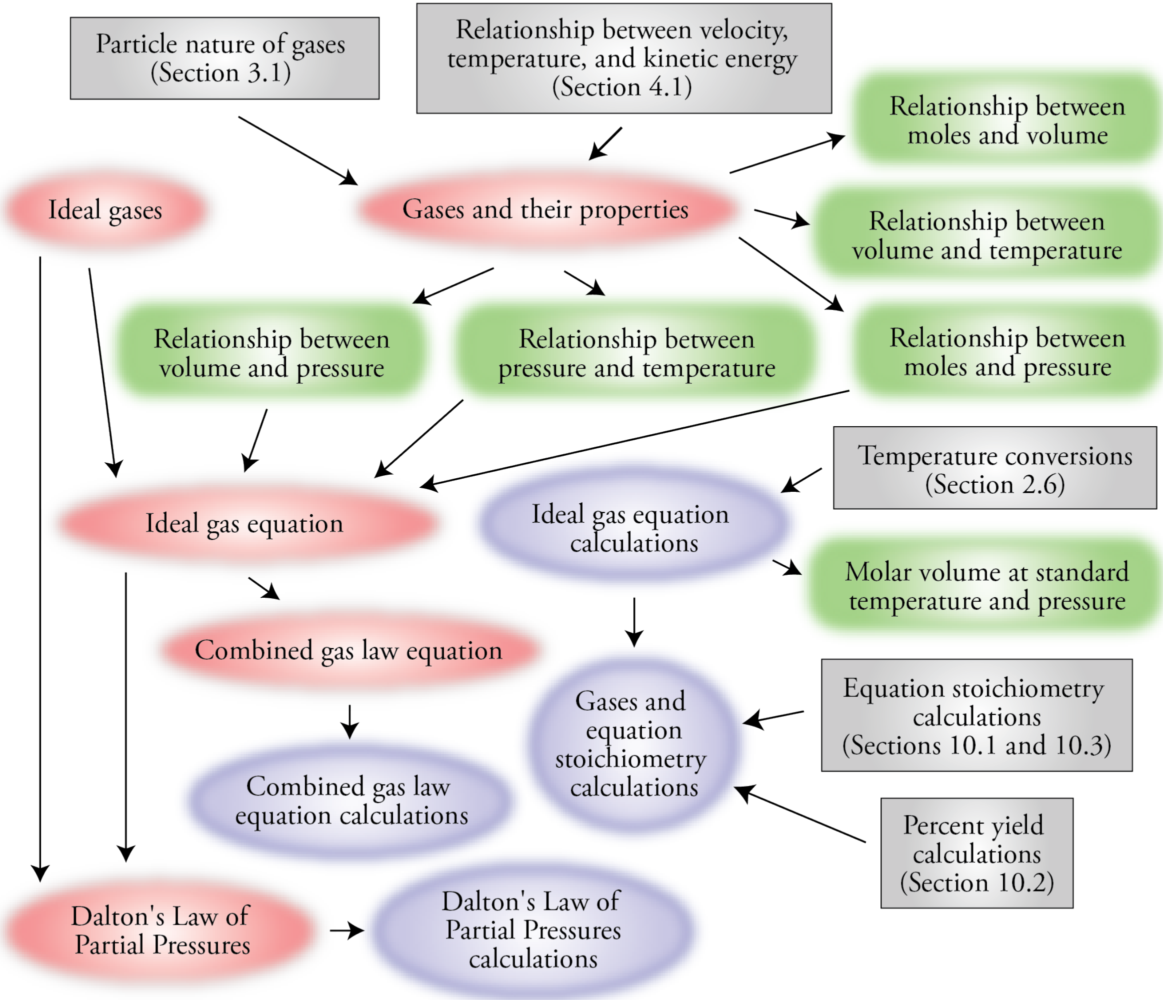
Chapter 11
Gases

An Introduction to Chemistry

by Mark Bishop

Chapter Map



Particle nature of gases (Section 3.1)

Relationship between velocity, temperature, and kinetic energy (Section 4.1)

Relationship between moles and volume

Ideal gases

Gases and their properties

Relationship between volume and temperature

Relationship between volume and pressure

Relationship between pressure and temperature

Relationship between moles and pressure

Temperature conversions (Section 2.6)

Ideal gas equation

Ideal gas equation calculations

Molar volume at standard temperature and pressure

Combined gas law equation

Gases and equation stoichiometry calculations

Equation stoichiometry calculations (Sections 10.1 and 10.3)

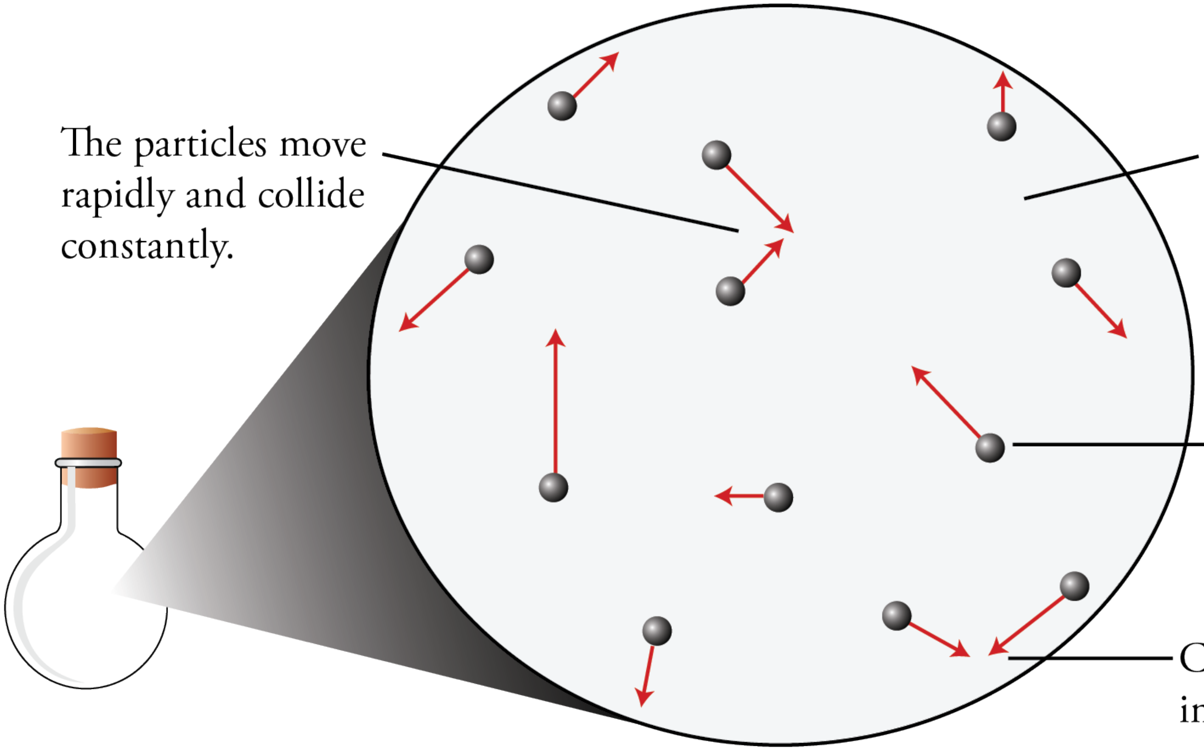
Combined gas law equation calculations

Percent yield calculations (Section 10.2)

Dalton's Law of Partial Pressures

Dalton's Law of Partial Pressures calculations

Gas



The particles move rapidly and collide constantly.

Particles occupy a small part of the total volume.

Little mutual attraction or repulsion between particles

Collisions cause changes in direction and velocity.

Gas Model

- Gases are composed of tiny, widely-spaced particles.
  - For a typical gas, the average distance between particles is about ten times their diameter.

Gas Model (cont.)

- Because of the large distance between the particles, the volume occupied by the particles themselves is negligible (approximately zero).
    - For a typical gas at room temperature and pressure, the gas particles themselves occupy about 0.1% of the total volume. The other 99.9% of the total volume is empty space. This is very different than for a liquid for which about 70% of the volume is occupied by particles.

Gas Model (cont.)

- The particles have rapid and continuous motion.
  - For example, the average velocity of a helium atom, He, at room temperature is over 1000 m/s (or over 2000 mi/hr). The average velocity of the more massive nitrogen molecules, N_2, at room temperature is about 500 m/s.
  - Increased temperature means increased average velocity of the particles.

Gas Model (cont.)

- The particles are constantly colliding with the walls of the container and with each other.
  - In a typical situation, a gas particle moves a very short distance between collisions. Oxygen, O_2, molecules at normal temperatures and pressures move an average of 10^{-7} m between collisions.

Gas Model (cont.)

- There is no net loss of energy in the collisions. A collision between two particles may lead to each particle changing its velocity and thus its energy, but the increase in energy by one particle is balanced by an equal decrease in energy by the other particle.

Ideal Gas Assumptions

- The particles are assumed to be point-masses, that is, particles that have a mass but occupy no volume.
- There are no attractive or repulsive forces at all between the particles.

Gas Pressure

- Pressure (*P*) = Force/Area
- Gas particles are constantly colliding with the walls of their container, and each time a particle hits a wall, it exerts a force against the wall.
- This force is proportional to its momentum, which is mass times velocity.
  - A more massive particle moving at a given velocity will exert more force against the wall.
  - A faster moving particle exerts more force against the wall than a slower moving particle of the same mass.

$$\text{Gas pressure} = \frac{\text{Force due to particle collisions with the walls}}{\text{Area of the walls}}$$

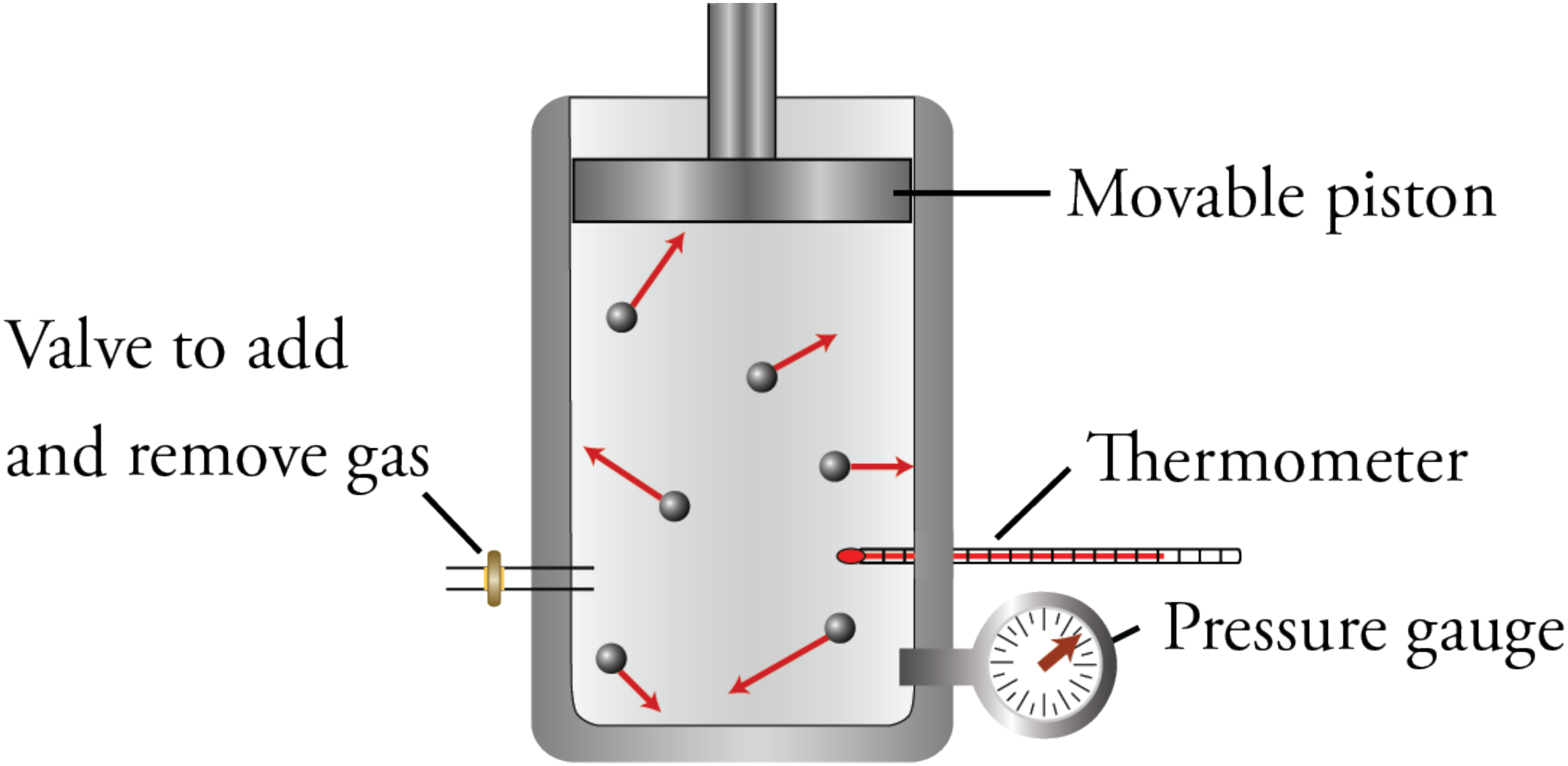
Gas Properties and their Units

- Pressure (P) = Force/Area
  - units
    - 1 atm = 101.325 kPa = 760 mmHg = 760 torr
    - 1 bar = 100 kPa = 0.9869 atm = 750.1 mmHg
- Volume (V)
  - unit usually liters (L)
- Temperature (T)
  - ? K = --- °C + 273.15
- Number of gas particles expressed in moles (n)

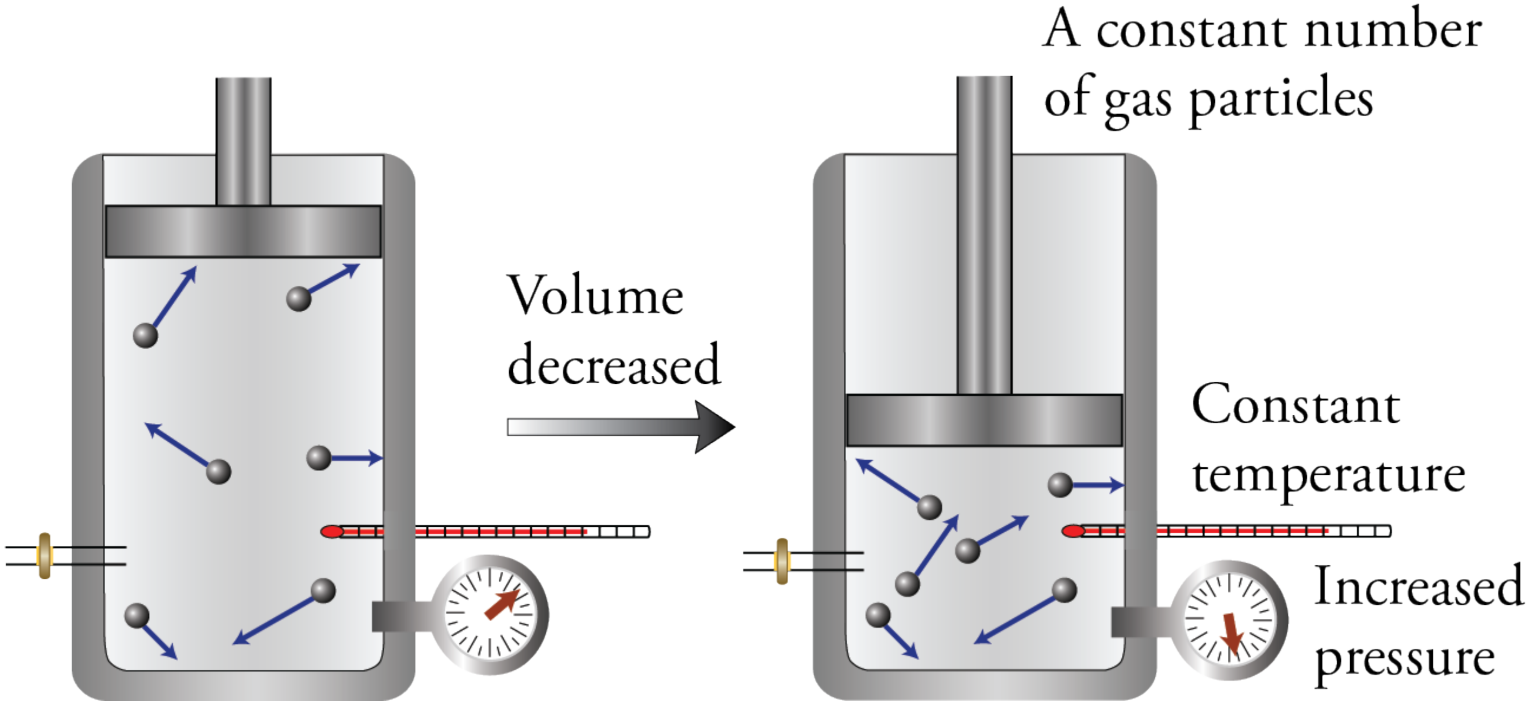
Gas Law Objectives

- For each of the following pairs of gas properties, (1) describe the relationship between them, (2) describe a simple system that could be used to demonstrate the relationship, and (3) explain the reason for the relationship.
    - V and P when n and T are constant
    - P and T when n and V are constant
    - V and T when n and P are constant
    - n and P when V and T are constant
    - n and V when P and T are constant

Apparatus for Demonstrating Relationships Between Properties of Gases



Decreased Volume Leads to Increased Pressure



$P \propto 1/V$ if n and T are constant

https://preparatorychemistry.com/Boyles_Law_Canvas.html

Relationship between P and V

Decreased volume

↓

$$\text{Increased } \frac{\text{number of gas particles}}{\text{volume of container}}$$

↓

Increased number of particles close to any area of wall

↓

$$\text{Increased } \frac{\text{number of collisions per second}}{\text{area of wall}}$$

↓

$$\text{Increased } \frac{\text{force due to collisions}}{\text{area of wall}}$$

↓

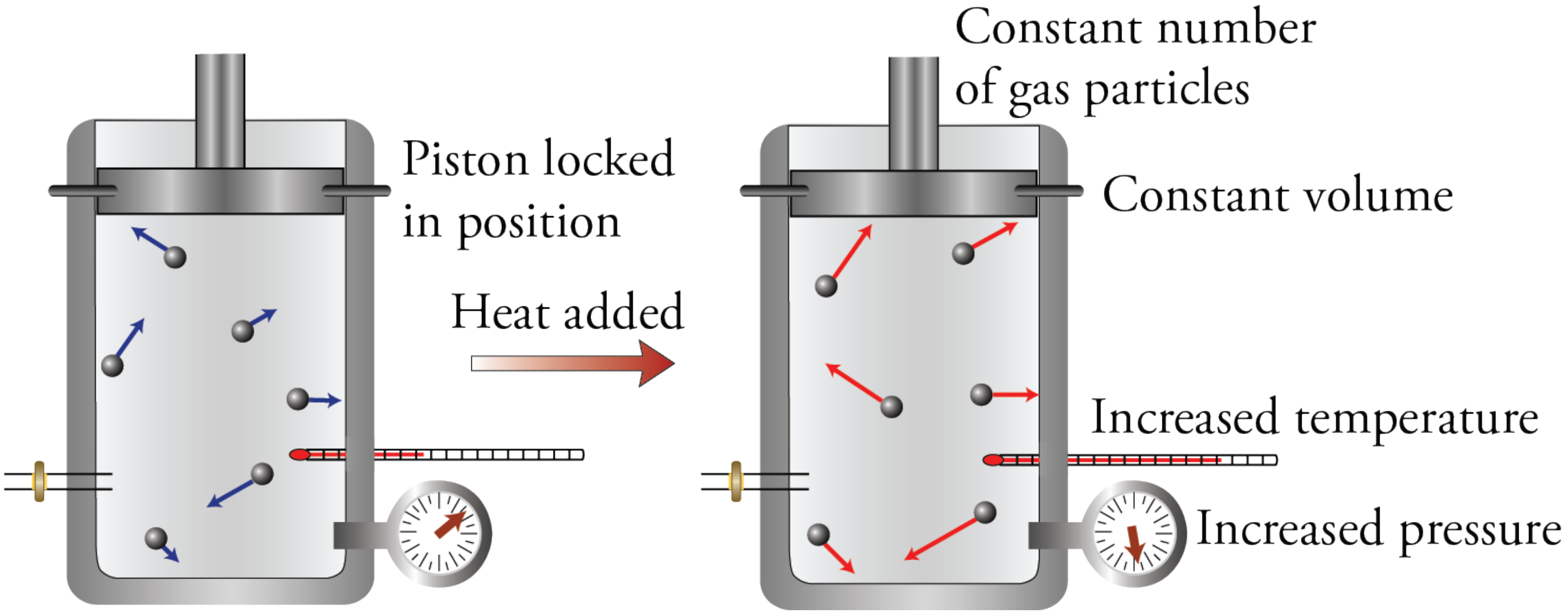
Increased gas pressure

Boyle’s Law

- The pressure of an ideal gas is inversely proportional to the volume it occupies if the moles of gas and the temperature are constant.

$$P \propto \frac{1}{V} \qquad \text{if } n \text{ and } T \text{ are constant}$$

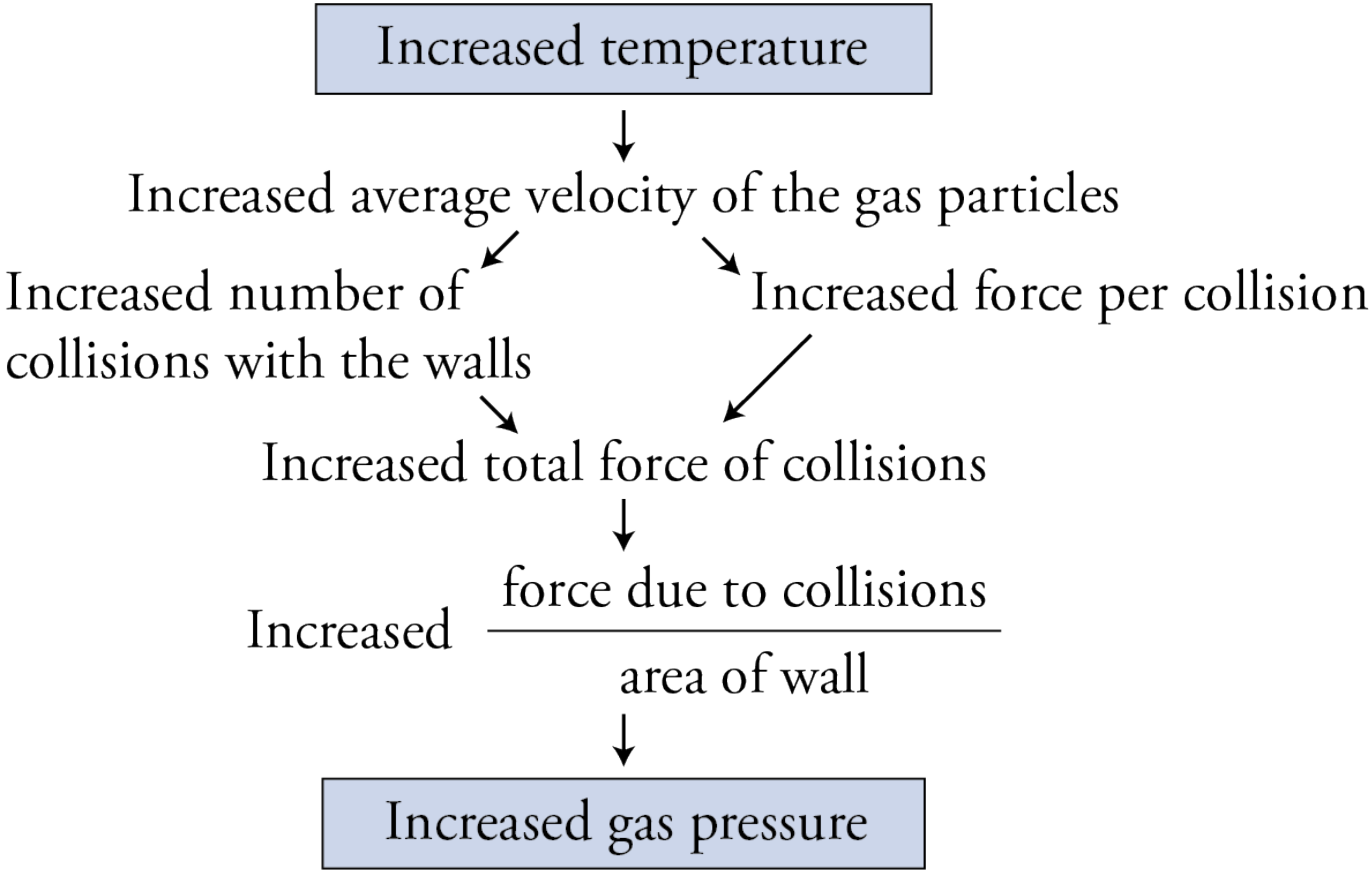
Increased Temperature Leads to Increased Pressure



$P \propto T$ if n and V are constant

https://preparatorychemistry.com/Gay_Lussacs_Law_Canvas.html

Relationship between P and T



Increased temperature

Increased average velocity of the gas particles

Increased number of collisions with the walls

Increased force per collision

Increased total force of collisions

Increased $\frac{\text{force due to collisions}}{\text{area of wall}}$

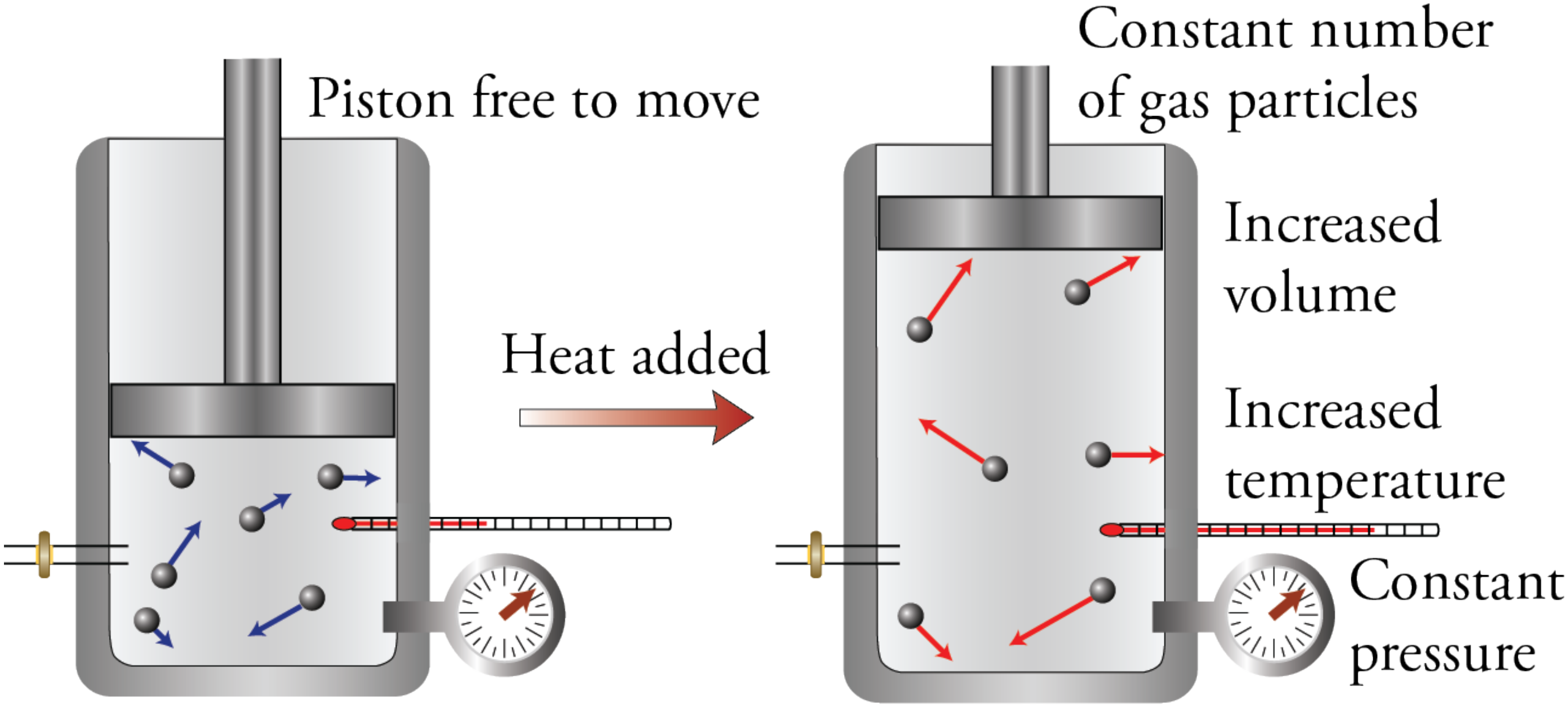
Increased gas pressure

Gay-Lussac’s Law

- The pressure of an ideal gas is directly proportional to the Kelvin temperature of the gas if the volume and moles of gas are constant.

$P \propto T$ if V and n are constant

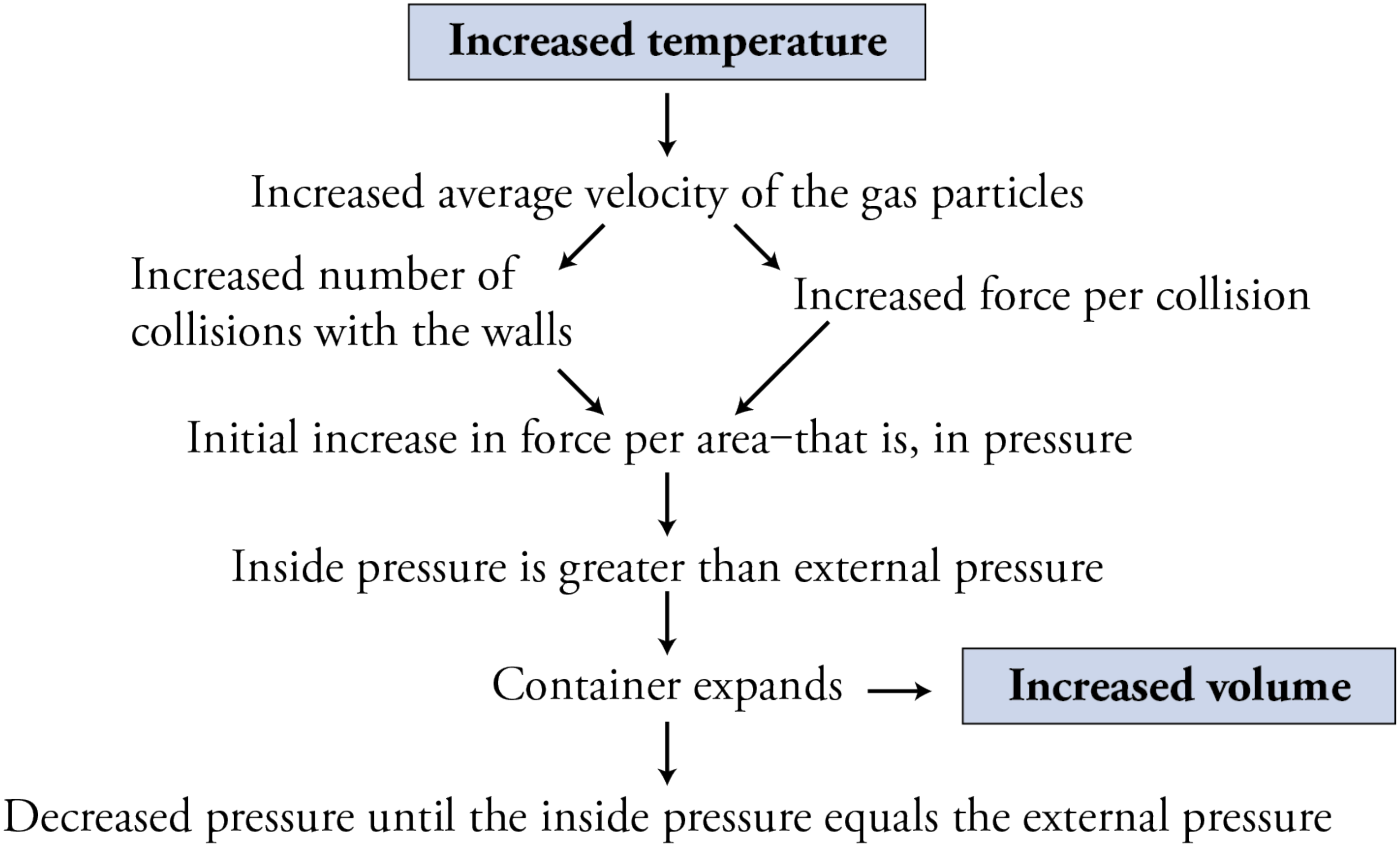
Increased Temperature Leads to Increased Volume



$V \propto T$ if n and P are constant

https://preparatorychemistry.com/Charles_Law_Canvas.html

Relationship between T and V



Increased temperature

↓

Increased average velocity of the gas particles

↙ ↘

Increased number of collisions with the walls

Increased force per collision

↘ ↙

Initial increase in force per area–that is, in pressure

↓

Inside pressure is greater than external pressure

↓

Container expands → **Increased volume**

↓

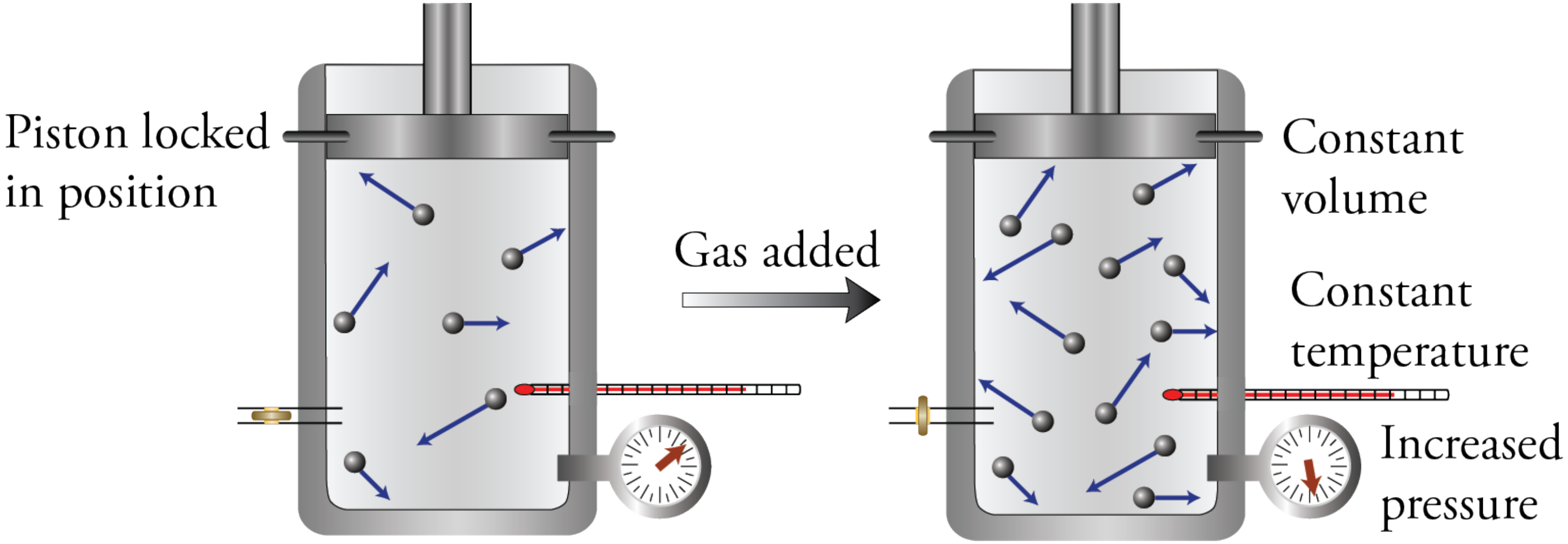
Decreased pressure until the inside pressure equals the external pressure

Charles’ Law

- For an ideal gas, volume and temperature described in kelvins are directly proportional if moles of gas and pressure are constant.

$V \propto T$ if n and P are constant

Increased Moles of Gas Leads to Increased Pressure



$P \propto n$ if T and V are constant

https://preparatorychemistry.com/Moles_Pressure_Law_Canvas.html

Relationship between n and P

Increased number of gas particles

↓

Increased number of collisions with the walls

↓

Increased total force of collisions

↓

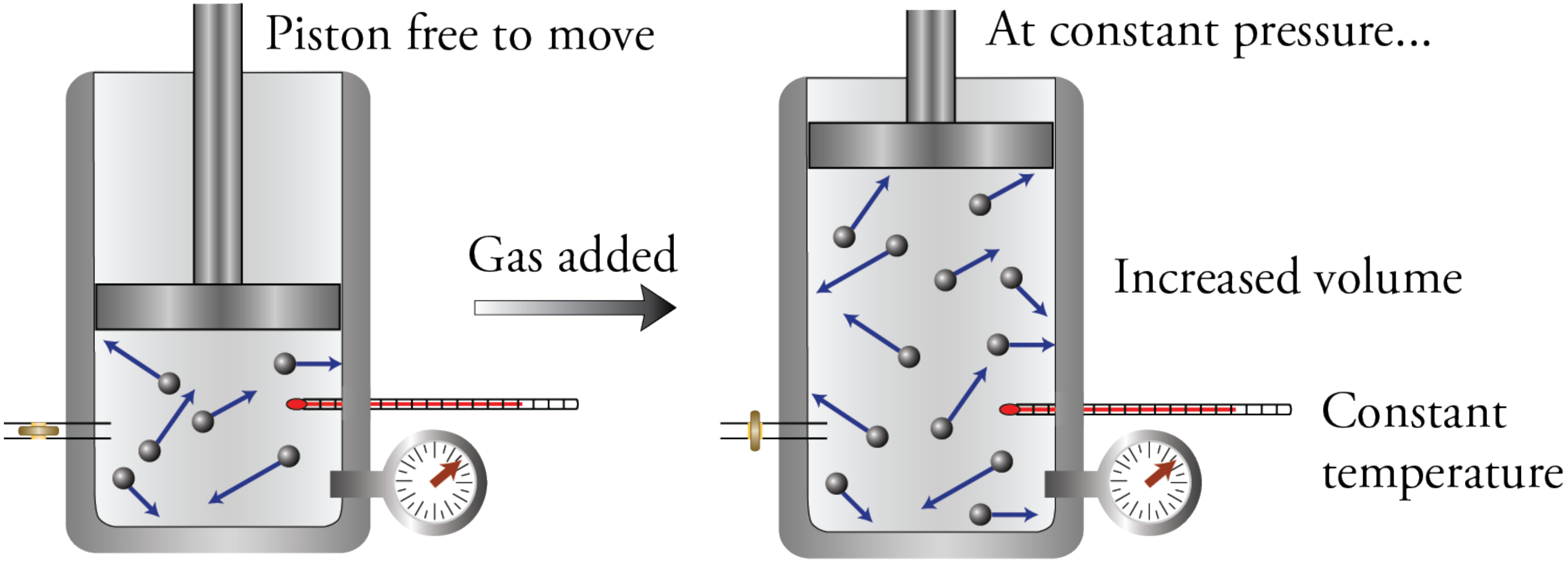
Increased gas pressure

Relationship Between Moles of Gas and Pressure

- If the temperature and the volume of an ideal gas are held constant, the moles of gas in a container and the gas pressure are directly proportional.

$P \propto n$ if T and V are constant

Increased Moles of Gas Leads to Increased Volume



V ∝ n   if T and P are constant

https://preparatorychemistry.com/Avogadros_Law_Canvas.html

Relationship between n and V

Increased number of gas particles

↓

Increased number of collisions with the walls

↓

Increased total force of collisions

↓

Initial increased in force per area - that is, in pressure

↓

Inside pressure is greater than external pressure

↓

Container expands ⟶ Increased volume

↓

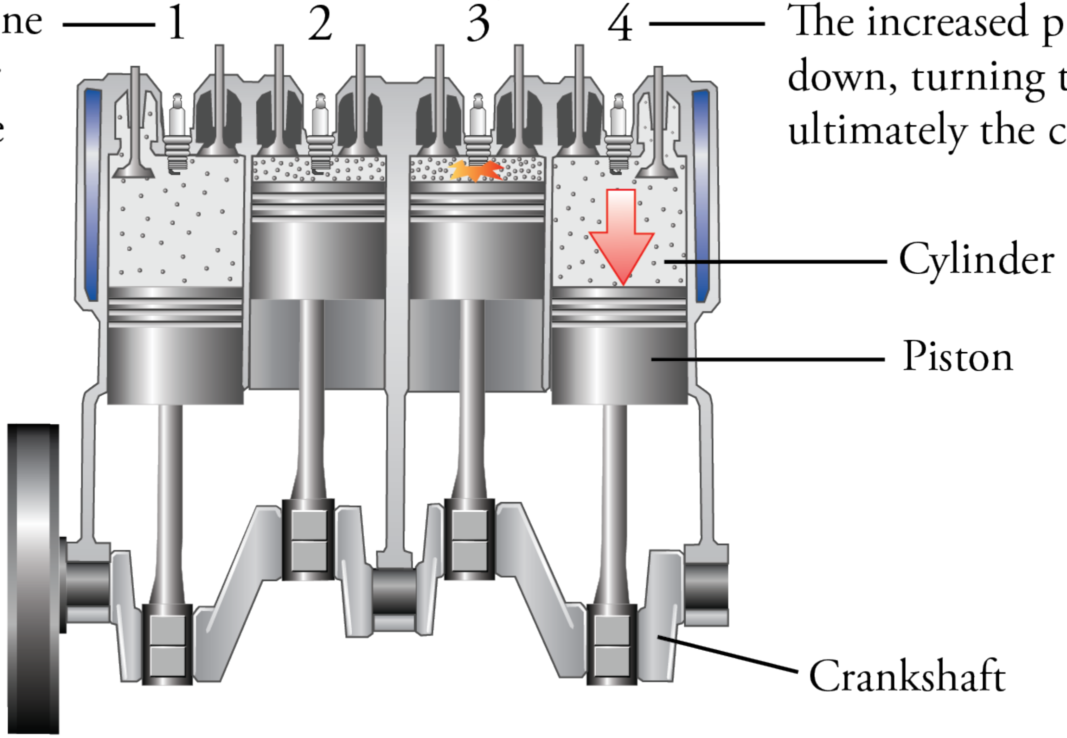
Decreased pressure until the inside pressure equals the external pressure

Avogadro’ s Law

- For an ideal gas, the volume and moles of gas are directly proportional if the temperature and pressure are constant.

$V \propto n$ if T and P are constant

Engines and Pressure



Gaseous gasoline mixed with air moves into the cylinder.

Decreased volume leads to increased gas pressure inside the cylinder.

The combustion of the gasoline leads to an increase in moles of gas, which also causes the gas pressure to increase.

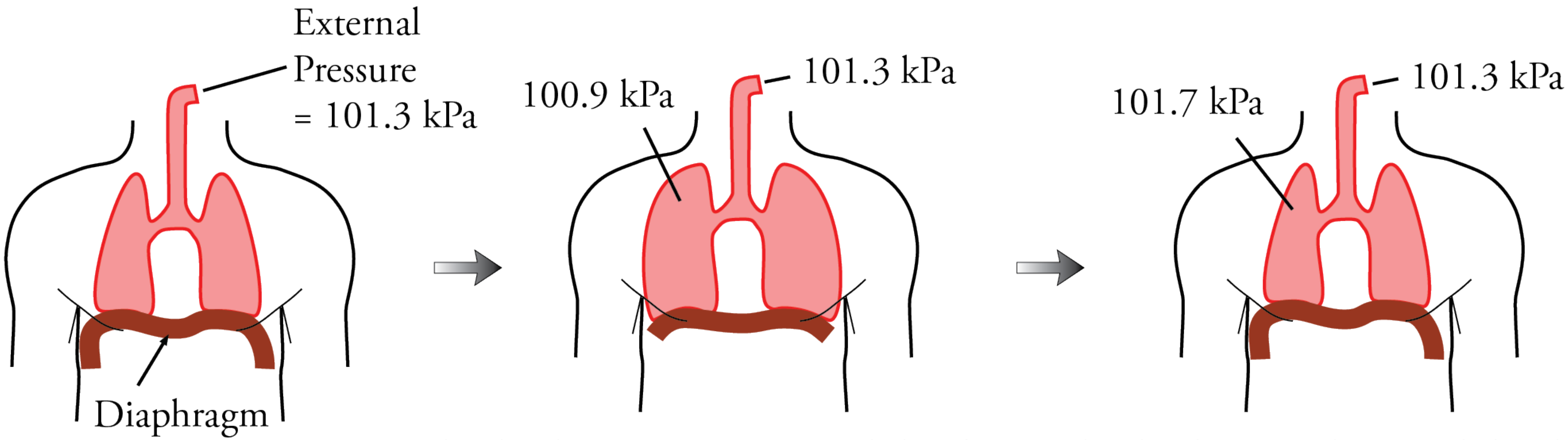
The reaction is exothermic, so the temperature of the product gases increases, contributing to the increased gas pressure.

The increased pressure pushes the piston down, turning the crankshaft and ultimately the car's wheels.

Sample reaction: $2C_8H_{18}(g) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g) + \text{energy}$

27 moles of gas → 34 moles of gas

Breathing



The diaphragm contracts, and the chest expands, causing the lungs to expand. The increased volume decreases the pressure in the lungs to below the external pressure, causing air to move into the lungs faster than it moves out.

The diaphragm relaxes and the chest returns to its original volume, causing the lung volume to decrease.
This increases the pressure in the lungs, causing air to move out of the lungs faster than it moves in.

Four types of Ideal Gas Problems

- Ideal gas equation problems.
- Combined gas equation problems
- Gas stoichiometry problems
- Dalton’s Law of partial pressures problems

Algebraic Approach to Problem Solving

- Write down the values you are given, assigning variables to each.
- Write the appropriate equation.
- Solve for the variable of your unknown.
- Plug in values given, including their units.
- Cancel units and if necessary, do unit conversions to get the units to cancel.
- Do the calculation and report your answer with the correct significant figures and unit.

Unit Check

- If your units cancel to yield the correct unit or units, you know that you…
  - used the correct equation,
  - did the algebra correctly,
  - and did all the necessary unit conversions.

Ideal Gas Equation Derivation

$P \propto n$ if T and V are constant

$P \propto T$ if n and V are constant

$P \propto \frac{1}{V}$ if n and T are constant

$$P \propto \frac{nT}{V}$$

$$P = (\text{a constant})\,\frac{nT}{V}$$

$$PV = n(\text{a constant})\,T$$

$$\boldsymbol{PV = nRT} \qquad R = \frac{0.082058\ \text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}} \text{ or } \frac{8.3145\ \text{L} \cdot \text{kPa}}{\text{K} \cdot \text{mol}}$$

Expanded Ideal Gas Equation

$$PV = nRT \qquad R = \frac{0.082058\ \text{L} \bullet \text{atm}}{\text{K} \bullet \text{mol}} \ \text{ or } \ \frac{8.3145\ \text{L} \bullet \text{kPa}}{\text{K} \bullet \text{mol}}$$

$$n = \text{moles} = \text{---}\ \text{g}\left(\frac{1\ \text{mol}}{\text{---}\ \cancel{\text{g}}}\right) = \frac{\text{grams}}{\dfrac{\text{grams}}{\text{mole}}} = \frac{\text{mass in grams}}{\text{molar mass}} = \frac{g}{M}$$

$$PV = \frac{g}{M}RT \qquad g = \text{mass} \qquad M = \text{molar mass}$$

Ideal Gas Equation Problems

- **Tip-off** – The usual tip-off that you can use the ideal gas equation to answer a question is that you are given three properties of a sample of gas and asked to calculate the fourth. A more general tip-off is that only one gas is mentioned, there's not chemical reaction mentioned, and there are no changing properties.

Ideal Gas Equation Problem Step 1

- **Step 1:** Assign variables to the values given and the value that is unknown. Use P for pressure, V for volume, n for moles, T for temperature, g for mass, and M for molar mass.

Ideal Gas Equation Problem Step 2

- **Step 2:** Write the appropriate form of the Ideal Gas Equation.
  - If the number of particles is given or desired in moles, use the most common form of the ideal gas equation.

$$PV = nRT \quad R = \frac{0.082058\ \mathrm{L \cdot atm}}{\mathrm{K \cdot mol}} \text{ or } \frac{8.3145\ \mathrm{L \cdot kPa}}{\mathrm{K \cdot mol}}$$

  - If mass or molar mass is given or desired, use the expanded form of the ideal gas equation.

$$PV = \frac{g}{M}RT \quad g = \text{mass} \quad M = \text{molar mass}$$

Ideal Gas Equation Problem Steps 3-6

- **Step 3:** Rearrange the equation to isolate the unknown.
- **Step 4:** Plug in the known values, including units. Be sure to use kelvin temperatures.
- **Step 5:** Make any necessary unit conversions and cancel your units.
- **Step 6:** Calculate your answer and report it to the correct significant figures and with the correct unit.

Example

- Most incandescent light bulbs contain argon, but krypton gas does a better job than argon of slowing the evaporation of the tungsten filament in the light bulb. Because of its higher cost, however, krypton is only used when longer life is considered to be worth the extra expense. How many moles of krypton gas must be added to a 175-mL incandescent light bulb to yield a gas pressure of 117 kPa at 21.6 ºC?

Example 1

How many moles of krypton gas must be added to a 175-mL incandescent light bulb to yield a gas pressure of 117 kPa at 21.6 ºC?

n = ? V = 175 mL   P = 117 kPa

T = 21.6 ºC + 273.15 = 294.8 K

$PV = nRT$

$$n = \frac{PV}{RT} = \frac{117\ \cancel{\text{kPa}}\ (175\ \cancel{\text{mL}})}{\left(\dfrac{8.3145\ \cancel{\text{L}}\cdot\cancel{\text{kPa}}}{\cancel{\text{K}}\cdot\text{mol}}\right) 294.8\ \cancel{\text{K}}} \left(\frac{1\ \cancel{\text{L}}}{10^3\ \cancel{\text{mL}}}\right)$$

$$= \mathbf{8.35 \times 10^{-3}\ mol}$$

Example 2

What is the volume of an incandescent light bulb that contains 1.196 g Kr at a pressure of 1.70 atm and a temperature of 97 ºC?

$V = ?$ $g = 1.196$ g P = 1.70 atm

$T = 97$ ºC + 273.15 = 370 K

$$PV = \frac{g}{M}RT \qquad g = \text{mass} \qquad M = \text{molar mass}$$

$$V = \frac{gRT}{PM} = \frac{1.196\,\cancel{\text{g}}\left(\dfrac{0.082058\ \text{L}\cdot\cancel{\text{atm}}}{\cancel{\text{K}}\cdot\cancel{\text{mol}}}\right)370\ \cancel{\text{K}}}{(1.70\ \cancel{\text{atm}})\ 83.80\,\dfrac{\cancel{\text{g}}}{\cancel{\text{mol}}}} = \mathbf{0.255\ L\ Kr}$$

Example 3

What is the density of krypton gas at 18.2 ºC and 762 mmHg?

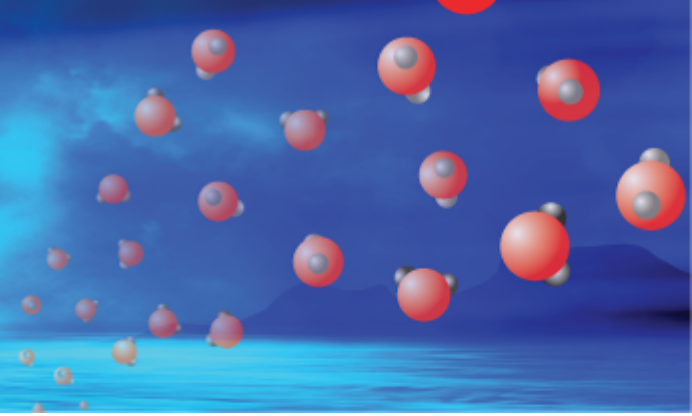
$$\frac{g}{V} = ? \quad T = 18.2\ ^{\circ}\text{C} + 273.15 = 291.4\text{ K} \quad P = 762\text{ mmHg}$$

$$PV = \frac{g}{M}RT \qquad g = \text{mass} \qquad M = \text{molar mass}$$

$$\frac{g}{V} = \frac{PM}{RT} = \frac{762\ \cancel{\text{mmHg}}\left(\frac{83.80\text{ g}}{1\ \cancel{\text{mol}}}\right)}{\left(\frac{0.082058\text{ L}\cdot\cancel{\text{atm}}}{\cancel{\text{K}}\cdot\cancel{\text{mol}}}\right)291.4\ \cancel{\text{K}}}\left(\frac{1\ \cancel{\text{atm}}}{760\ \cancel{\text{mmHg}}}\right)$$

$$= 3.51\text{ g/L}$$

Combined Gas Law Equation Derivation



$$PV = nRT \qquad \frac{PV}{nT} = R$$

$$P_1V_1 = n_1RT_1 \quad \text{so} \quad \frac{P_1V_1}{n_1T_1} = R$$

$$P_2V_2 = n_2RT_2 \quad \text{so} \quad \frac{P_2V_2}{n_2T_2} = R$$

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$$

Combined Gas Law Equation Problems

- **Tip-off –** The problem requires calculating a value for a gas property that has changed. In other words, you are asked to calculate a new pressure, temperature, moles, or volume of gas given sufficient information about the initial and final of the other properties. A more general tip-off is that only one gas is mentioned, there’s no chemical reaction mentioned, and there are changing properties.

Combined Gas Law Equation Problem Steps 1 and 2

- **Step 1:** Assign the variables P, T, n, and V to the values you are given and to the unknown value. Use the subscripts 1 and 2 to show initial or final conditions.
- **Step 2:** Write out the combined gas law equation, and eliminate the variables for any constant properties. (You can assume that the properties not mentioned in the problem remain constant.)

Combined Gas Law Equation Problem Steps 3-6

- **Step 3:** Rearrange the equation to isolate the unknown property.
- **Step 4:** Plug in the values for the given properties.
- **Step 5:** Make any necessary unit conversions and cancel your units.
- **Step 6:** Calculate your answer and report it with the correct units and significant figures.

Example 4

- A helium weather balloon is filled in Monterey, California, on a day when the atmospheric pressure is 102 kPa and the temperature is 18 ºC. Its volume under these conditions is 1.6×10^4 L. Upon being released, it rises to an altitude where the temperature is -8.6 ºC, and its volume increases to 4.7×10^4 L. Assuming that the internal pressure of the balloon equals the atmospheric pressure, what is the pressure at this altitude?

Example 4

A helium weather balloon is filled in Monterey, California, on a day when the atmospheric pressure is 102 kPa and the temperature is 18 ºC. Its volume under these conditions is 1.6 × 10^4 L. Upon being released, it rises to an altitude where the temperature is -8.6 ºC, and its volume increases to 4.7 × 10^4 L. Assuming that the internal pressure of the balloon equals the atmospheric pressure, what is the pressure at this altitude?

P_1 = 102 kPa T_1 = 18 ºC + 273.15 = 291 K V_1 = 1.6 × 10^4 L

P_2 = ? T_2 = -8.6 ºC + 273.15 = 264.6 K V_1 = 4.7 × 10^4 L

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2} \qquad \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \qquad P_2 = \frac{P_1 V_1 T_2}{V_2 T_1}$$

$$P_2 = P_1 \left(\frac{T_2}{T_1}\right)\left(\frac{V_1}{V_2}\right) = 103 \text{ kPa} \left(\frac{264.6 \cancel{\text{K}}}{291 \cancel{\text{K}}}\right)\left(\frac{1.6{\times}10^4 \cancel{\text{L}}}{4.7{\times}10^4 \cancel{\text{L}}}\right) = \mathbf{32\ kPa}$$

Measurable Property and Moles

Measurable property	⇌	Moles
Mass for (s) and (l)	Molar mass ⇌	Moles of pure substance
Volume of solution for (aq)	Molarity ⇌	Moles of solute
Volume of gas for (g)	3 ways ⇌	Moles of gas

Conversion between moles and volume of gas

- Using molar volume at STP (only for STP, which is rare)
- Using the Ideal Gas Equation

$$n = \frac{PV}{RT} \qquad V = \frac{nRT}{P}$$

- R as a conversion factor

$$\frac{\mathrm{K{\bullet}mol}}{8.3145\ \mathrm{L{\bullet}kPa}} \quad \text{or} \quad \frac{8.3145\ \mathrm{L{\bullet}kPa}}{\mathrm{K{\bullet}mol}}$$

$$\frac{\mathrm{K{\bullet}mol}}{0.082058\ \mathrm{L{\bullet}atm}} \quad \text{or} \quad \frac{0.082058\ \mathrm{L{\bullet}atm}}{\mathrm{K{\bullet}mol}}$$

Standard Temperature and Pressure (STP) and Molar Volume

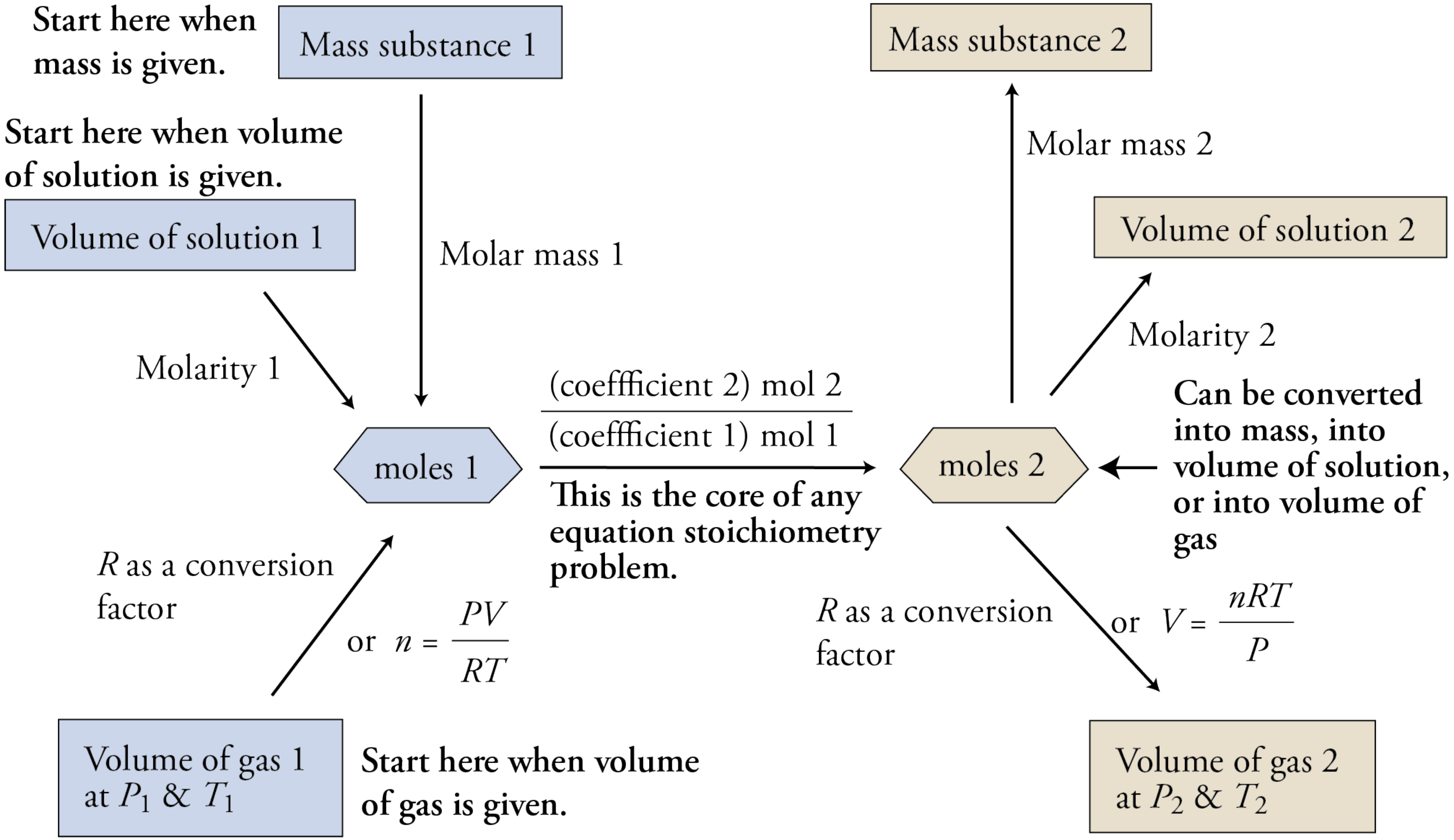
- **Common:** 0 ºC (273.15 K) and 1 atm

$$\frac{V}{n} = \frac{RT}{P} = \frac{\left(\dfrac{8.3145\ \text{L}\cdot\cancel{\text{kPa}}}{\cancel{\text{K}}\cdot\text{mol}}\right)(273.15\ \cancel{\text{K}})}{101.325\ \cancel{\text{kPa}}} = \left(\frac{22.414\ \text{L}}{1\ \text{mol}}\right)_{\text{STP}}$$

- **Correct:** 0 ºC (273.15 K) and 1 bar (100 kPa)

$$\frac{V}{n} = \frac{RT}{P} = \frac{\left(\dfrac{8.3145\ \text{L}\cdot\cancel{\text{kPa}}}{\cancel{\text{K}}\cdot\text{mol}}\right)(273.15\ \cancel{\text{K}})}{100\ \cancel{\text{kPa}}} = \left(\frac{22.711\ \text{L}}{1\ \text{mol}}\right)_{\text{STP}}$$

Equation Stoichiometry



Start here when mass is given.

Mass substance 1

Molar mass 1

Start here when volume of solution is given.

Volume of solution 1

Molarity 1

moles 1

$$\frac{(\text{coeffficient 2})\ \text{mol 2}}{(\text{coeffficient 1})\ \text{mol 1}}$$

This is the core of any equation stoichiometry problem.

moles 2

R as a conversion factor

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

Volume of gas 1 at P_1 & T_1

Start here when volume of gas is given.

Mass substance 2

Molar mass 2

Volume of solution 2

Molarity 2

Can be converted into mass, into volume of solution, or into volume of gas

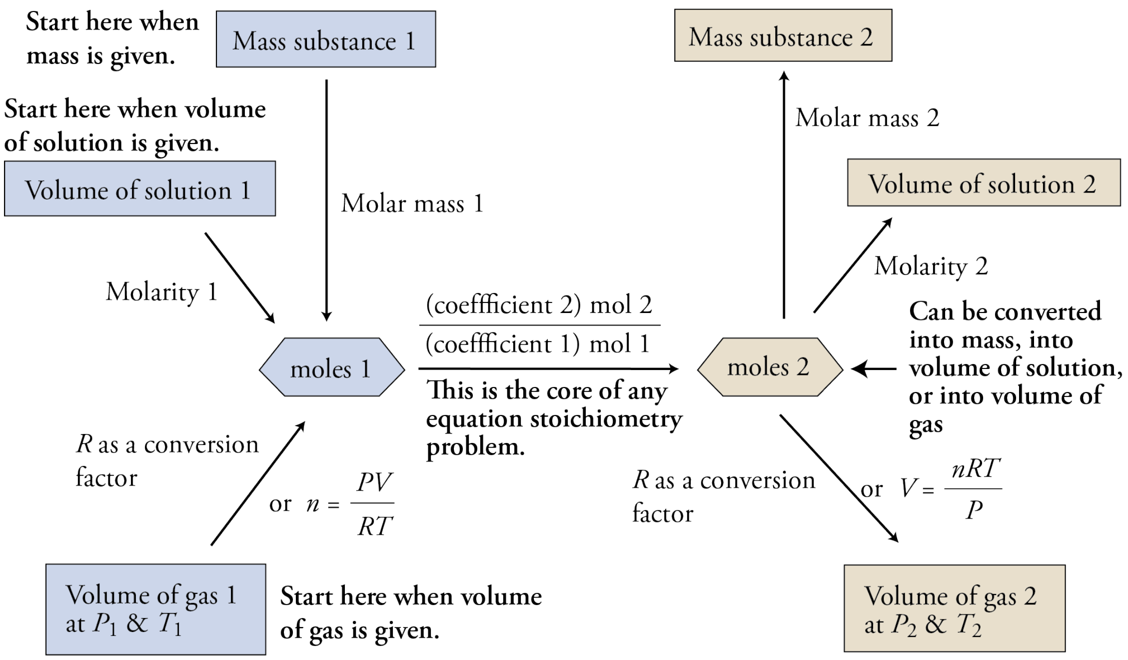
R as a conversion factor

or $V = \frac{nRT}{P}$

Volume of gas 2 at P_2 & T_2

Example 1

Iron is combined with carbon in a series of reactions to form pig iron, which is about 4.3% carbon. The first step in this process is the reaction of carbon with oxygen to form carbon monoxide. For this reaction, what is the maximum volume of carbon monoxide at 105 kPa and 35 ºC that could form from the conversion of 8.74×10^5 L of oxygen at 99.4 kPa and 27 ºC?



Example 1

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- Conversion from units of one substance to units of another substance, both involved in a chemical equation, so it’s equation stoichiometry.
- Write a balanced equation.

$$2C + O_2 \rightarrow 2CO$$

Example 1

For this reaction, what is the maximum volume of carbon monoxide at 105 kPa and 35 ºC that could form from the conversion of 8.74×10^5 L of oxygen at 99.4 kPa and 27 ºC?

$$2C + O_2 \rightarrow 2CO$$

- Convert from the units of substance 1 you have to moles substance 1.

$$n_{O_2} = \frac{PV}{RT} = \frac{99.4\ \cancel{\text{kPa}}\,(8.74 \times 10^5\ \cancel{\text{L}})}{\left(\dfrac{8.3145\ \cancel{\text{L}}\bullet\cancel{\text{kPa}}}{\cancel{\text{K}}\bullet\text{mol}}\right)(300\ \cancel{\text{K}})} = 3.48 \times 10^4\ \text{mol}\ O_2$$

Example 1

For this reaction, what is the maximum volume of carbon monoxide at 105 kPa and 35 ºC that could form from the conversion of 8.74×10^5 L of oxygen at 99.4 kPa and 27 ºC?

$$2C + O_2 \rightarrow 2CO$$

$$n_{O_2} = \frac{PV}{RT} = \frac{99.4\ \cancel{\text{kPa}}\,(8.74 \times 10^5\ \cancel{\text{L}})}{\left(\dfrac{8.3145\ \cancel{\text{L}}\bullet\cancel{\text{kPa}}}{\cancel{\text{K}}\bullet\text{mol}}\right)(300\ \cancel{\text{K}})} = 3.48 \times 10^4\ \text{mol } O_2$$

- Convert from moles of substance 1 to moles substance 2.

$$?\ \text{mol CO} = 3.48 \times 10^4\ \cancel{\text{mol } O_2}\left(\frac{2\ \text{mol CO}}{1\ \cancel{\text{mol } O_2}}\right) = 6.96 \times 10^4\ \text{mol CO}$$

Example 1

For this reaction, what is the maximum volume of carbon monoxide at 105 kPa and 35 ºC that could form from the conversion of 8.74×10^5 L of oxygen at 99.4 kPa and 27 ºC?

$$2C + O_2 \rightarrow 2CO$$

$$? \text{ mol CO} = 3.48 \times 10^4 \cancel{\text{mol } O_2} \left(\frac{2 \text{ mol CO}}{1 \cancel{\text{mol } O_2}} \right) = 6.96 \times 10^4 \text{ mol CO}$$

- Convert from the moles of substance 2 to units of substance 2 that you want.

$$V_{CO} = \frac{nRT}{P} = \frac{6.96 \times 10^4 \cancel{\text{mol}} \left(\frac{8.3145 \text{ L} \cdot \cancel{\text{kPa}}}{\cancel{\text{K}} \cdot \cancel{\text{mol}}} \right) 308 \cancel{\text{K}}}{105 \cancel{\text{kPa}}} = \mathbf{1.70 \times 10^6 \text{ L CO}}$$

Example 1

For this reaction, what is the maximum volume of carbon monoxide at 105 kPa and 35 ºC that could form from the conversion of 8.74 × 10^5 L of oxygen at 99.4 kPa and 27 ºC?

$$2C + O_2 \rightarrow 2CO$$

- Or unit analysis, using R as a conversion factor.

$$?\ \mathrm{L\ CO} = 8.74 \times 10^5\ \mathrm{L\ O_2}$$

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$$?\ \mathrm{L\ CO} = 8.74 \times 10^5\ \cancel{\mathrm{L}}\ \mathrm{O_2}\left(\frac{\cancel{\mathrm{K}} \bullet \mathrm{mol}}{8.3145\ \cancel{\mathrm{L}} \bullet \cancel{\mathrm{kPa}}}\right)\left(\frac{99.4\ \cancel{\mathrm{kPa}}}{300\ \cancel{\mathrm{K}}}\right)$$

$$?\ \mathrm{L\ CO} = 8.74 \times 10^5\ \cancel{\mathrm{L}}\ \cancel{\mathrm{O_2}}\left(\frac{\cancel{\mathrm{K}} \bullet \cancel{\mathrm{mol}}}{8.3145\ \cancel{\mathrm{L}} \bullet \cancel{\mathrm{kPa}}}\right)\left(\frac{99.4\ \cancel{\mathrm{kPa}}}{300\ \cancel{\mathrm{K}}}\right)\left(\frac{2\ \mathrm{mol\ CO}}{1\ \cancel{\mathrm{mol}}\ \cancel{\mathrm{O_2}}}\right)$$

Example 1

For this reaction, what is the maximum volume of carbon monoxide at 105 kPa and 35 ºC that could form from the conversion of 8.74×10^5 L of oxygen at 99.4 kPa and 27 ºC?

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$$?\ \mathrm{L\ CO} = 8.74 \times 10^5\ \cancel{\mathrm{L}}\ \cancel{\mathrm{O_2}} \left(\frac{\cancel{\mathrm{K}} \cdot \cancel{\mathrm{mol}}}{8.3145\ \cancel{\mathrm{L}} \cdot \cancel{\mathrm{kPa}}}\right)\left(\frac{99.4\ \cancel{\mathrm{kPa}}}{300\ \cancel{\mathrm{K}}}\right)\left(\frac{2\ \cancel{\mathrm{mol}}\ \mathrm{CO}}{1\ \cancel{\mathrm{mol}}\ \cancel{\mathrm{O_2}}}\right)\left(\frac{8.3145\ \mathrm{L} \cdot \cancel{\mathrm{kPa}}}{\cancel{\mathrm{K}} \cdot \cancel{\mathrm{mol}}}\right)\left(\frac{308\ \cancel{\mathrm{K}}}{105\ \cancel{\mathrm{kPa}}}\right)$$

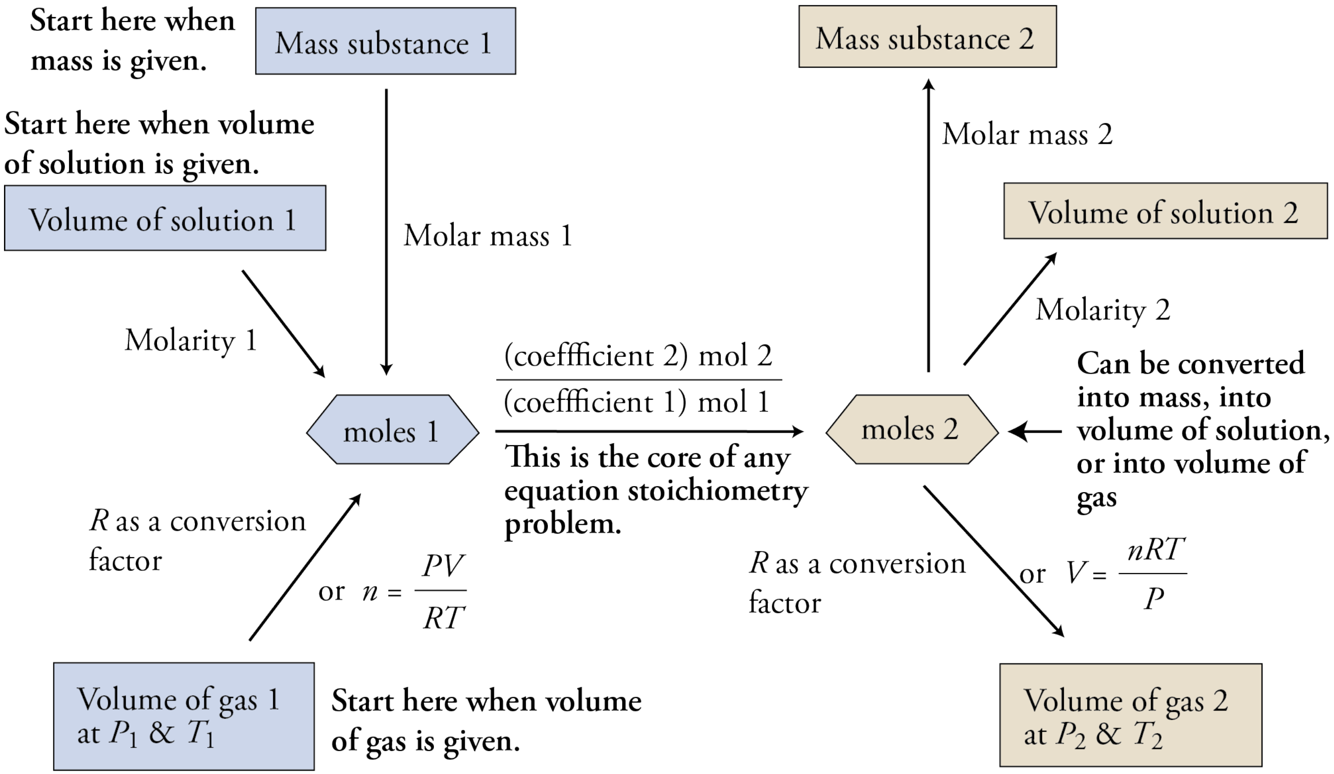
$$?\ \mathrm{L\ CO} = 8.74 \times 10^5\ \cancel{\mathrm{L}}\ \cancel{\mathrm{O_2}} \left(\frac{\cancel{\mathrm{K}} \cdot \cancel{\mathrm{mol}}}{8.3145\ \cancel{\mathrm{L}} \cdot \cancel{\mathrm{kPa}}}\right)\left(\frac{99.4\ \cancel{\mathrm{kPa}}}{300\ \cancel{\mathrm{K}}}\right)\left(\frac{2\ \cancel{\mathrm{mol}}\ \mathrm{CO}}{1\ \cancel{\mathrm{mol}}\ \cancel{\mathrm{O_2}}}\right)\left(\frac{8.3145\ \mathrm{L} \cdot \cancel{\mathrm{kPa}}}{\cancel{\mathrm{K}} \cdot \cancel{\mathrm{mol}}}\right)\left(\frac{308\ \cancel{\mathrm{K}}}{105\ \cancel{\mathrm{kPa}}}\right)$$

$$= \mathbf{1.70 \times 10^6\ L\ CO}$$

Example 2

In the reaction of carbon with oxygen to form carbon monoxide, what minimum volume of oxygen at STP (273.15 K and 1 bar) is necessary to convert 125 Mg of carbon to carbon monoxide?

$$2C + O_2 \rightarrow 2CO$$



Start here when mass is given.

Mass substance 1

Start here when volume of solution is given.

Volume of solution 1

Molar mass 1

Molarity 1

moles 1

$\frac{\text{(coefficient 2) mol 2}}{\text{(coefficient 1) mol 1}}$

This is the core of any equation stoichiometry problem.

R as a conversion factor

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

Volume of gas 1 at P_1 & T_1

Start here when volume of gas is given.

Mass substance 2

Molar mass 2

Volume of solution 2

Molarity 2

moles 2

Can be converted into mass, into volume of solution, or into volume of gas

R as a conversion factor

or $V = \frac{nRT}{P}$

Volume of gas 2 at P_2 & T_2

Example 2

- In the reaction of carbon with oxygen to form carbon monoxide, what minimum volume of oxygen at STP (273.15 K and 1 bar) is necessary to convert 125 Mg of carbon to carbon monoxide?

$$2C + O_2 \rightarrow 2CO$$

$$?\ \mathrm{L\ O_2} = 125\ \cancel{\mathrm{Mg}}\ \cancel{\mathrm{C}} \left(\frac{10^6\ \cancel{\mathrm{g}}}{1\ \cancel{\mathrm{Mg}}}\right)\left(\frac{1\ \cancel{\mathrm{mol\ C}}}{12.011\ \cancel{\mathrm{g}}\ \cancel{\mathrm{C}}}\right)\left(\frac{1\ \cancel{\mathrm{mol\ O_2}}}{2\ \cancel{\mathrm{mol\ C}}}\right)\left(\frac{22.711\ \mathrm{L\ O_2}}{1\ \cancel{\mathrm{mol\ O_2}}}\right)_{\mathrm{STP}}$$

$$= \mathbf{1.18\times10^5\ L\ O_2}$$

$$?\ \mathrm{L\ O_2} = 125\ \cancel{\mathrm{Mg}}\ \cancel{\mathrm{C}} \left(\frac{10^6\ \cancel{\mathrm{g}}}{1\ \cancel{\mathrm{Mg}}}\right)\left(\frac{1\ \cancel{\mathrm{mol\ C}}}{12.011\ \cancel{\mathrm{g}}\ \cancel{\mathrm{C}}}\right)\left(\frac{1\ \cancel{\mathrm{mol}}\ \mathrm{O_2}}{2\ \cancel{\mathrm{mol\ C}}}\right)\left(\frac{8.3145\ \mathrm{L}\bullet\cancel{\mathrm{kPa}}}{\cancel{\mathrm{K}}\bullet\cancel{\mathrm{mol}}}\right)\left(\frac{273.15\ \cancel{\mathrm{K}}}{1\ \cancel{\mathrm{bar}}}\right)\left(\frac{1\ \cancel{\mathrm{bar}}}{100\ \cancel{\mathrm{kPa}}}\right)$$

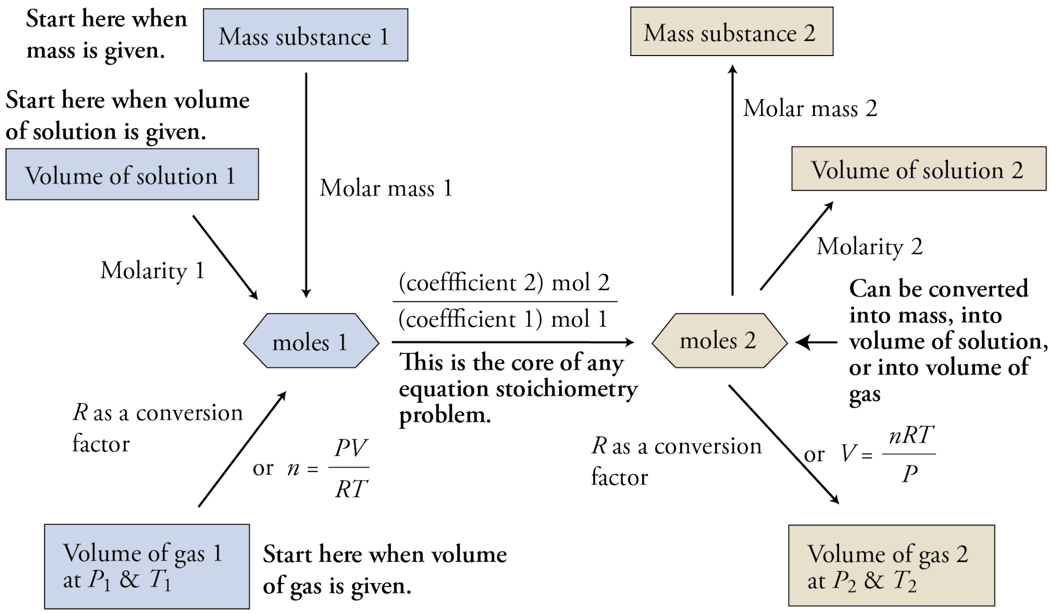
$$= \mathbf{1.18\times10^5\ L\ O_2}$$

Example 3

Sodium hypochlorite, NaOCl, found in household bleaches, can be made from a reaction using chlorine gas and aqueous sodium hydroxide:

$$Cl_2(g) + 2NaOH(aq) \rightarrow NaOCl(aq) + NaCl(aq) + H_2O(l)$$

What minimum volume of chlorine gas at 101.4 kPa and 18.0 °C must be used to react with all the sodium hydroxide in 3525 L of 12.5 M NaOH?



Start here when mass is given.

Mass substance 1

Start here when volume of solution is given.

Volume of solution 1

Molar mass 1

Molarity 1

moles 1

$\frac{\text{(coeffficient 2) mol 2}}{\text{(coeffficient 1) mol 1}}$

This is the core of any equation stoichiometry problem.

R as a conversion factor

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

Volume of gas 1 at P_1 & T_1

Start here when volume of gas is given.

Mass substance 2

Molar mass 2

Volume of solution 2

Molarity 2

moles 2

Can be converted into mass, into volume of solution, or into volume of gas

R as a conversion factor

or $V = \frac{nRT}{P}$

Volume of gas 2 at P_2 & T_2

Example 3

$$Cl_2(g) + 2NaOH(aq) \rightarrow NaOCl(aq) + NaCl(aq) + H_2O(l)$$

What minimum volume of chlorine gas at 101.4 kPa and 18.0 ºC must be used to react with all the sodium hydroxide in 3525 L of 12.5 M NaOH?

$$? \text{ L } Cl_2 = 3525 \cancel{\text{L NaOH soln}} \left(\frac{\qquad\qquad}{1 \cancel{\text{L NaOH soln}}} \right)$$

$$? \text{ L } Cl_2 = 3525 \cancel{\text{L NaOH soln}} \left(\frac{12.5 \text{ mol NaOH}}{1 \cancel{\text{L NaOH soln}}} \right)$$

$$? \text{ L } Cl_2 = 3525 \cancel{\text{L NaOH soln}} \left(\frac{12.5 \cancel{\text{mol NaOH}}}{1 \cancel{\text{L NaOH soln}}} \right) \left(\frac{1 \text{ mol } Cl_2}{2 \cancel{\text{mol NaOH}}} \right)$$

Example 3

$$Cl_2(g) + 2NaOH(aq) \rightarrow NaOCl(aq) + NaCl(aq) + H_2O(l)$$

What minimum volume of chlorine gas at 101.4 kPa and 18.0 ºC must be used to react with all the sodium hydroxide in 3525 L of 12.5 M NaOH?

$$?\ \mathrm{L\ Cl_2} = 3525\ \cancel{\mathrm{L\ NaOH\ soln}} \left(\frac{12.5\ \cancel{\mathrm{mol\ NaOH}}}{1\ \cancel{\mathrm{L\ NaOH\ soln}}}\right)\left(\frac{1\ \cancel{\mathrm{mol}}\ \mathrm{Cl_2}}{2\ \cancel{\mathrm{mol\ NaOH}}}\right)\left(\frac{8.3145\ \mathrm{L{\bullet}kPa}}{\mathrm{K{\bullet}}\cancel{\mathrm{mol}}}\right)$$

$$?\ \mathrm{L\ Cl_2} = 3525\ \cancel{\mathrm{L\ NaOH\ soln}} \left(\frac{12.5\ \cancel{\mathrm{mol\ NaOH}}}{1\ \cancel{\mathrm{L\ NaOH\ soln}}}\right)\left(\frac{1\ \cancel{\mathrm{mol}}\ \mathrm{Cl_2}}{2\ \cancel{\mathrm{mol\ NaOH}}}\right)\left(\frac{8.3145\ \mathrm{L{\bullet}}\cancel{\mathrm{kPa}}}{\cancel{\mathrm{K}}{\bullet}\cancel{\mathrm{mol}}}\right)\left(\frac{291.0\ \cancel{\mathrm{K}}}{101.4\ \cancel{\mathrm{kPa}}}\right)$$

$$?\ \mathrm{L\ Cl_2} = 3525\ \cancel{\mathrm{L\ NaOH\ soln}} \left(\frac{12.5\ \cancel{\mathrm{mol\ NaOH}}}{1\ \cancel{\mathrm{L\ NaOH\ soln}}}\right)\left(\frac{1\ \cancel{\mathrm{mol}}\ \mathrm{Cl_2}}{2\ \cancel{\mathrm{mol\ NaOH}}}\right)\left(\frac{8.3145\ \mathrm{L{\bullet}}\cancel{\mathrm{kPa}}}{\cancel{\mathrm{K}}{\bullet}\cancel{\mathrm{mol}}}\right)\left(\frac{291.0\ \cancel{\mathrm{K}}}{101.4\ \cancel{\mathrm{kPa}}}\right)$$

$$= 5.26\times10^5\ \mathrm{L\ Cl_2}$$

Gas Pressure

- Gas pressure is the total force due to particle collisions with the walls of the container at an instant in time divided by the area of the walls.
- The total force of collisions is determined by
  - The number of collisions
  - The average force per collision

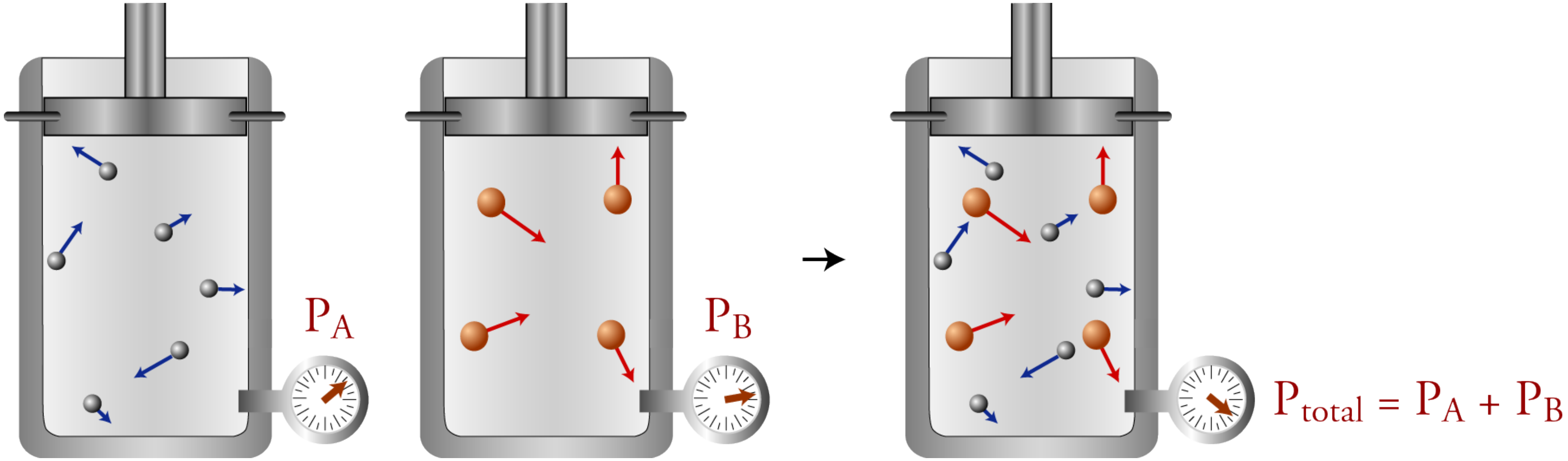
Force per Collision

- The force per collision is proportional to the momentum of the colliding particle.
- Momentum is mass times velocity.
- At a constant temperature, the mass and average velocity are constant, so the average momentum is constant, so the average force per collision is constant.

Ideal Gas Assumptions

- The particles are assumed to be point-masses, that is, particles that have a mass but occupy no volume.
- There are no attractive or repulsive forces at all between the particles.

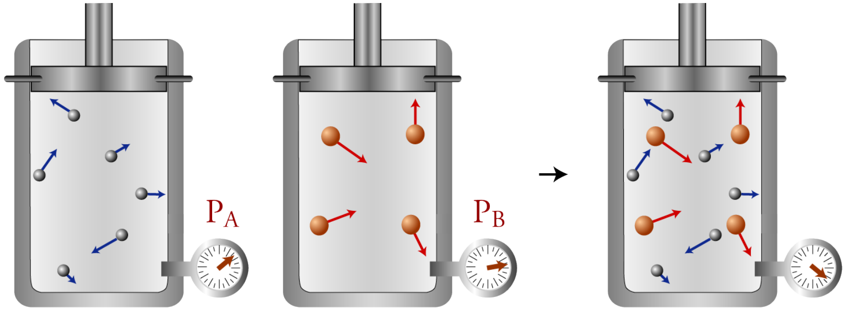
Mixture of Gases



- Partial pressure can be defined as the portion of the total pressure that one gas in a mixture of gases contributes.
- For ideal gases, this is the pressure that a gas would have if it were alone in the container.

Dalton’s Law of Partial Pressures

Dalton’s Law of Partial Pressures states that the total pressure of a mixture of gases is equal to the sum of the partial pressures of all the gases.



$$P_{\mathrm{total}} = P_{\mathrm{A}} + P_{\mathrm{B}}$$

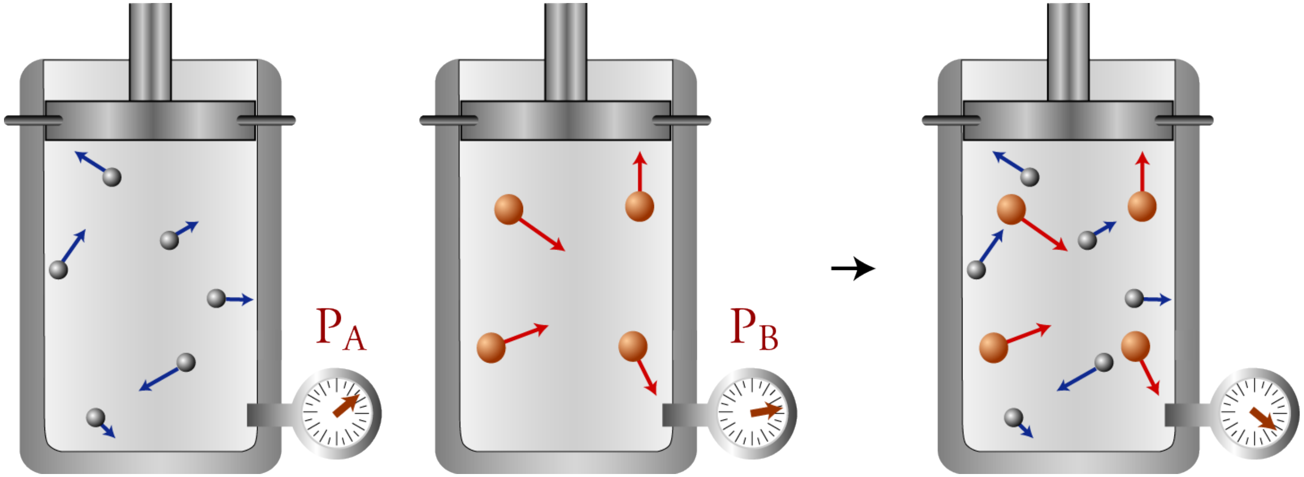
$$P_{\mathrm{A}} = \frac{n_{\mathrm{A}} R T_{\mathrm{A}}}{V_{\mathrm{A}}} \qquad P_{\mathrm{B}} = \frac{n_{\mathrm{B}} R T_{\mathrm{B}}}{V_{\mathrm{B}}}$$

$$P_{\mathrm{total}} = \frac{n_{\mathrm{A}} R T_{\mathrm{A}}}{V_{\mathrm{A}}} + \frac{n_{\mathrm{B}} R T_{\mathrm{B}}}{V_{\mathrm{B}}}$$

$$V_{\mathrm{A}} + V_{\mathrm{B}} = V \qquad T_{\mathrm{A}} + T_{\mathrm{B}} = T$$

$$P_{\mathrm{total}} = \frac{n_{\mathrm{A}} R T}{V} + \frac{n_{\mathrm{B}} R T}{V} = (n_{\mathrm{A}} + n_{\mathrm{B}}) \frac{RT}{V}$$

Dalton’s Law of Partial Pressures



$P_{total} = P_A + P_B$

$$P_{total} = \Sigma P_{partial} \quad \text{or} \quad P_{total} = (\Sigma n_{\text{each gas}})\frac{RT}{V}$$

Dalton’s Law of Partial Pressures Problems

- **Tip-off –** The problem involves a mixture of gases and no chemical reaction. You are asked to calculate a value for one of the variables in the equations below, and you are given (directly or indirectly) values for the other variables.

Dalton’s Law of Partial Pressures Problem Steps 1 & 2

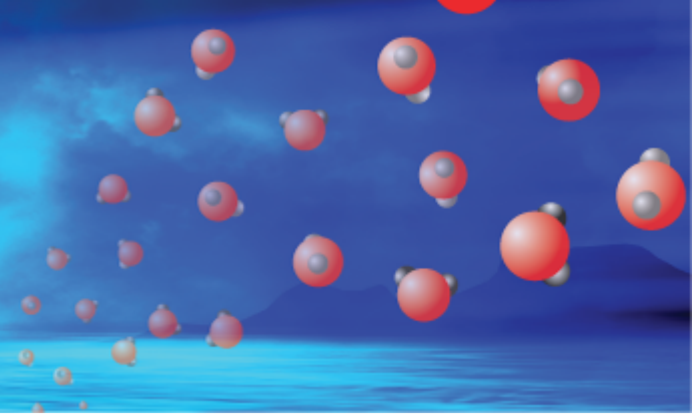
- **Step 1:** Assign variables to the values that are given and the value that is unknown.
- **Step 2:** From the following equations, choose the one that best fits the variables assigned in Step 1.

$$P_{\text{total}} = \Sigma P_{\text{partial}} \quad \text{or} \quad P_{\text{total}} = (\Sigma n_{\text{each gas}})\frac{RT}{V}$$

Dalton’s Law of Partial Pressures Problem Steps 3-6

- **Step 3:** Rearrange the equation to solve for your unknown.
- **Step 4:** Plug in the values for the given properties.
- **Step 5:** Make sure that the equation yields the correct units. Make any necessary unit conversions.
- **Step 6:** Calculate your answer and report it with the correct units and significant figures.

Example 1



A typical “neon light” contains neon gas mixed with argon gas. If the total pressure of the mixture of gases is 1.30 kPa and the partial pressure of neon gas is 0.27 kPa, what is the partial pressure of the argon gas?

P_T = 1.30 kPa P_{Ne} = 0.27 kPa P_{Ar} = ?

$$P_T = P_{Ne} + P_{Ar}$$

$$P_{Ar} = P_T - P_{Ne} = 1.30\ \text{kPa} - 0.27\ \text{kPa} = 1.03\ \text{kPa}$$

Example 2

If 6.3 mg of Ar and 1.2 mg Ne are added to a 375-mL tube at 291 K, what is the total pressure of the gases in kilopascals?

$V = 375\ \text{mL} \qquad T = 291\ \text{K} \qquad P = ?$

$$n_{\text{Ar}} = ?\ \text{mol Ar} = 6.3\ \cancel{\text{mg Ar}} \left(\frac{1\ \cancel{\text{g}}}{10^3\ \cancel{\text{mg}}}\right)\left(\frac{1\ \text{mol Ar}}{39.948\ \cancel{\text{g Ar}}}\right) = 0.00016\ \text{mol Ar}$$

$$n_{\text{Ne}} = ?\ \text{mol Ne} = 1.2\ \cancel{\text{mg Ne}} \left(\frac{1\ \cancel{\text{g}}}{10^3\ \cancel{\text{mg}}}\right)\left(\frac{1\ \text{mol Ne}}{20.1797\ \cancel{\text{g Ne}}}\right) = 0.000059\ \text{mol Ne}$$

$$P_{\text{total}} = (\Sigma n_{\text{each gas}})\frac{RT}{V} \qquad\qquad P_{\text{total}} = (n_{\text{Ar}} + n_{\text{N}_2})\frac{RT}{V}$$

$$P_{\text{total}} = (0.00016\ \cancel{\text{mol Ar}} + 0.000059\ \cancel{\text{mol Ne}}) \left(\frac{\left(\frac{8.3145\ \cancel{\text{L}}\cdot\text{kPa}}{\cancel{\text{K}}\cdot\cancel{\text{mol}}}\right)(291\ \cancel{\text{K}})}{375\ \cancel{\text{mL}}}\right)\left(\frac{10^3\ \cancel{\text{mL}}}{1\ \cancel{\text{L}}}\right)$$

$$= \mathbf{1.4\ kPa}$$

Tip-offs for Gas Problems

- **Ideal Gas Equation:** 1 gas, no chemical reaction, and no changing properties.
- **Combined Gas Law Equation:** 1 gas, no chemical reaction, changing properties.
- **Gas Stoichiometry:** converting from one substance to another, both in a chemical reaction, and one or both are gases.
- **Dalton's Law of Partial Pressures:** 2 or more gases, no changing properties, no chemical reaction.