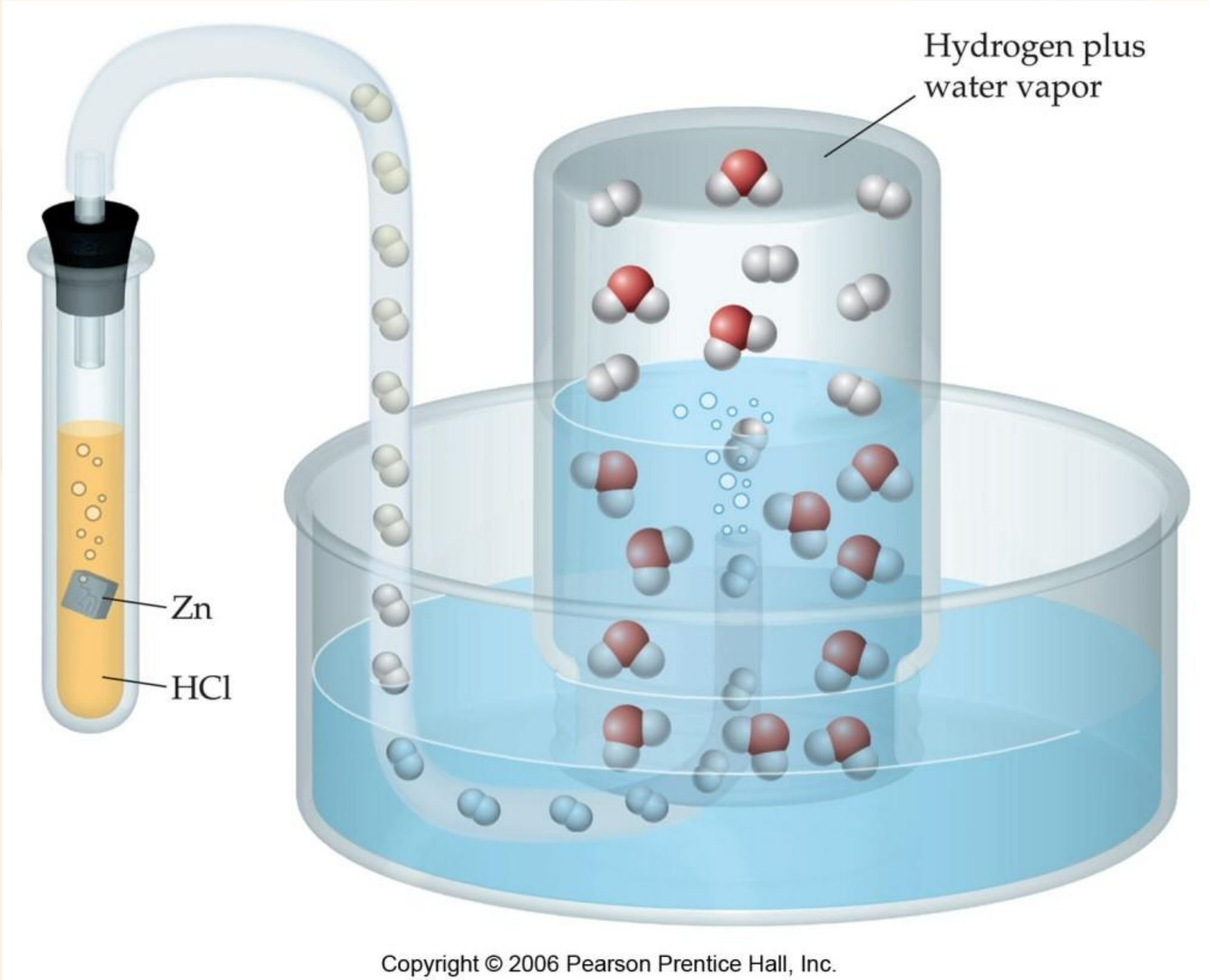
Collecting Gases

- gases are often collected by having them displace water from a container
- the problem is that since water evaporates, there is also water vapor in the collected gas
- the partial pressure of the water vapor, called the **vapor pressure**, depends only on the temperature
  - so you can use a table to find out the partial pressure of the water vapor in the gas you collect
- if you collect a gas sample with a total pressure of 758 mmHg at 25°C, the partial pressure of the water vapor will be 23.8 mmHg – so the partial pressure of the dry gas will be 734 mmHg

Vapor Pressure of Water



Temp., °C	Pressure, mmHg
10	9.2
20	17.5
25	23.8
30	31.8
40	55.3
50	92.5
60	149.4
70	233.7
80	355.1



Zn metal reacts with HCl(aq) to produce $H_2(g)$.

The gas flows through the tube and bubbles into the jar, where it displaces the water in the jar.

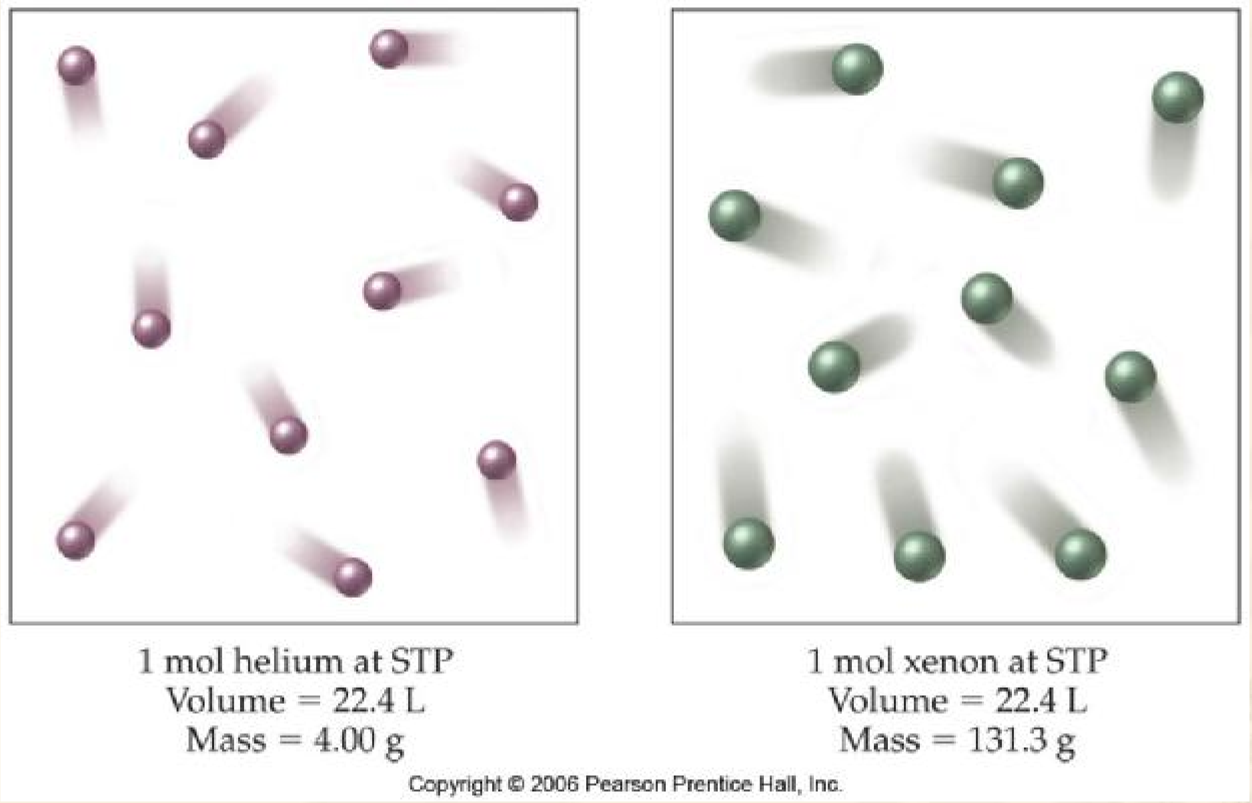
Because water evaporates, some water vapor gets mixed in with the H_2.

Reactions Involving Gases

- the principles of reaction stoichiometry from Chapter 8 can be combined with the Gas Laws for reactions involving gases
- in reactions of gases, the amount of a gas is often given as a Volume
  - instead of moles
  - as we've seen, must state pressure and temperature
- the Ideal Gas Law allows us to convert from the volume of the gas to moles; then we can use the coefficients in the equation as a mole ratio

Example 11.11: Gases in Chemical Reactions

Molar Volume



There is so much empty space between molecules in the gas state, the volume of the gas is not effected by the size of the molecules, (under ideal conditions).

Example 11.12: Using Molar Volume in Calculations

Example:

- How many grams of water will form when 1.24 L of H_2 at STP completely reacts with O_2?

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

Example:
How many grams of water will form when 1.24 L of H_2 at STP completely reacts with O_2?

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

- Write down the given quantity and its units.

Given: 1.24 L H_2 @ STP

Example:

How many grams of water will form when 1.24 L of H_2 at STP completely reacts with O_2?

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

Information

Given: 1.24 L H_2

- Write down the quantity to find and/or its units.

Find: mass H_2O, g

Example:

How many grams of water will form when 1.24 L of H_2 at STP completely reacts with O_2?

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

Information

Given: 1.24 L H_2

Find: g H_2O

CF: 1 mol H_2O = 18.02 g

2 mol $H_2O \equiv$ 2 mol H_2

1 mol $H_2 \equiv$ 22.4 L

SM: L $\rightarrow$ mol $H_2 \rightarrow$ mol $H_2O \rightarrow$ g H_2O

- Apply the Solution Map:

$$1.24\ \mathrm{L\ H_2} \times \frac{1\ \mathrm{mol\ H_2}}{22.4\ \mathrm{L\ H_2}} \times \frac{2\ \mathrm{mol\ H_2O}}{2\ \mathrm{mol\ H_2}} \times \frac{18.02\ \mathrm{g\ H_2O}}{1\ \mathrm{mol\ H_2O}} = 0.988\ \mathrm{g\ H_2O}$$

Example:

How many grams of water will form when 1.24 L of H_2 at STP completely reacts with O_2?

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g)$$

Information

Given: 1.24 L H_2

Find: g H_2O

CF: 1 mol H_2O = 18.02 g

2 mol $H_2O \equiv$ 2 mol H_2

1 mol $H_2 \equiv$ 22.4 L

SM: L $\rightarrow$ mol $H_2 \rightarrow$ mol $H_2O \rightarrow$ g H_2O

- Check the Solution:

$$1.24\text{ L } H_2 = 0.988\text{ g } H_2O$$

The units of the answer, g H_2O, are correct.
It is hard to judge the magnitude with so many variables, but with less than 22.4 L we have less than 1 mole of H_2 - with a mole ratio 2:2 we should expect to make less than 1 mole of H_2O