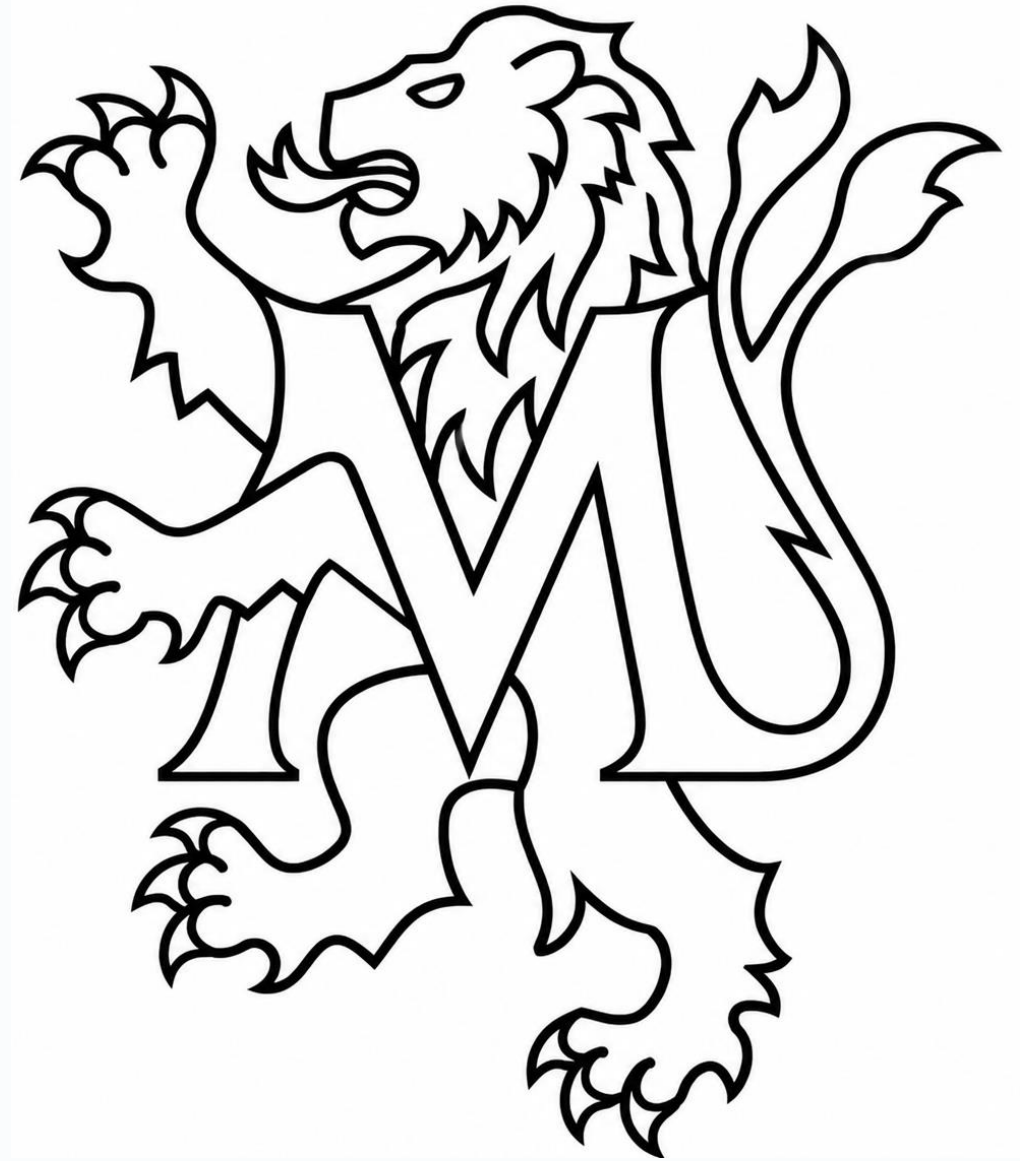


Gases

M-CATS · General Chemistry Chapter 8



Need to know before the session

Pressure comes from collisions with the container.

$PV = nRT$ relates pressure, volume, moles, and temperature.

Partial pressure depends on mole fraction.

At the same temperature, lighter molecules move faster.

Today's topics

Gas properties, pressure, and units

Ideal gas law and proportional relationships

Mixtures, solubility, and molecular motion

Real gases and MCAT practice

Warm-up 1 · A shrinking container

A sealed gas occupies 4.0 L at 1.0 atm.

Temperature stays constant. Volume falls to 2.0 L.

What is the new pressure?

A 0.50 atm B 1.0 atm C 2.0 atm D 4.0 atm

Warm-up 2 · Helium versus oxygen

He and O₂ have the same temperature.

Which comparison is correct?

A Same average kinetic energy, faster He

B Same speed C More kinetic energy in He D Faster O₂

Warm-up 3 · A gas mixture

A mixture contains 1 mol O_2 and 3 mol N_2 .

The total pressure is 2.0 atm.

What is the partial pressure of O_2 ?

A 0.25 atm B 0.50 atm C 1.0 atm D 1.5 atm

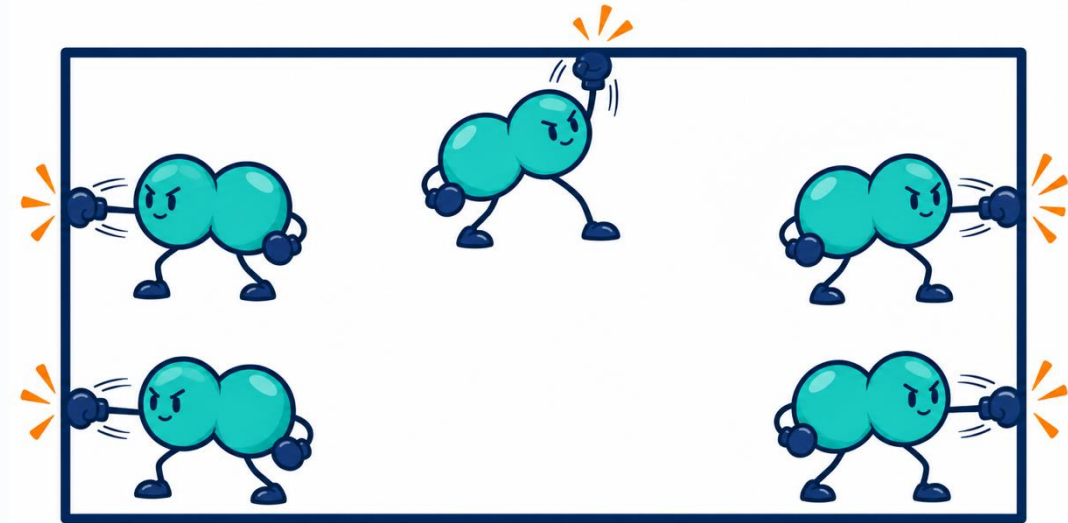
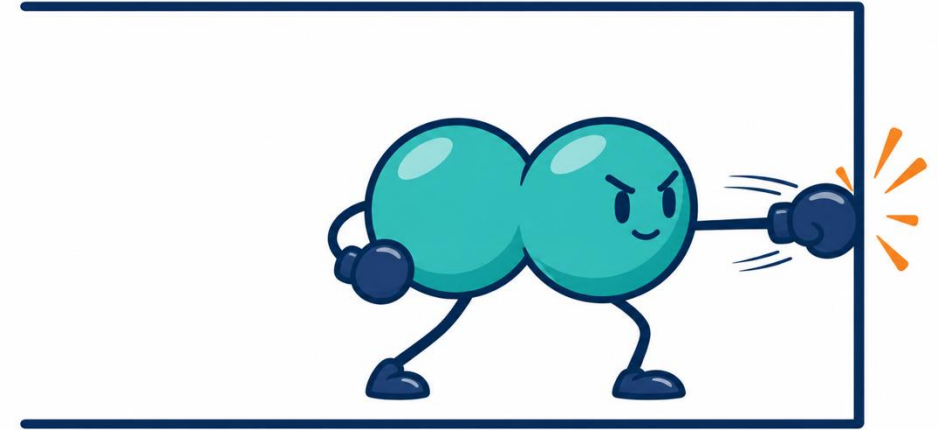
Pressure and gas properties

$$P = F / A$$

Gases fill their containers and compress easily.

Collisions with the walls produce pressure.

More frequent or stronger collisions raise pressure.



Units and STP

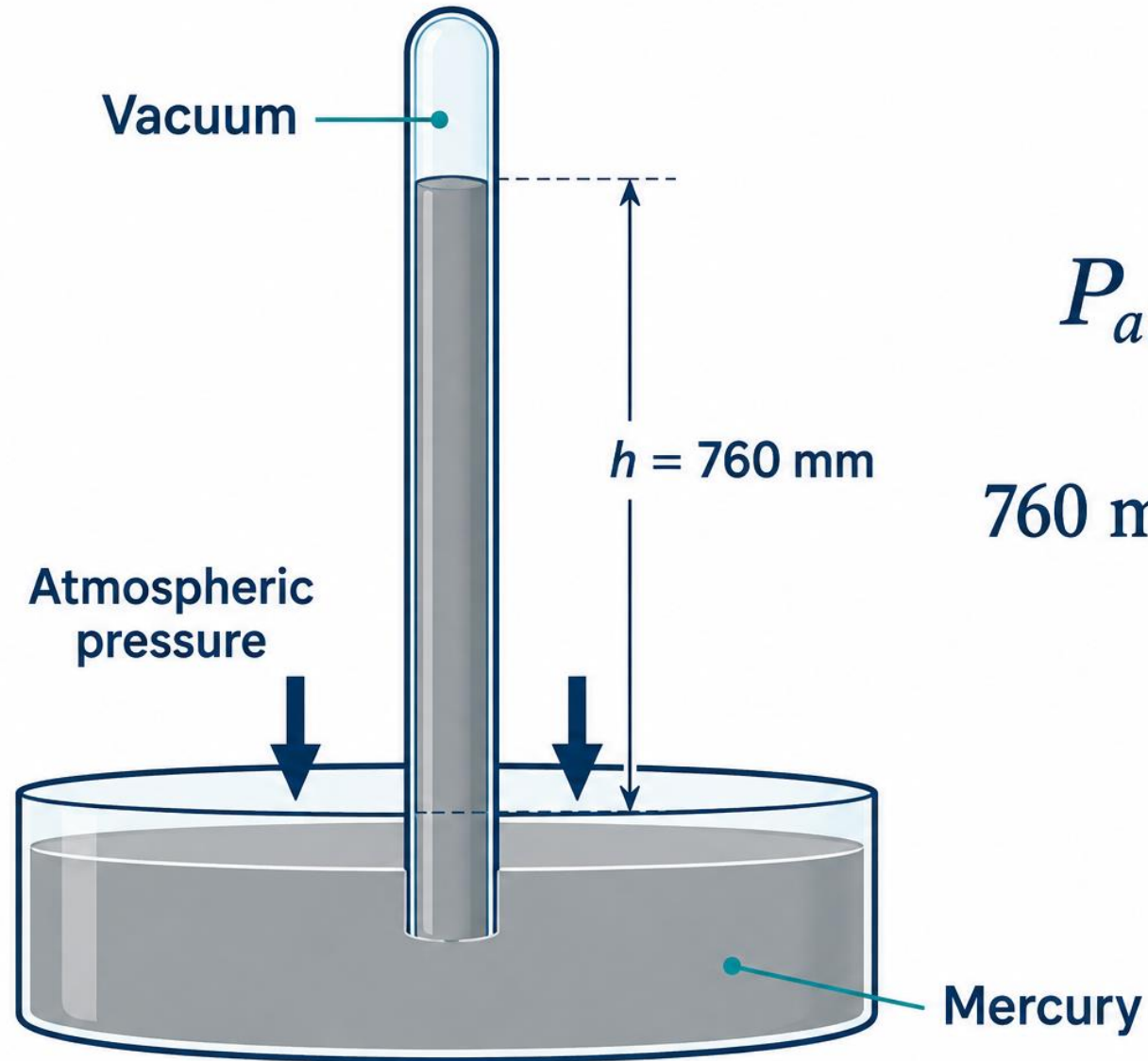
$$1 \text{ atm} = 760 \text{ mmHg} \approx 101.3 \text{ kPa}$$

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

MCAT convention: STP = 273.15 K and 1 atm

At this STP, 1 mol of ideal gas occupies $\approx 22.4 \text{ L}$.

Barometers

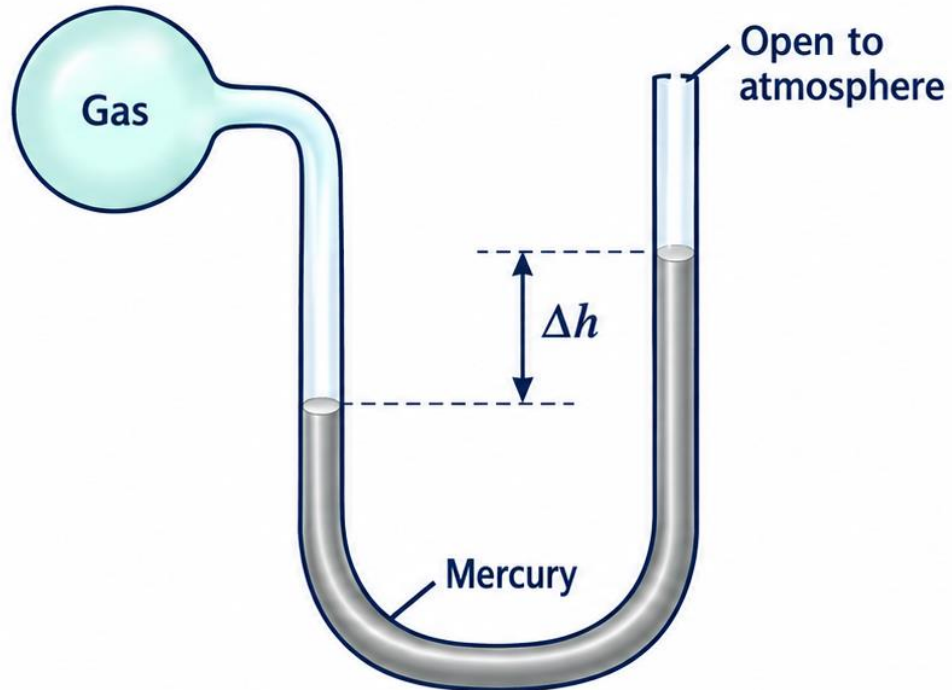


$$P_{atm} = \rho gh$$

$$760 \text{ mmHg} = 1 \text{ atm}$$

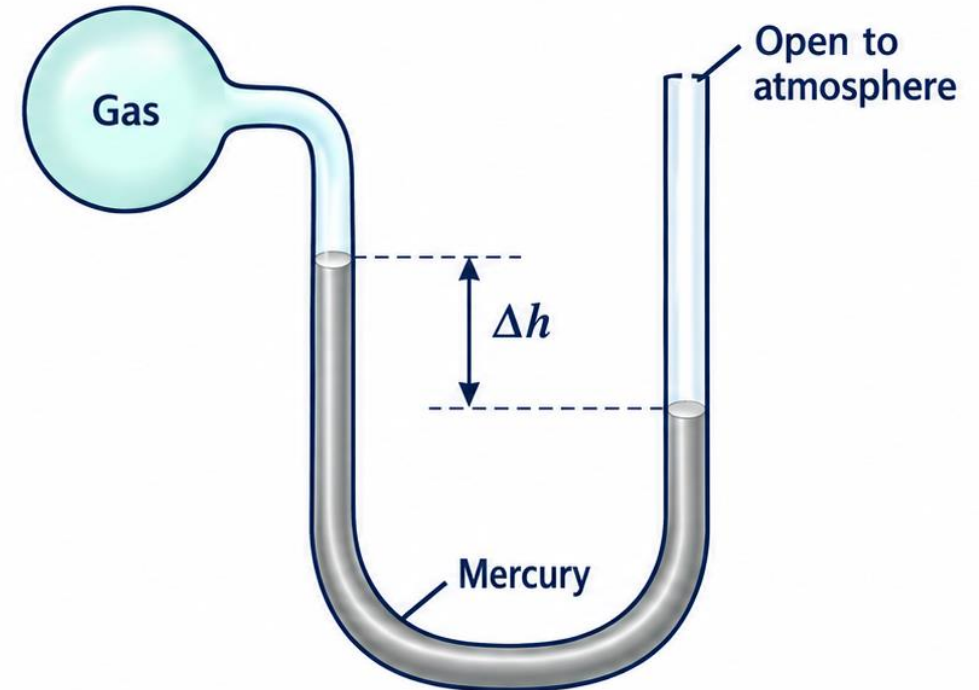
Open-end manometers

Gas pressure > atmospheric pressure



$$P_{\text{gas}} = P_{\text{atm}} + \Delta P$$

Gas pressure < atmospheric pressure



$$P_{\text{gas}} = P_{\text{atm}} - \Delta P$$

$$\Delta P = \rho g \Delta h$$

The ideal gas law

$$PV = nRT$$

P = pressure V = volume n = amount in mol

T = absolute temperature in kelvin

$$R = 0.0821 \text{ L}\cdot\text{atm}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$$

$$R = 8.314 \text{ J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1} \text{ for SI calculations}$$

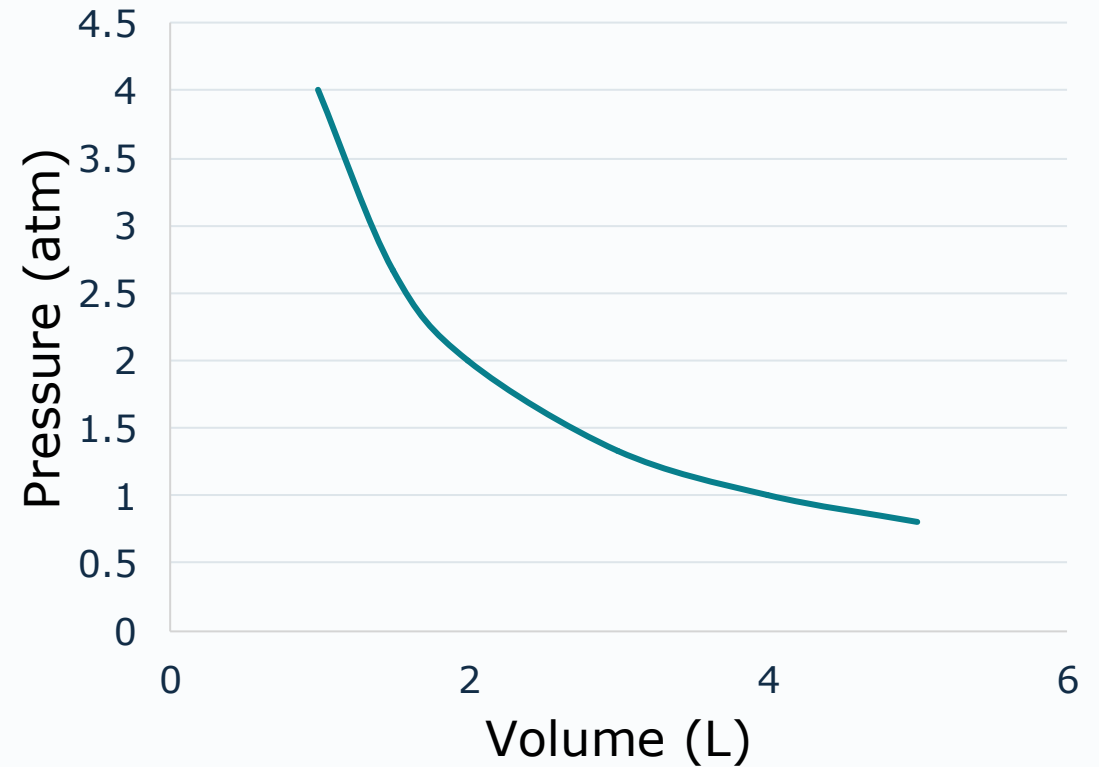
Boyle's law

$$P_1V_1 = P_2V_2$$

Constant n and T

Smaller volume gives higher pressure.

Doubling V halves P .



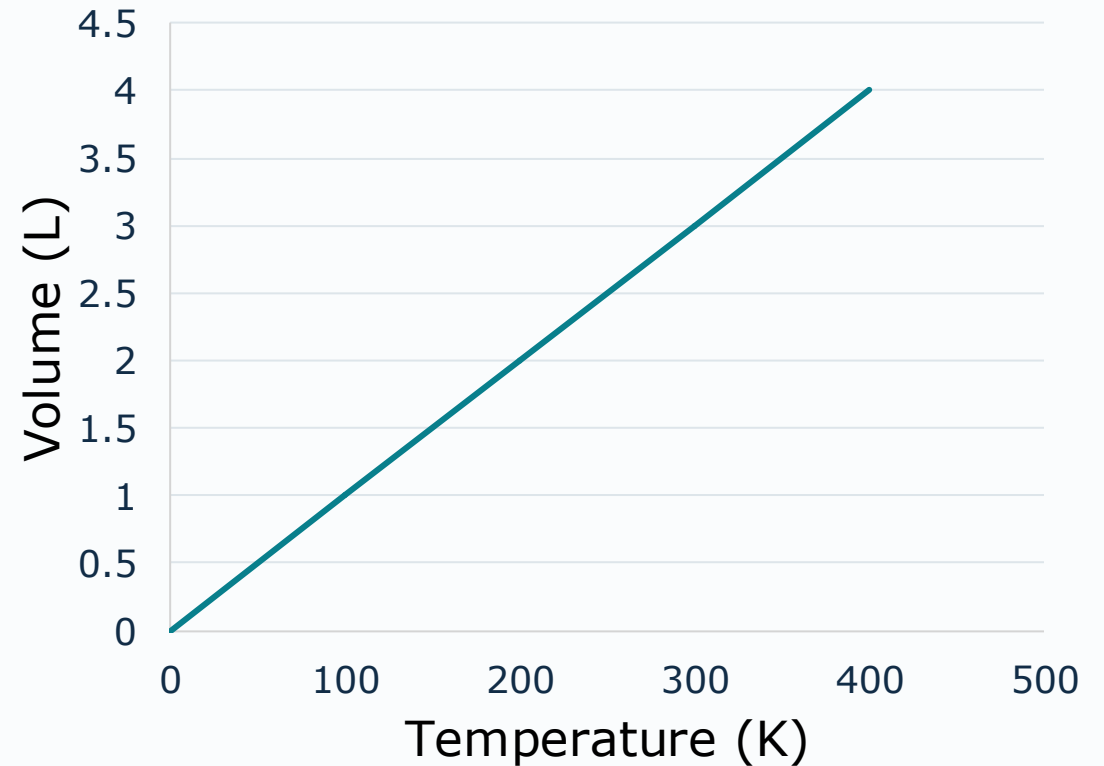
Temperature relationships

$$V_1/T_1 = V_2/T_2$$

Charles: V/T is constant at fixed n , P .

Gay-Lussac: P/T is constant at fixed n , V .

A temperature ratio uses kelvin.



Avogadro and the combined law

Avogadro: $V_1/n_1 = V_2/n_2$ at fixed P and T

Combined law: $P_1V_1/T_1 = P_2V_2/T_2$ at fixed n

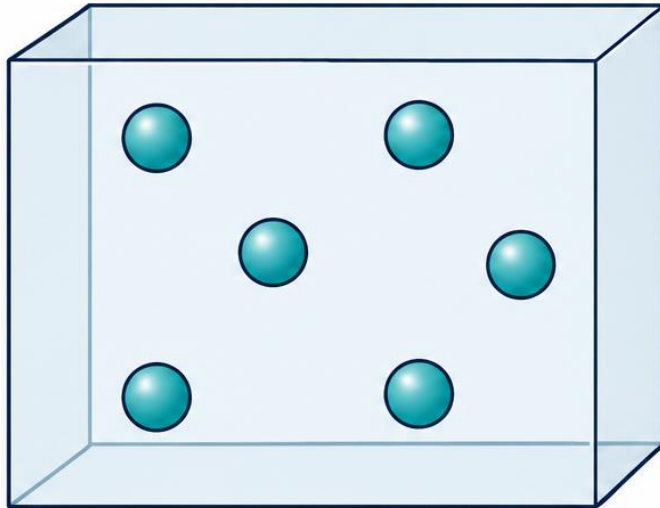
If moles change, return to $PV = nRT$.

Gas density and molar mass

Same P , V , and T

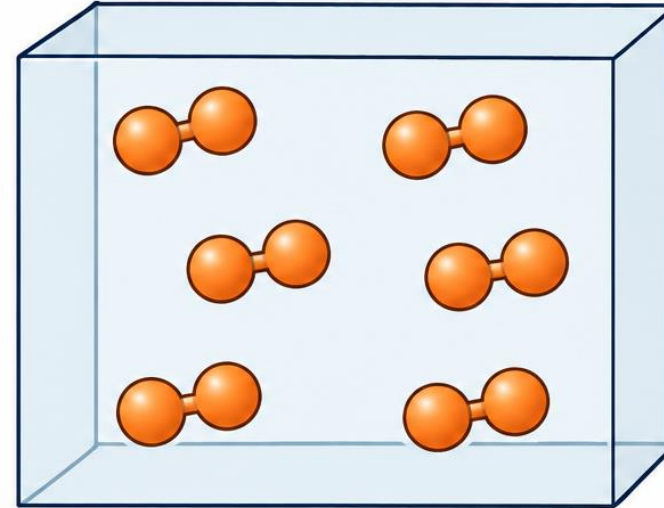
Same number of gas particles

Helium (He)



$M = 4 \text{ g/mol}$
Lower density

Oxygen (O_2)



$M = 32 \text{ g/mol}$
8× the density

$$d = \frac{PM}{RT}$$

At fixed P and T , density is proportional to molar mass.

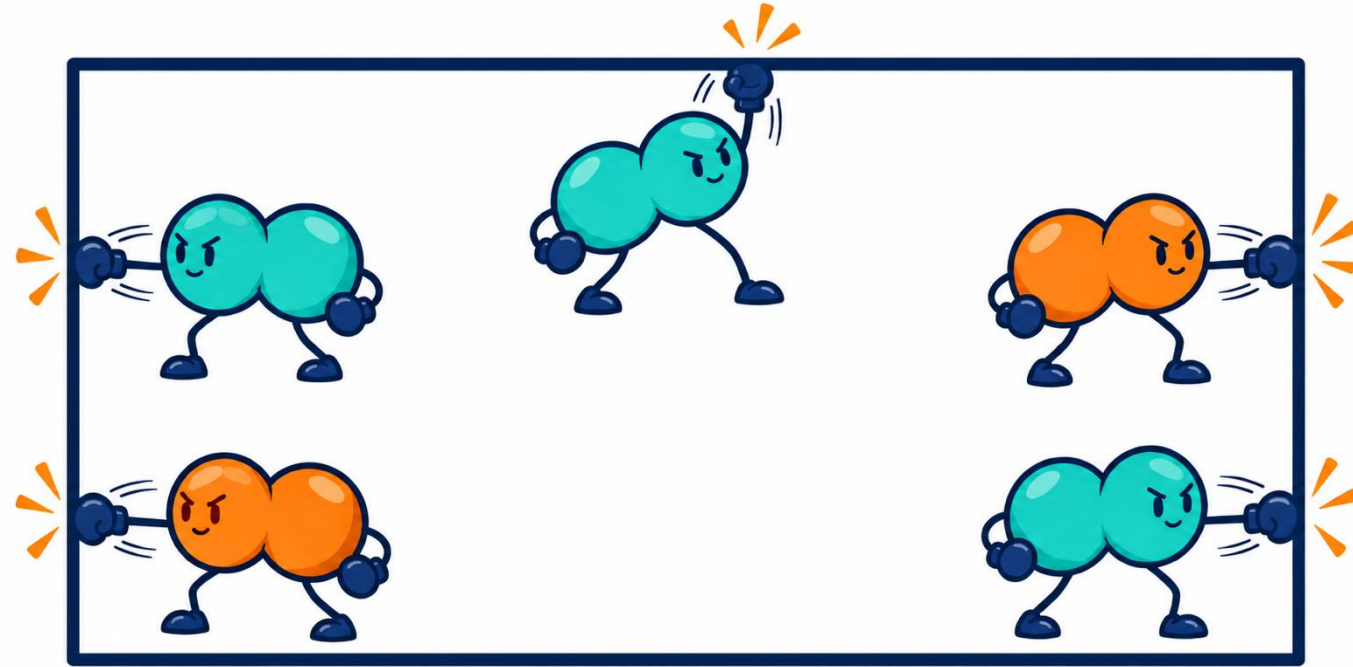
Dalton's law

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

$$X_i = n_i / n_{\text{total}}$$

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

Each gas contributes part of the total pressure.



Gas collected over water

The sample contains the collected gas plus water vapor.

$$P_{\text{gas}} = P_{\text{total}} - P_{\text{H}_2\text{O}}$$

Water-vapor pressure depends on temperature.

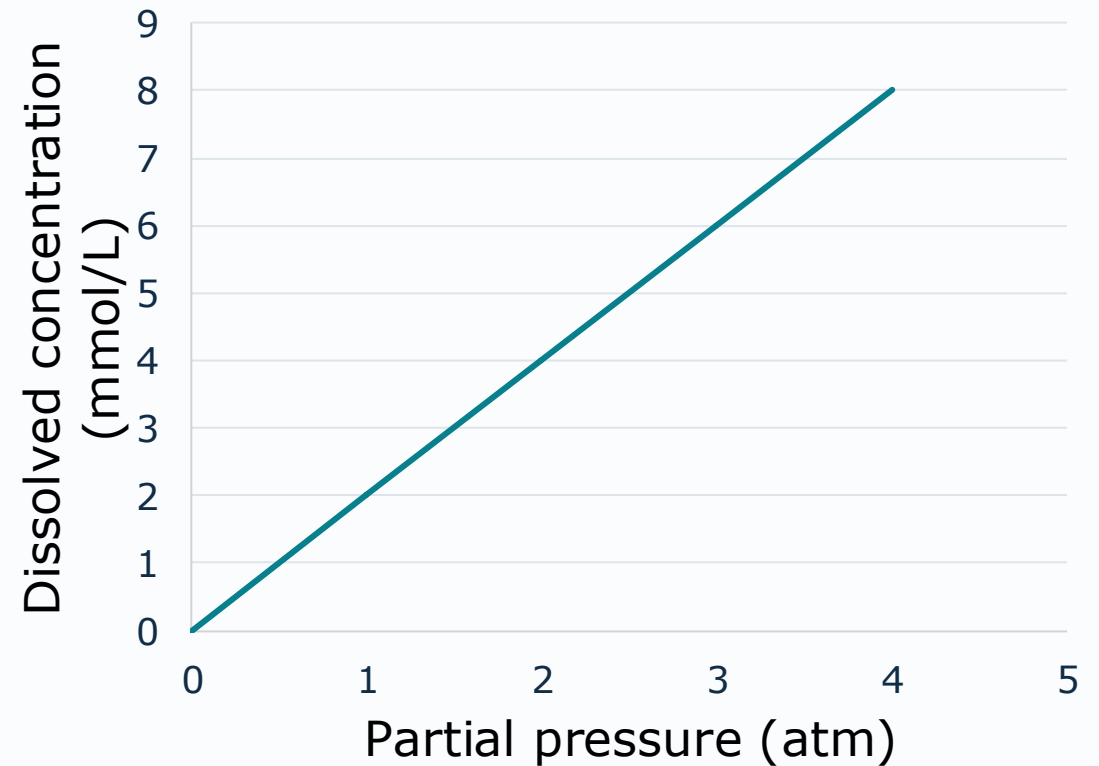
Henry's law

$$C = kP_{\text{gas}}$$

Higher gas partial pressure increases dissolved gas.

k depends on the gas, solvent, and temperature.

Many gases dissolve less as temperature rises.



Kinetic molecular theory

Particles have negligible volume in the ideal model.

The ideal model ignores attractions between particles.

Particles move randomly and collide elastically.

Average translational kinetic energy depends only on T .

Molecular speeds

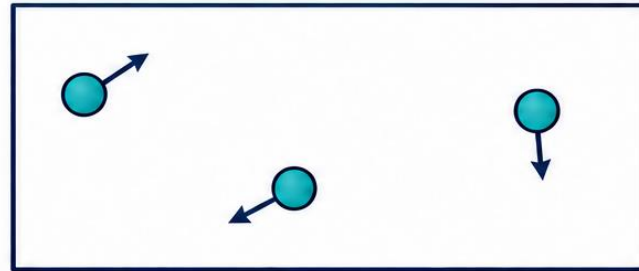
$$u_{\text{rms}} = \sqrt{(3RT / M)}$$

Higher T increases molecular speed.

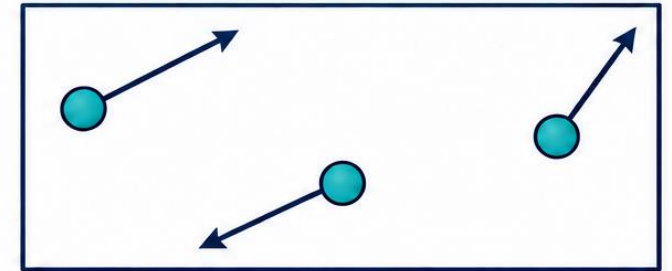
At the same T, lighter molecules move faster.

Same gas: higher temperature means faster motion

Cooler



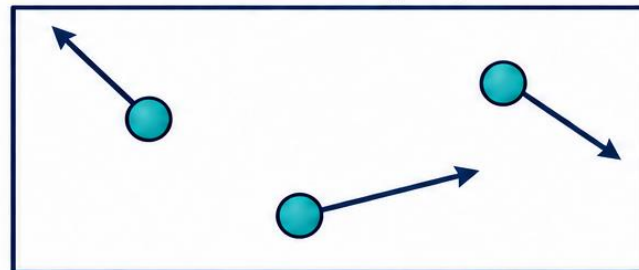
Hotter



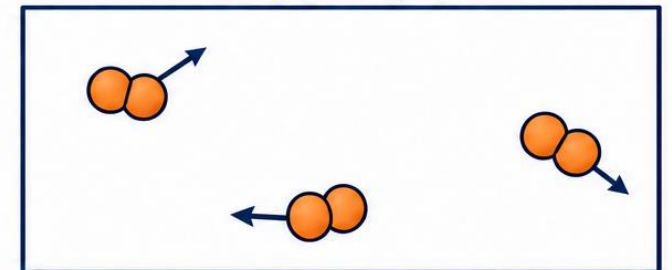
Higher temperature = greater average kinetic energy

Same temperature: lighter particles move faster

Helium (He)



Oxygen (O₂)



Same average kinetic energy

$$KE = \frac{1}{2}mv^2 = \frac{3}{2}k_B T$$

Average kinetic energy

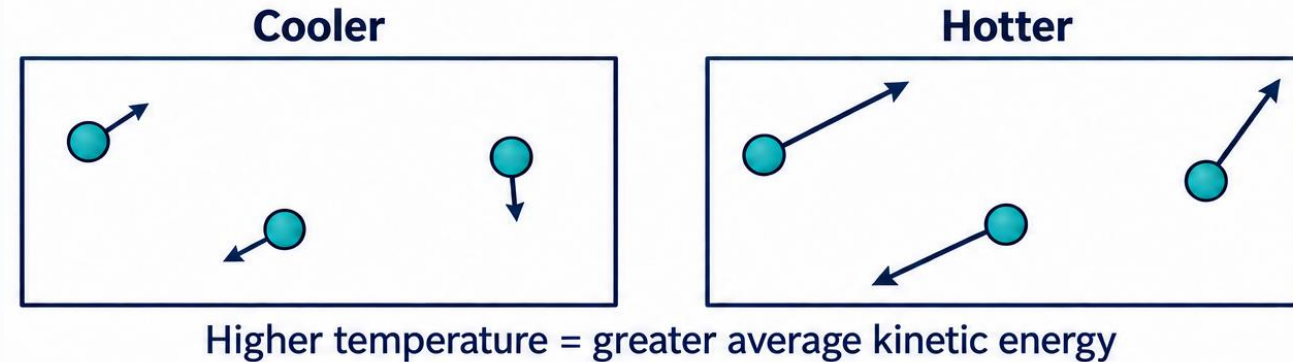
$$\langle KE \rangle = \frac{3}{2} k_B T$$

Same T gives the same average KE for every ideal gas.

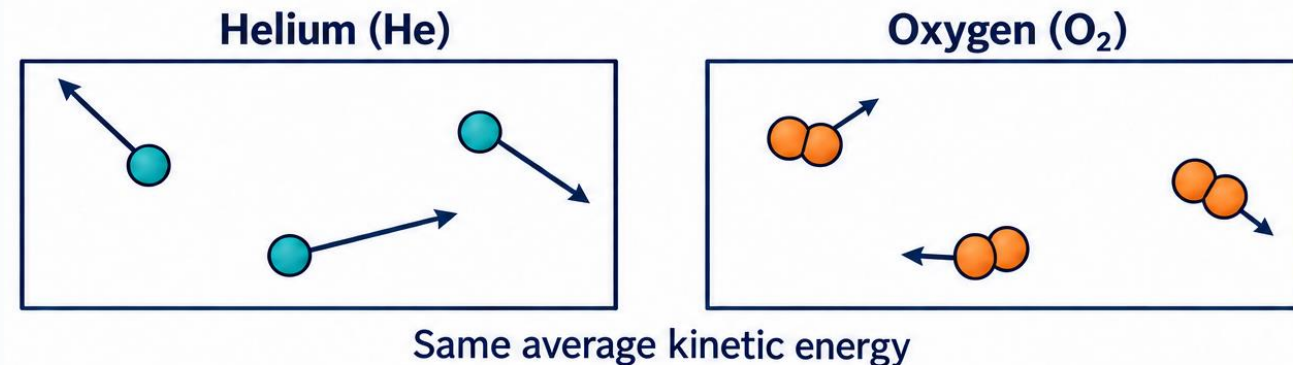
A lighter molecule moves faster at the same average KE.

Doubling T doubles average KE, but multiplies RMS speed by $\sqrt{2}$.

Same gas: higher temperature means faster motion



Same temperature: lighter particles move faster



$$KE = \frac{1}{2}mv^2 = \frac{3}{2}k_B T$$

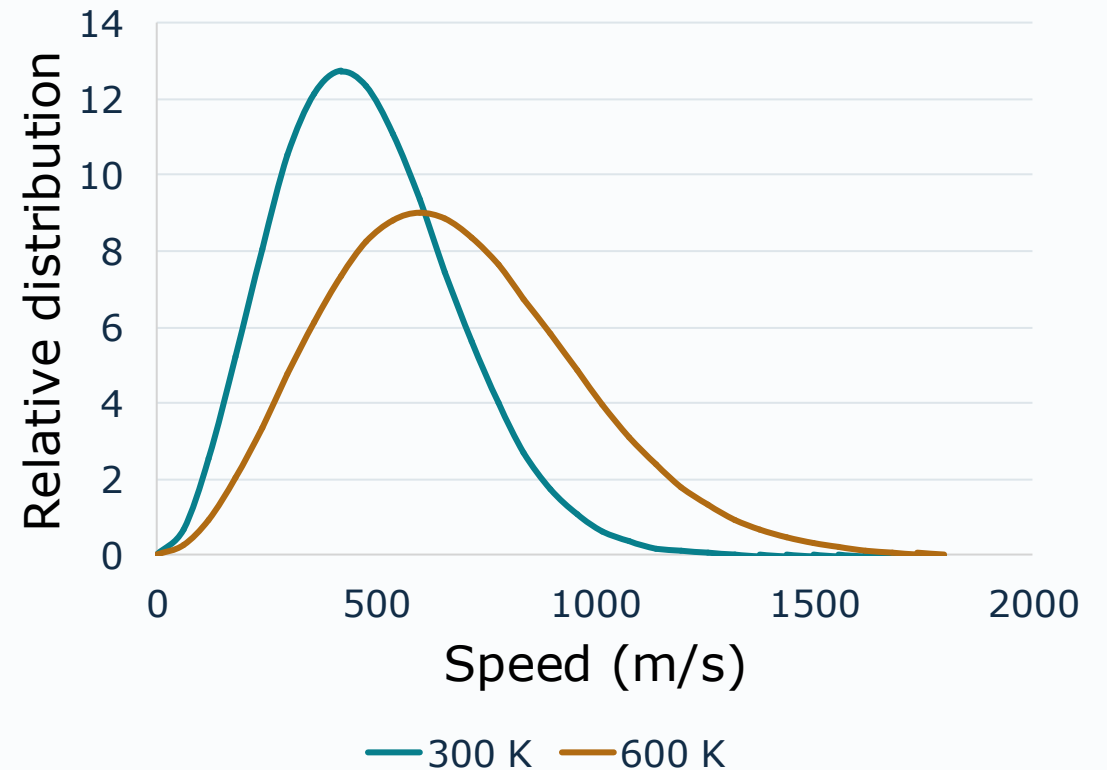
The speed distribution

Higher temperature

Molecules have a range of speeds.

Heating shifts the peak right and lowers it.

The area stays constant for the same number of molecules.



Diffusion and effusion

$$r_1/r_2 = \sqrt{(M_2/M_1)}$$

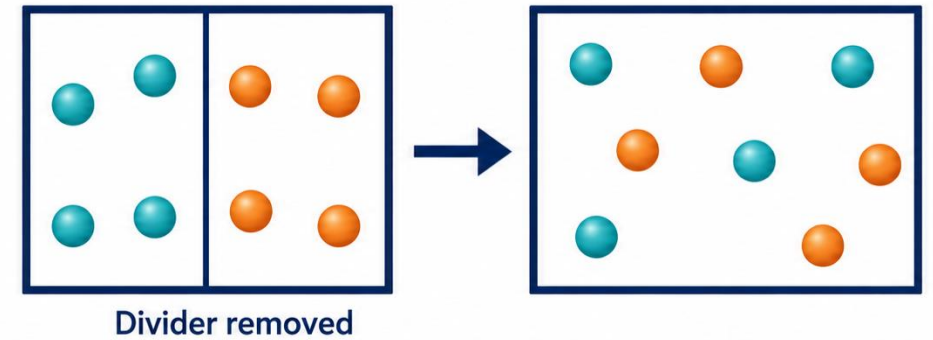
Diffusion: gas spreads and mixes through space.

Effusion: gas escapes through a tiny opening.

Lighter gas effuses faster at the same temperature.

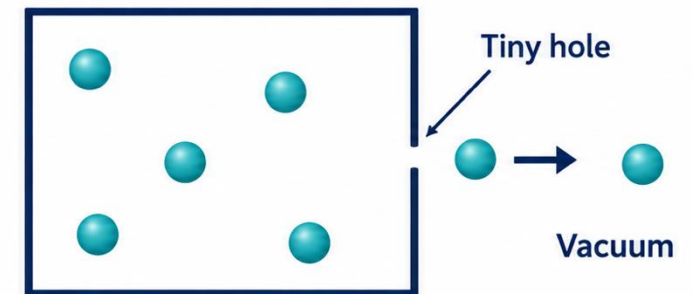
He versus O₂: rate ratio = $\sqrt{(32/4)} \approx 2.8$

Diffusion



Gases spread and mix

Effusion



Gas escapes through a tiny opening

Most probable, mean, and RMS speeds

Most probable < Mean < RMS

Most probable: $v = \sqrt{2RT/M}$

Mean: $v = \sqrt{8RT/\pi M}$

RMS: $v = \sqrt{3RT/M}$

Graham's law · Rate versus time

$$t_1/t_2 = r_2/r_1 = \sqrt{(M_1/M_2)}$$

Compare the time for equal amounts of gas to escape.

Faster effusion means less time.

He/O₂ rate ratio ≈ 2.8 , so the time ratio ≈ 0.35 .

If O₂ takes 28 s, He takes about 10 s under comparable conditions.

Interactive activity · What changes?

Teams predict P , V , and molecular speed.

- 1 Compress a sealed syringe at constant T .
- 2 Heat a rigid sealed tank.
- 3 Add moles to a flexible balloon at fixed P and T .

Mini-passage · Gas collection

A student collects O₂ over water at 300 K.

The sample occupies 2.46 L at 1.00 atm total pressure.

Water-vapor pressure at 300 K is 0.030 atm.

Use $R = 0.082 \text{ L}\cdot\text{atm}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$. Assume ideal gases.

Passage questions

- 1 How many moles of O_2 were collected?
- 2 All O_2 is dried, then sealed in 1.23 L at 600 K. What is P?
- 3 How does its rms speed compare with the speed at 300 K?

Explain which equation or relationship you used.

Real gases

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

Low pressure and high temperature favor ideal behavior.

Attractions can lower pressure below the ideal prediction.

Finite particle volume matters at very high pressure.

Real-gas check

Which sample should behave most ideally?

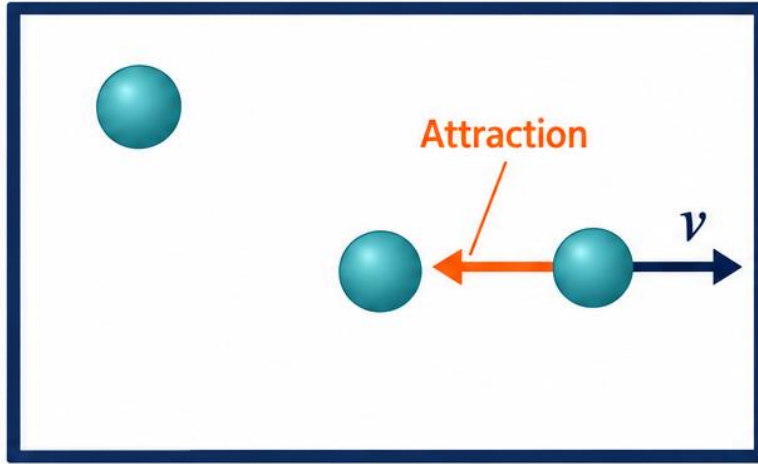
A High P, low T B Low P, high T

C High P, high T D Low P, low T

If attractions dominate, is P_{real} above or below P_{ideal} ?

Van der Waals corrections

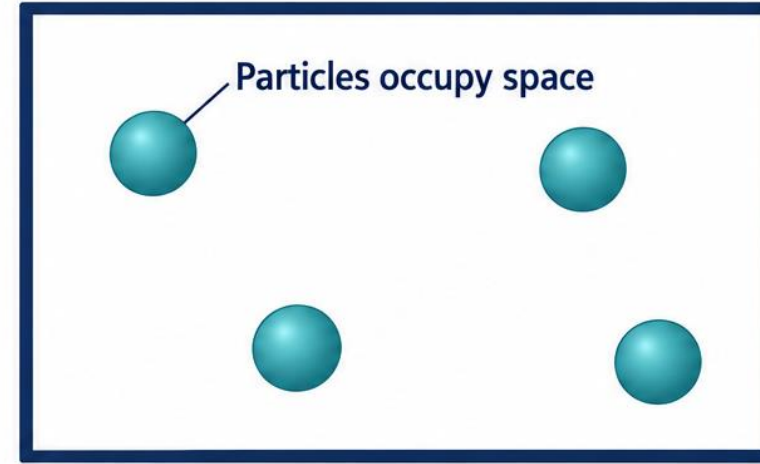
a: Attractions



Attractions reduce pressure on the walls

$$P + \frac{an^2}{V^2}$$

b: Particle volume



Less space is available for molecular motion

$$V - nb$$

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

a corrects attractions. b corrects excluded volume.

Why physicians learn gas laws

Ventilation: expanding thoracic volume helps lower pressure.

Gas mixtures: oxygen fraction and total pressure set partial pressure.

Dissolved gas: Henry's law links equilibrium concentration to pressure.

Formula summary

$$PV = nRT \quad d = PM/RT \quad T(K) = T(^{\circ}C) + 273.15$$

$$P_1V_1/T_1 = P_2V_2/T_2 \quad (\text{fixed } n)$$

$$P_i = X_iP_{\text{total}} \quad C = kP_i$$

$$u_{\text{rms}} = \sqrt{(3RT/M)} \quad r_1/r_2 = \sqrt{(M_2/M_1)}$$