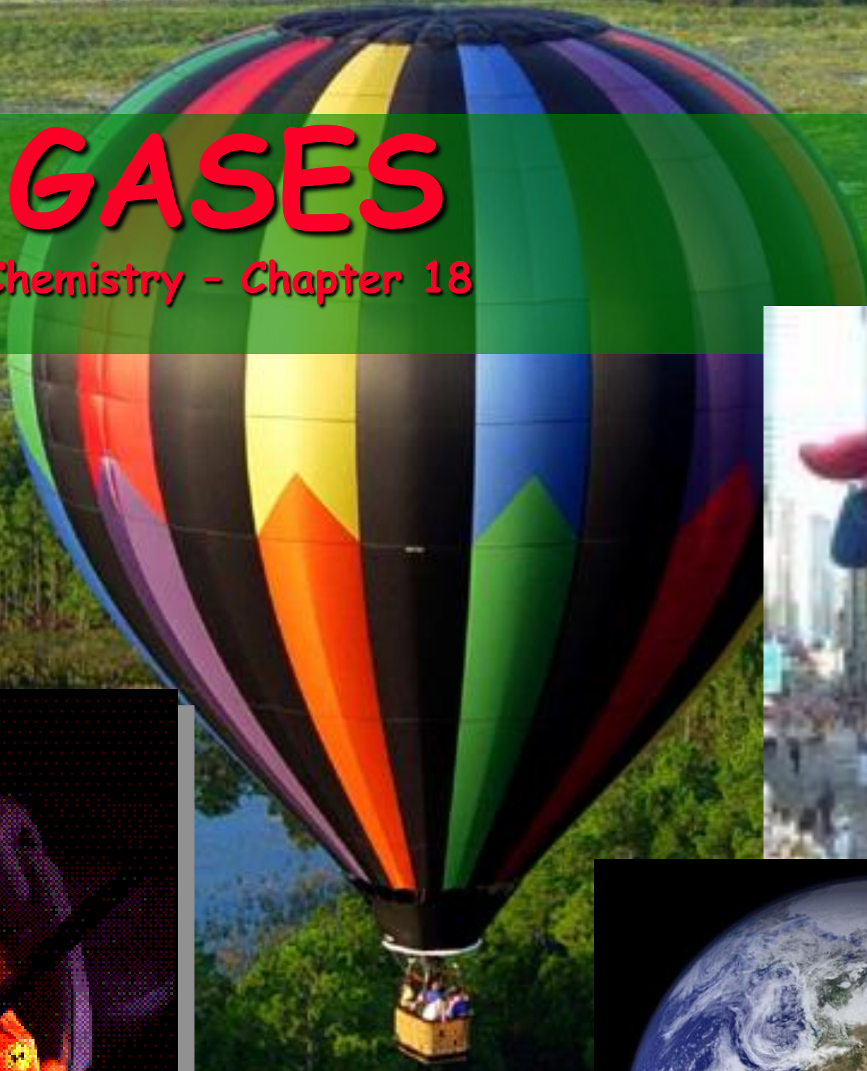
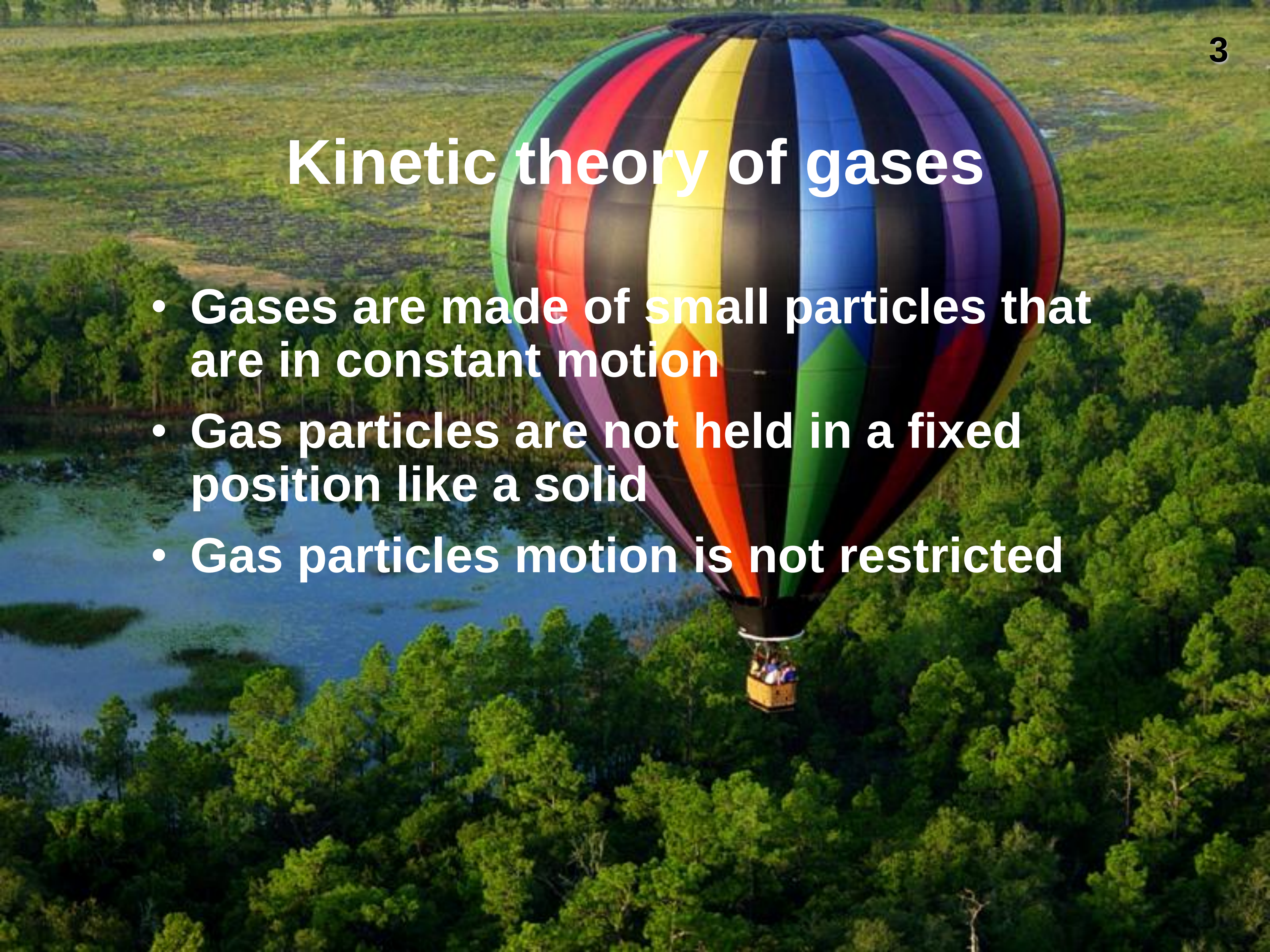
GASES

Chemistry – Chapter 18



Importance of Gases

- Airbags fill with N_2 gas in an accident.
- Gas is generated by the decomposition of sodium azide, NaN_3.
- $2\ NaN_3 \ \text{--->}\ 2\ Na + 3\ N_2$



Kinetic theory of gases

- Gases are made of small particles that are in constant motion
- Gas particles are not held in a fixed position like a solid
- Gas particles motion is not restricted

General Properties of Gases

- There is a lot of "free" space in a gas.
- Gases can be expanded infinitely.
- Gases fill containers uniformly and completely.
- Gases diffuse and mix rapidly.

Ideal Gas

- Gases are like point masses within a volume of mostly empty space
  - A point mass has no volume or diameter
  - Point masses do not exist
- So and imaginary gas with point masses is called an ideal gas
  - An ideal gas only exist theoretically
- Temperature and pressure effect the volume of a gas

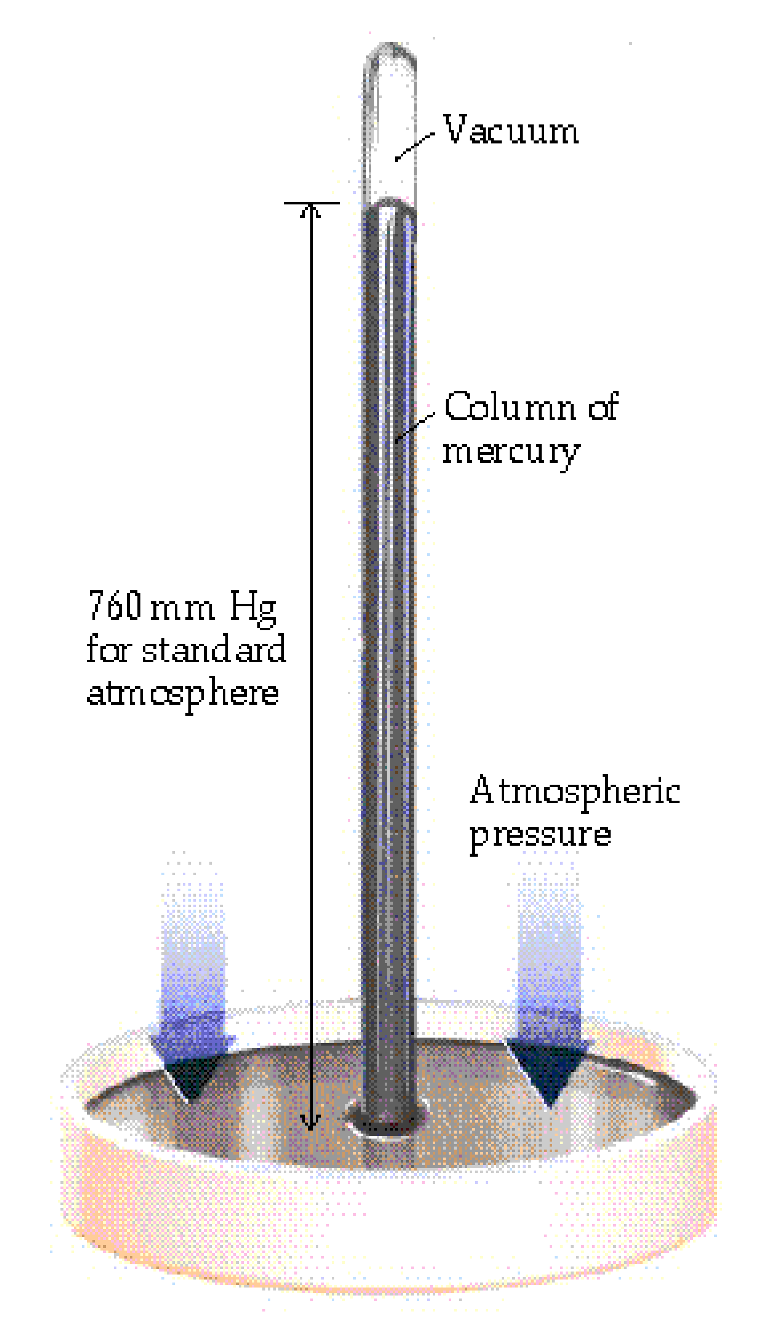
Properties of Gases

Gas properties can be modeled using math. Model depends on—

- V = volume of the gas (dm^3) or (m^3)
- T = temperature (K)
  - ALL temperatures in the entire chapter MUST be in Kelvin!!! No Exceptions!
- n = amount (moles)
- P = pressure (atmospheres)

Pressure

- Pressure of air is measured with a **BAROMETER** (developed by Torricelli in 1643)

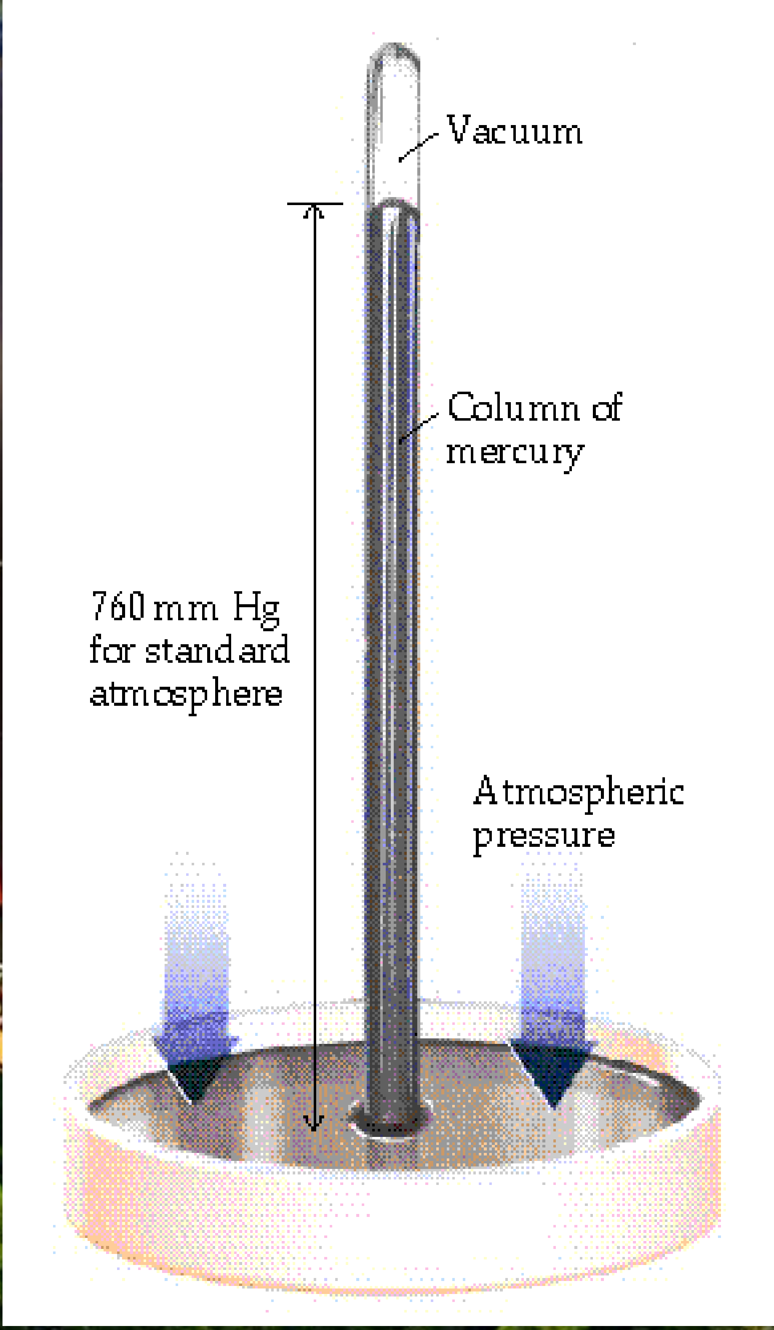


Pressure

Hg rises in tube until force of Hg (down) balances the force of atmosphere (pushing up). (Just like a straw in a soft drink)

P of Hg pushing down related to

- Hg density
- column height



Pressure

Column height measures
Pressure of atmosphere

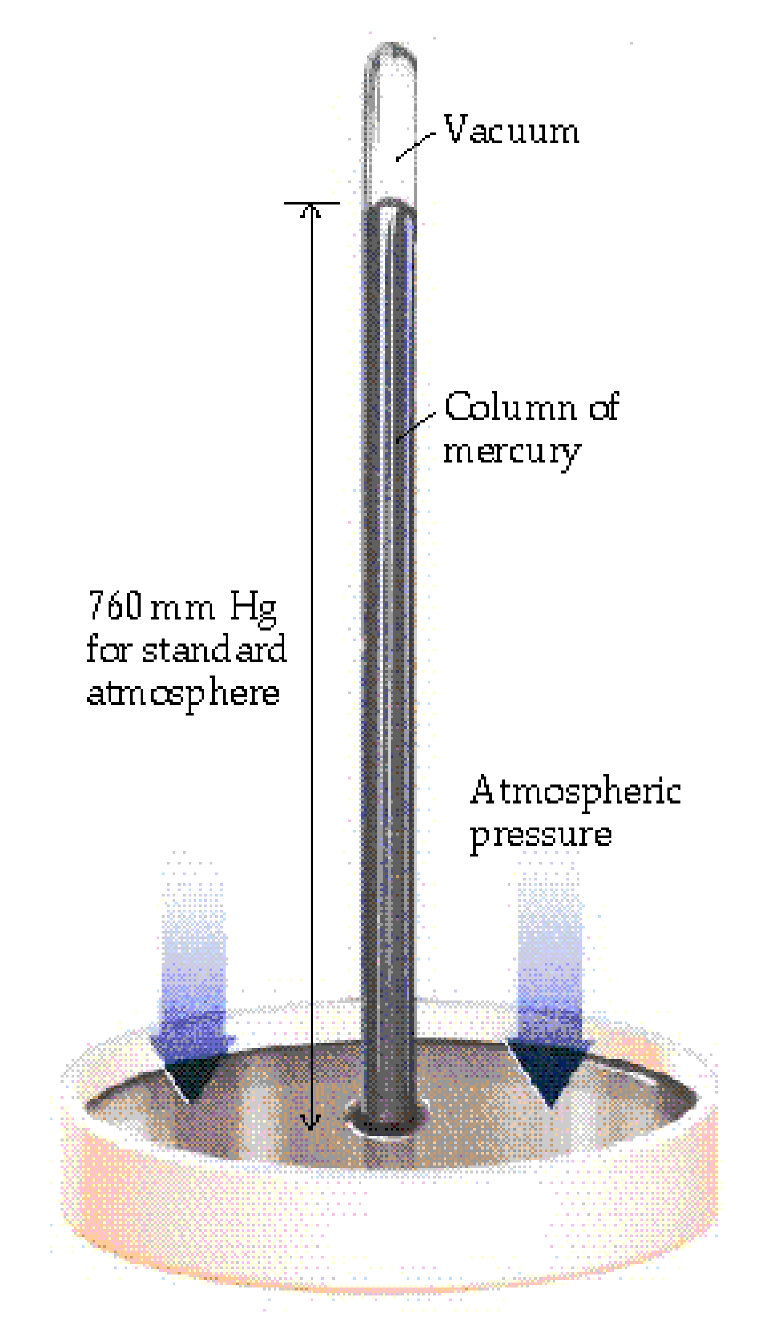
- 1 standard atmosphere (atm) *

= 760 mm Hg (or torr) *

= 29.92 inches *

= 14.7 pounds/in^2 (psi)

= 101.3 kPa (SI unit is PASCAL) *
= about 34 feet of water!



Standard conditions

- Standard Temperature is 0 °C 273 K
- Standard pressure 101.325 kPa, 760 mm, or 1 atmosphere
- STP is Standard Temperature Pressure

And now, we pause for this commercial message from STP

OK, so it's really not THIS kind of STP…

STP in chemistry stands for Standard Temperature and Pressure

STP allows us to compare amounts of gases between different pressures and temperatures

Standard Pressure = 1 atm (or an equivalent)

Standard Temperature = 0 deg C (273 K)

Boyle's Law

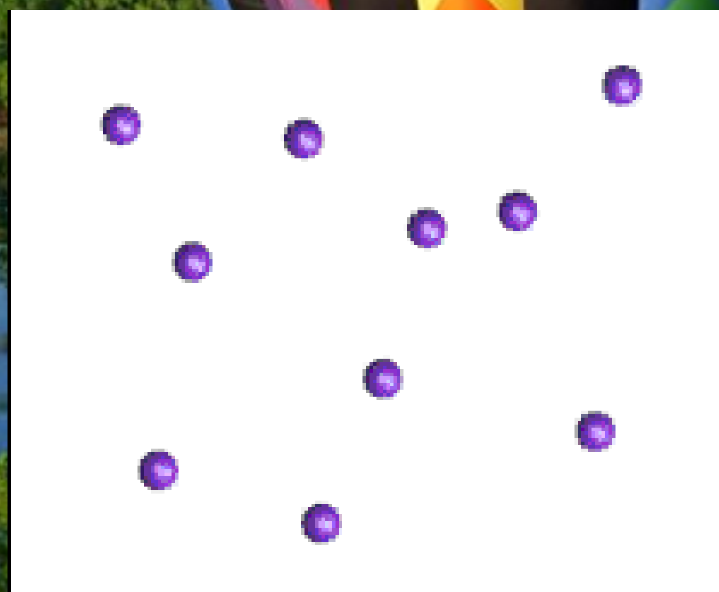
P α 1/V

This means Pressure and Volume are INVERSELY PROPORTIONAL if moles and temperature are constant (do not change). For example, P goes up as V goes down.

$P_1V_1 = P_2 V_2$

Robert Boyle (1627-1691). Son of Early of Cork, Ireland.

Boyle's Law and Kinetic Molecular Theory

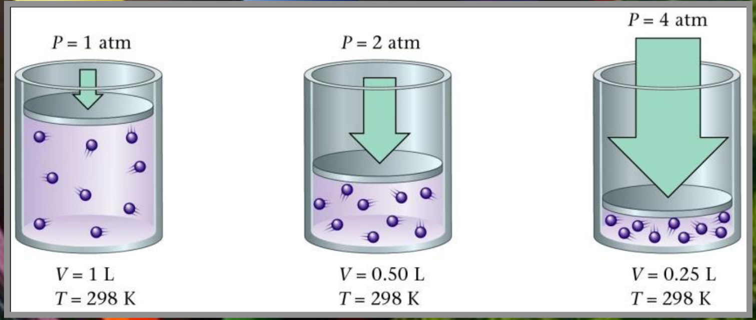


P proportional to 1/V

Boyle's Law

A bicycle pump is a good example of Boyle's law.

As the volume of the air trapped in the pump is reduced, its pressure goes up, and air is forced into the tire.



Dalton's Law of Partial Pressures

$$2\ H_2O_2\ (l) \ \text{--->}\ 2\ H_2O\ (g) \ +\ O_2\ (g)$$

0.32 atm   0.16 atm

What is the total pressure in the flask?

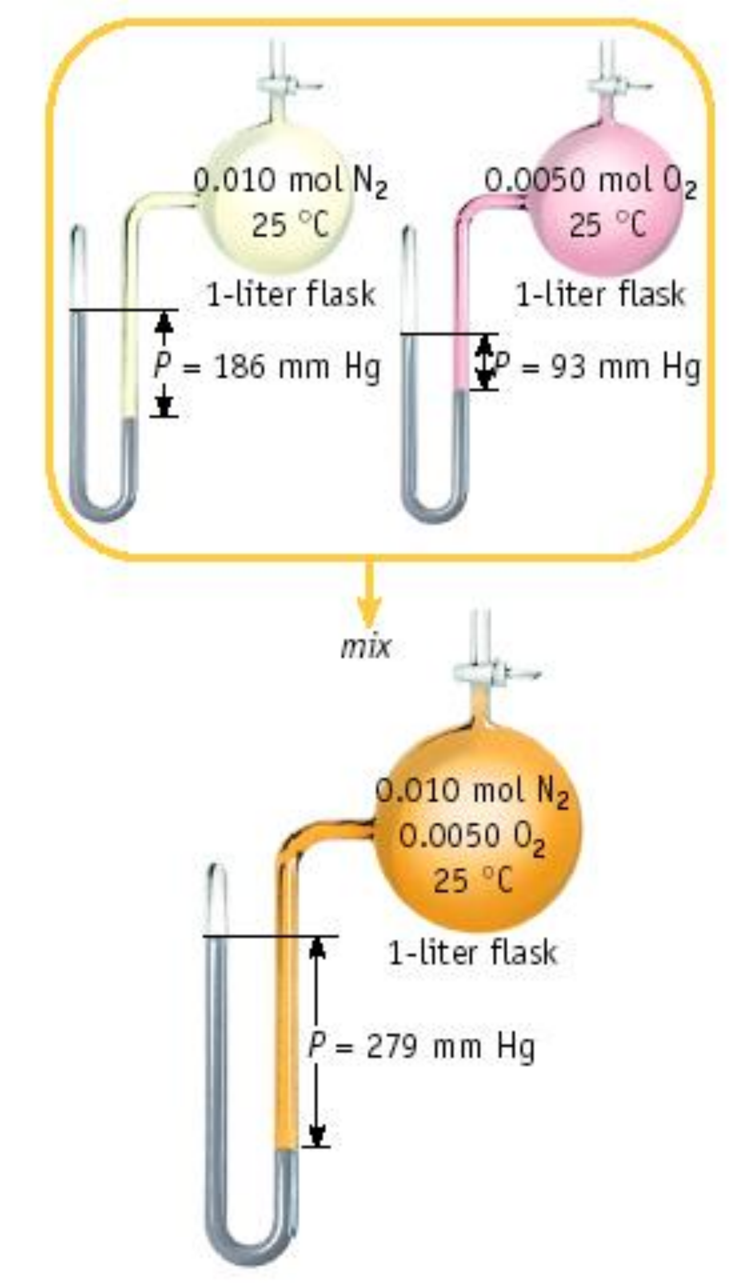
P_{total} in gas mixture = $P_A + P_B + ...$

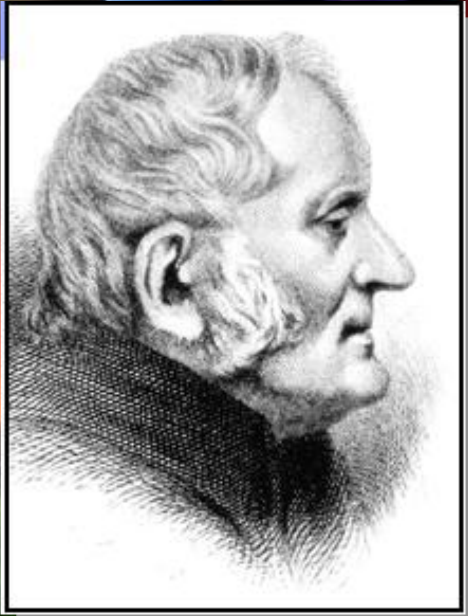
Therefore,

$P_{total} = P_{H_2O} + P_{O_2} = 0.48$ atm

Dalton's Law: total P is sum of PARTIAL pressures.

Dalton's Law





John Dalton
1766-1844

Charles's Law

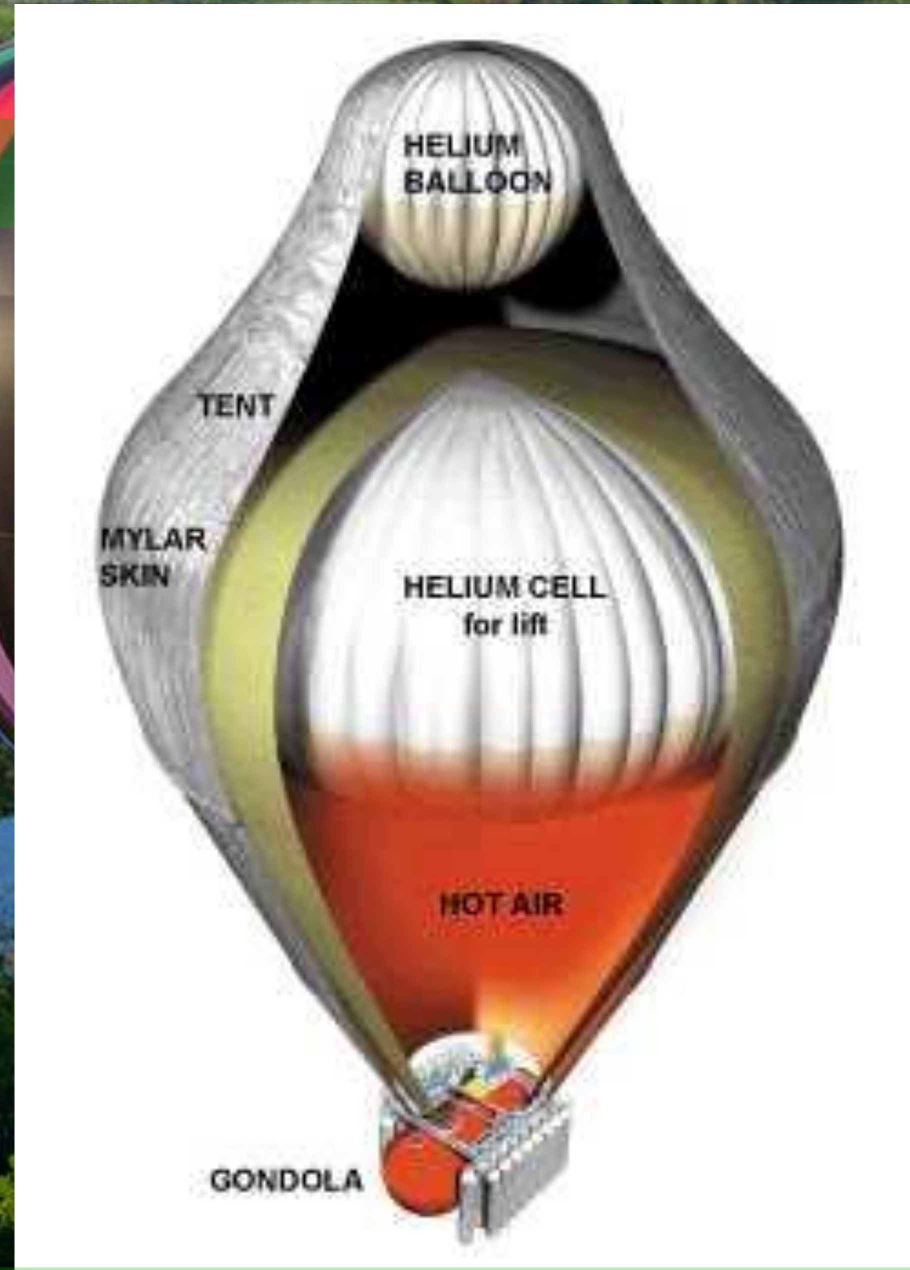
- If n and P are constant, then V α T
- V and T are directly proportional.

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

- If one temperature goes
volume goes up!

Jacques Charles (1746-1823). Isolated boron and studied gases. Balloonist.

Charles's original balloon



Modern long-distance balloon

Charles and Kelvin Scale

- Jacques Charles found that if you started at 0°C any gas would double if the temperature was raised to 273 °C
- Increase in 1 °C will cause volume to increase by 1/273
- That would suggest that a gas would have no volume at -273°C
- Increase temperature means an increase volume
- V= k T , or k = V/T , so $V_1/T_1 = V_2/T_2$

Gay-Lussac's Law

If n and V are constant, then P α T

P and T are directly proportional.

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

- If one temperature goes up, pressure goes up!
- If you increase temperature, the # of collisions with the side of the container increase causing increase presure

Joseph Louis Gay-Lussac (1778-1850)

Combined Gas Law

- Since they are all related to each other, we can combine them into a single equation.

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

Combined Gas Law

If you should only need one of the other gas laws, you can cover up the item that is constant and you w l get that gas law!

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

Boyle's Law

Charles' Law

Gay-Lussac's Law

Combined Gas Law Problem

A sample of helium gas has a volume of 0.180 L, a pressure of 0.800 atm and a temperature of 29°C. What is the new temperature(°C) of the gas at a volume of 90.0 mL and a pressure of 3.20 atm?

Set up Data Table

P_1 = 0.800 atm	V_1 = 180 mL	T_1 = 302 K
P_2 = 3.20 atm	V_2= 90 mL	T_2 = ??

Calculation

P_1 = 0.800 atm V_1 = 180 mL T_1 = 302 K

P_2 = 3.20 atm V_2 = 90 mL T_2 = ??

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

$$P_1 V_1 T_2 = P_2 V_2 T_1$$

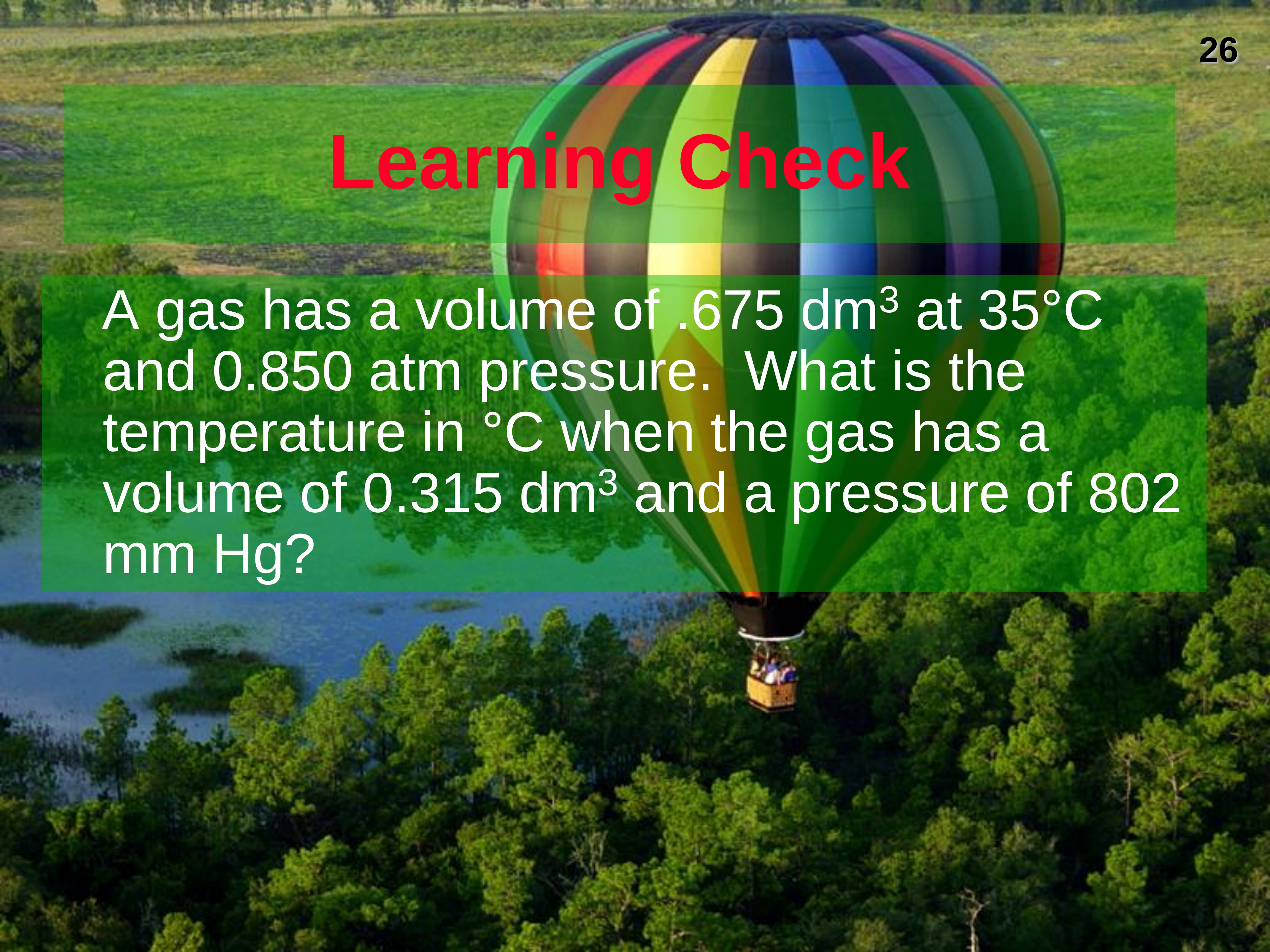
$$T_2 = \frac{P_2 V_2 T_1}{P_1 V_1}$$

$$T_2 = \frac{3.20 \text{ atm} \times 90.0 \text{ mL} \times 302 \text{ K}}{0.800 \text{ atm} \times 180.0 \text{ mL}} = 604 \text{ K}$$

T_2 = 604 K - 273 = 331 °C

Solve using unit Cancel

- First take what you have and solve for what you don't have and
- Use the gas laws at each step
- $$302\ \text{K} \times \frac{0.090\ \text{dm}^3}{0.180\ \text{dm}^3} \times \frac{3.20\ \text{atm}}{0.80\ \text{atm}} = 604\ \text{K}$$



Learning Check

A gas has a volume of .675 dm^3 at 35°C and 0.850 atm pressure. What is the temperature in °C when the gas has a volume of 0.315 dm^3 and a pressure of 802 mm Hg?

One More Practice Problem

A balloon has a volume of 785 mL on a fall day when the temperature is 21°C. In the winter, the gas cools to 0°C. What is the new volume of the balloon?

Try This One

A sample of neon gas used in a neon sign has a volume of 15 dm^3 at STP. What is the volume (dm^3) of the neon gas at 2.0 atm and –25°C?

HONORS only

GAS DIFFUSION AND EFFUSION

- **diffusion** is the gradual mixing of molecules of different gases.
- **effusion** is the movement of molecules through a small hole into an empty container.

Diffusion and Graham's Law

- Diffusion is the random scattering of gas particles
- Each gas diffuse at a different rate when kinetic energy is constant
- Graham's Law of diffusion - The relative rates at which two gases under identical conditions of temperature and pressure will pass through a small hole vary inversely as the square roots of the molecular masses of the gas

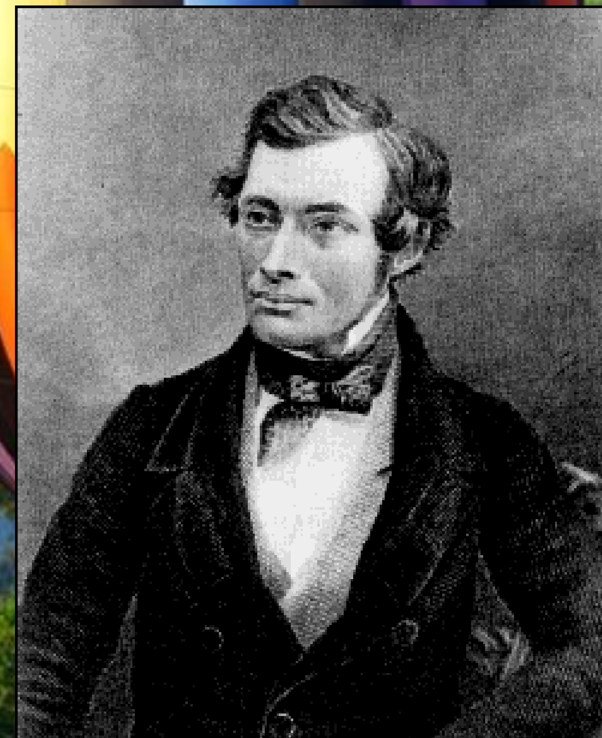
HONORS only

GAS DIFFUSION AND EFFUSION

Graham's law governs effusion and diffusion of gas molecules.

$$\frac{\text{Rate for A}}{\text{Rate for B}} = \sqrt{\frac{\text{M of B}}{\text{M of A}}}$$

Rate of effusion is inversely proportional to its molar mass.



Thomas Graham, 1805-1869. Professor in Glasgow and London.

HONORS only

GAS DIFFUSION AND EFFUSION

Molecules effuse thru holes in a rubber balloon, for example, at a rate (= moles/time) that is

- proportional to T
- inversely proportional to M.

Therefore, He effuses more rapidly than O_2 at same T.

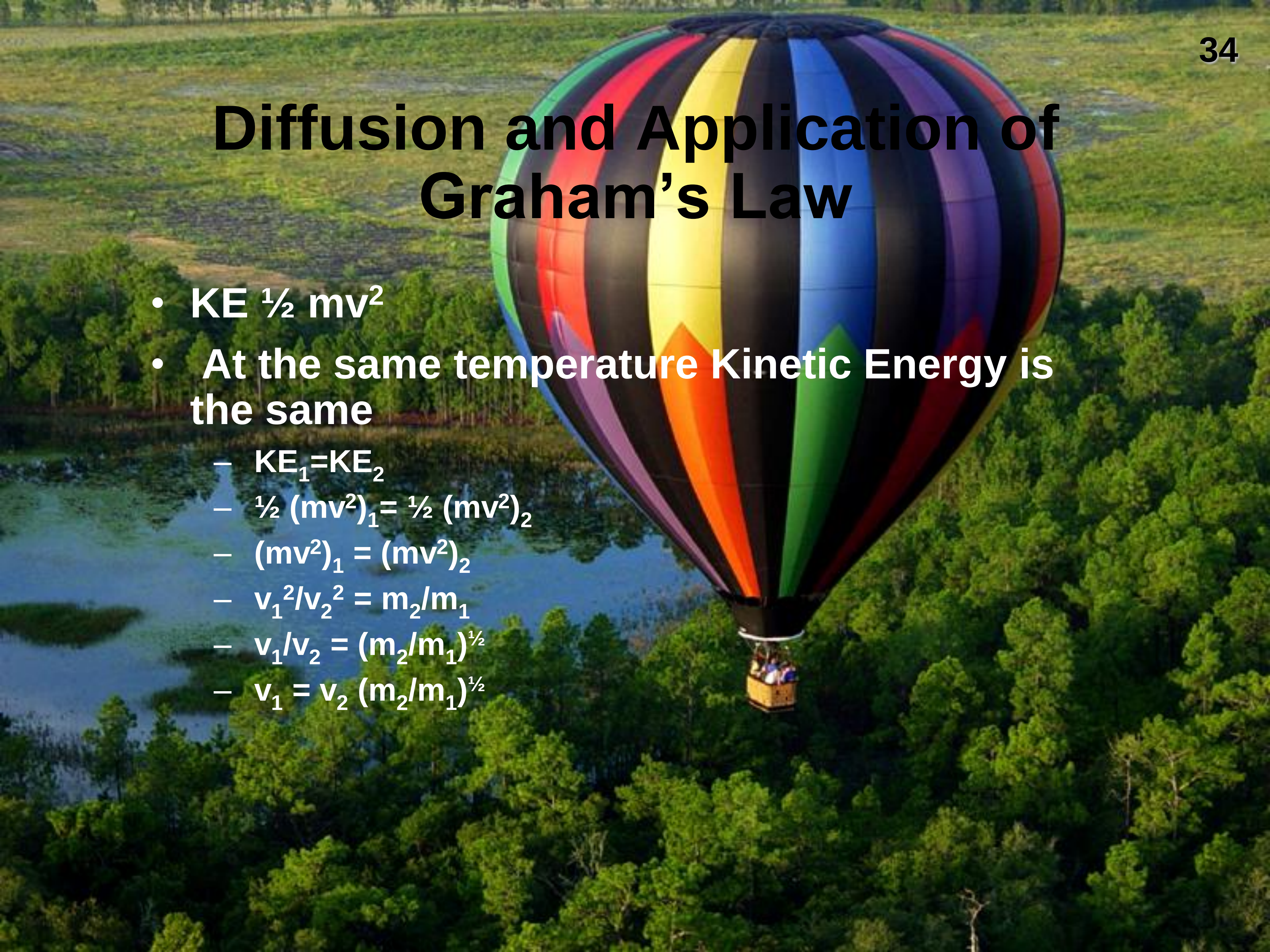
Gas Diffusion

relation of mass to rate of diffusion

- HCl and NH_3 diffuse from opposite ends of tube.
- Gases meet to form NH_4Cl
- HCl heavier than NH_3
- Therefore, NH_4Cl forms closer to HCl end of tube.

Gaseous diffusion of $NH_3(g)$ and HCl(g)

Diffusion and Application of Graham's Law



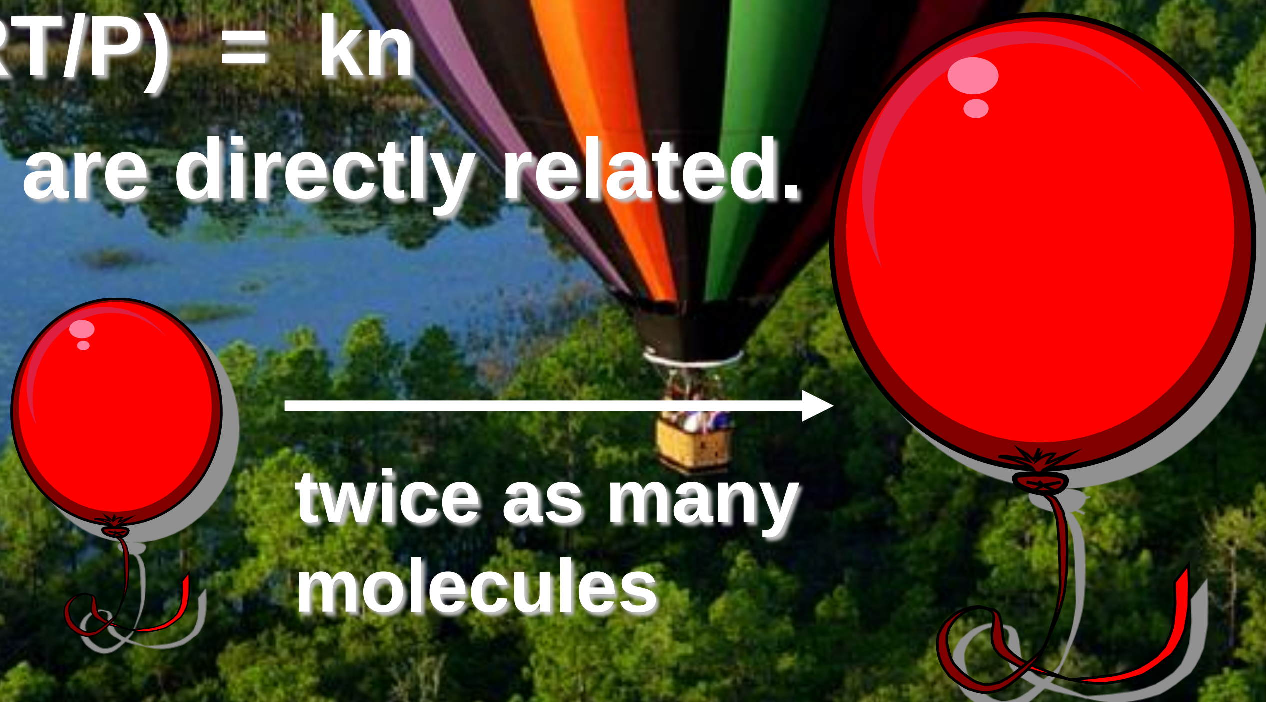
- KE $\frac{1}{2}$ mv^2
- At the same temperature Kinetic Energy is the same
  - $KE_1 = KE_2$
  - $\frac{1}{2}(mv^2)_1 = \frac{1}{2}(mv^2)_2$
  - $(mv^2)_1 = (mv^2)_2$
  - $v_1^2/v_2^2 = m_2/m_1$
  - $v_1/v_2 = (m_2/m_1)^{1/2}$
  - $v_1 = v_2\,(m_2/m_1)^{1/2}$

Avogadro's Hypothesis

Equal volumes of gases at the same T and P have the same number of molecules.

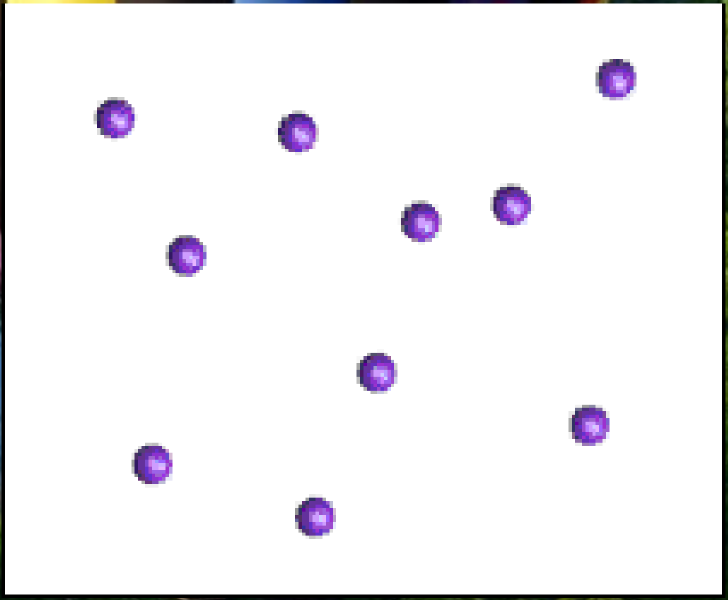
$V = n\ (RT/P) = kn$

V and n are directly related.



Avogadro's Hypothesis and Kinetic Molecular Theory

The gases in this experiment are all measured at the same T and V.



P proportional to n

IDEAL GAS LAW

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

- Brings together gas properties.
- Can be derived from experiment and theory.
- BE SURE YOU KNOW THIS EQUATION!

Using PV = nRT

P = Pressure

V = Volume

T = Temperature

N = number of moles

R is a constant, called the **Ideal Gas Constant**

Instead of learning a different value for R for all the possible unit combinations, we can just **memorize one** value and **convert the units to match R.**

$$R = 0.0821 \; \frac{L \cdot atm}{Mol \cdot K}$$

Using PV = nRT

How much N_2 is required to fill a small room with a volume of 960 cubic feet (27,000 L) to 745 mm Hg at 25 ºC?

Solution

1. Get all data into proper units

V = 27,000 L

T = 25 ºC + 273 = 298 K

P = 745 mm Hg (1 atm/760 mm Hg)
= 0.98 atm

And we always know R, 0.0821 L atm / mol K

Using PV = nRT

How much N_2 is req'd to fill a small room with a volume of 960 cubic feet (27,000 L) to P = 745 mm Hg at 25 ºC?

Solution

2. Now plug in those values and solve for the unknown.

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

$$n = \frac{(0.98 \text{ atm})(2.7 \times 10^{4} \text{ L})}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(298 \text{ K})}$$

$n = 1.1 \times 10^3$ mol (or about 30 kg of gas)

Learning Check

Dinitrogen monoxide (N_2O), laughing gas, is used by dentists as an anesthetic. If 2.86 mol of gas occupies a 20.0 L tank at 23°C, what is the pressure (mm Hg) in the tank in the dentist office?

Learning Check

A 5.0 L cylinder contains oxygen gas at 20.0°C and 735 mm Hg. How many grams of oxygen are in the cylinder?

Deviations from Ideal Gas Law

- Real molecules have **volume**.

The ideal gas consumes the entire amount of available volume. It does not account for the volume of the molecules themselves.

- There are **intermolecular forces**.

An ideal gas assumes there are no attractions between molecules. Attractions slow down the molecules and reduce the amount of collisions.

  - Otherwise a gas could not condense to become a liquid.

Gases in the Air

The % of gases in air		Partial pressure (STP)
78.08%	N_2	593.4 mm Hg
20.95%	O_2	159.2 mm Hg
0.94%	Ar	7.1 mm Hg
0.03%	CO_2	0.2 mm Hg

$$P_{AIR} = P_{N_2} + P_{O_2} + P_{Ar} + P_{CO_2} = 760 \text{ mm Hg}$$

Total Pressure 760 mm Hg

Health Note

When a scuba diver is several hundred feet under water, the high pressures cause N_2 from the tank air to dissolve in the blood. If the diver rises too fast, the dissolved N_2 will form bubbles in the blood, a dangerous and painful condition called "the bends". Helium, which is inert, less dense, and does not dissolve in the blood, is mixed with O_2 in scuba tanks used for deep descents.

Collecting a gas "over water"

- Gases, since they mix with other gases readily, must be collected in an environment where mixing can not occur. The easiest way to do this is under water because water displaces the air. So when a gas is collected "over water", that means the container is filled with water and the gas is bubbled through the water into the container. Thus, the pressure inside the container is from the gas AND the water vapor. This is where Dalton's Law of Partial Pressures becomes useful.

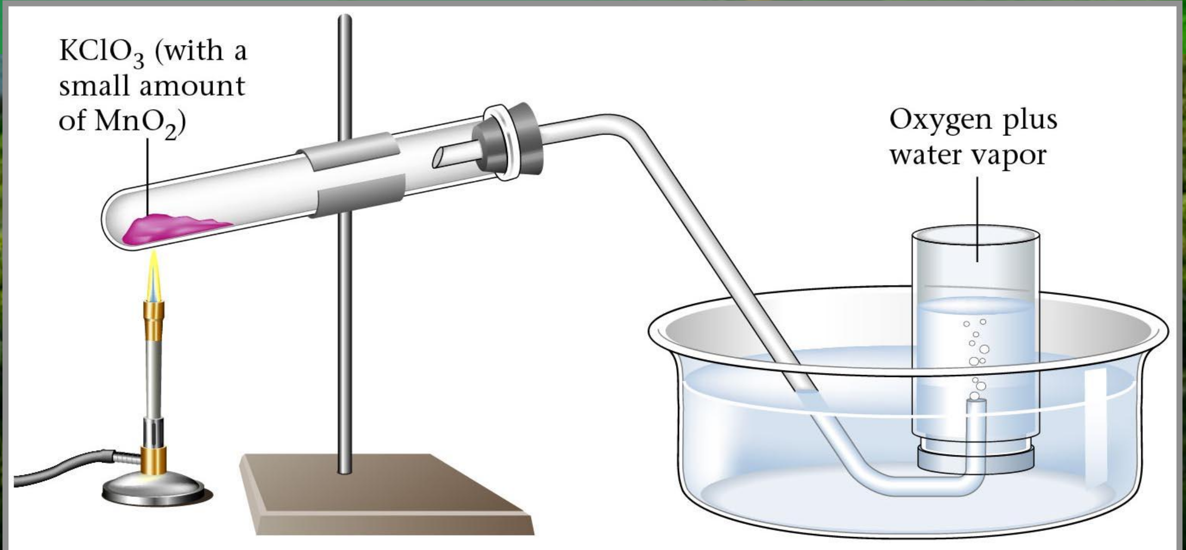


Table of Vapor Pressures for Water

Temperature, °C	Pressure, mmHg	Temperature, °C	Pressure, mmHg
0	4.6	27	26.7
5	6.5	28	28.3
10	9.2	29	30.0
11	9.8	30	31.8
12	10.5	35	42.2
13	11.2	40	55.3
14	12.0	45	71.9
15	12.8	50	92.5
16	13.6	55	118.0
17	14.5	60	149.4
18	15.5	65	187.5
19	16.5	70	233.7
20	17.5	75	289.1
21	18.7	80	355.1
22	19.8	85	433.6
23	21.1	90	525.8
24	22.4	95	633.9
25	23.8	100	760.0
26	25.2	105	906.1

Solve This!

A student collects some hydrogen gas over water at 20 degrees C and 768 torr. What is the pressure of the gas?

Temperature, °C	Pressure, mmHg	Temperature, °C	Pressure, mmHg
0	4.6	27	26.7
5	6.5	28	28.3
10	9.2	29	30.0
11	9.8	30	31.8
12	10.5	35	42.2
13	11.2	40	55.3
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24	22.4	95	633.9
25	23.8	100	760.0
26	25.2	105	906.1

768 torr – 17.5 torr = 750.5 torr

GAS DENSITY

22.4 L of ANY gas AT STP = 1 mole

Gases and Stoichiometry

$2\ H_2O_2\ (l)$ ---> $2\ H_2O\ (g)\ +\ O_2\ (g)$

Decompose 1.1 g of H_2O_2 in a flask with a volume of 2.50 L. What is the volume of O_2 at STP?

Bombardier beetle uses decomposition of hydrogen peroxide to defend itself.

Gases and Stoichiometry

$2\ H_2O_2\ (l)\ \text{--->}\ 2\ H_2O\ (g)\ +\ O_2\ (g)$

Decompose 1.1 g of H_2O_2 in a flask with a volume of 2.50 L. What is the volume of O_2 at STP?

Solution

1.1 g H_2O_2	1 mol H_2O_2	1 mol O_2	22.4 L O_2
	34 g H_2O_2	2 mol H_2O_2	1 mol O_2

= 0.36 L O_2 at STP

Gas Stoichiometry: Practice!

A. What is the volume at STP of 4.00 g of CH_4?

B. How many grams of He are present in 8.0 L of gas at STP?

What if it's NOT at STP?

- 1. Do the problem like it was at STP. (V_1)
- 2. Convert from STP (V_1, P_1, T_1) to the stated conditions (P_2, T_2)

Try this one!

How many L of O_2 are needed to react 28.0 g NH_3 at 24°C and 0.950 atm?

$$4\ NH_3(g) + 5\ O_2(g) \longrightarrow 4\ NO(g) + 6\ H_2O(g)$$