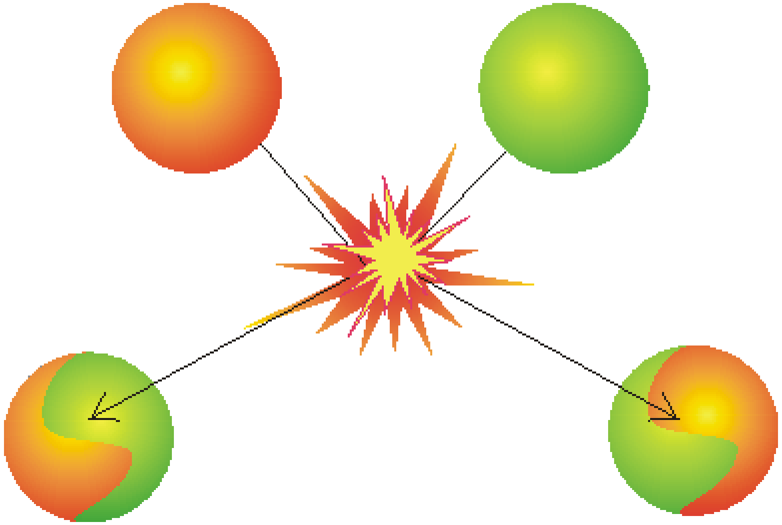
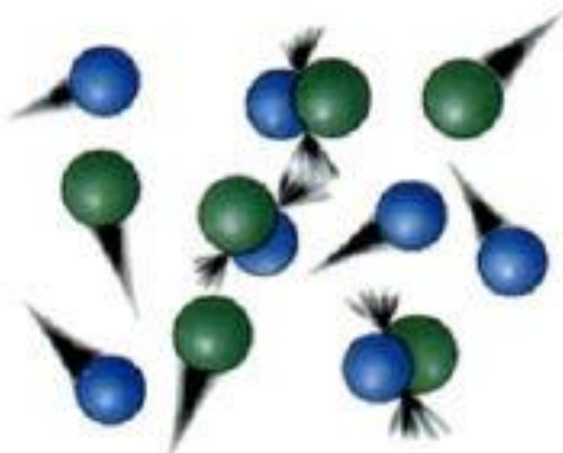
Gases

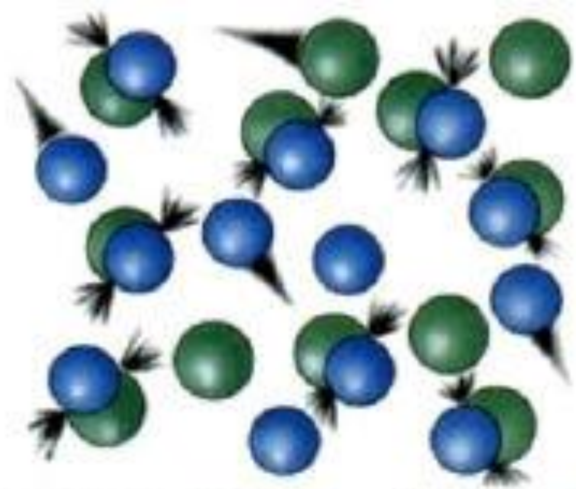
Gases



T ↑ $\Rightarrow$ move faster
Kinetic energy ↑



Low concentration = Few collisions

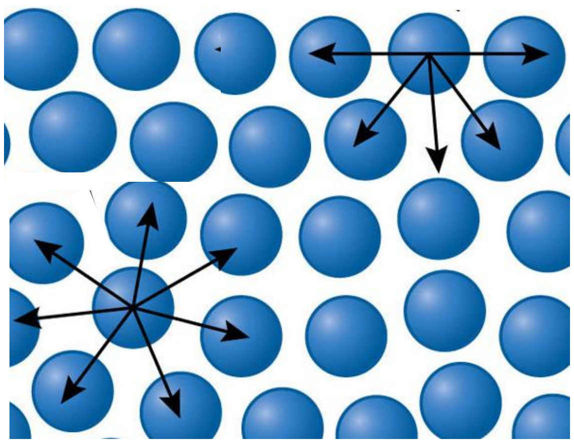


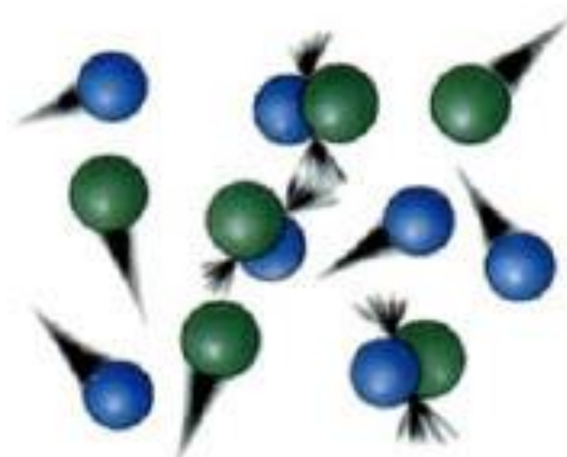
High concentration = More collisions

Gases

Physical sate of matter depends on:

Attractive forces ⟷ Kinetic energy



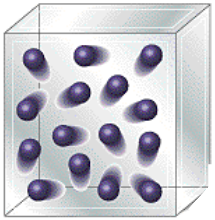


Brings molecules together

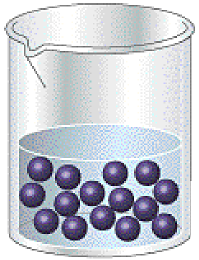
Keeps molecules apart

Gases

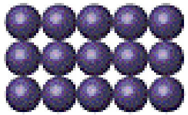
Gas → High kinetic energy (move fast) → Low attractive forces



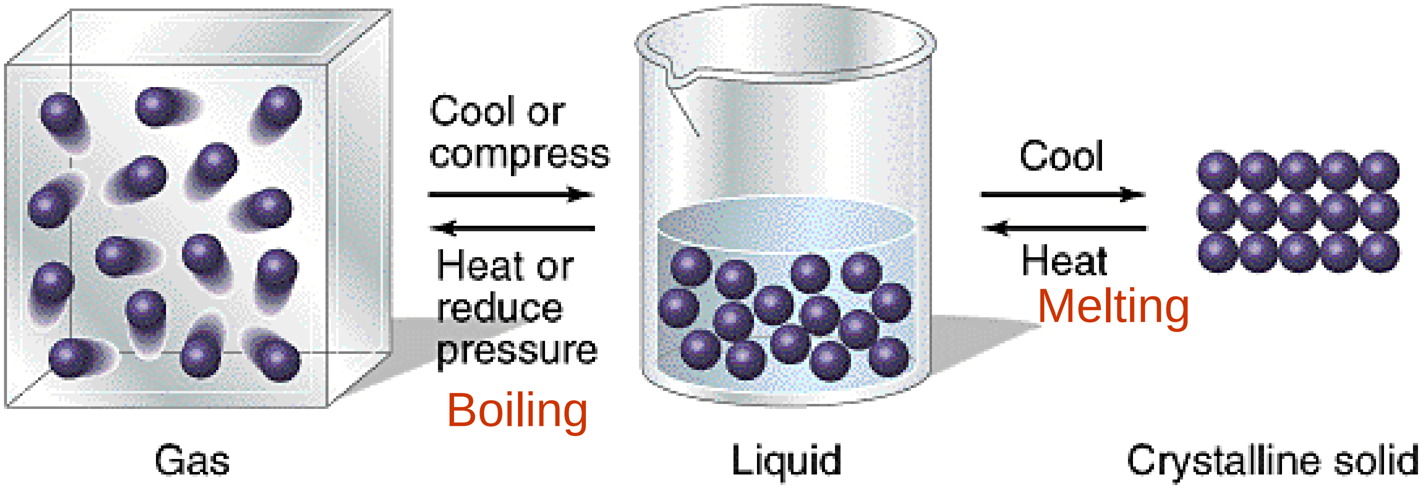
Liquid → Medium kinetic energy (move slow) → medium attractive forces



Solid → Low kinetic energy (move slower) → High attractive forces



Physical Changes



Change of states

Ideal Gases

Kinetic molecular theory:

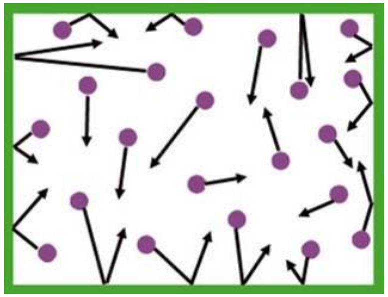
1. **Particles move in straight lines, randomly.**
2. **Average Kinetic energy of particles depends on temperature.**
3. **Particles collide and change direction (they may exchange kinetic energies). Their collisions with walls cause the pressure.**
4. **Gas particles have no volume.**
5. **No attractive forces (or repulsion) between gas particles.**
6. **More collision = greater pressure.**

In reality, no gas is ideal (all gases are real).

At low pressure (around 1 atm or lower) and at 0°C or higher,
we can consider real gases as ideal gases.

Pressure (P)

$$\text{Pressure (P)} = \frac{\text{Force (F)}}{\text{Area (A)}}$$



F: constant

A ↓ ➡ P ↑

A: constant

F ↑ ➡ P ↑

Atmosphere (atm)
Millimeters of mercury (mm Hg)
torr
in. Hg
Pascal

1.000 atm = 760.0 mm Hg
= 760.0 torr
= 101,325 pascals
= 29.92 in. Hg

Pressure (P)

At STP: Standard Temperature & Pressure

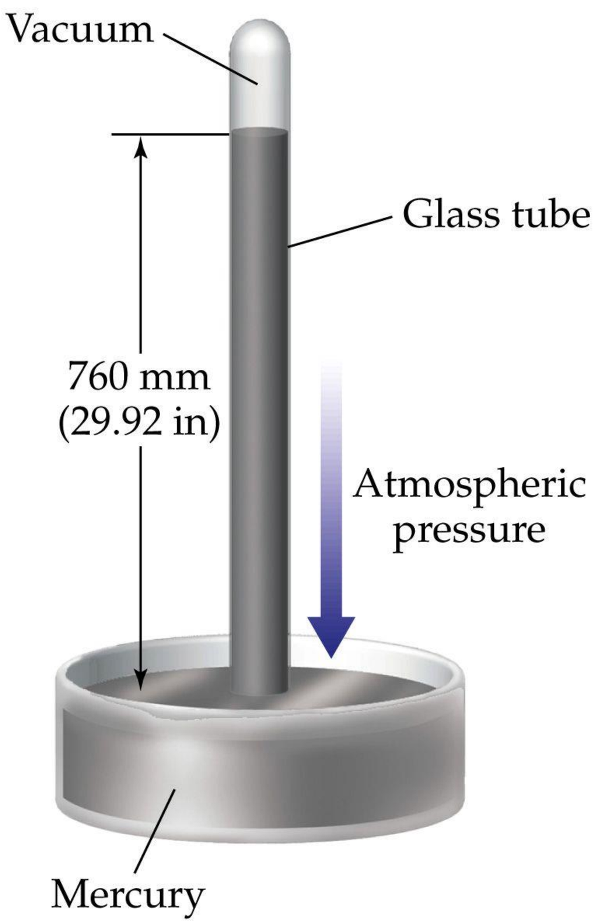
1 standard atmosphere = 1.000 atm = 760.0 mm Hg = 760.0 torr = 101,325 pa

T = 0.00ºC

Pounds per square inch (psi)

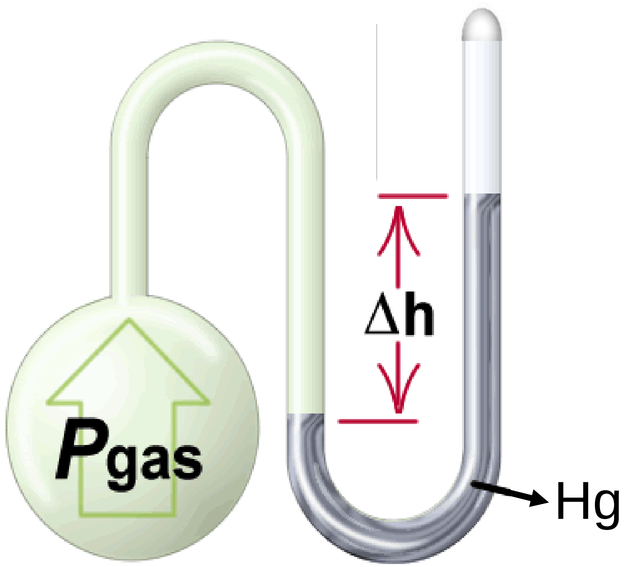
1.000 atm = 14.69 psi

Pressure (P)



barometer

atmospheric pressure



manometer

pressure of gas in a container

Boyle's Law

Robert Boyle

Boyle's law: m,T: constant

$P \quad 1/\alpha \quad V \Longrightarrow PV = k$ (a constant)

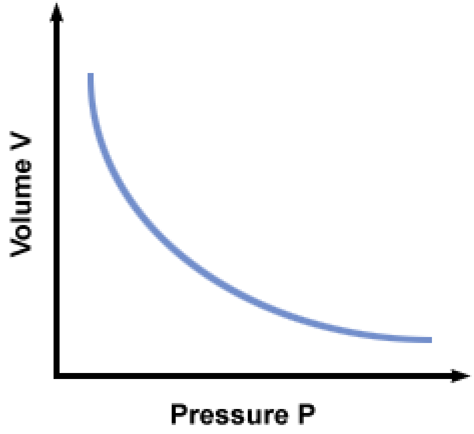
P_1V_1 = k (a constant)

P_2V_2 = k (a constant)

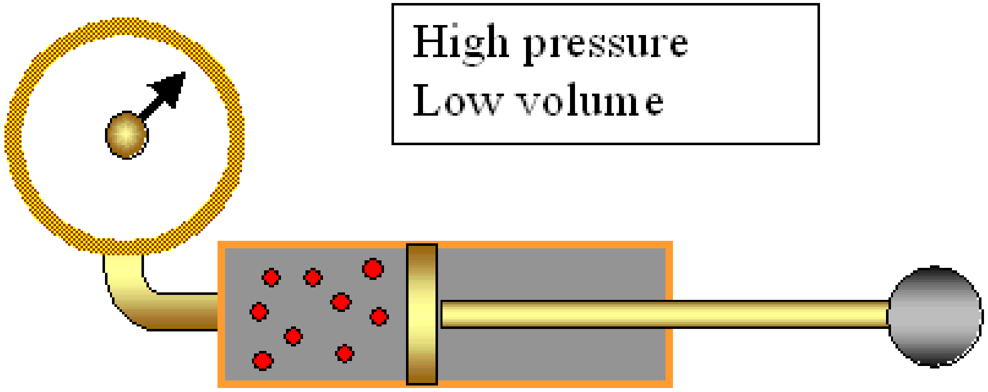
$$\Longrightarrow \boxed{P_1V_1 = P_2V_2}$$

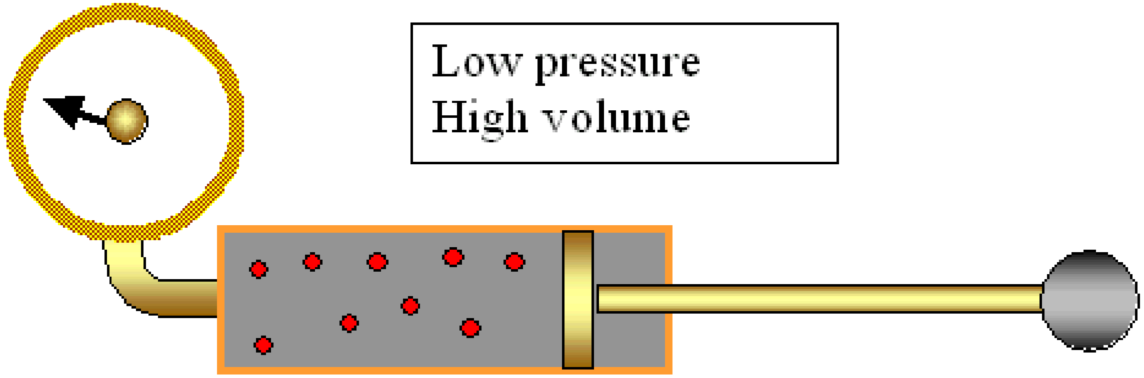
$$P_2 = \frac{P_1V_1}{V_2} \qquad V_2 = \frac{P_1V_1}{P_2}$$

Boyle's Law



decreasing the volume of a gas sample means increasing the pressure.





Charles's Law

Jacques Charles

Charles's law: m,P: constant

$$T \ \alpha \ V \quad \Longrightarrow \quad \frac{V}{T} = k \text{ (a constant)}$$

$$\frac{V_1}{T_1} = \text{k (a constant)}$$

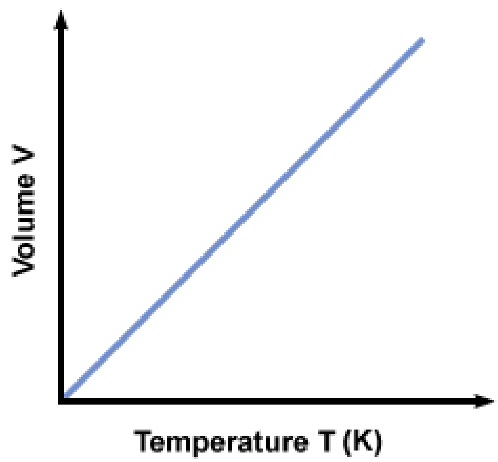
$$\frac{V_2}{T_2} = \text{k (a constant)}$$

$$\Longrightarrow \quad \boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$$

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

Charles's Law

gas-properties_en.jar



an increase in temperature at constant pressure results in an increase in volume.

Gay-Lussac's Law

Gay-Lussac's law: m,V: constant

$$P \;\alpha\; T \quad\Longrightarrow\quad \frac{P}{T} = k \text{ (a constant)}$$

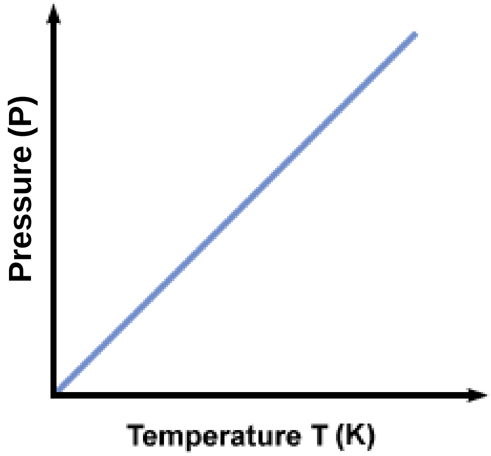
$$\frac{P_1}{T_1} = k \text{ (a constant)}$$

$$\frac{P_2}{T_2} = k \text{ (a constant)}$$

$$\Longrightarrow \quad \boxed{\frac{P_1}{T_1} = \frac{P_2}{T_2}}$$

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

Gay-Lussac's Law



Combined Gas Law

Combined gas law:

m (or n): constant

$$\frac{PV}{T} = k \text{ (a constant)}$$

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

Practice:

- A sample of gas occupies 249 L at 12.1 mm Hg pressure. The pressure is changed to 654 mm Hg; calculate the new volume of the gas. (Assume temperature and moles of gas remain the same).

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

T constant $\rightarrow T_1 = T_2$

$$P_1V_1 = V_2P_2$$

$$(12.1 \text{ mm Hg})(249 \text{ L}) = V_2(654 \text{ mm Hg})$$

$$\mathbf{V_2 = 4.61\ L}$$

Practice:

- A 25.2 mL sample of helium gas at 29°C is heated to 151°C. What will be the new volume of the helium sample? (Assume pressure remains the same).

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

P constant → $P_1 = P_2$

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

$$\frac{25.2\ \mathrm{mL}}{(273 + 29\ \mathrm{K})} = \frac{V_2}{(273 + 151\ \mathrm{K})}$$

$$\mathbf{V_2 = 35\ mL}$$

Practice:

- A nitrogen gas sample at 1.2 atm occupies 14.2 L at 25°C. The sample is heated to 34°C as the volume is decreased to 8.4 L. What is the new pressure?

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

Avogadro's Law

Avogadro's law: P,T: constant

$$V \;\alpha\; n \quad \Longrightarrow \quad \frac{V}{n} = k \text{ (a constant)}$$

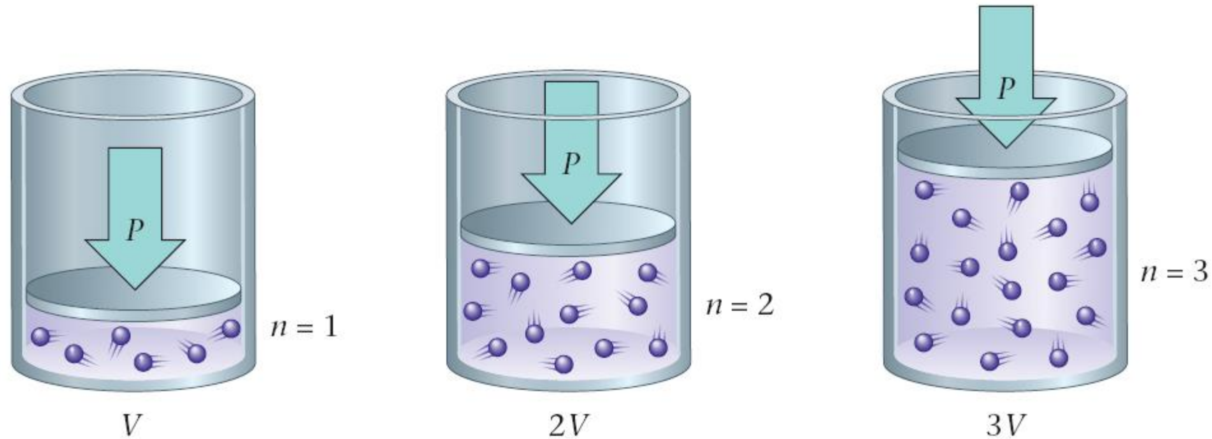
$$\frac{V_1}{n_1} = k \text{ (a constant)}$$

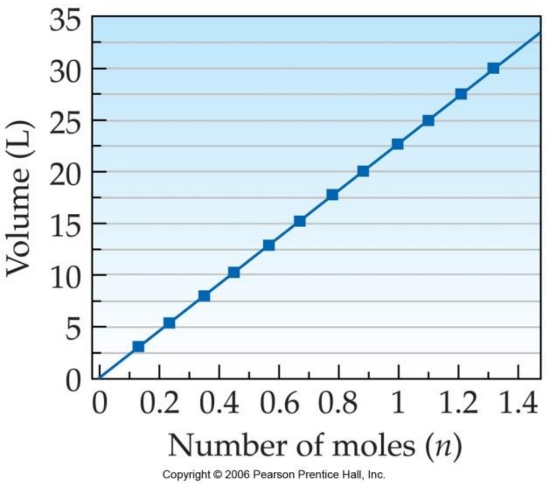
$$\frac{V_2}{n_2} = k \text{ (a constant)}$$

$$\Longrightarrow \quad \boxed{\frac{V_1}{n_1} = \frac{V_2}{n_2}}$$

Avogadro’s Law

increasing the moles of gas at constant pressure means more volume is needed to hold the gas.





gas-properties_en.jar

Avogadro's Law

Practice:

- If 0.105 mol of helium gas occupies a volume 2.35 L at a certain temperature and pressure, what volume would 0.337 mol of helium occupy under the same conditions?

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

$$\frac{2.35\ \mathrm{L}}{0.105\ \mathrm{mol}} = \frac{V_2}{0.337\ \mathrm{mol}}$$

$$\mathbf{V_2 = 7.54\ L}$$

Ideal Gas Law

Ideal gas law:

$$\boxed{PV = nRT}$$

n: number of moles (mol)
R: universal gas constant
V: volume (L)
P: pressure (atm)
T: temperature (K)

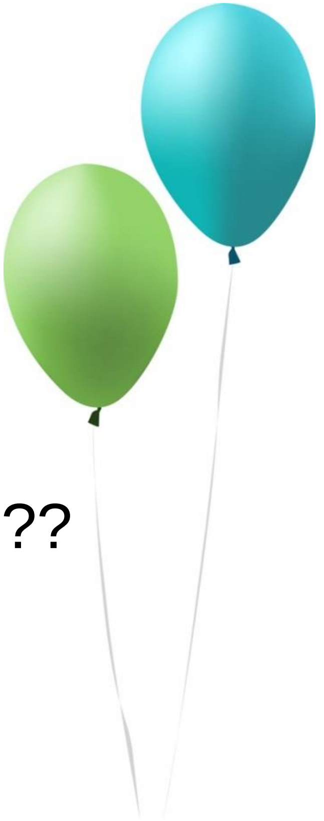
Standard Temperature and Pressure (STP)

T = 0.00°C (273 K)
P = 1.000 atm

$\Rightarrow$ 1 mole → V = 22.4 L

$$R = \frac{PV}{nT} = \frac{(1.000 \text{ atm})\,(22.4 \text{ L})}{(1 \text{ mol})\,(273 \text{ K})} = 0.082106 \; \frac{\text{L.atm}}{\text{mol.K}}$$

Ideal Gas Law



Practice:

- Given three of the variables in the Ideal Gas Law, calculate the fourth, unknown quantity.
- P = 0.98 atm, n = 0.1021 mol, T = 302 K.   V = ???

$$\mathrm{PV = nRT}$$

$$(0.98\ \mathrm{atm})\mathrm{V} = (0.1021\ \mathrm{mol})(0.08206\ \frac{\mathrm{L\ atm}}{\mathrm{mol\ K}})(302\ \mathrm{K})$$

$$\mathbf{V = 2.6\ L}$$

Ideal Gas Law

Practice:

- At what temperature will a 1.00 g sample of neon gas exert a pressure of 500. torr in a 5.00 L container?

$$\mathrm{PV = nRT}$$

$$1.00\ \mathrm{g} \times \frac{1\ \mathrm{mol}}{20.18\ \mathrm{g}} = 0.0496\ \mathrm{mol}$$

$$500.\ \mathrm{torr} \times \frac{1\ \mathrm{atm}}{760\ \mathrm{torr}} = 0.658\ \mathrm{atm}$$

$$(0.658\ \mathrm{atm})\ (5.00\ \mathrm{L}) = (0.0496\ \mathrm{mol})(0.08206\ \tfrac{\mathrm{L\ atm}}{\mathrm{mol\ K}}) \times \mathrm{T}$$

$$\mathrm{T} = 809\ \mathrm{K}$$

At-Home Practice

1a. A 5.00 g sample of neon gas at 25.0 °C is injected into a rigid container. The measured pressure is 1.204 atm. What is the volume of the container?

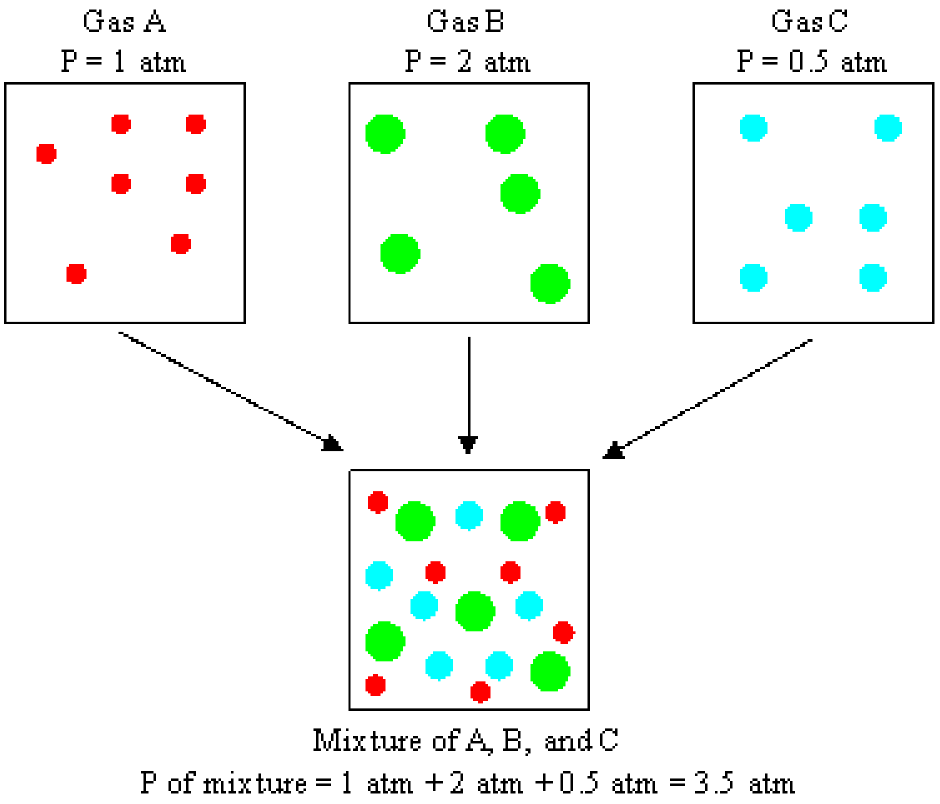
1b. The neon sample is then heated to 200.0 °C. What is the new pressure?

2. An initial gas sample of 1 mole has a volume of 22 L. The same type of gas is added until the volume increases to 44 L. How many moles of gas were <u>added</u>?

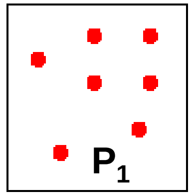
Dalton's Law

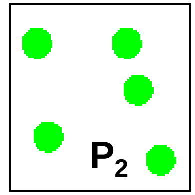
Dalton's law of partial pressure:

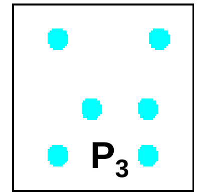
$$P_T = P_1 + P_2 + P_3 + \ldots$$



Dalton's Law





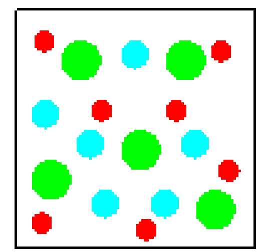


$$P_1 = \frac{n_1 RT}{V} \qquad P_2 = \frac{n_2 RT}{V} \qquad P_3 = \frac{n_3 RT}{V}$$

$$P_T = P_1 + P_2 + P_3$$

$$P_T = n_1(RT/V) + n_2(RT/V) + n_3(RT/V)$$

$$P_T = (n_1 + n_2 + n_3)\ (RT/V)$$



$$\boxed{P_T = n_{total} \left(\frac{RT}{V}\right)}$$

Dalton's Law

For a mixture of ideal gases, the total number of moles is important.
(not the identity of the individual gas particles)

1.75 mol He

V=5L
T=20°C

P_T = 8.4 atm

0.75 mol H_2
0.75 mol He
0.25 mol Ne

1.75 mol

V=5L
T=20°C

P_T = 8.4 atm

1.00 mol N_2
0.50 mol O_2
0.25 mol Ar

1.75 mol

V=5L
T=20°C

P_T = 8.4 atm

1. Volume of the individual gas particles must not be very important.

2. Forces among the particles must not be very important.