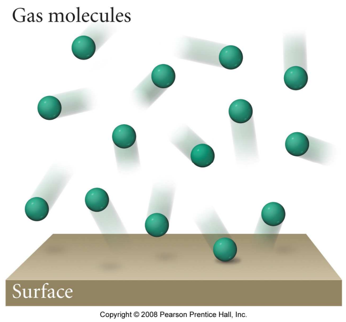
Chapter 5: Gases

Gases: fluid or rigid?

compressible or incompressible?

Pressure: force per unit area

$$= \frac{F}{A}$$



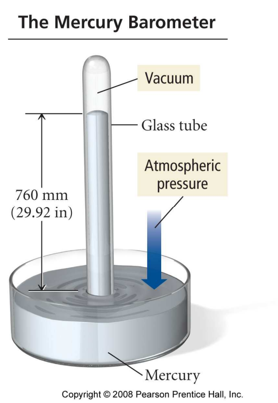
Units of pressure:

atmosphere (atm): 1 atm = average atmospheric pressure at sea level

millimeters mercury (mmHg): 1 atm = 760 mmHg

1 mmHg = 1 torr

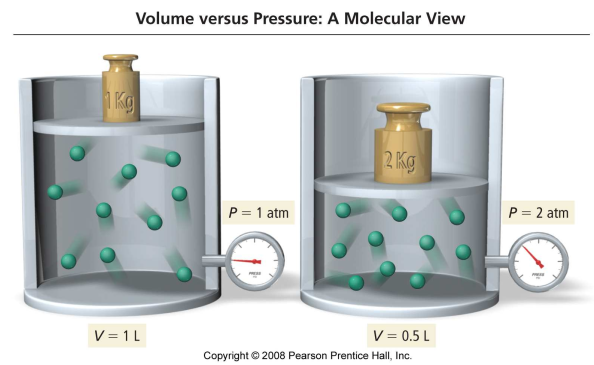
pascal (Pa): SI unit of pressure
I atm = 101,325 Pa

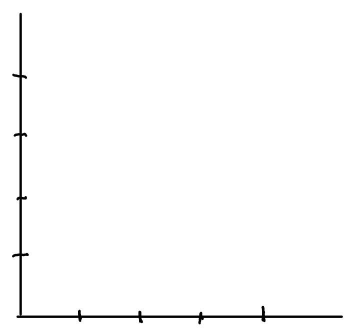


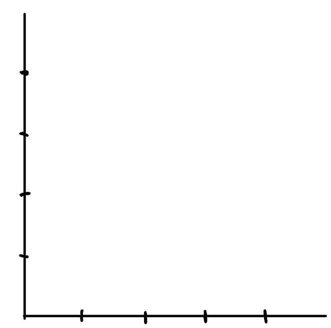
Simple gas laws

Simple gas laws (sometimes called empirical gas laws): relationships between pressure (P), volume (V), absolute temperature (T), and amount in moles (n)

Boyle's Law: relationship of P and V, and the compressibility of gases







Boyle's law:

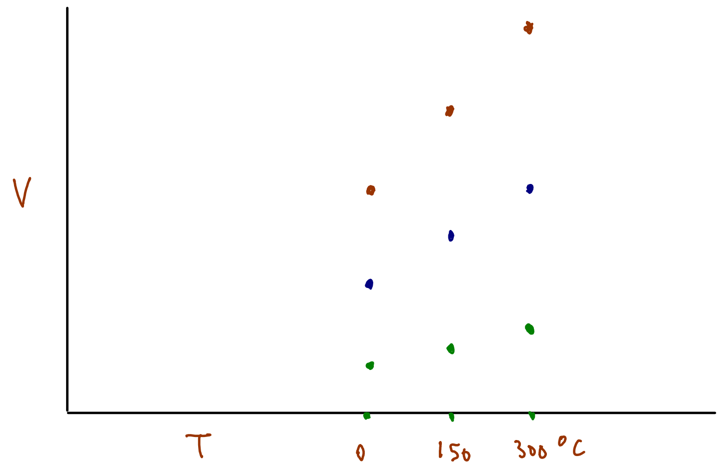
Boyle's law

$V = \text{constant} \times \frac{1}{P}$

6.2 L air at 760 mmHg is compressed to 4.4 L. What is the new pressure (assuming constant temperature)?

Charles's Law

Charles's Law relates *V* and *T* for a fixed amount of gas



Charles's law:

A balloon has a volume of 1.0 L at 298 K. What is its volume at 77.4 K?

Avogadro's Law

Avogadro's law relates amount (in mol) with volume, assuming constant temperature and pressure

Avogadro's law:

Ideal Gas Law

Boyle's law:

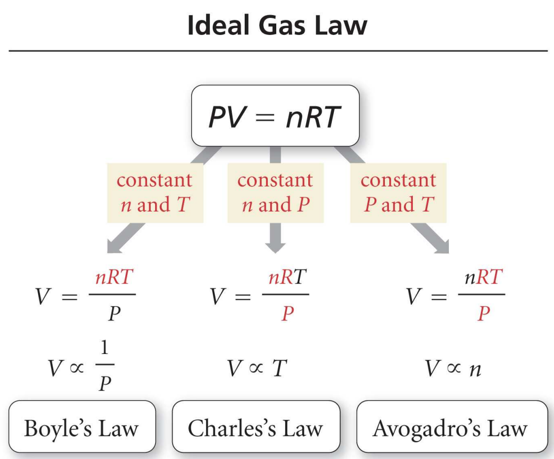
Charles's law:

Avogadro's law:

Combined:

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

<u>Ideal gas law</u>:



Ideal gas law problem

A 438 L gas cylinder contains 0.885 kg O_2(*g*) at 21.0 ^oC. What is the pressure (in atm and mmHg) inside this container?

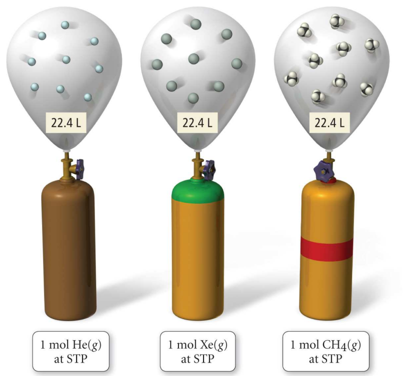
Molar volume at STP

Standard temperature and pressure (STP):

- $T = 0\ ^oC = 273$ K
- $P = 1.00$ atm

At STP, solve $PV = nRT$ for V and for $n = 1.00$ mol:

V = 22.4 L for 1.00 mol of **any** gas at STP.



Density of a gas

$$d = \frac{\text{mass}}{\text{volume}} = \frac{\text{molar mass}}{\text{molar volume}}$$

What is the density of oxygen gas at STP?

At other pressures or temperatures…

$n =$

What is the density of $O_2(g)$ at 125 ^oC and 740 mmHg?

Molar mass of a gas

Given: mass, volume, pressure, temperature
Calculate: molar mass (g/mol)

$PM_m = dRT \qquad M_m = \frac{dRT}{P}$

Calculate the molar mass of a gas sample with mass of 0.582 g, volume of 213 mL, pressure of 754 mmHg, and temperature of 100.0 ^oC.

Mixtures of gases

Gas molecules act independently!

For any mixture of gases, the total pressure is the sum of the pressures that each gas exerts (partial pressures)

$P_{total} = P_a + P_b + P_c + \ldots$ (Dalton's law of partial pressures)

Mole fraction = $\frac{n_a}{n_{total}} = \chi_a$

In one container with constant *V* and *T*,

$$\frac{n_a}{n_{total}} = \frac{P_a}{P_{total}} = \chi_a \qquad P_a = \chi_a P_{total}$$

Air is 78% nitrogen by volume. What is the partial pressure of nitrogen if the atmospheric pressure is 755 mmHg that day?

Mole fraction and partial pressure

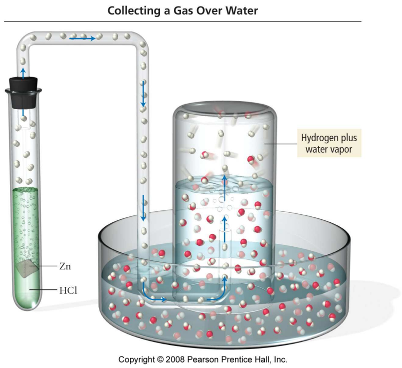
A mixture of 5 g H_2, 8 g N_2, and 10 g O_2 has a total pressure of 7 atm. What is P_{H2}, P_{N2}, and P_{O2}?

Collecting a gas over water

Partial pressures work in the ideal gas law too…

$P_aV = n_aRT$ for one gas (a) in a mixture of gases

Collection of a gas over water



When H_2 is produced by the reaction, the gas collected is a mixture of H_2 and water vapor.

Vapor pressure of water:

max P_{H2O} when a gas is saturated with water vapor

T	P_{H2O} (mmHg)
20 °C	17.55
23 °C	21.10
25 °C	23.78

$P_{total} = P_{H2O} + P_{H2}$

Collecting gases over water

If you collect 19.0 mL of gas over water from a reaction that produces $H_2(g)$, how many moles of hydrogen gas were collected? Atmospheric pressure = 760 mmHg, T = 23 °C.

Stoichiometry and gas laws

With known *P, V,* and *T,* you can calculate...

From a balanced chemical equation you can construct a...

$CO_2(g) + 2\ LiOH(s) \rightarrow Li_2CO_3(s) + H_2O(l)$ (a CO_2 scrubber)

How many L of CO_2 (at 1 atm and 293 K) can be removed by 10.0 g LiOH?

On your own, how many kg LiOH are necessary to remove the CO_2 produced by 2 men for 2 days? (Assume 12 breaths per minute, 0.50 L air per breath, 5.0% of exhaled air's volume is CO_2)

Kinetic molecular theory

Simplest model for behavior of gases:

- Gases made of particles (atoms/molecules) with nearly negligible volumes (mostly empty space)
- Average kinetic energy is proportional to absolute temperature
- Particles fly in a straight line until they collide with another particle or wall of container
- Collisions are completely elastic (energy is exchanged, not lost)

Boyle's law

Charles's law

Avogadro's law

Dalton's law

Temperature and velocity

Average molecular velocity (u_{RMS})
(root mean square is a type of average)

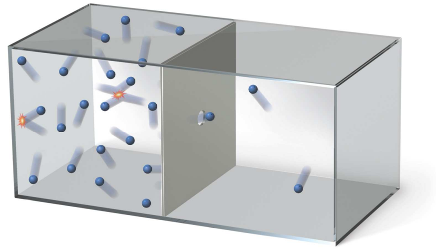
What is the root mean square velocity of nitrogen molecules at 298 K?

Be very careful of units in these calculations!!

R = 8.314 (kg · m^2)/(s^2 · K · mol) (This is R in SI units)

Effusion and diffusion

Effusion: escape of gas through a pinhole into a vacuum



Diffusion: gas spreading in a container or mixing with other gases

Both are related to molar mass of gas

Effusion/diffusion rate

Which balloon would shrink faster? One filled with He or N_2? By what ratio?

Real Gases, the Van der Waals Equation

The ideal gas law is accurate when:

- The volume of gas particles is small compared to the space between them
- The attractive forces between gas molecules are not significant

These assumptions are valid around STP, but they are not valid:

- **when the pressure is much higher than 1 atm.**
  With higher pressure (several hundred atm), the molecules themselves take up a significant amount of the sample's volume. Ideal gas law predicts the molecules have **no** volume, so the actual volume will be:

  To correct for this, *V* becomes:

- **when the temperature is much lower than 298 K.**
  With low temperatures, the molecules begin to stick together more, reducing the number of collisions with the walls of the container. The actual pressure will be:

  To correct for this, *P* becomes:

Van der Waals Equation: