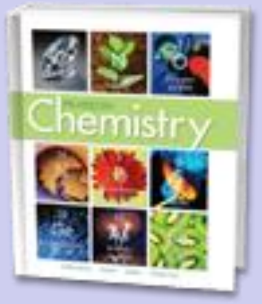


Chapter 14
The Behavior of Gases

Properties of Gases

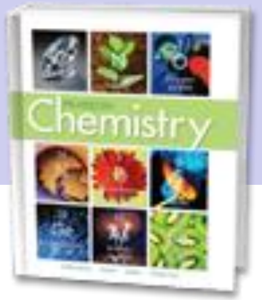
The Gas Laws

Ideal Gases

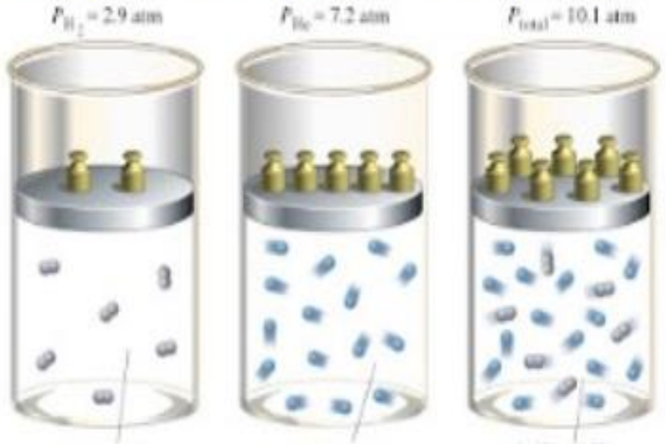
Lesson Objectives

By the end of this lesson, you should be able to:

- State Boyle's law, Charles's law, and Gay-Lussac's law and apply these laws to calculate the relationships between volume, temperature, and pressure.
- Derive the combined gas law from Boyle's law, Charles's law, and Gay-Lussac's law and calculate for pressure, volume, or temperature.
- Define partial pressure and apply Dalton's law of partial pressures to describe the composition of gases.
- *State the Ideal Gas Law ($PV = nRT$) and calculate moles of a gas, understanding the difference between ideal and real gases.*



Gas Laws Chapter 14B

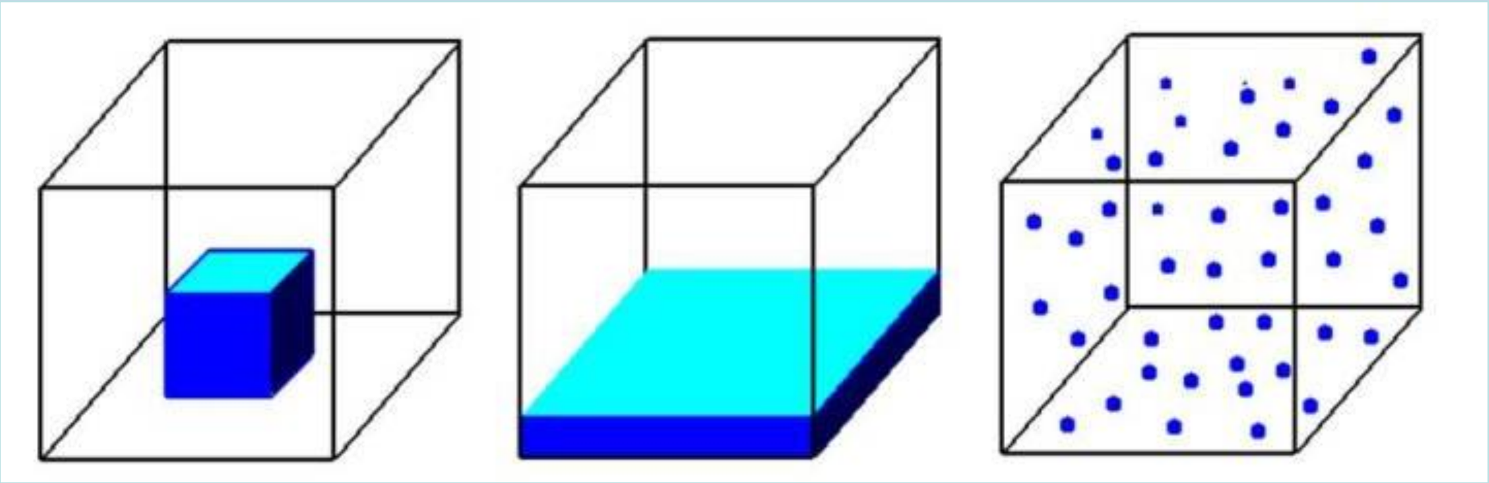


Topics:

1. The Behavior of Gases

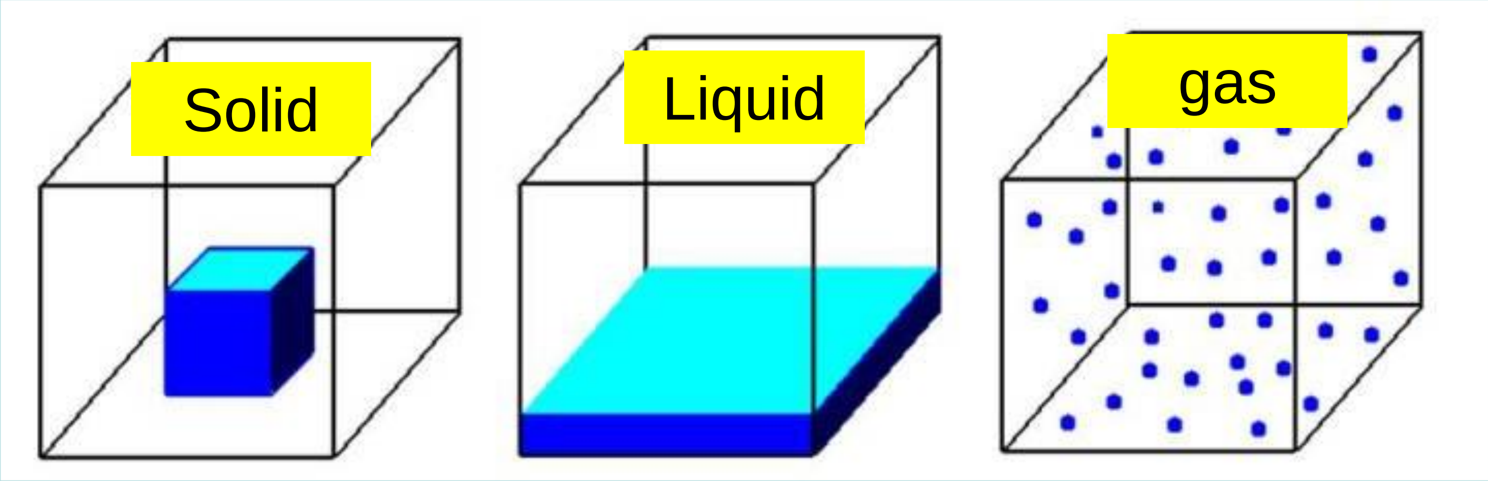
Objectives:

1. State Boyle's law, Charles's law, Gay-Lussac's law, and Avogadro's law and apply these laws to calculate the relationships between volume, temperature, moles, and pressure.
2. Derive the combined gas law from Boyle's law, Charles's law, and Gay-Lussac's law and calculate for pressure, volume, or temperature.
3. Define partial pressure and apply Dalton's law of partial pressures to describe the composition of gases.
4. State the Ideal Gas Law (PV = nRT) and calculate moles of a gas, understanding the difference between ideal and real gases.
5. Understand Graham's Law which states that gases of lower molar mass diffuse and effuse faster than gases of higher molar mass.



Label the three states of matter in the diagram & describe their shape, volume, density, molecular attractions, compressibility, expansivity, & what might affect it … in relation to each other.

	Shape	Volume	Density	Molecular Attractions	Compress-ibility	Expansiv-ity	What might affect it
Solid							
Liquid							
Gas							



Label the three states of matter in the diagram & describe their shape, volume, density, molecular attractions, compressibility, expansivity, & what might affect it … in relation to each other.

	Shape	Volume	Density	Molecular Attractions	Compress-ibility	Expansiv-ity	What might affect it
Solid	definite	definite	greatest	strongest	little	little	little
Liquid	Takes the shape container	definite	middle	middle	little	little	little
Gas	indefinite	fills container	least	weakest	great	great	T, P

Review of Pressure

Watch the video on Learning CTR Online:

https://screencast-o-matic.com/watch/cFeY3gDvx1

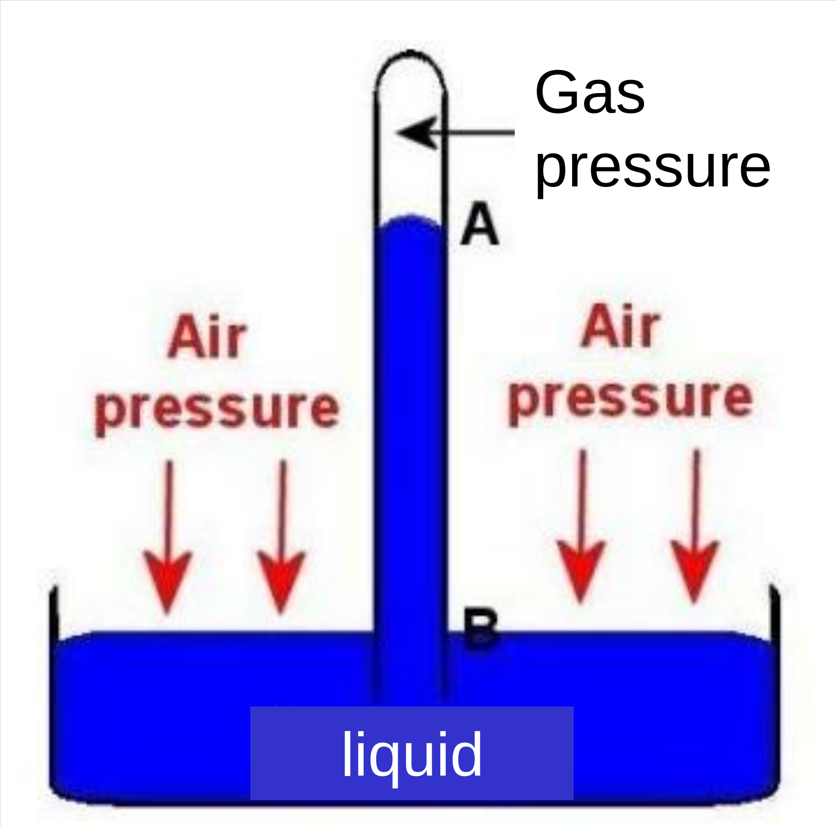
Pressure ctr (5:32)

- PSI (Breaking a Board using Atmospheric Pressure)
- Manometer Readings

Measuring Gas Pressure

Torricelli developed the "barometer" from *baro*, meaning *weight* and *meter* meaning *measure*.

He also coined the unit "Torr" which is equal to mm Hg.



What is pressure & its units?

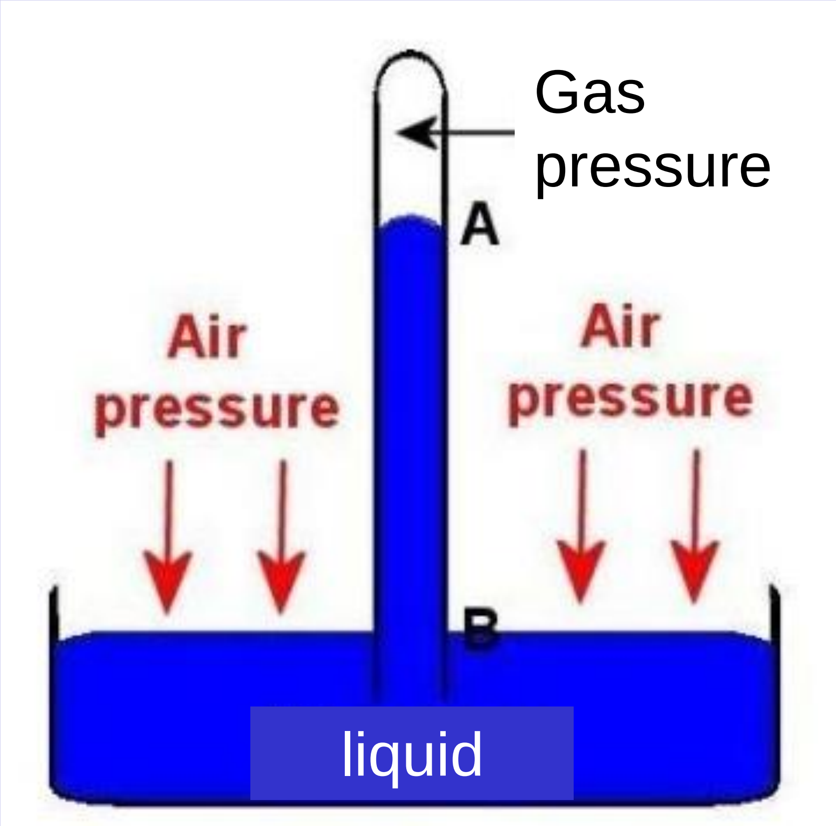
What is STP and its units?

Measuring Gas Pressure

WARM-UP

Torricelli developed the "barometer" from *baro*, meaning *weight* and *meter* meaning *measure*.

He also coined the unit "Torr" which is equal to mm Hg.



What is pressure & its units?

Force/Area → N/m^2

What is STP and its units?

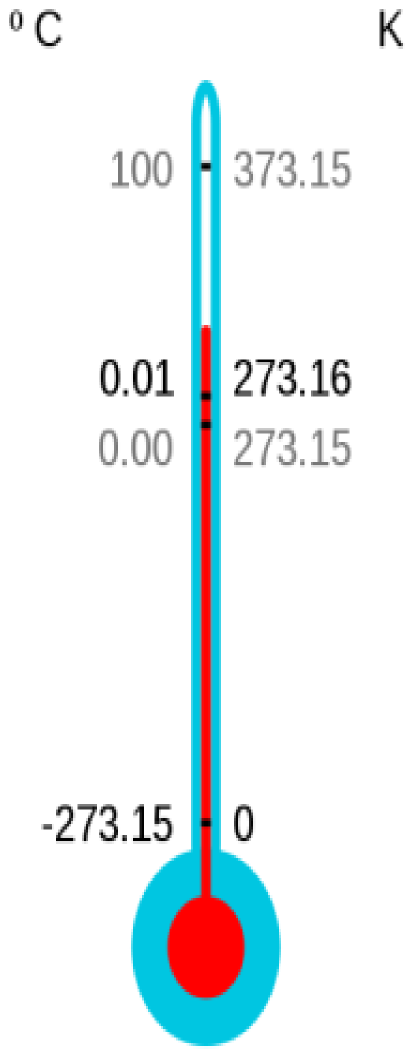
1 atm = 101.3 kPa = 760 mm Hg

0º C = 273.15 K

Standard Temperature and Pressure

Standard Temperature and Pressure (STP) standard conditions used to compare gas measurements; defined as an absolute temperature of **273.15 K** and a pressure of **101.325 kPa**

- Scientists use absolute temperature (K) in all gas calculations.

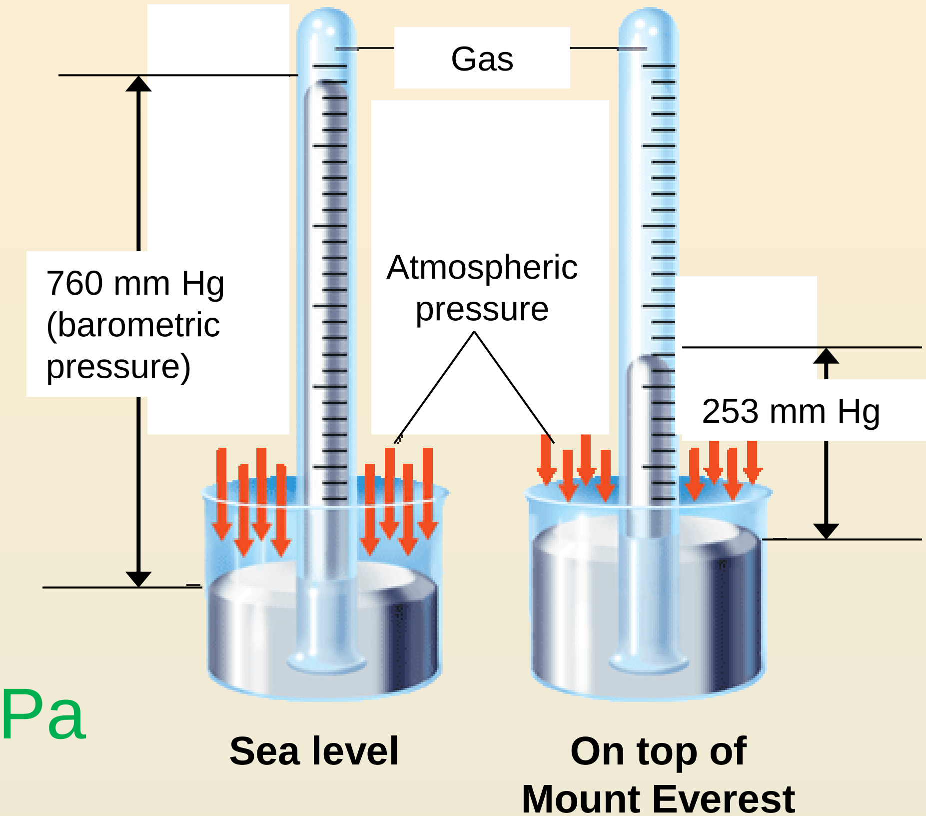


Gas Pressure

Two older units of pressure are still commonly used:

millimeters of mercury (mm Hg) and atmospheres (atm).

One **standard atmosphere (atm)** is the pressure required to support 760 mm of mercury in a mercury barometer at 25° C.



1 atm = 760 mm Hg = 101.3 kPa

Standard temperature & pressure (STP)

0º C or 273.15 K 101.3 kPa, 760 mm Hg, or 1 atm

The pressure in mm Hg inside a vacuum is 0 mm Hg.

Collapsing Can Experiment

Click on the link: (0:47)

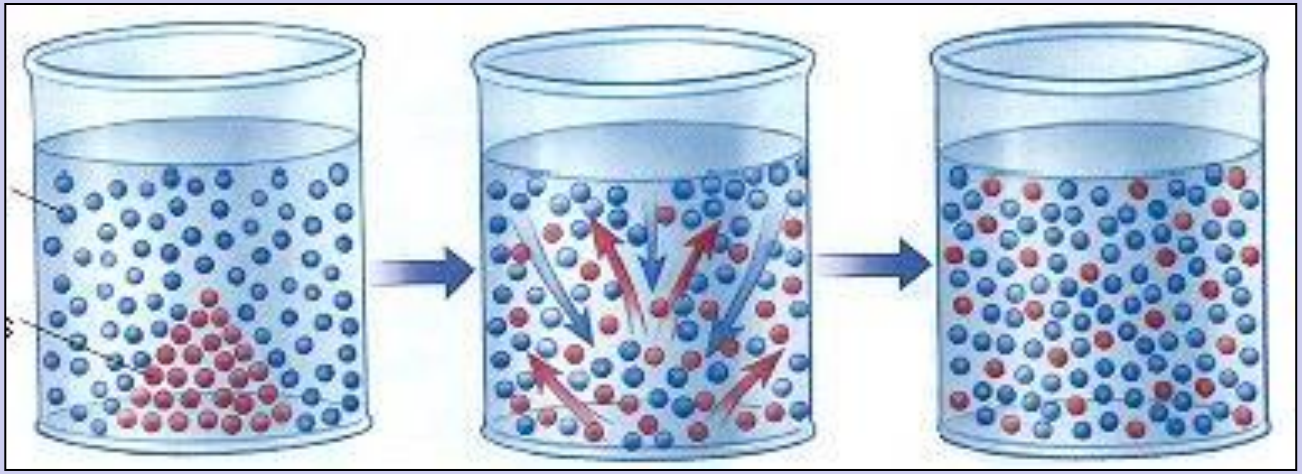
https://screencast-o-matic.com/watch/cFQ6XyqEwm

Why did the can collapse?

See the link on Learning CTR Online:

"FUN Pressure Activities" Labs

Gases exhibit:

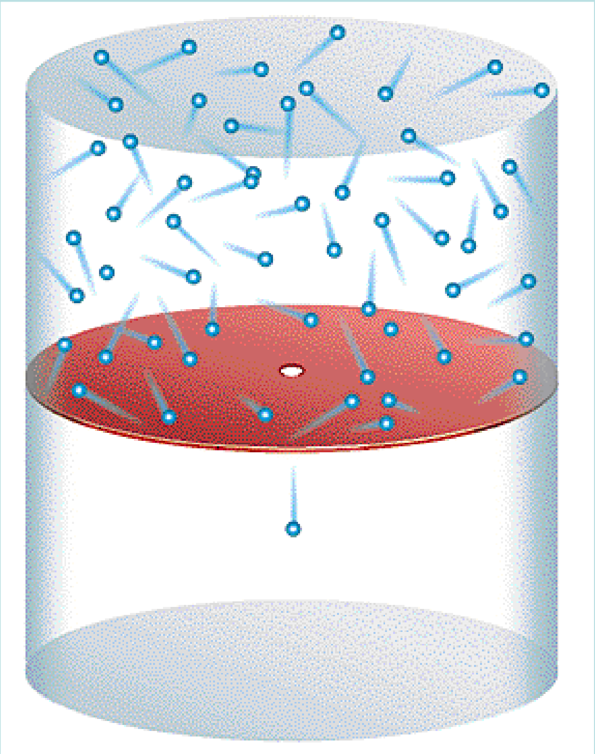


Diffusion

the gradual mixing of molecules of different gases from HIGH to LOW concentration.

Effusion

the movement of molecules through a small hole into an empty container.



Ideal Gas Law

$$PV = nRT$$

P = pressure in atm, mm Hg, or kPa

V = volume in liters

n = moles

R = ideal gas constant based on units

= *0.0821 liter•atm / mol•K*

= *62.396 liter•mm Hg / mol•K*

= *8.317 liter•kPa / mol•K*

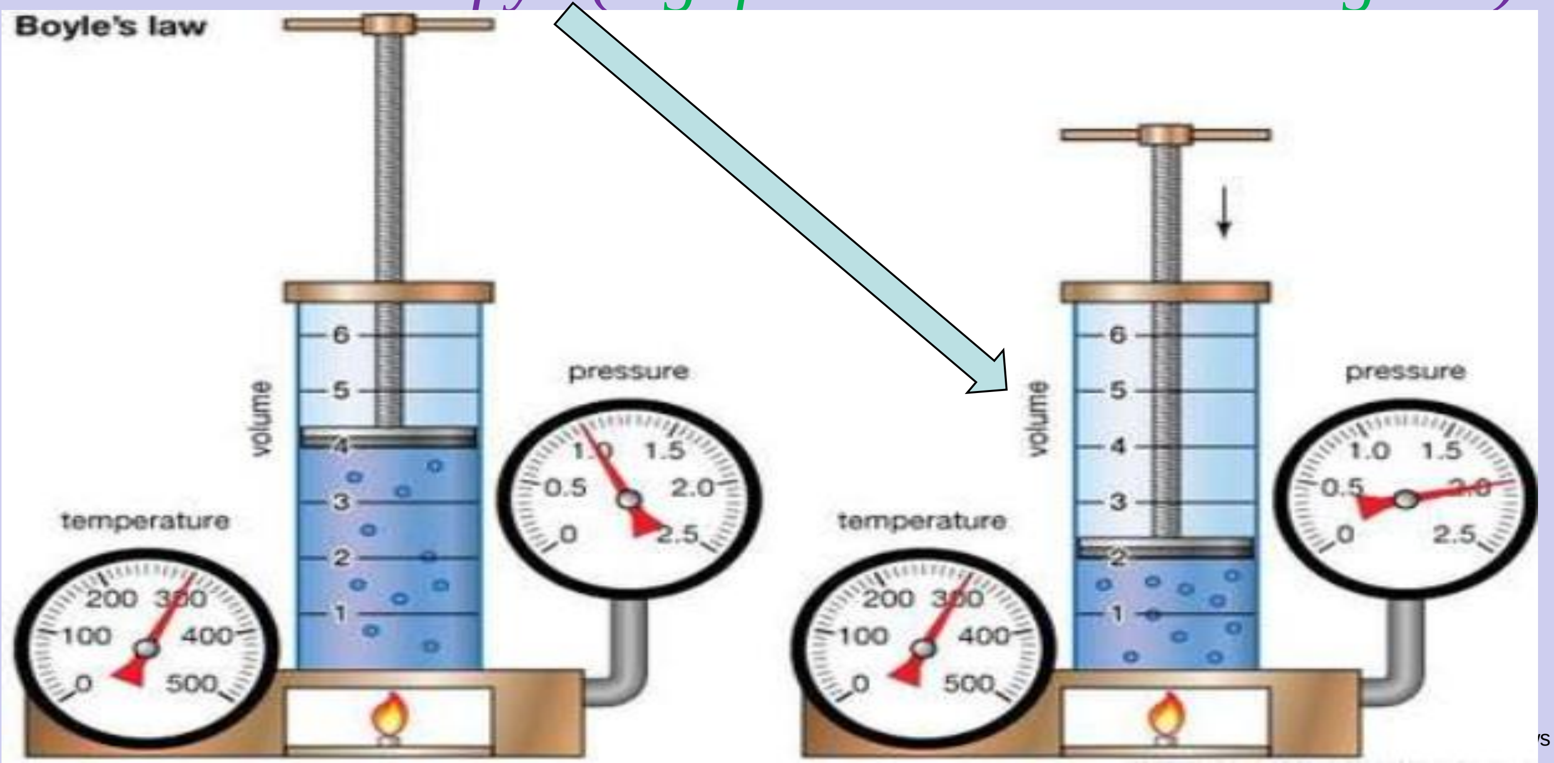
T = temperature in Kelvins

Boyle's Law

Boyle's Law describes the relationship between the **volume** of a gas and its **pressure** *at a constant temperature and* number of moles.

A gas can be compressed to take up less space:

The higher the pressure applied to the gas, the smaller the volume it will occupy. (e.g. piston chamber in engines)



Boyle's Law

The higher the pressure applied to the gas, the smaller the volume it will occupy.

Example

Bubbles in a fish tank increase in size as they move toward the surface of the water experiencing LESS pressure.

Bubbles at the bottom of the tank are smaller, experiencing more pressure than bubbles up higher in the water.

Boyle's Law

Robert Boyle (1627-1691) decided to experiment with only pressure and volume of a gas, so he held the temperature and number of moles of the gas constant.

$$PV = \cancel{nRT} \rightarrow PV = k$$

What mathematical relationship does this show?

Boyle's Law

Robert Boyle (1627-1691) decided to experiment with only pressure and volume of a gas, so he held the temperature and number of moles of the gas constant.

$PV =$ ~~nRT~~ → $PV = k$

Mathematically: P $\frac{1}{\propto}$ V OR P ∞ 1/V

(*take away the constants*)

At the same conditions of temperature and moles:

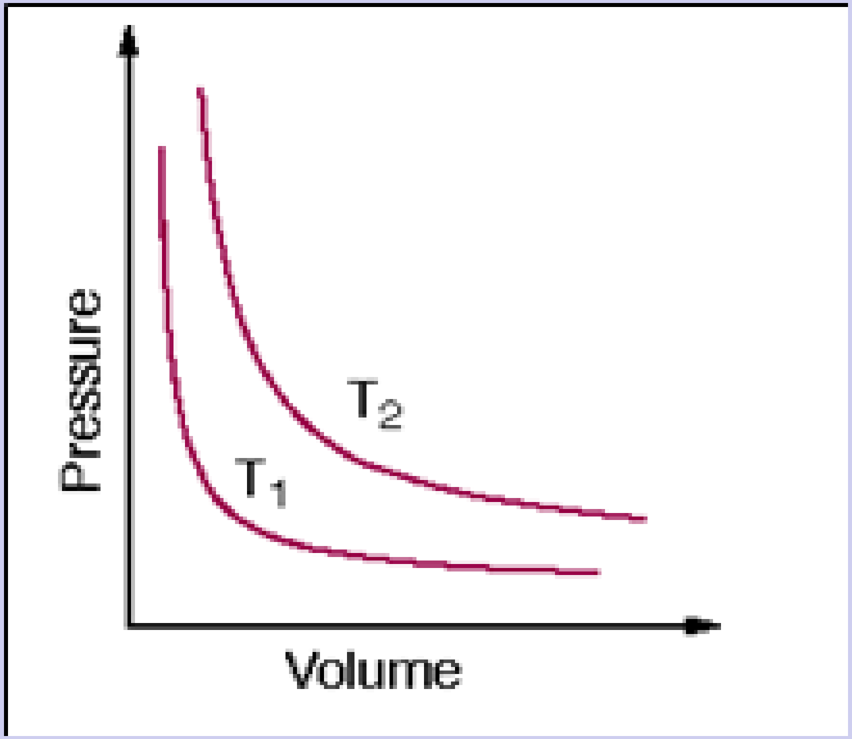
$P_1V_1 = k$, $P_2V_2 = k$

Thus: $$P_1V_1 = P_2V_2$$

Boyle's Law

Pressure is inversely proportional to volume:

$$P_1V_1 = P_2V_2$$



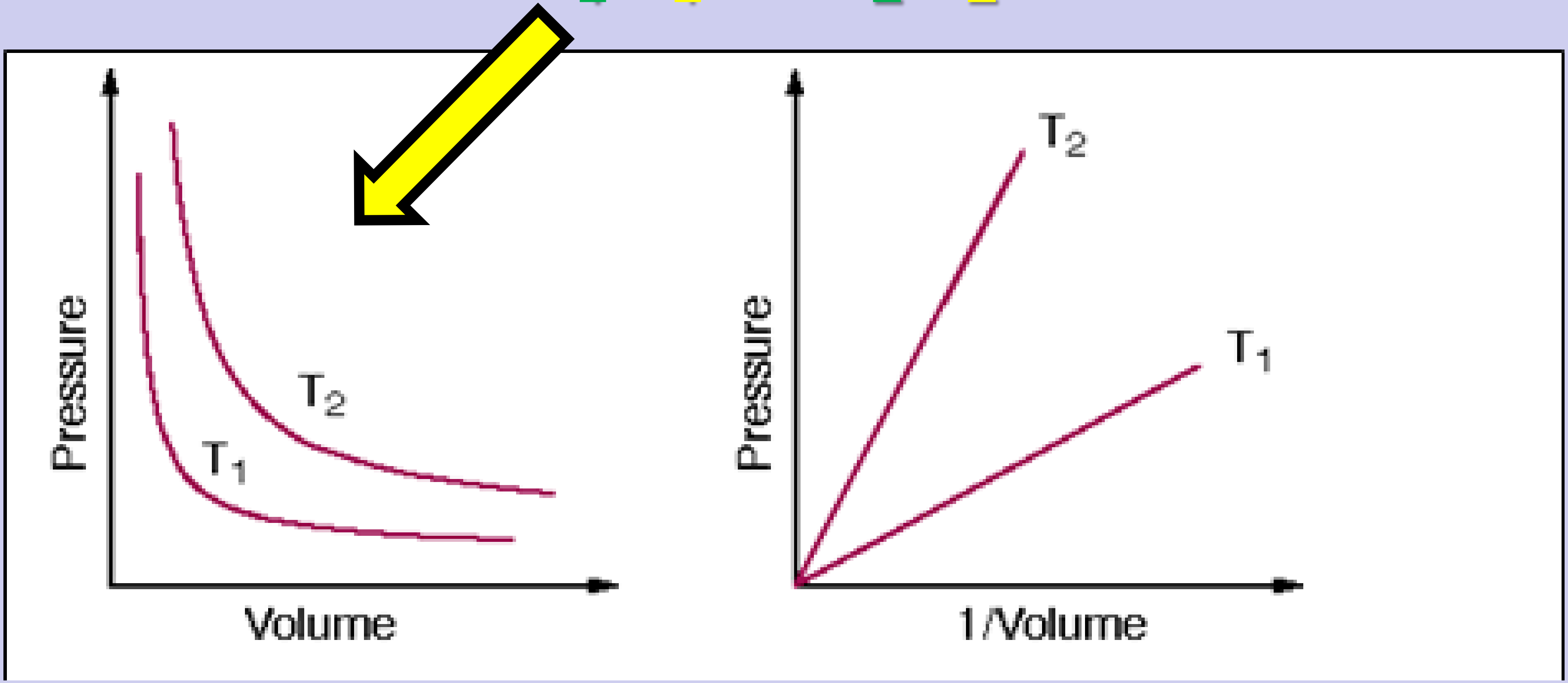
P ↑ V ↓

Each line on the graph is a line of constant temperature, and is called an isotherm.

Boyle's Law

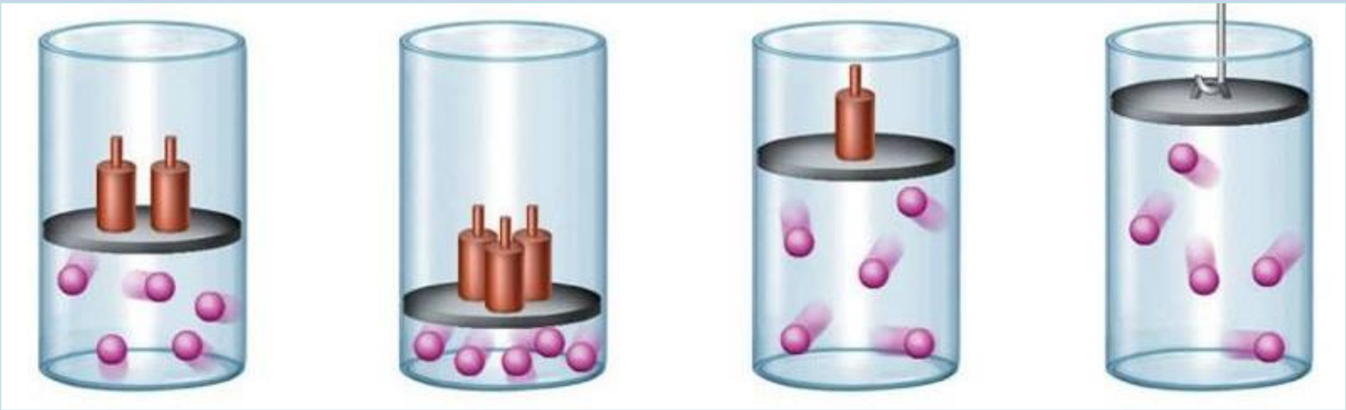
Pressure is inversely proportional to volume:

$$P_1V_1 = P_2V_2$$



Each line on the graph is a line of constant temperature, and is called an isotherm. In a plot of P versus 1/V, we see that the isotherms are straight lines with constant slopes.

Boyle's Law

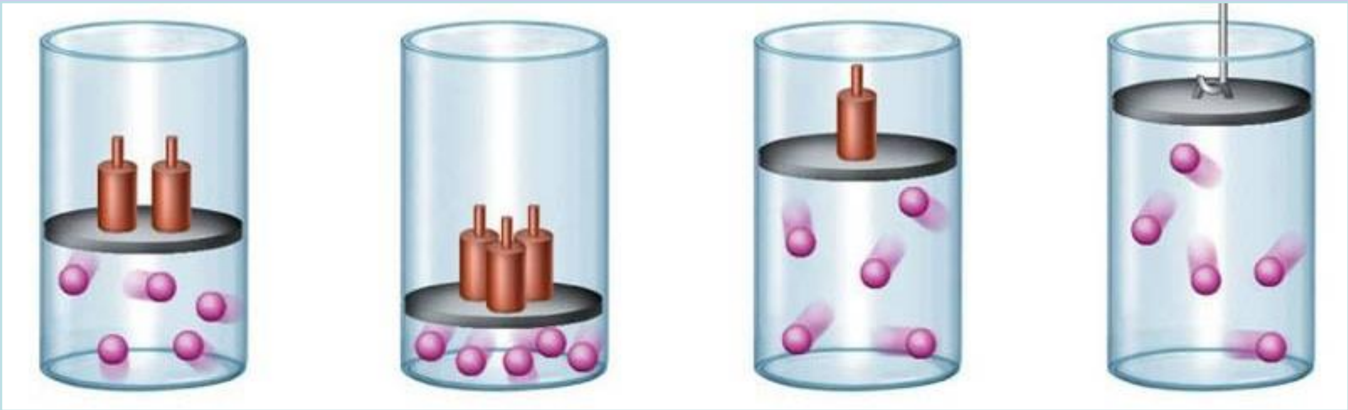


Describe the relationship between pressure and volume in the cylinders. Remember pressure is related to collisions.

Find the volume of a gas which occupies 200. ml at 2.6 atm when it is at standard pressure.

QUICK CHECK

Boyle's Law



Describe the relationship between pressure and volume in the cylinders. Remember pressure is related to collisions.

$P \infty 1/V$ … as pressure > volume < or vice versa
When molecules are confined to a smaller space (less volume), there are more collisions (more pressure)

Find the **volume** (V_2) of a gas which occupies 200. ml (V_1) at 2.6 atm (P_1) when it is at standard pressure (P_2).

$P_1V_1 = P_2V_2$ … solve for V_2: $= P_1V_1 / P_2$

V_2 = 200. ml x (2.6 atm/1 atm) = <u>520 ml</u> (notice P ↓ V ↑)

Charles' Law

Jacques Charles (1746-1823) experimented with only temperature and volume of a gas, so he held the pressure and number of moles of the gas constant.

$$\cancel{P}V = \cancel{nR}T \rightarrow V/T = k$$

What mathematical relationship does this show?

Charles' Law

Jacques Charles (1746-1823) experimented with only temperature and volume of a gas, so he held the pressure and number of moles of the gas constant.

~~P~~V = ~~nR~~T → V/T = k

Mathematically: $V \propto T$ *(take away the constants)*

At the same conditions of pressure and moles:

$V_1 = kT_1, \quad V_2 = kT_2$

STRATEGY

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

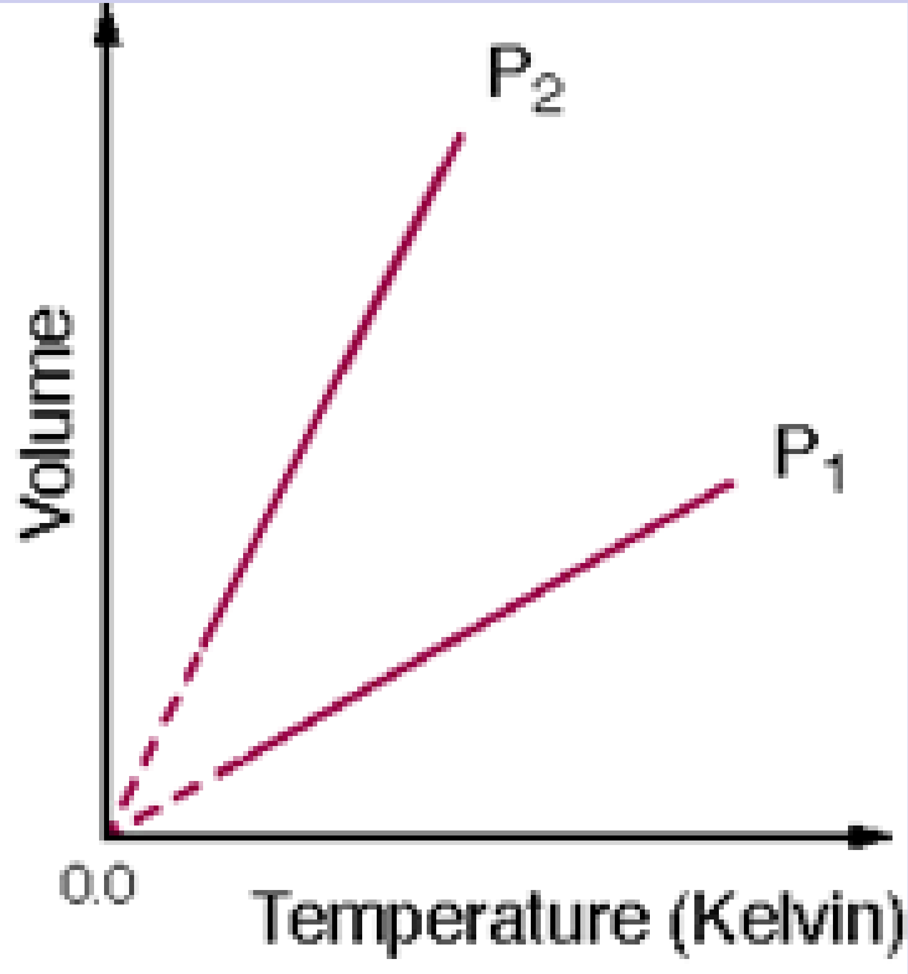
Thus: $V_1/T_1 = V_2/T_2$

Charles' Law

Temperature is directly proportional to volume:

$$V_1/T_1 = V_2/T_2$$

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



V ↑ T ↑

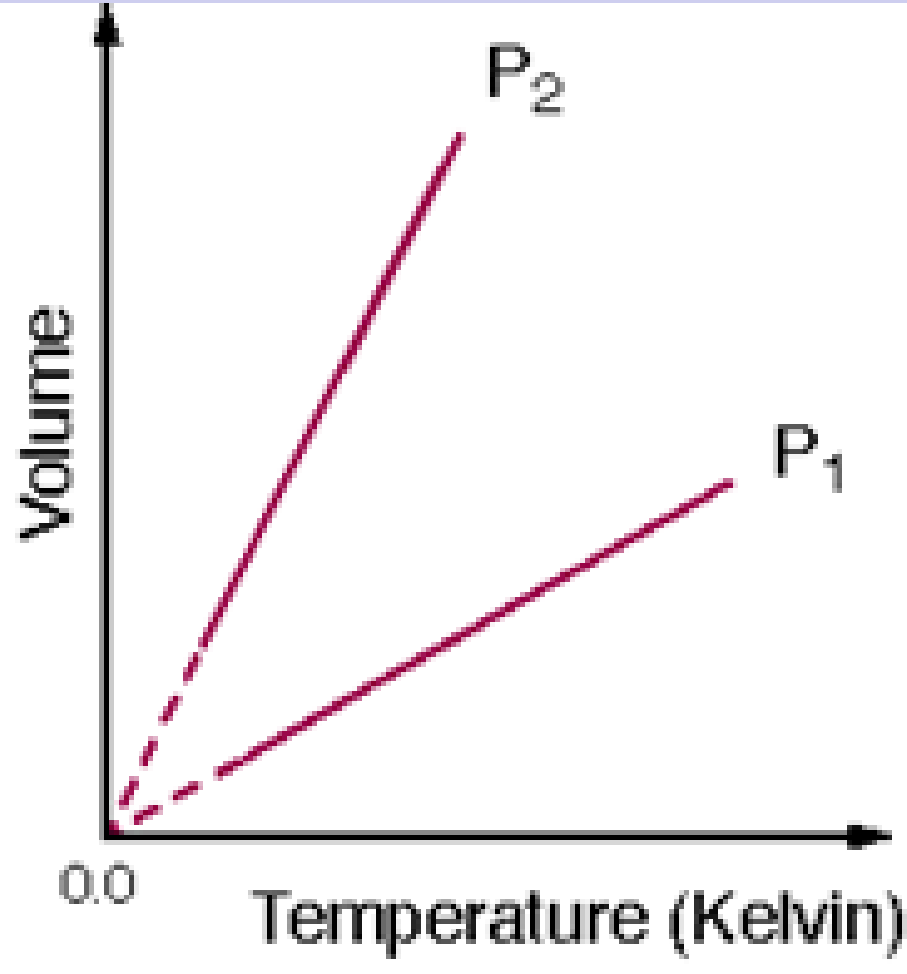
Absolute temperature

Charles' Law

Temperature is directly proportional to volume:

$$V_1/T_1 = V_2/T_2$$

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



In the plot of V-vs-T, the direct relationship is linear, with an intercept of absolute zero on the **Kelvin** scale.

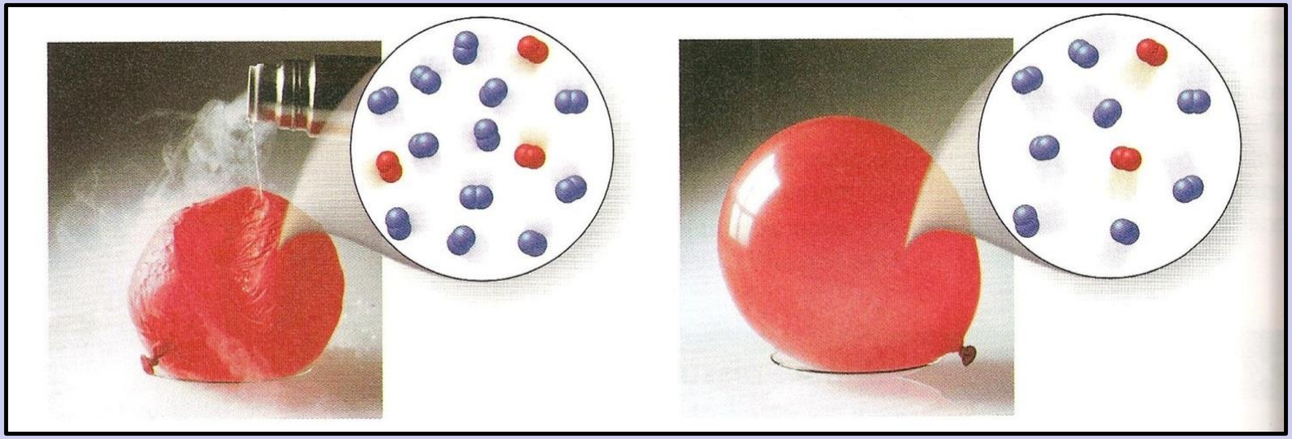
The slope depends on the pressure and the amount of gas.

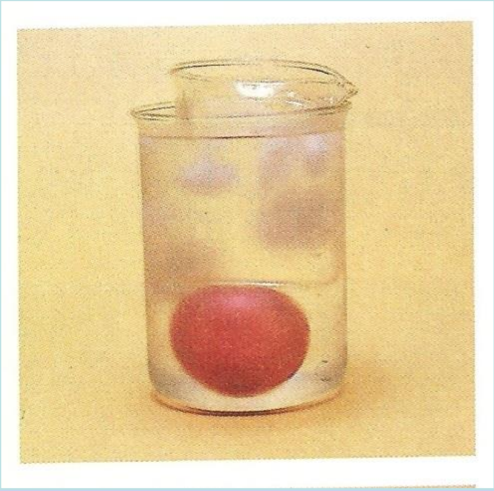
The **dashed lines** on the graph are to represent the fact that you can not achieve absolute zero.

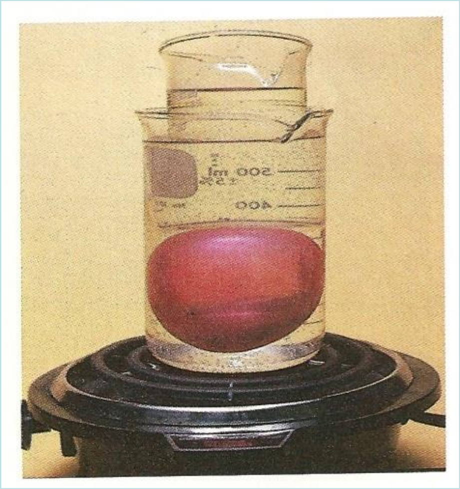
Kelvin includes avg KE & volume.

Charles' Law

Liquid nitrogen poured on a balloon causes it to shrink.





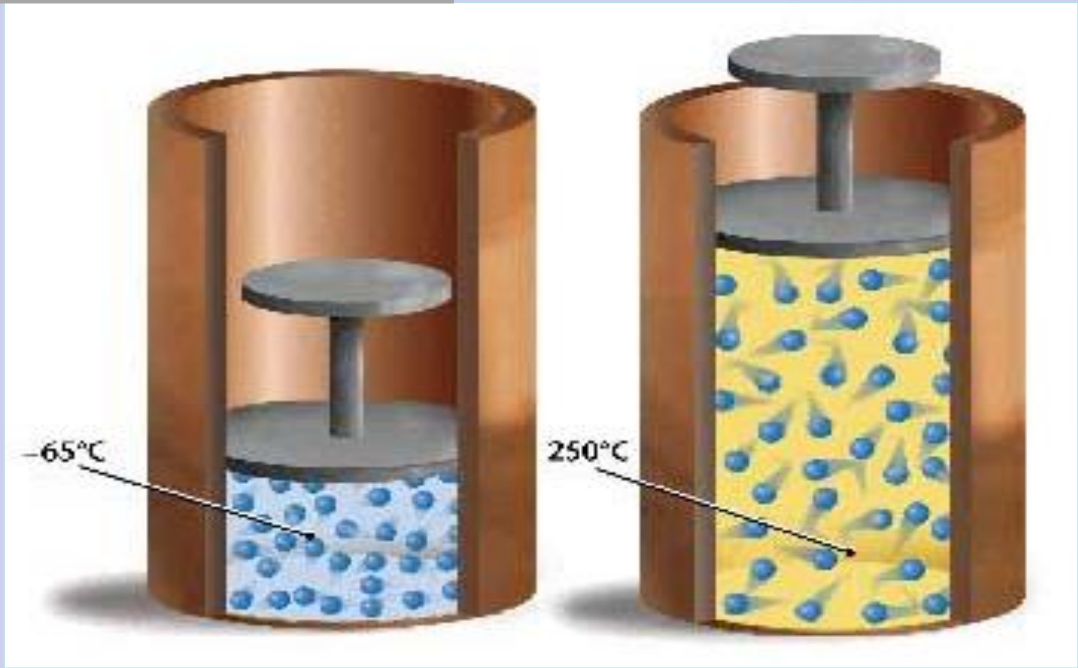


A balloon is smaller in a beaker of ice water than in a beaker of boiling water. The volume of the air in balloon increases with increasing temperature.

Heated gas expands, as in a hot air balloon.

Charles' Law

Describe the relationship between temperature and volume in the cylinders. Remember temperature is related to average kinetic energy of the molecules.



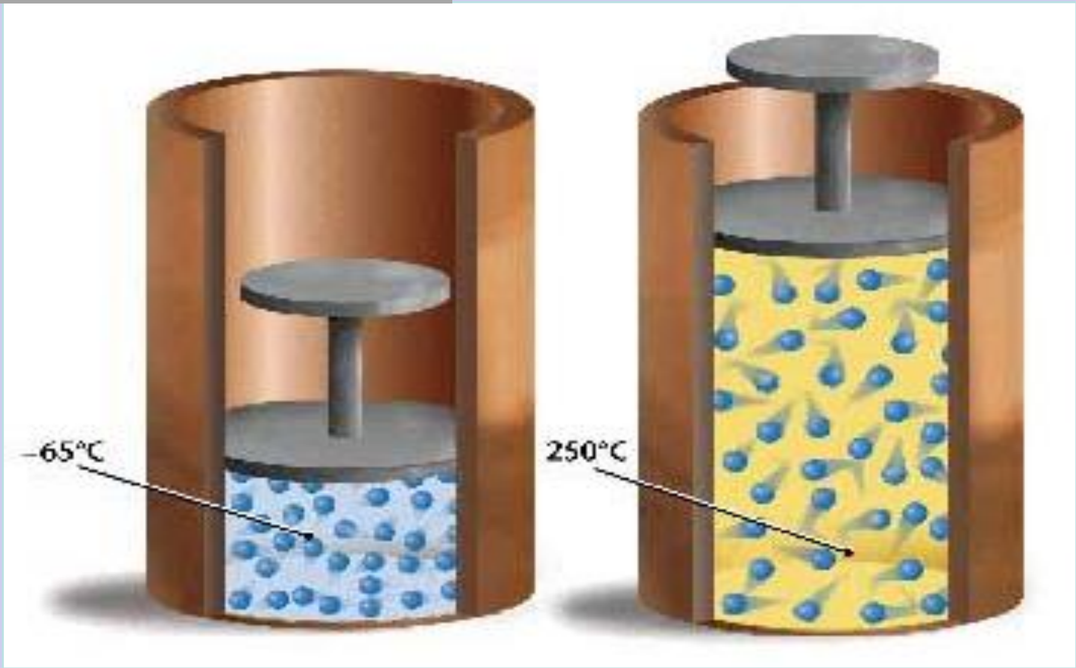
A scientist has an initial volume of 100. ml of a gas at 25.0º C and constant pressure. If the temperature changes to 100.º C, what is the new volume of the gas?

QUICK CHECK

Charles' Law

Describe the relationship between temperature and volume in the cylinders. Remember temperature is related to average kinetic energy of the molecules.

The motion of the molecules (T) > as volume > to keep pressure constant.



A scientist has an initial volume of 100. ml (V1) of a gas at 25.0º C (T1) and constant pressure. If the temperature changes to 100.º C (T2) , what is the new **volume** (V2) of the gas?

V_1 = 100. ml T_1 = 25.0 C = 298 K T_2 = 100. C = 373 K

$V_1/T_1 = V_2/T_2$ $V_2 = V_1T_2 / T_1$ V_2 = (100. ml)(373 K) / 298 K

V_2 = 125 ml   (notice V ↑ T ↑)

Gay-Lussac's Law

Gay-Lussac (1746-1823) experimented with only temperature and pressure of a gas, so he held the volume and number of moles of the gas constant.

$$P\cancel{V} = \cancel{nR}T \rightarrow P/T = k$$

What mathematical relationship does this show?

Gay-Lussac's Law

Gay-Lussac (1746-1823) experimented with only temperature and pressure of a gas, so he held the volume and number of moles of the gas constant.

$$P\cancel{V} = \cancel{nR}T \rightarrow P/T = k$$

Mathematically: $P \infty T$ (*take away the constants*)

At the same conditions of volume and moles:

$P_1 = kT_1$, $P_2 = kT_2$

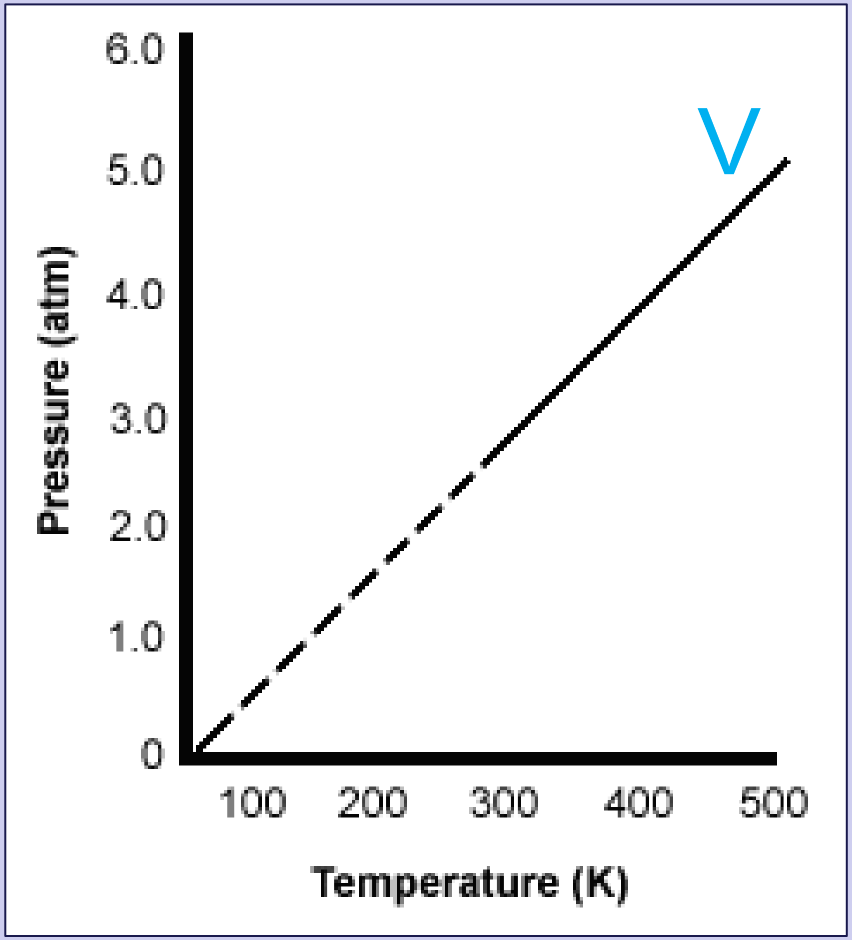
Thus: $$P_1/T_1 = P_2/T_2$$

STRATEGY

Gay-Lussac's Law

Pressure is directly proportional to temperature:

$$P_1/T_1 = P_2/T_2$$



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

P ↑ T ↑

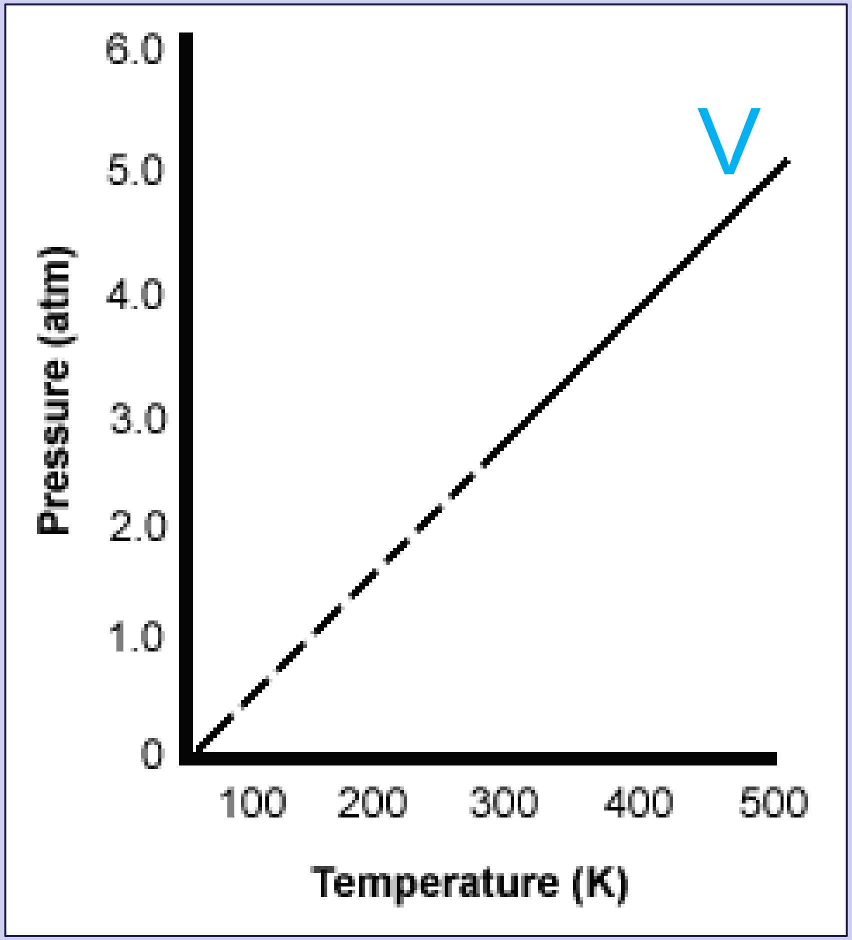
Absolute temperature

Gay-Lussac's Law

Pressure is directly proportional to temperature:

$$P_1/T_1 = P_2/T_2$$

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



In the plot of P-vs-T, the direct relationship is linear.

The slope depends on the volume and the amount of gas.

The **dashed lines** on the graph are to represent the fact that you can not achieve absolute zero.

Kelvin includes avg KE & volume.

THINK ABOUT IT

Gay-Lussac’s Law

Using Gay-Lussac’s law, explain when you need to check the tire pressure more often: in the summer or the winter.

The temperature of a sample of gas in a steel tank at 30.0 kPa is increased from –100.0˚C to 25.0˚ C. What is the final pressure inside the tank?

THINK ABOUT IT

Gay-Lussac's Law

Using Gay-Lussac's law, explain when you need to check the tire pressure more often: in the summer or the winter.

As the tires rotate, the temperature of the air in the tires increases due to friction on the road. The pressure of the air also increases (molecules move faster, more collisions).

On a hot summer day, pressure in a car tire increases over what it would be on a cold winter day. Tire pressure often drops significantly in the winter.

The temperature of a sample of gas in a steel tank at 30.0 kPa (P1) is increased from –100.0˚C (T1) to 25.0˚ C (T2). What is the final pressure (P2) inside the tank?

$T_1 = -100.0\ C = 173\ K \quad T_2 = 25.0\ C = 298\ K$

$P_1/T_1 = P_2/T_2 \qquad P_2 = P_1T_2 / T_1 \qquad P_2 = (30.0\ kPa)(298\ K)/(173\ K)$

$\underline{P_2 = 51.7\ kPa}$ (notice T ↑ P ↑)

Charles' Law & Gay-Lussac's Law

Watch the video on Learning CTR Online:

http://somup.com/cFX10cn17m

Charles' Law & Gay-Lussac's Law ctr (3:12)

- Volume vs. Temperature
- Pressure vs. Temperature

Avogadro's Law

Amedeo Avogadro (1776-1856) experimented with only volume and moles of a gas, so he held the pressure and temperature of the gas constant.

$$\cancel{P}V = n\cancel{RT} \rightarrow V = kn \rightarrow k = V/n$$

What mathematical relationship does this show?

Avogadro's Law

Amedeo Avogadro (1776-1856) experimented with only volume and moles of a gas, so he held the pressure and temperature of the gas constant.

$$\cancel{P}V = n\cancel{RT} \rightarrow V = kn \rightarrow k = V/n$$

Mathematically: $V \propto n$ (*take away the constants*)

At the same conditions of pressure and temperature:

$V_1 = kn_1$, $V_2 = kn_2$

Thus: $V_1/n_1 = V_2/n_2$

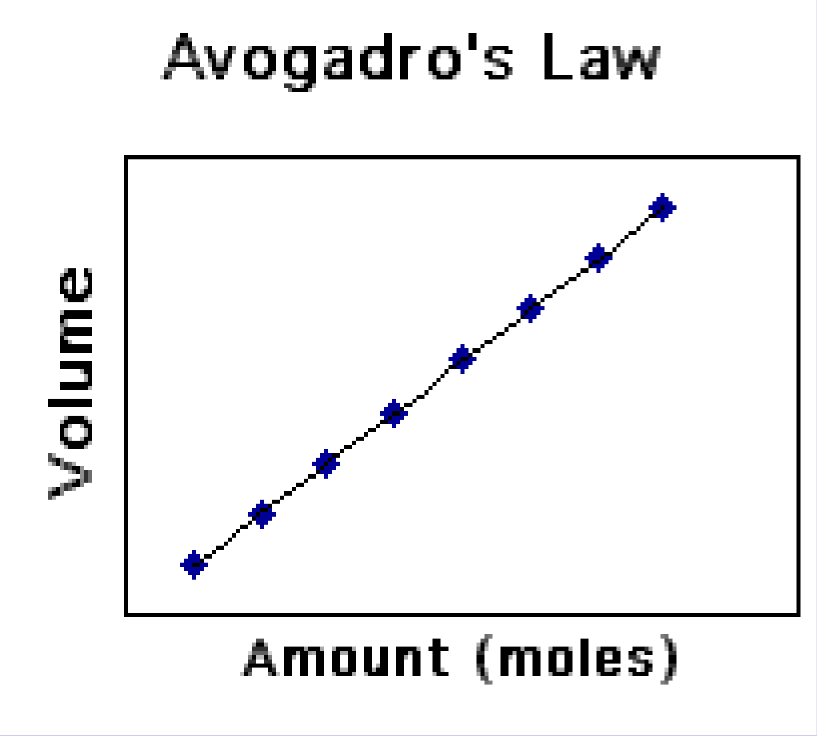
$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Avogadro's Law

Volume is directly proportional to moles:

$$V_1/n_1 = V_2/n_2$$

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



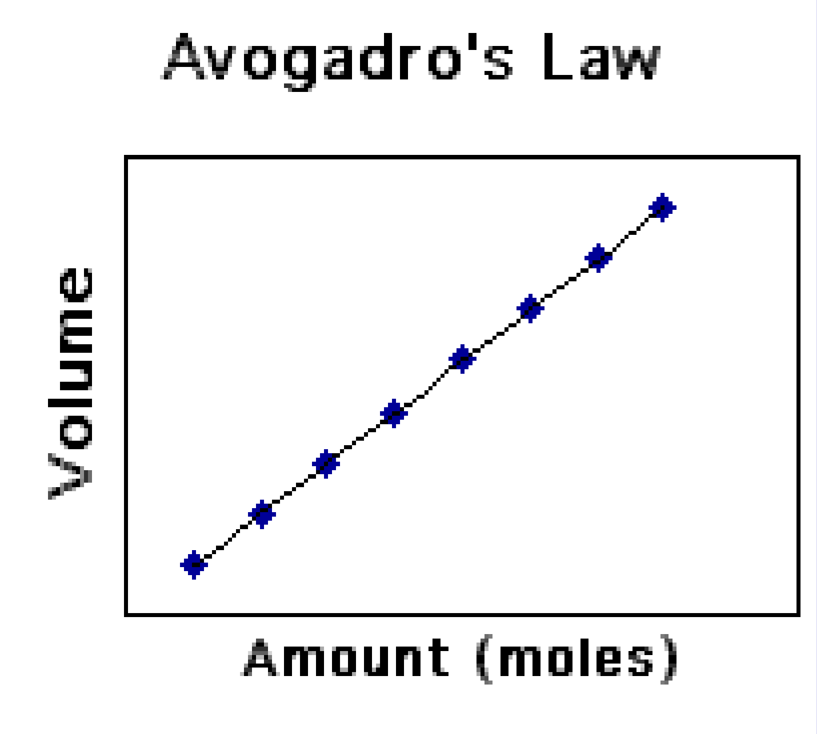
V ↑ n ↑

Avogadro's Law

Volume is directly proportional to moles:

$$V_1/n_1 = V_2/n_2$$

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



In the plot of V-vs-n, the direct relationship is linear.

For a gas:

1 mol = # molecules / 22.4 L

molar volume = 22.4 L

$N_A = 6.02 \times 10^{23}$ particles, etc.

Avogadro's Law

ENRICHMENT

How many molecules of oxygen gas (O_2) are in 60.0 liters @ STP?

Describe the volume and moles in each balloon:

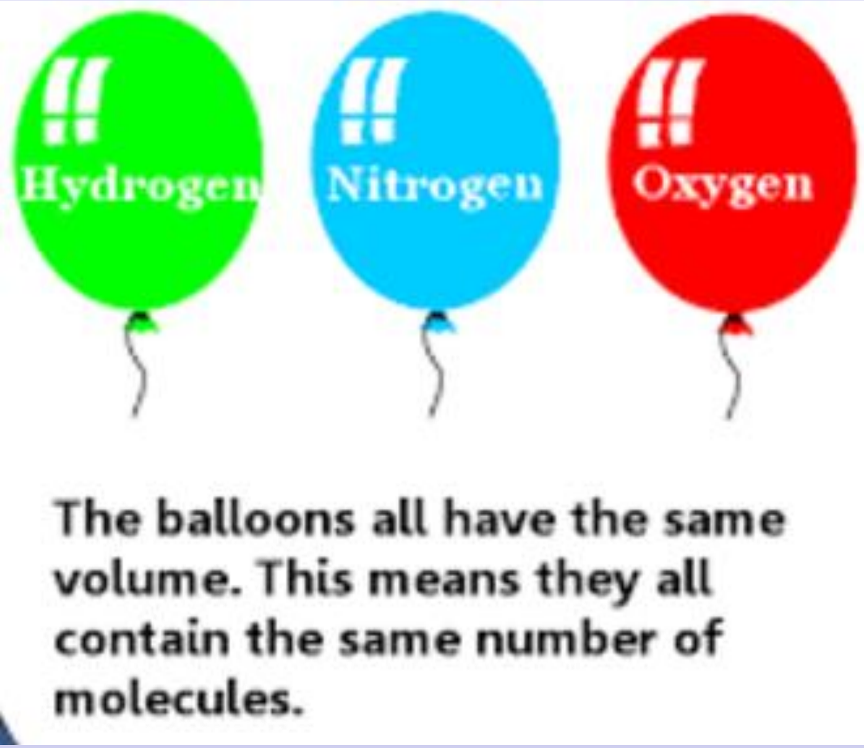
If the balloons all have the same volume, how do their moles compare?

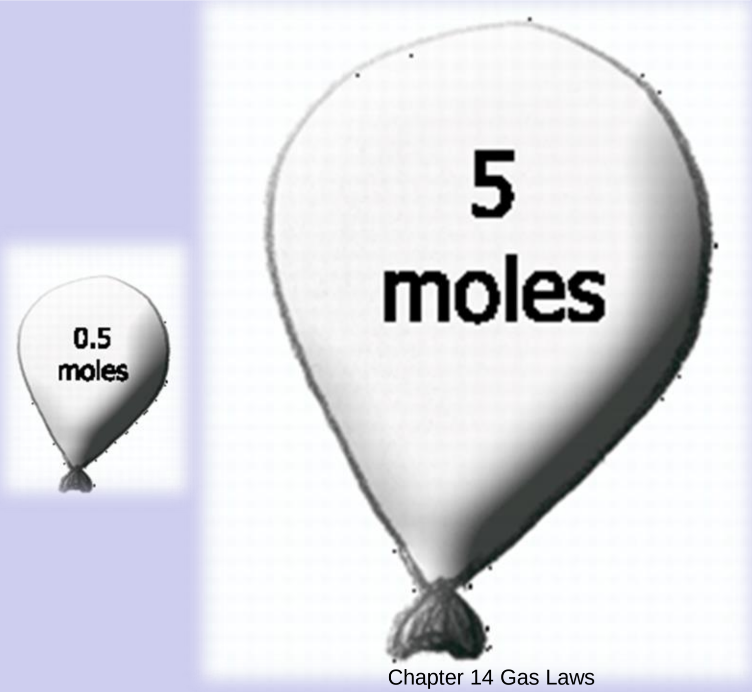
Avogadro's Law

How many molecules of oxygen gas (O_2) are in 60.0 liters @ STP?

60.0 L x 1 mol / 22.4 L x 6.02 x 10^{23}/1 mol = <u>1.61 x 10^{24} molecules</u>

The volume and moles increased from left to right as a direct proportion.





Combined gas law:

For a fixed quantity of gas, pressure & volume vary inversely while temperature varies directly with pressure & volume

Boyle's law

$$P_1V_1 = P_2V_2$$

Charles's law

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

Gay-Lussac's law

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

What would it look like if we combined all three gas laws?

Combined gas law:

For a fixed quantity of gas, pressure & volume vary inversely while temperature varies directly with pressure & volume

Boyle's law

$$P_1V_1 = P_2V_2$$

Charles's law

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

Gay-Lussac's law

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

What would it look like if we combined all three gas laws?

Remember: $V \propto 1 / P$ inverse; $V \propto T$ direct; $P \propto T$ direct

Therefore, $V \propto T / P$ combine; $PV \propto T$ rearrange; **$PV / T = k$** equation

Combined gas law:

For a fixed quantity of gas, pressure & volume vary inversely while temperature varies directly with pressure & volume

Boyle's law

$$P_1V_1 = P_2V_2$$

Charles's law

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

Gay-Lussac's law

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

Combined gas law

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

Combined Gas Law

You *must* have temperature in Kelvin:

1) K includes volume; 2) you can't divide by zero (e.g. 0º C).

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

Cover up Pressure (*it remains constant*) … what do you have?

Cover up Volume (*it remains constant*) … what do you have?

Cover up Temperature (*it remains constant*) … what do you have?

THINK ABOUT IT

STRATEGY

Combined Gas Law

You *must* have temperature in Kelvin:
1) K includes volume; 2) you can't divide by zero (e.g. 0º C).

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

Cover up **Pressure** (*it remains constant*) … what do you have?

Charles' Law

Direct Relationship

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

Cover up **Temperature** (*it remains constant*) … what do you have?

Boyle's Law

Inverse Relationship

$P_1V_1 = P_2V_2$

Cover up **Volume** (*it remains constant*) … what do you have?

Gay-Lussac's Law

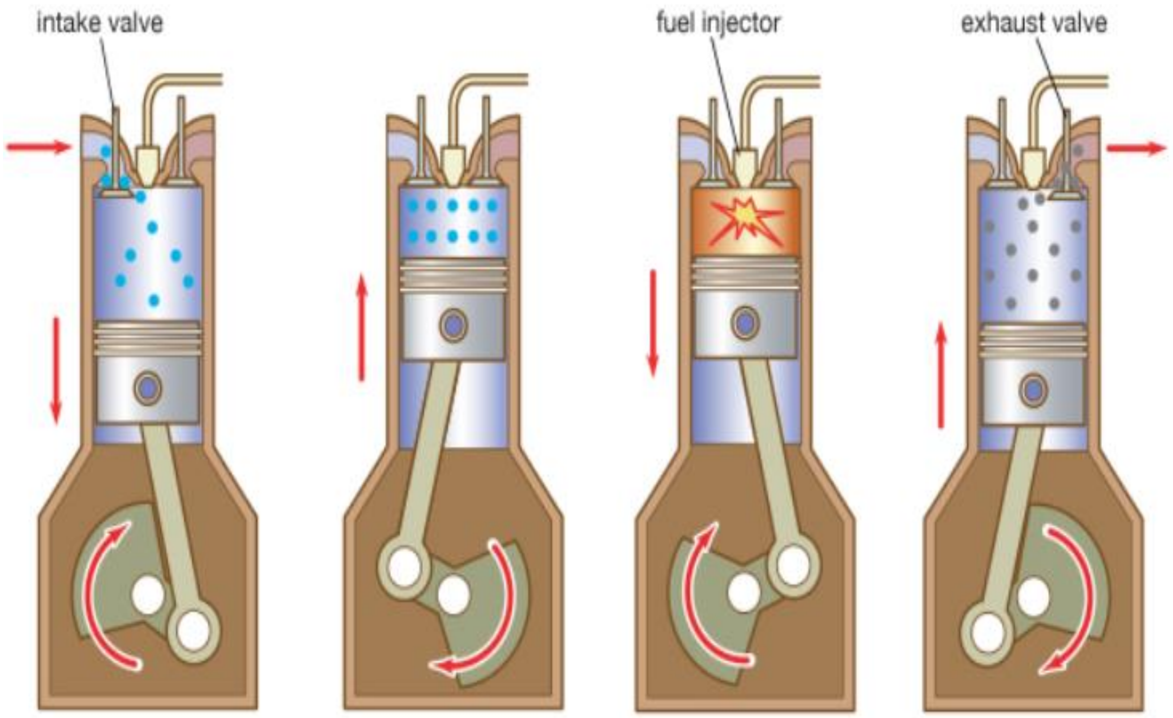
Direct Relationship

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

Real World Application of the Combined Gas Law

Internal Combustion Engines

Operate on the principle of taking a volume (V) of gas in, compressing it (P), igniting it (T), then pushing the exhaust out.



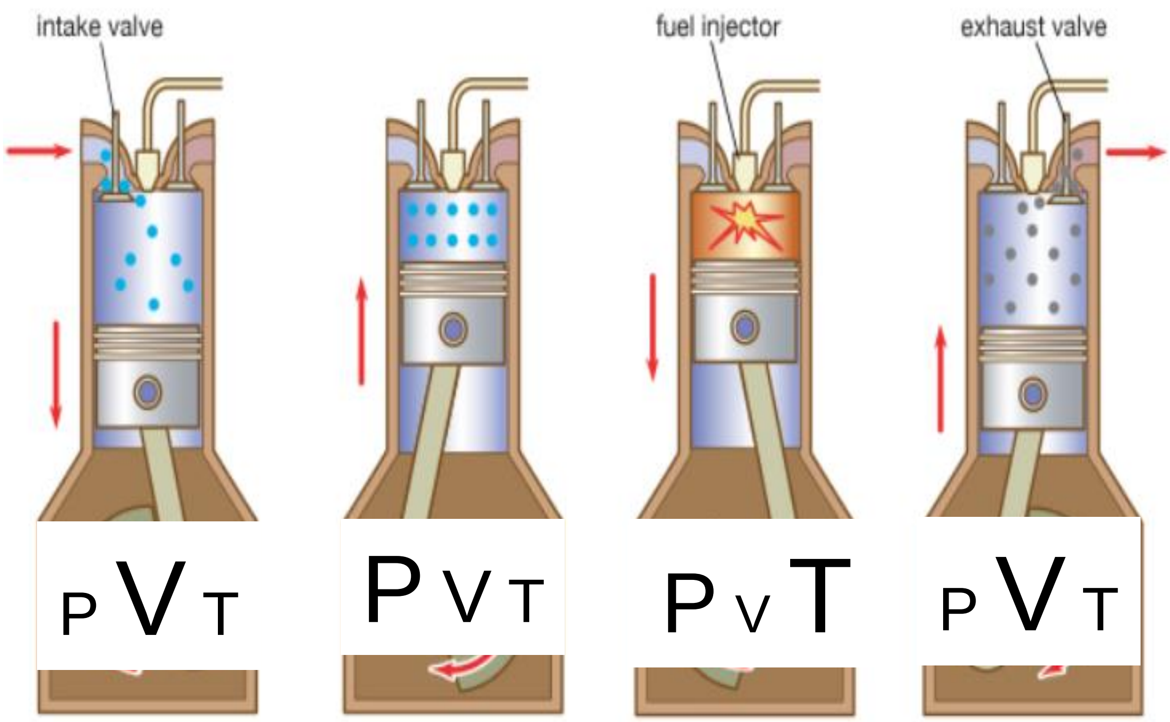
1. The air-fuel mixture enters the piston chamber.
2. The air-fuel mixture is compressed (piston moves up)
3. The fuel is ignited.
4. The hot, expanding gases force the piston down and the exhaust gases are expelled from the chamber.

Real World Application of the Combined Gas Law

Internal Combustion Engines Notice P, V, T

The volume (V), pressure (P), and temperature (T) interactively vary:

1. The air-fuel mixture enters the piston chamber.
2. The air-fuel mixture is compressed (piston moves up)
3. The fuel is ignited.



4. The hot, expanding gases force the piston down and the exhaust gases are expelled from the chamber.

Joule-Thomson Effect

ENRICHMENT

Spraying a gas from an aerosol can:

Pressure (P) of gas decreases as it goes from being compressed in the container to being released into the atmosphere. Volume (V) increases substantially.

Kinetic energies of gas molecules decrease. Gas temperature (T) therefore decreases as it is sprayed.

Real World Application of the Combined Gas Law

- Open your hand, palm in, and place near your mouth (~1-2 inches away). Open your mouth wide and blow.
- Repeat the same procedure EXCEPT whistle into your palm.
- Make your observations and explain what happened.

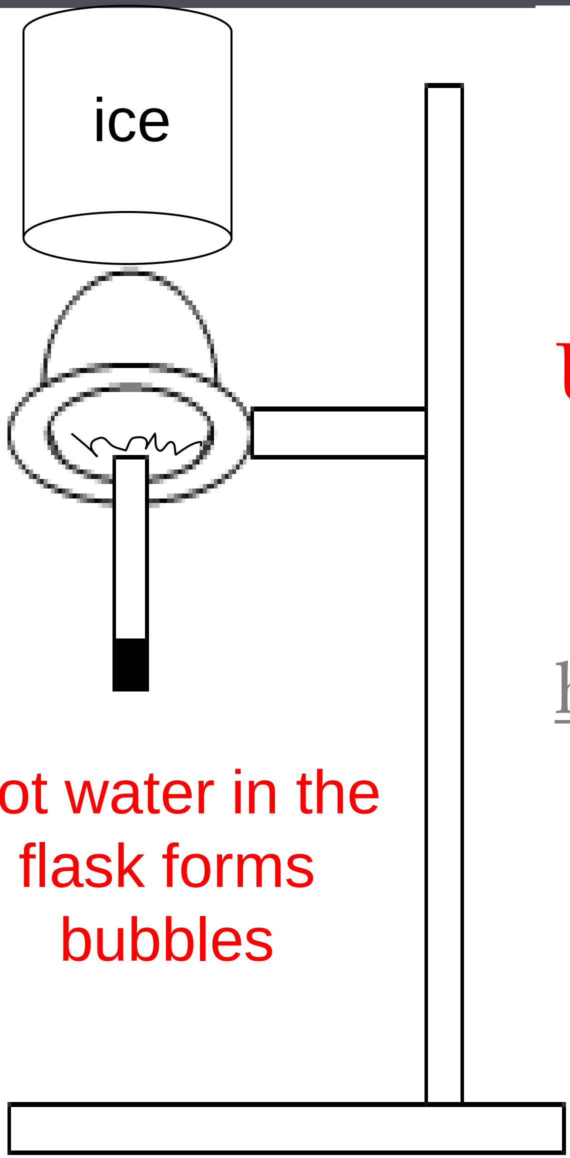
THINK ABOUT IT

Real World Application of the Combined Gas Law

- Open your hand, palm in, and place near your mouth (~1-2 inches away). Open your mouth wide and blow.
- Repeat the same procedure EXCEPT whistle into your palm.
- Notice with mouth wide, it feels warmer, but when whistling it feels cooler (air expands).
- *When whistling, the pressure (P) was relatively high leaving the mouth; Volume (V) INcreased drastically (expansion), Decreasing the perceived temperature (T).*

THINK ABOUT IT

What is the relationship between pressure, temperature, and volume of a gas?

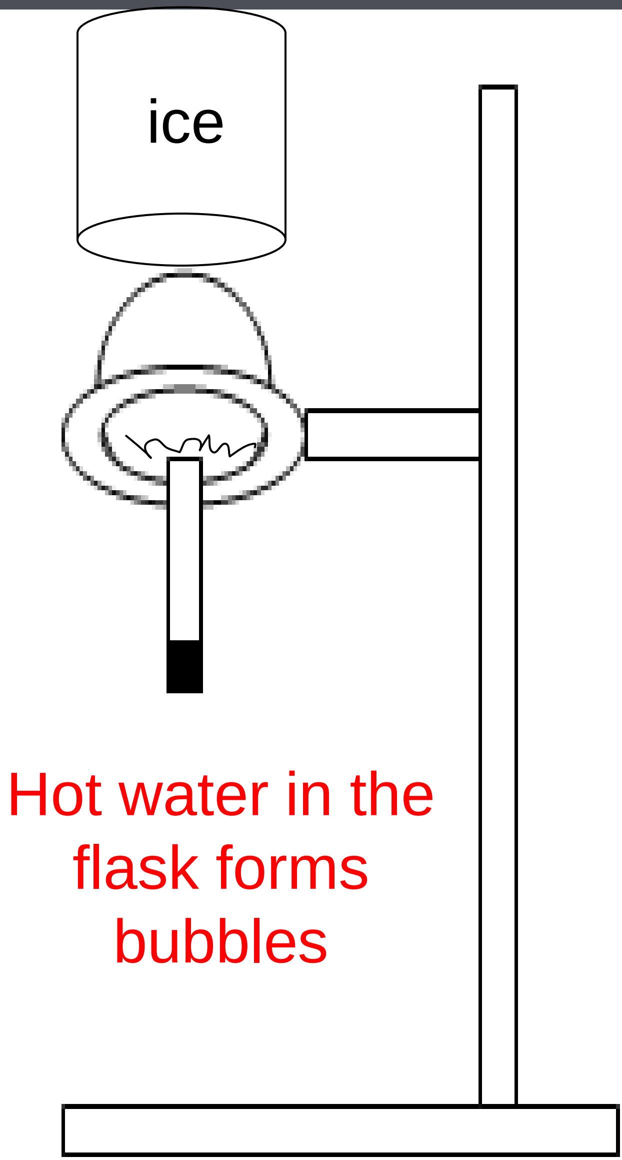


Hot water in the flask forms bubbles

Using ice to make water boil!

http://somup.com/cFXiDHn19g (3:08)

What is the relationship between pressure, temperature, and volume of a gas?



Using **ice to make water boil!**

http://somup.com/cFXiDHn19g (3:08)

- Ice cools the gas inside the flask (above the water).
- Cooling the gas decreases the pressure of the gas: P ∞ T
- The gas pressure is lower so it matches the vapor pressure of the water. Thus, boiling.
- Volume (V) changes slightly before equilibrium.

A weather balloon is filled with helium before it is launched. On the ground, the gas has a temperature of 25 °C and a pressure of 1.0 atm. The volume of the balloon is 4.5 m^3. As the balloon rises, however, the pressure and temperature change to 0.30 atm and 15 °C. What is the new volume of the balloon?

TRY IT

A weather balloon is filled with helium before it is launched. On the ground, the gas has a temperature of 25 °C and a pressure of 1.0 atm. The volume of the balloon is 4.5 m³. As the balloon rises, however, the pressure and temperature change to 0.30 atm and 15 °C. What is the new volume of the balloon?

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

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

TRY IT

T_1 = 25 C + 273 K = 298 K

P_1 = 1.0 atm

V_1 = 4.5 m^3

T_2 = 15 C + 273 K = 288 K

P_2 = 0.30 atm

V_2 = ?

V_2 = (1.0 atm)(4.5 m^3)(288 K) / (298 K)(0.30 atm)

V_2 = 14 m^3 (2 sig figs)

Ideal Gases

Ideal gases are imaginary gases that perfectly fit all of the assumptions of the kinetic molecular theory; i.e.,

- Gases consist of tiny particles that are far apart relative to their size.
- There are no forces of attraction or repulsion between the gas particles.
- Collisions between gas particles and between particles and the walls of the container are elastic collisions, so no kinetic energy is lost in those collisions.

There is no gas for which these assumptions are true.

"Kinetic Theory" video (2:01)

Kinetic Theory ctr

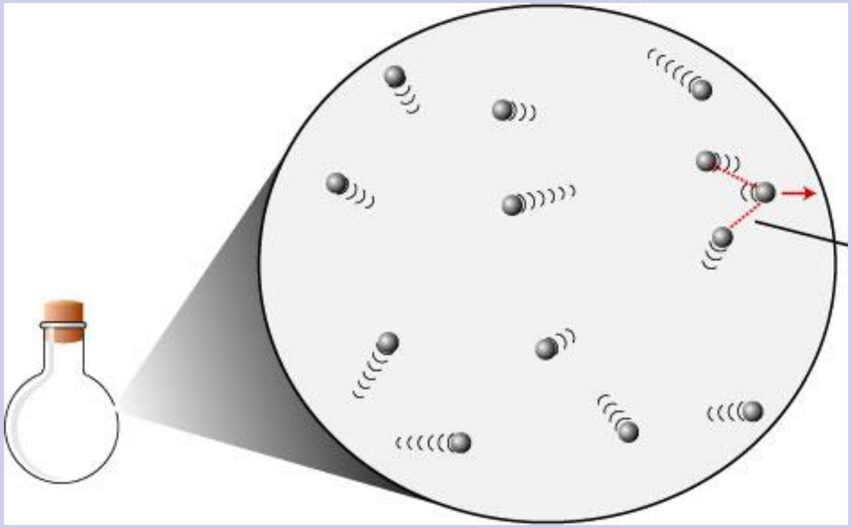
Ideal Gases & Real Gases

A gas will *behave* in an ideal fashion (PV = nRT) if its temperature is at or greater than 273 K (0^o C), and its pressure is near or less than 1.00 atm (101.3 kPa).

Standard Temperature and Pressure.

273 K or higher

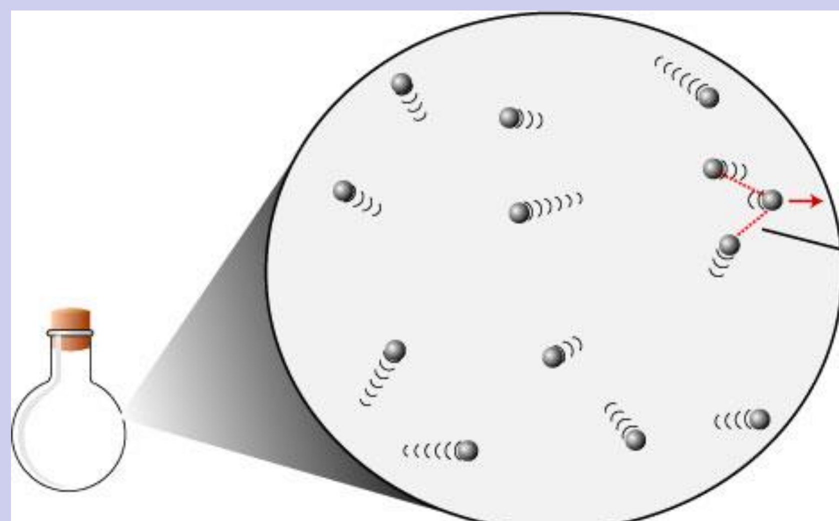
101.3 kPa or lower



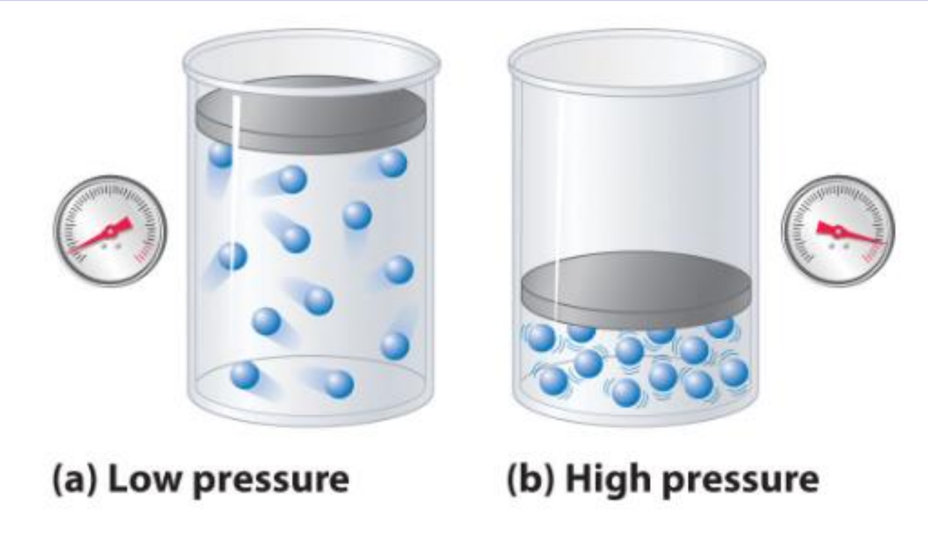
Standard Temperature and Pressure.

- Boyle's Law, Charles' Law, Gay-Lussac's Law and Avogadro's Law work at STP.

- They do not work when conditions do not match STP.



Ideal Gases & Real Gases



When pressure is HIGH as in image (b), the volume that the gas molecules occupy is significant compared to the region in which they are contained.

Particles in such real gases have volume and there are attractions between the particles.

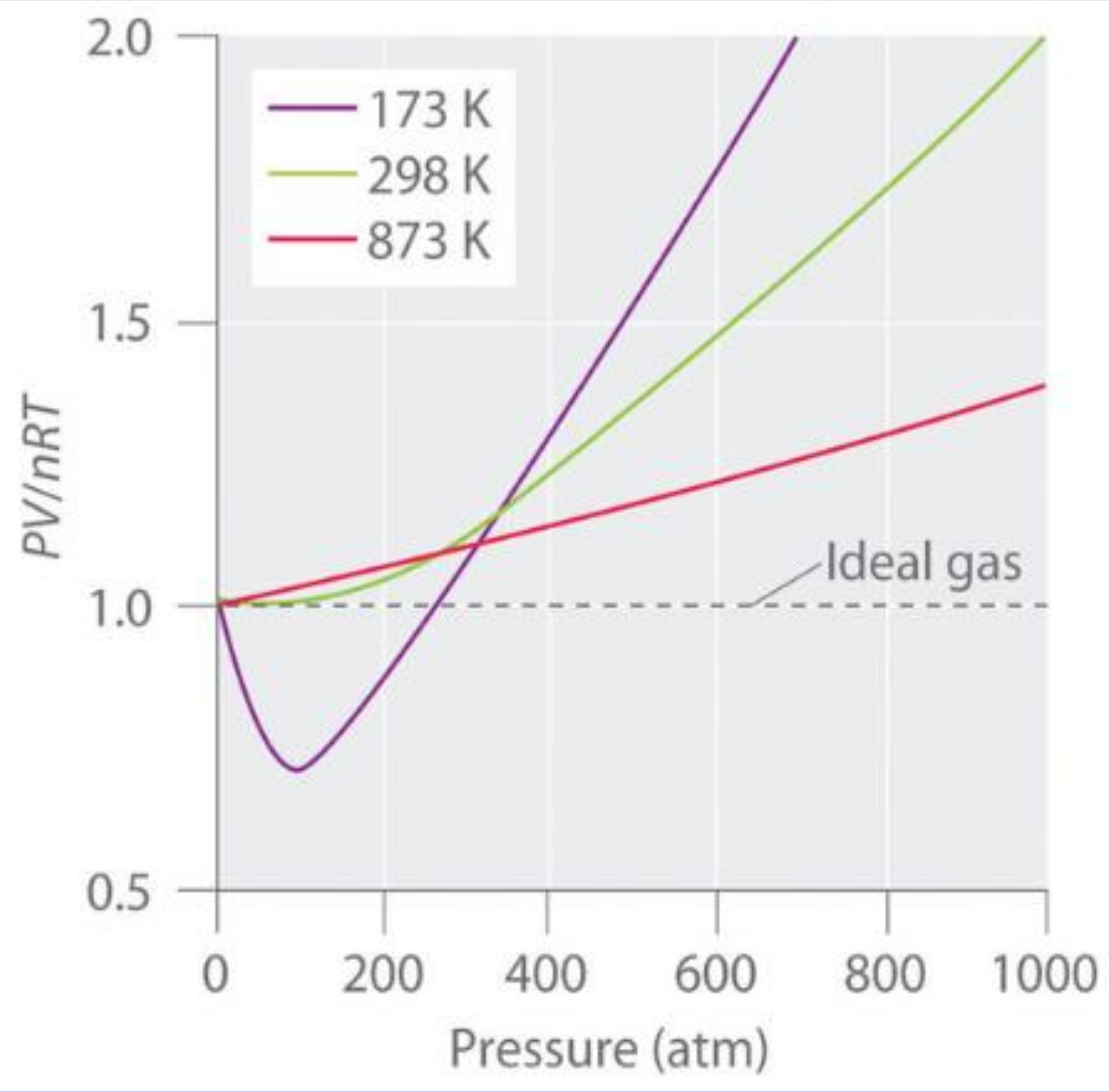
Because of these attractions, a gas can condense, or even solidify, when it is compressed or cooled.

Therefore LOW temperatures also cause deviations from the ideal gas law.

Ideal Gases & Real Gases

Ideal gas behavior exists at STP (dotted line).

LOW temperatures (molecules move slower … more attractions) and HIGH pressures (more collisions in less volume) show significant deviations from the ideal gas law.



Ideal Gases

Boyle's Law : volume, V $1/\propto$ pressure, P

Charles' Law: volume, V $\propto$ temperature, T

Gay-Lussac's Law: Pressure, P $\propto$ temperature, T

Avogadro's Law: volume, V $\propto$ number of moles, n, of gas

So, we could say: $V \propto nT/P$

To change a proportionality relationship into an equality, just requires multiplying by a constant, R.

$$V = R\,(nT/P) \quad \text{or} \quad V = RnT/P$$

We rearrange this equation to form:

$$PV = nRT$$

Ideal Gas Law Problem

Determine the number of moles of air present in 1.35 L at 100. kPa and 17.0° C. (R = *8.317 liter•kPa / mol•K).*

Ideal Gas Law Problem

Determine the number of moles of air present in 1.35 L (V) at 100. kPa (P) and 17.0 ºC (T). (R = *8.317 liter•kPa / mol•K).*

T = 17.0º C + 273. = 290. K

$$PV = nRT$$

$$n = \frac{P \times V}{R \times T}$$

$$n = \frac{(100.\ \text{kPa})(1.35\ \text{L})}{(8.317\ \text{liter}\cdot\text{kPa} / \text{mol}\cdot\text{K})(290.\ \text{K})}$$

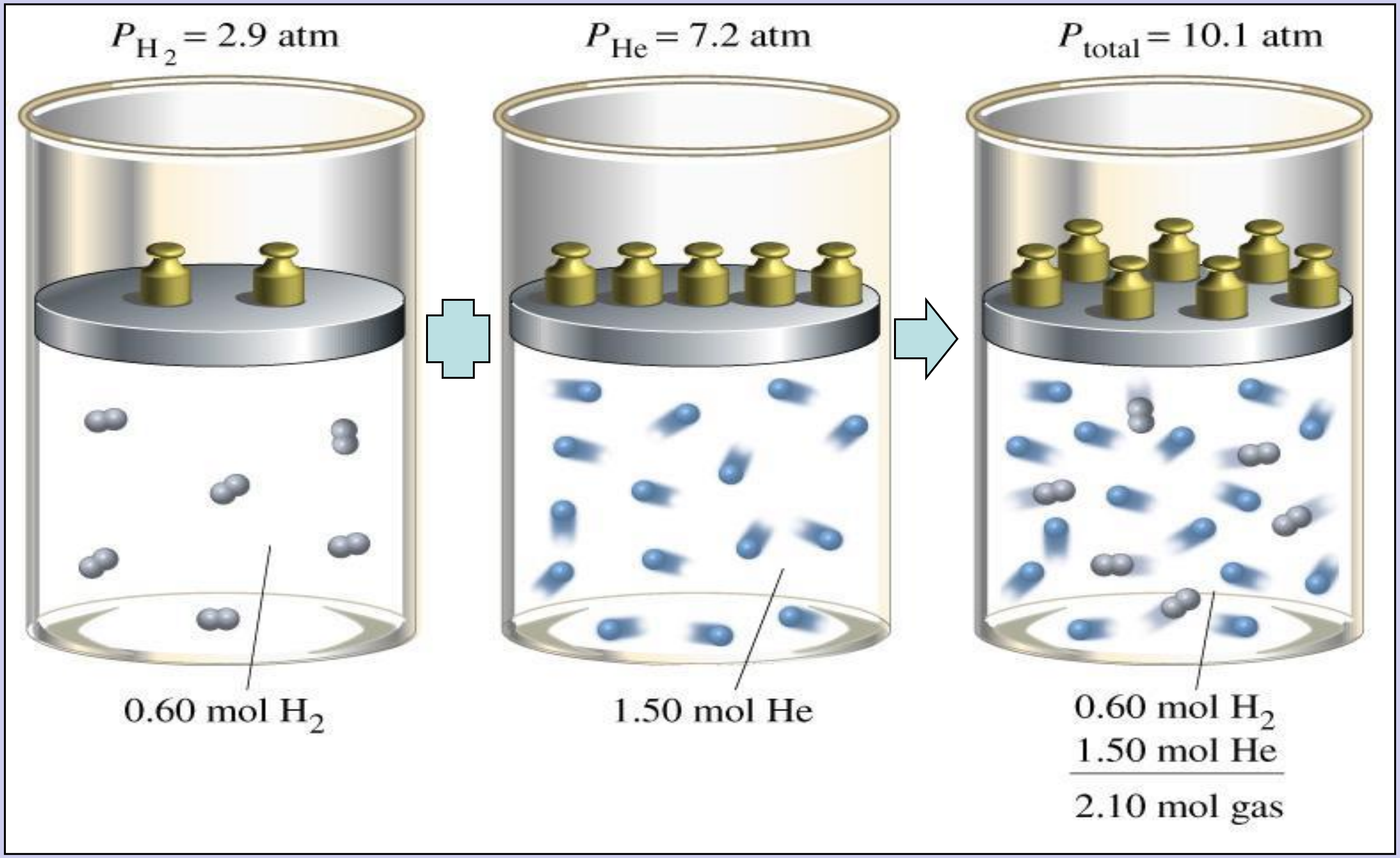
n = 0.0564 mol

Dalton's Law of Partial Pressures

When two or more ideal gases are mixed together, the total pressure of the mixture is equal to the sum of the pressures of each individual gas.

$$P_{Total} = P_1 + P_2 + P_3 + \ldots$$

The pressure of an ideal gas does not depend on the identity of the gas.
It depends only on the quantity of that gas!

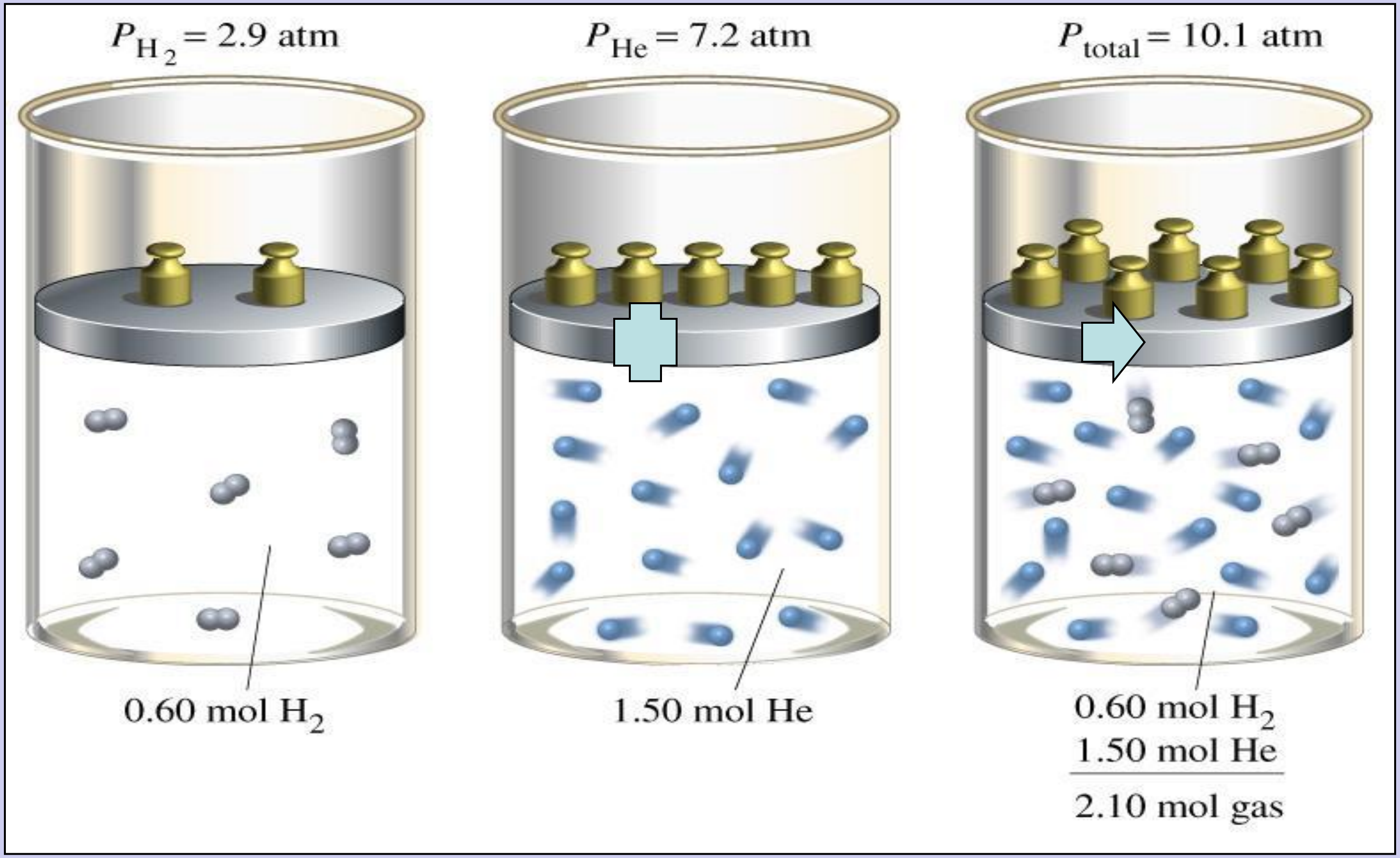


Dalton's Law of Partial Pressures

Mole Fraction

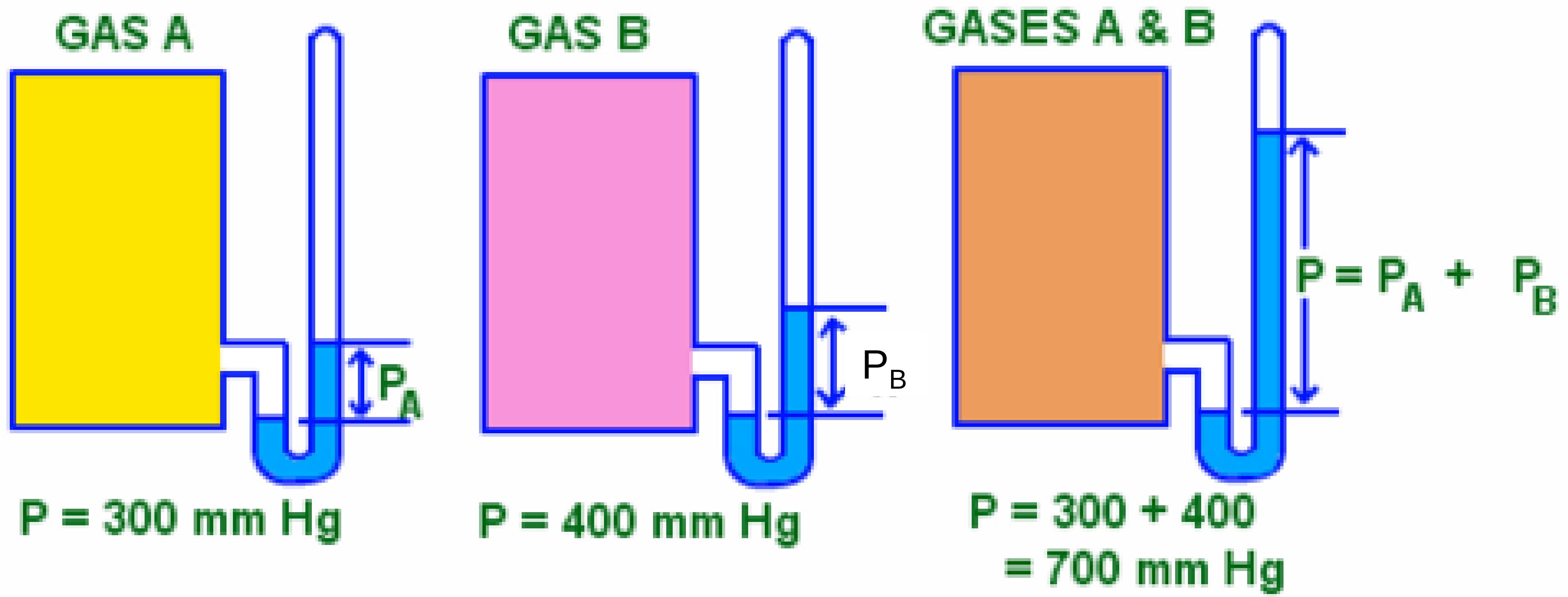
0.60 mol / 2.10 mol H_2 (g)

1.50 mol / 2.10 mol He (g)



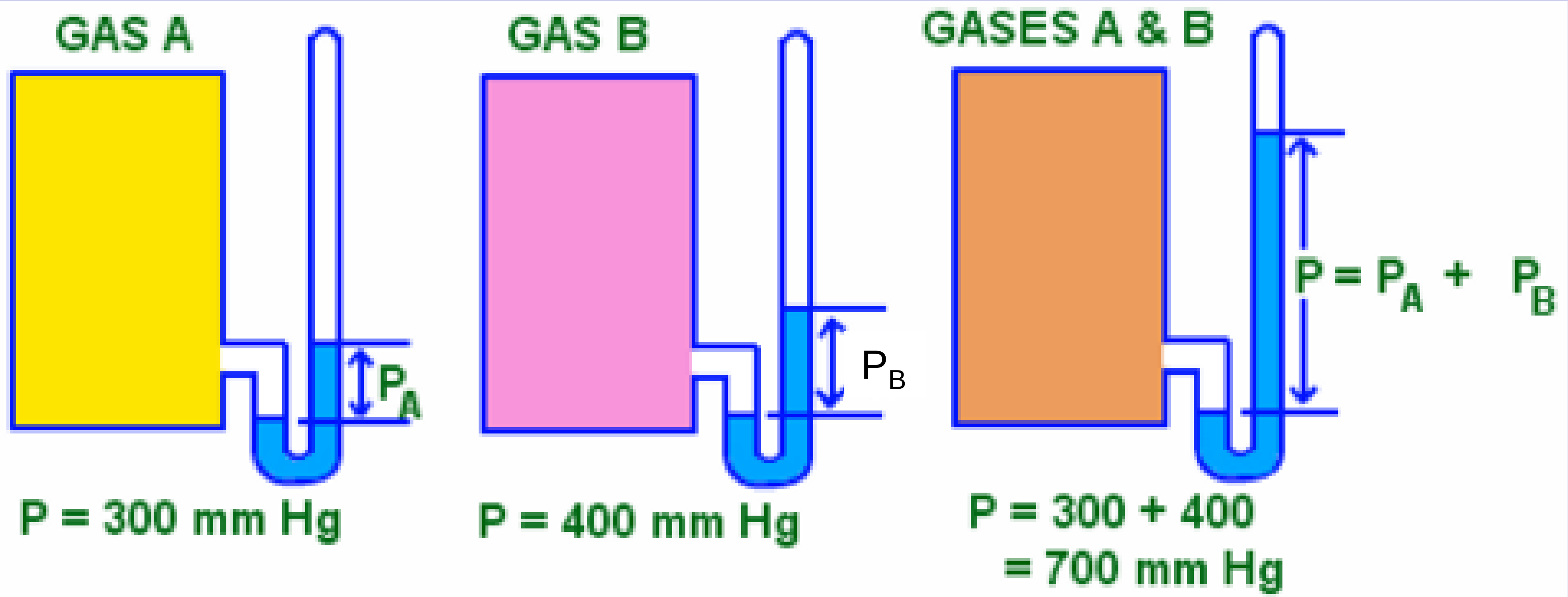
Dalton's Law of Partial Pressures

What is the total pressure of the gases when combined?



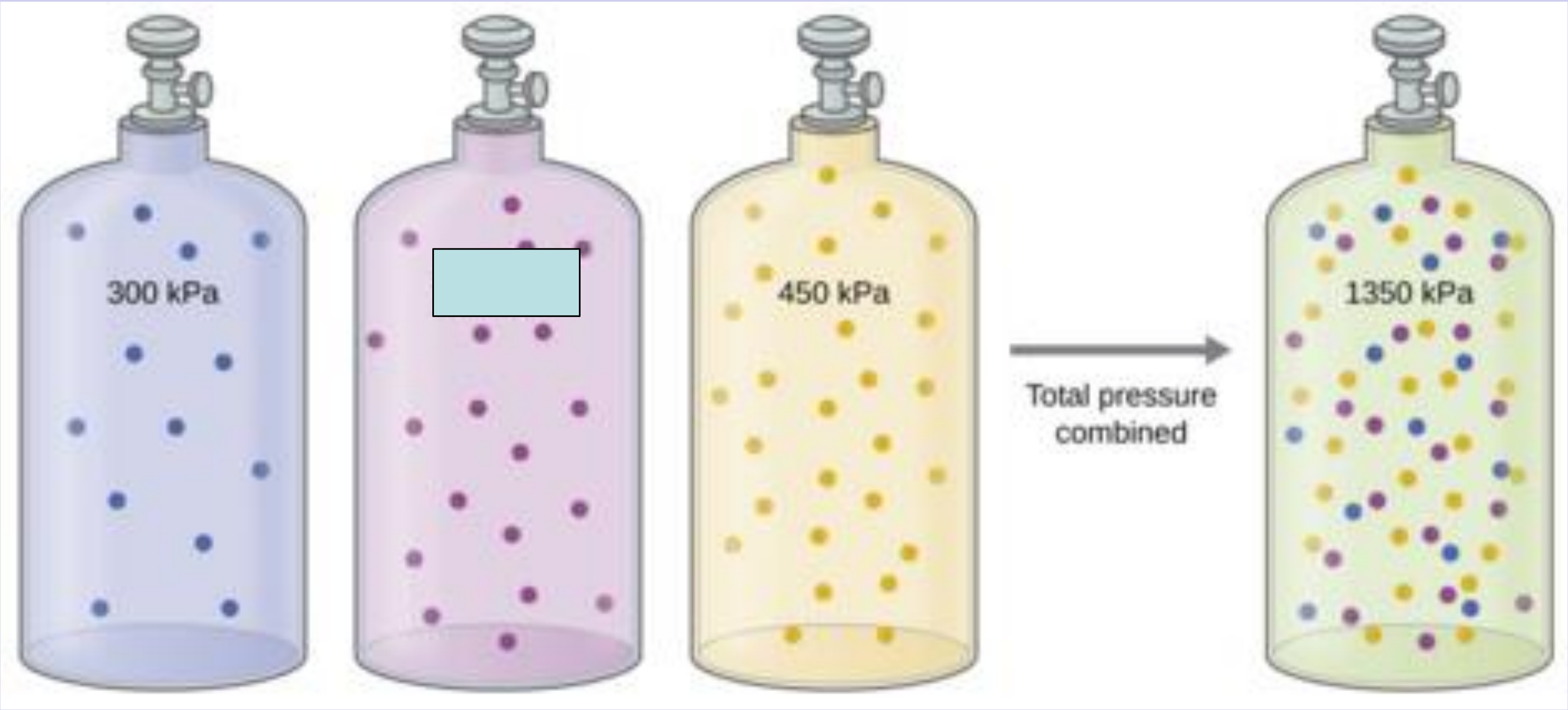
Dalton's Law of Partial Pressures

What is the total pressure of the gases when combined?



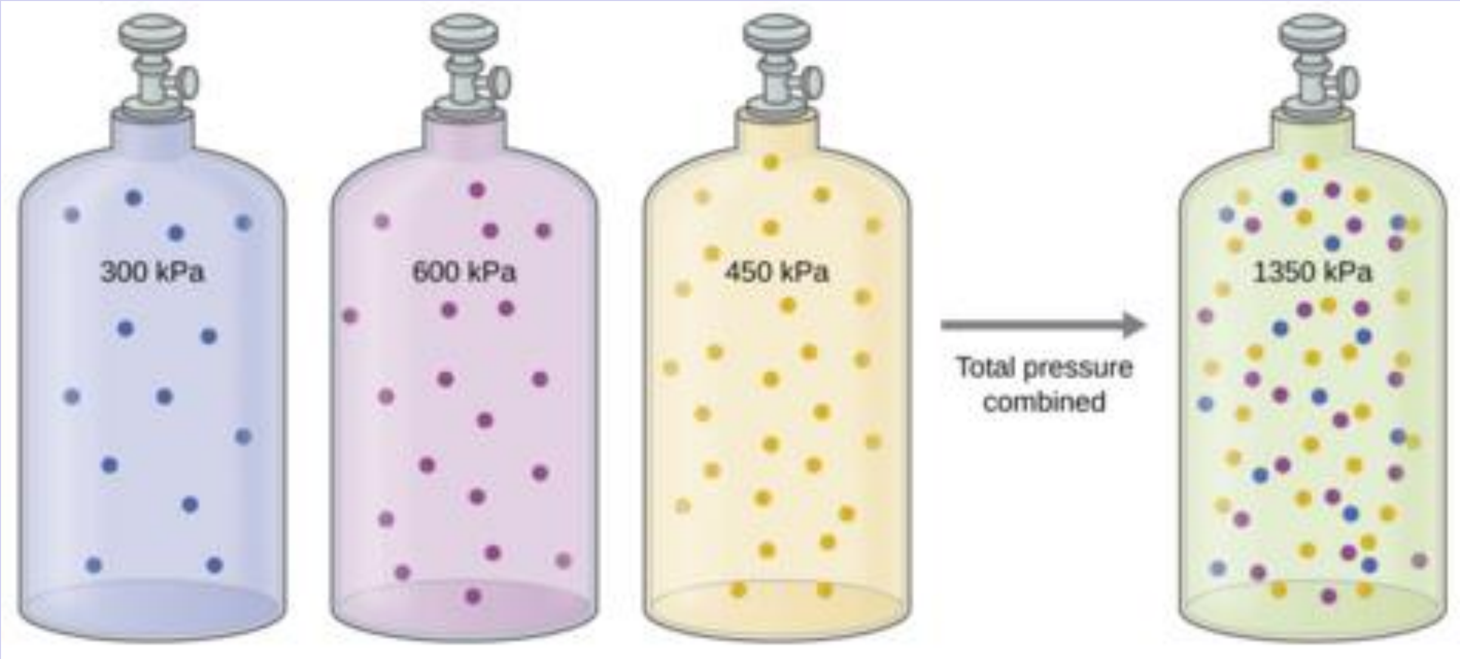
Dalton's Law of Partial Pressures

What is the pressure of the gas in the middle container?



Dalton's Law of Partial Pressures

What is the pressure of the gas in the middle container?



$$P_T = P_1 + P_2 + P_3$$

$$P_T - P_1 - P_3 = P_2$$

$$P_2 = 1350 – 450 – 300$$

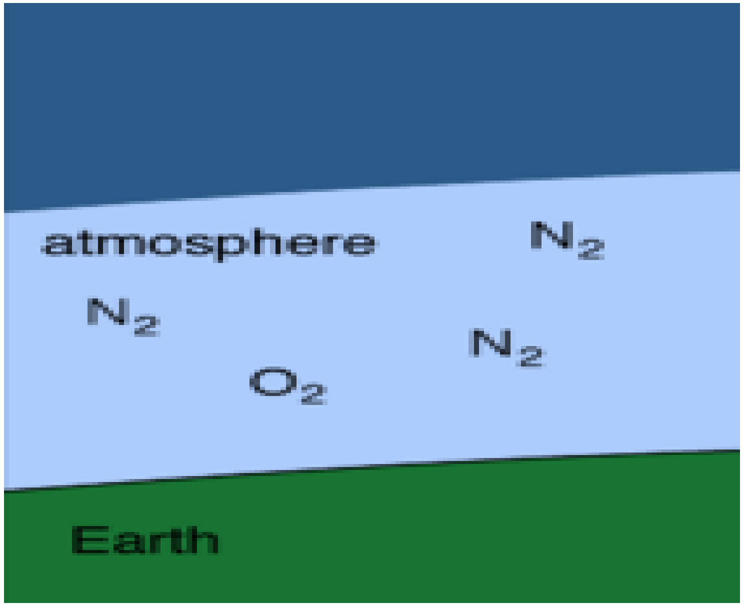
$$P_2 = 600\text{ kPa}$$

Partial pressure:

The fraction of the total pressure exerted by a mix of gases that is contributed by an individual gas

Partial pressure:

- refers to pressure of one gas in a mixture of gases
- the pressure of a single gas as if it occupied the container by itself
- the sum of all partial pressures equals total pressure of the mixture



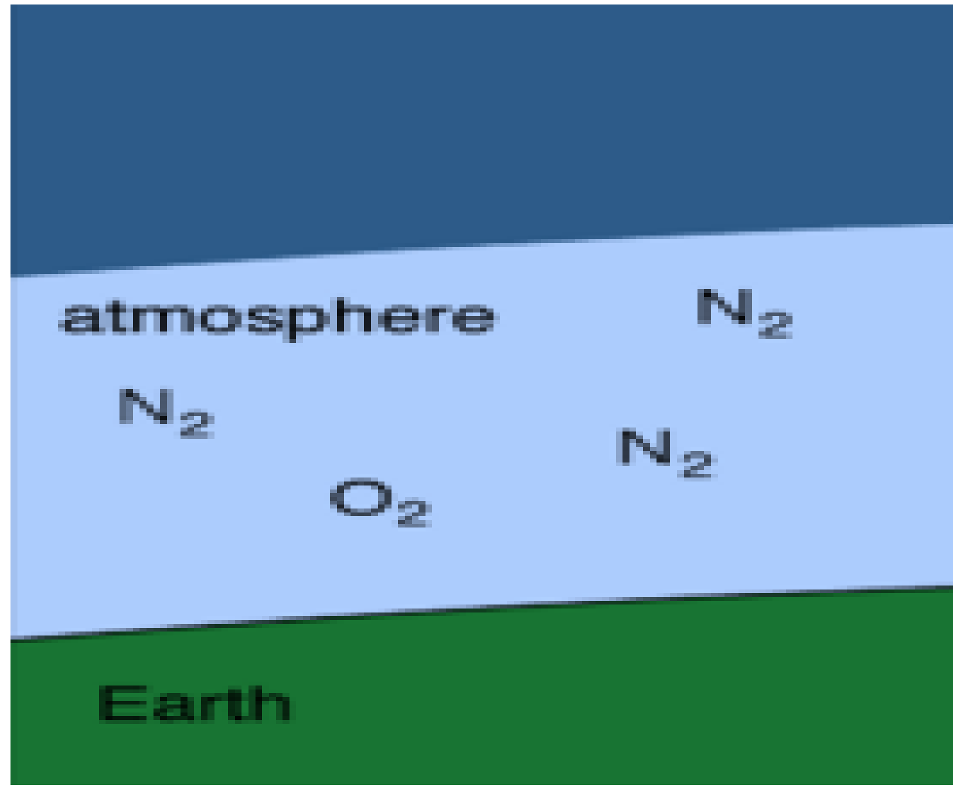
Partial pressure applies when there are two or more gases in the same container. Each gas contributed to the total pressure of the system.

Partial pressure:

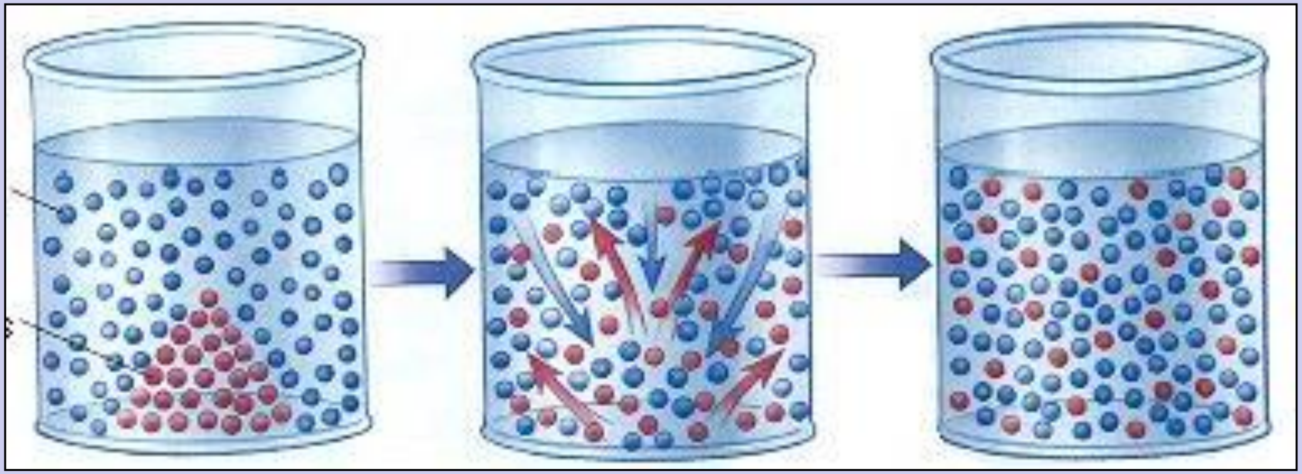
The fraction of the total pressure exerted by a mix of gases that is contributed by an individual gas

Air

N_2 (g):	78.08%	~593 torr
O_2 (g):	20.95%	~159 torr
Ar (g)	0.93%	~7 torr
mixture:	100%	760 torr



Gases exhibit:

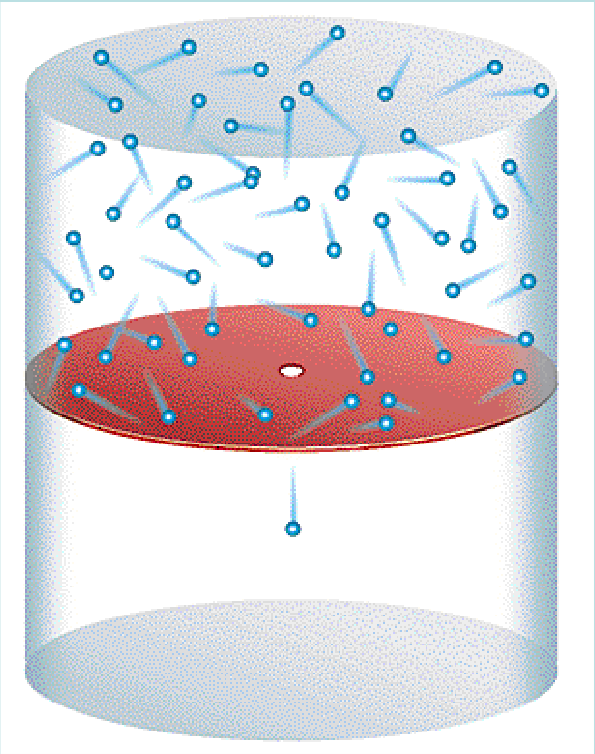


Diffusion

the gradual mixing of molecules of different gases from HIGH to LOW concentration.

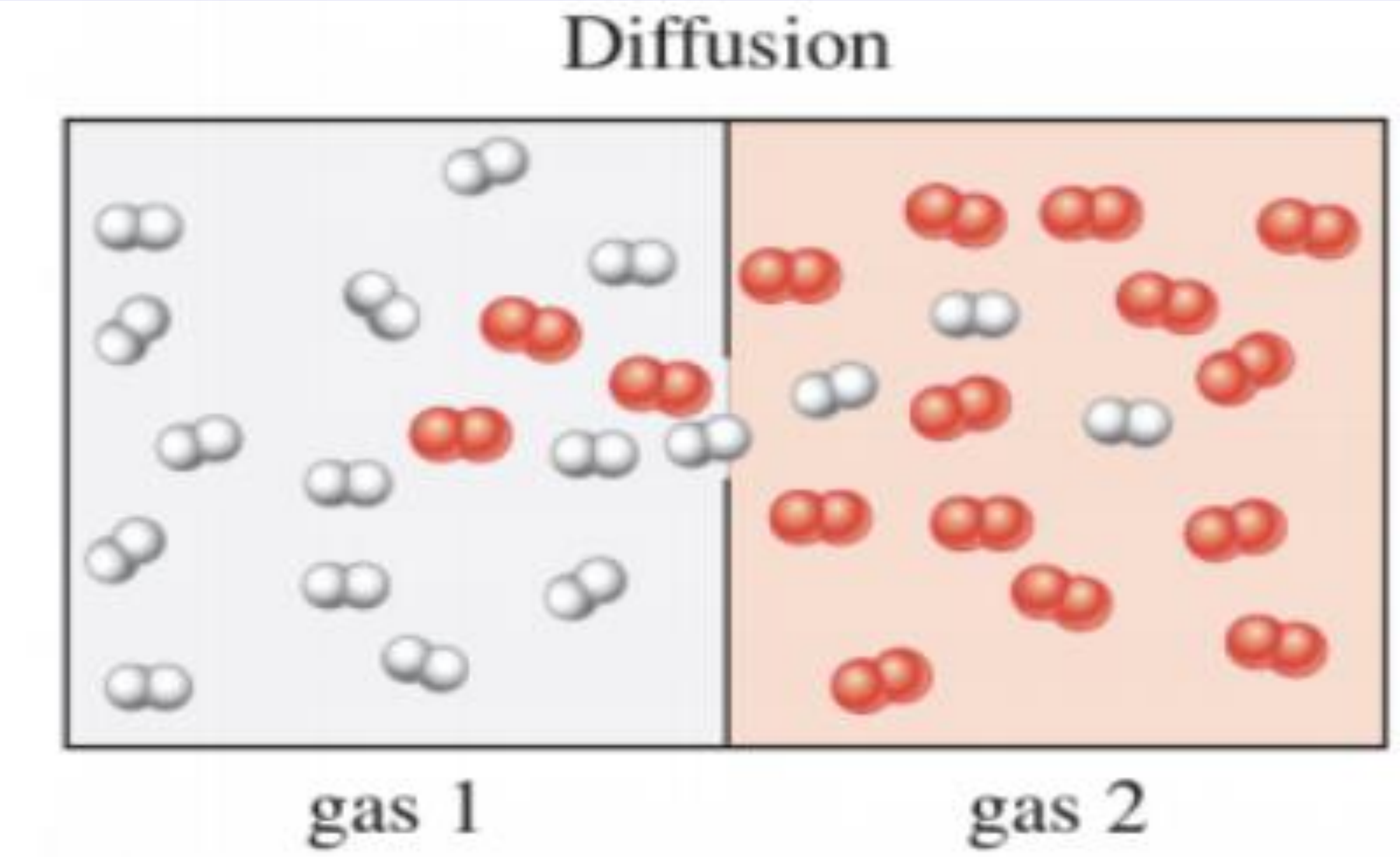
Effusion

the movement of molecules through a small hole into an empty container.



Graham's Law

Gases diffuse at different rates.



Effusion

Graham's Law

A gas escapes through a tiny hole in its container.

Gases of **lower molar mass** diffuse and effuse **faster** than gases of higher molar mass.

Graham's law of effusion

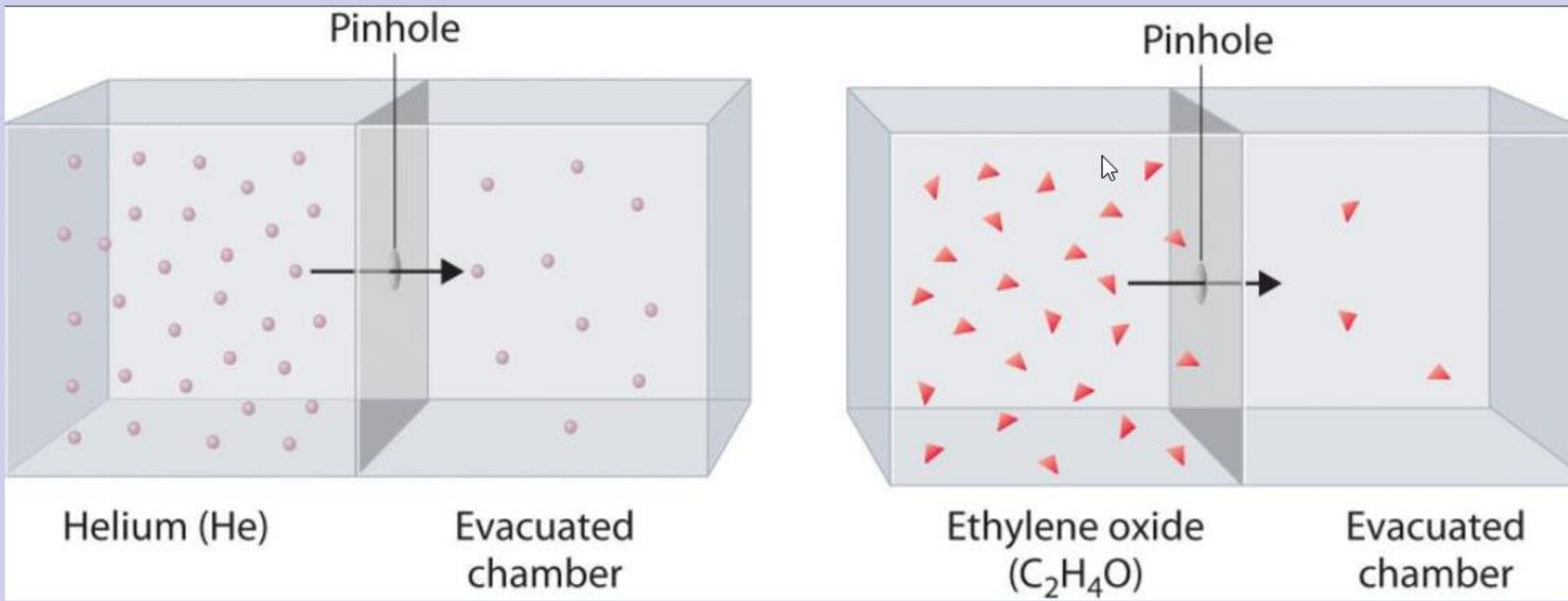
the rate of effusion of a gas is inversely proportional to the square root of the gas's molar mass.

The equation below is ENRICHMENT

$$\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$$

Graham's Law

Helium gas (less massive) Effuses much quicker than Ethylene oxide gas (more massive)



Graham's Law

Which of the following gas particles will diffuse fastest if all of these gases are at the same temperature and pressure?

SO_2 **N_2O**

Hg **Cl_2**

- Ammonia gas (NH_3) & Hydrochloric Acid (HCl)
- Place a cotton ball soaked in each compound at opposite ends of the cylinder.
- A precipitate (smoke ring) will form due to the reaction of ammonia and hydrochloric acid.
- Where will the smoke ring form?

17 g/mole 36 g/mole

Graham's Law

N_2O will diffuse fastest of these gases at the same temperature and pressure because its molar mass is least.

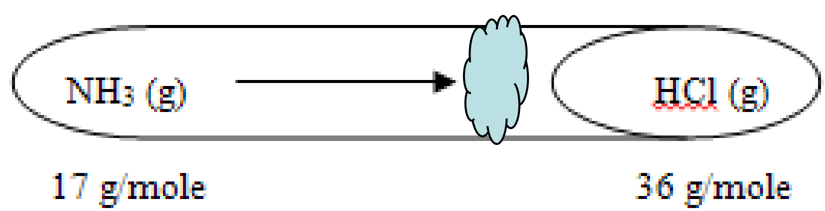
SO_2 (64 g/mol)

Hg (~201 g/mol)

N_2O (44 g/mol) … least mass

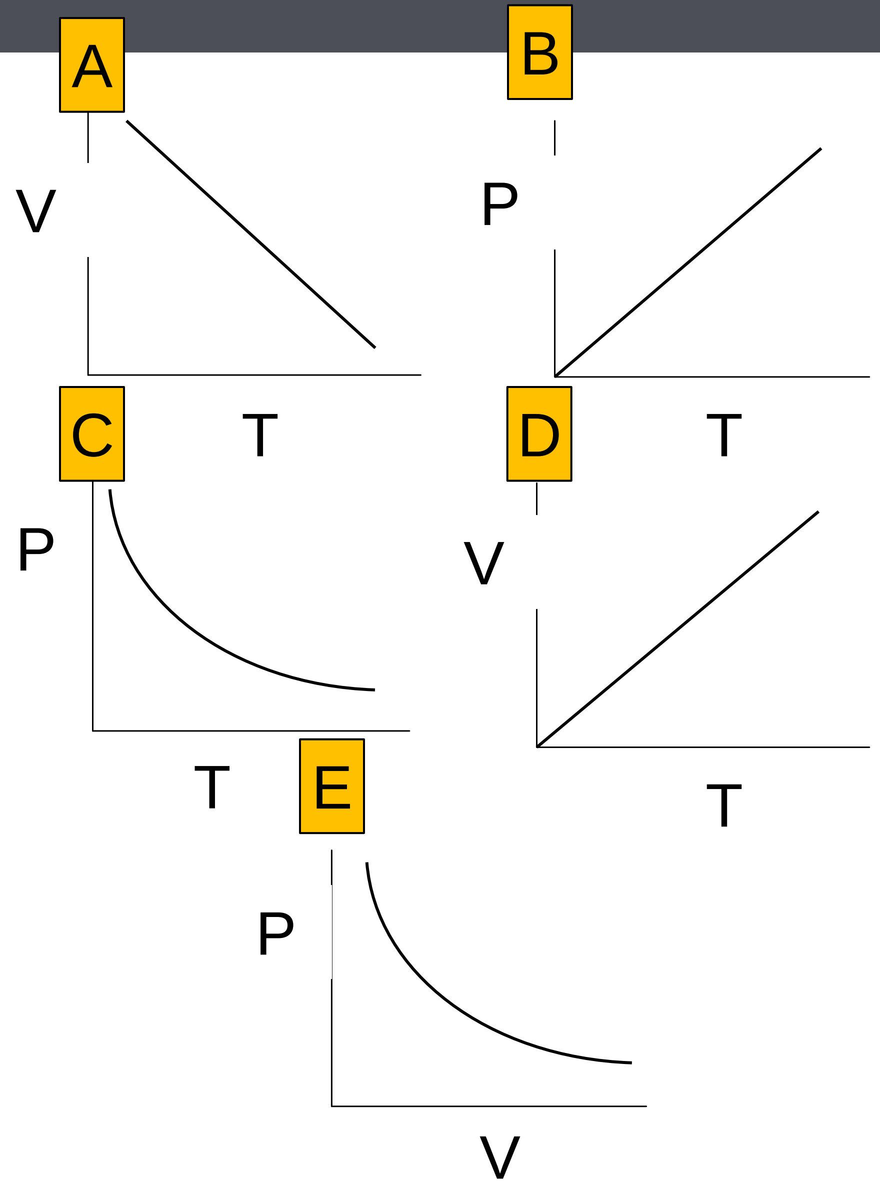
Cl_2 (~71 g/mol)

- Ammonia gas (NH_3) & Hydrochloric Acid (HCl)
- Place a cotton ball soaked in each compound at opposite ends of the cylinder.
- A precipitate (smoke ring) will form due to the reaction of ammonia and hydrochloric acid.
- A smoke ring should form closer to the HCl flask (*HCl has a GMM of 36 g/mole while NH_3 has a GMM of 17 g/mole)*



Practice matching graphed relationships to the gas laws.

Match one graph shown at right to the gas laws named below



Charles's law

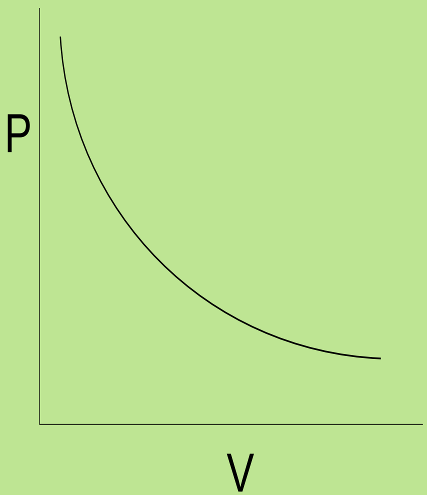
Gay-Lussac's law

Boyle's law

Three Gas Laws

E

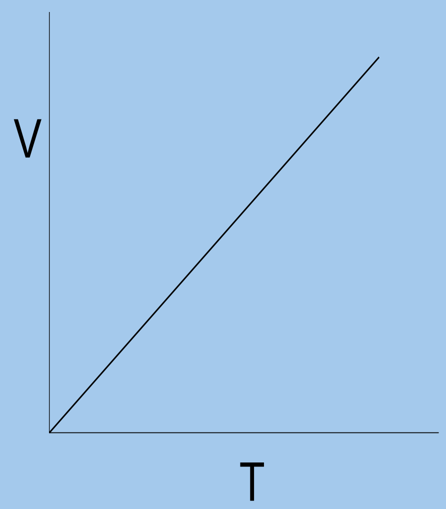
Boyle's law



$$P_1V_1 = P_2V_2$$

D

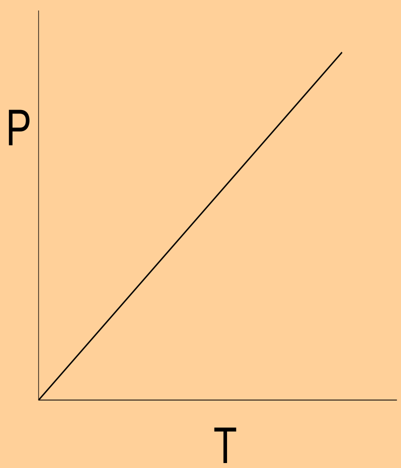
Charles's law



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

B

Gay-Lussac's law



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

A gas is stored in a 3.00 liter container at a pressure of 456 torr. If the temperature stays constant, what will the pressure of the gas be if the volume is decreased to 0.75 liters?

Torr = mm Hg

V_1 P_1

A gas is stored in a 3.00 liter container at a pressure of 456 torr. If the temperature stays constant, what will the pressure of the gas be if the volume is decreased to 0.75 liters?

P_2 V_2

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

But temperature stays constant,
so denominators cancel out, leaving

Boyle's Law → $P_1V_1 = P_2V_2$

$$P_2 = P_1V_1 / V_2$$

$$P_2 = (456 \text{ torr}) \times (3.00 \text{ L}) / 0.75 \text{ L}$$

= 1800 torr **… 2 sig figs**

= 1.8×10^3 torr

A balloon contains 30.0 L of helium gas at 103 kPa. What is the volume of the helium when the balloon rises to an altitude where the pressure is only 25.0 kPa? (Assume that the temperature remains constant.)

Boyle's Law Problem 2

A balloon contains 30.0 L (V_1) of helium gas at 103 kPa (P_1) . What is the **volume** (V_2) of the helium when the balloon rises to an altitude where the pressure is only 25.0 kPa (P_2)? (Assume that the temperature remains constant.)

$$P_1 \times V_1 = P_2 \times V_2$$

Isolate V_2 by dividing both sides by P_2:

$$\frac{P_1 \times V_1}{P_2} = \frac{\cancel{P_2} \times V_2}{\cancel{P_2}}$$

$$V_2 = \frac{V_1 \times P_1}{P_2}$$

$$V_2 = \frac{30.0 \text{ L} \times 103 \cancel{\text{kPa}}}{25.0 \cancel{\text{kPa}}}$$

$$V_2 = 1.24 \times 10^2 \text{ L}$$

A sample of neon gas occupies a volume of 677 mL at 134 kPa. What is the pressure of the sample if the volume is decreased to 642 mL? [Assume constant temperature]

Boyle's Law Problem 3

A sample of neon gas occupies a volume of 677 mL (V_1) at 134 kPa (P_1). What is the **pressure** (P_2) of the sample if the volume (V_2) is decreased to 642 mL? [Assume constant temperature]

$$P_1 \times V_1 = P_2 \times V_2$$

$$P_2 = \frac{V_1 \times P_1}{V_2}$$

$$P_2 = \frac{677\ \cancel{mL} \times 134\ kPa}{642\ \cancel{mL}}$$

$$P_2 = 141\ \text{kPa}$$

A balloon inflated in a room at 24.0^oC has a volume of 4.00 L. The balloon is then heated to a temperature of 58.0^oC. What is the new volume if the pressure remains constant?

Charles' Law Problem 1

A balloon inflated in a room at 24.0°C (T_1) has a volume of 4.00 L (V_1). The balloon is then heated to a temperature of 58.0°C (T_2). What is the new **volume** (V_2) if the pressure remains constant?

$T_1 = 24.0^oC + 273 = 297\ K$ $T_2 = 58.0^oC + 273 = 331\ K$

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

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

$$V_2 = \frac{4.00\ L \times 331\ K}{297\ K}$$

$$V_2 = 4.46\ L$$

What is the temperature of a 2.3 L balloon if it shrinks to a volume of 0.632 L when it is dipped into liquid nitrogen at a temperature of 77 K?

Charles' Law Problem 2

What is the **temperature** (T_1) of a 2.3 L (V_1) balloon if it shrinks to a volume of 0.632 L (V_2) when it is dipped into liquid nitrogen at a temperature of 77 K (T_2)?

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

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

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

$$T_1 = \frac{2.3\ L \times 77\ K}{0.632\ L}$$

$$T_1 = 280\ K$$

Aerosol cans carry labels warning not to incinerate (burn) the cans or store them above a certain temperature. The gas in a used aerosol can is at a pressure of 103 kPa at 25.0^oC. If the can is thrown onto a fire, what will the pressure be when the temperature reaches 928^oC?

Gay-Lussac's Law Problem 1

Aerosol cans carry labels warning not to incinerate (burn) the cans or store them above a certain temperature. The gas in a used aerosol can is at a pressure of 103 kPa (P_1) at 25.0°C (T_1). If the can is thrown onto a fire, what will the **pressure** (P_2) be when the temperature reaches 928°C (T_2)?

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

$$T_1 = 25^{o}C + 273 = 298\ K$$

$$T_2 = 928^{o}C + 273 = 1201\ K$$

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

$$P_2 = \frac{103\ kPa \times 1201\ \cancel{K}}{298\ \cancel{K}}$$

$$P_2 = 415\ kPa$$

$$P_2 = 4.15 \times 10^2\ kPa$$

A pressure cooker containing a roast and some water starts at 298 K and 101 kPa. The cooker is heated, and the pressure increases to 136 kPa. What is the final temperature inside the cooker?

Gay-Lussac's Law Problem 2

A pressure cooker containing a roast and some water starts at 298 K (T_1) and 101 kPa (P_1). The cooker is heated, and the pressure increases to 136 kPa (P_2). What is the final **temperature** (T_2) inside the cooker?

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

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

$$T_2 = \frac{136\ \text{kPa} \times 298\ \text{K}}{101\ \text{kPa}}$$

$$T_2 = 401\ \text{K}$$

The volume of a gas-filled balloon is 30.0 L at 313 K and 153 kPa pressure. What would the volume be at standard temperature and pressure (STP)?

Combined Gas Law Problem 1

The volume of a gas-filled balloon is 30.0 L (V_1) at 313 K (T_1) and 153 kPa (P_1) pressure. What would the **volume** (V_2) be at standard temperature (T_2) and pressure (STP) (P_2) ?

STP
273 K
101.3 kPa

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

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

$$V_2 = \frac{30.0\ \text{L} \times 153\ \text{kPa} \times 273\ \text{K}}{101.3\ \text{kPa} \times 313\ \text{K}}$$

$$V_2 = 39.5\ \text{L}$$

Which of the following equations could be used to correctly calculate the final temperature of a gas?

A. $T_2 = \dfrac{V_1 \times P_1 \times V_2}{P_2 \times T_1}$

B. $T_2 = \dfrac{V_2 \times P_2 \times T_1}{V_1 \times P_1}$

C. $T_2 = \dfrac{V_2 \times P_1 \times T_1}{V_1 \times P_2}$

D. $T_2 = \dfrac{V_1 \times P_1 \times T_1}{V_2 \times P_2}$

Combined Gas Law Problem 2

Which of the following equations could be used to correctly calculate the final temperature of a gas?

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

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

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

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

At 34.0^{o}C, the pressure inside a nitrogen-filled tennis ball with a volume of 0.148 L is 212 kPa. How many moles of nitrogen gas are in the tennis ball? (R = 8.317 L·kPa/Mol·K)

Ideal Gas Law Problem 1

At 34.0°C (**T**) , the pressure inside a nitrogen-filled tennis ball with a volume of 0.148 L (**V**) is 212 kPa (**P**) . How many **moles** (**n**) of nitrogen gas are in the tennis ball? (R = 8.317 L·kPa/Mol·K)

$T = 34^oC + 273 = 307\ K$

$$PV = nRT$$

$$n = \frac{P \times V}{R \times T}$$

$$n = \frac{212\ \text{kPa} \times 0.148\ \text{L}}{8.31\ (\text{L}\cdot\text{kPa})\ /\ (\text{K}\cdot\text{mol}) \times 307\ \text{K}}$$

$$n = 1.23 \times 10^{-2}\ \text{mol}\ N_2$$

Air contains oxygen, nitrogen, carbon dioxide, & trace amounts of other gases. What is the partial pressure of oxygen (P_{O_2}) at 101.30 kPa of total pressure if the partial pressures of nitrogen, carbon dioxide, and other gases are 79.10 kPa, 0.040 kPa, and 0.94 kPa, respectively?

Dalton's Law of Partial Pressures 1

Air contains oxygen, nitrogen, carbon dioxide, and trace amounts of other gases. What is the **partial pressure** of oxygen *(P_{O_2})* at 101.30 kPa *(P_T)* of total pressure if the partial pressures of nitrogen, carbon dioxide, and other gases are 79.10 kPa *(P_{N_2})* , 0.040 kPa *(P_{CO_2})* , and 0.94 kPa *(P_{Other})* , respectively?

Start with Dalton's law of partial pressures.

$$P_{total} = P_{O_2} + P_{N_2} + P_{CO_2} + P_{others}$$

Rearrange Dalton's law to isolate P_{O_2}.

$$P_{O_2} = P_{total} - (P_{N_2} + P_{CO_2} + P_{others})$$

Substitute the values for the total and the known partial pressures, and solve the equation.

$$P_{O_2} = 101.30 \text{ kPa} - (79.10 \text{ kPa} + 0.040 \text{ kPa} + 0.94 \text{ kPa})$$

$$P_{O_2} = 21.22 \text{ kPa}$$

A tank used by scuba divers has a P_{total} of 2.21×10^4 kPa. If P_{N_2} is 1.72×10^4 kPa and P_{O_2} is 4.641×10^3 kPa, what is the partial pressure of any other gases in the scuba tank (P_{other})?

Dalton's Law of Partial Pressures 2

A tank used by scuba divers has a P_{total} of 2.21×10^4 kPa. If P_{N_2} is 1.72×10^4 kPa and P_{O_2} is 4.641×10^3 kPa, what is the **partial pressure** of any **other gases** in the scuba tank (P_{other})?

$$P_{total} = P_{O_2} + P_{N_2} + P_{others}$$

$$P_{others} = P_{total} - (P_{N_2} + P_{O_2})$$

$$P_{others} = 2.21 \times 10^4 \text{ kPa} - (1.72 \times 10^4 \text{ kPa} + 4.641 \times 10^3 \text{ kPa})$$

$$P_{others} = 2.59 \times 10^2 \text{ kPa}$$