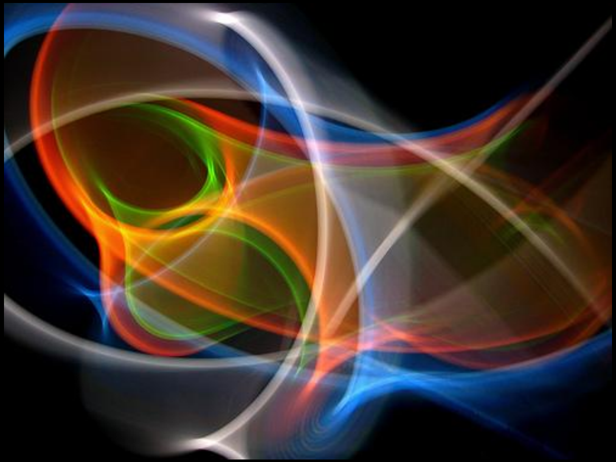



Gases

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State of Matter - Gas (Vapor)

Form	Fluid (Flows)
Compressibility	Very High
Shape	Variable (Fills Closed Container)
Volume	Variable (Fills Closed Container)
Particle Movement	Random, Independent

Example: Steam

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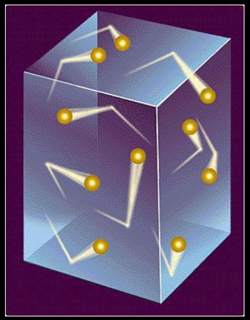
Kinetic (Moving) Theory of Gases

Gases are composed of molecules in constant motion
Gas molecules move in random directions
Molecules of a gas collide frequently with each other & with vessel walls
(why gases mix to uniformity & fill all portions of the containment vessel)

Gas molecules move with an average velocity at a given temperature.
(the average energy of molecules in a gas is the same for all substances)

Distance between gas molecules >> than size of the individual molecules
(why gases can be compressed)

Molecules are perfectly elastic … no energy is lost when molecules collide
(If not-elastic, the temperature of a gas mix would always decrease with time)



Works for "Ideal" Gases:

No molecular interactions 'cause of large separation
Volume of gas molecules insignificant

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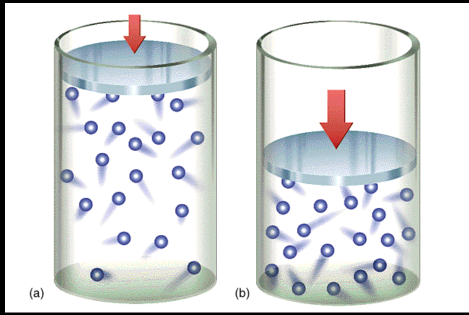
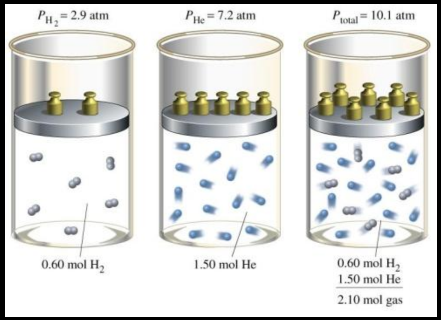
Molecular Explanation of Gas Properties

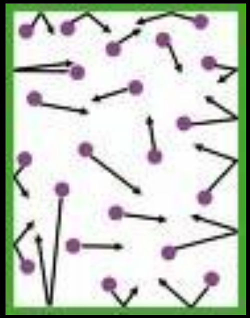
Property	Gas Molecules:
Compressibility	Widely spaced
Low density	Widely spaced
Mixable	Widely spaced In constant, random motion
Fills container	In constant, random motion
Uniform pressure	In constant, random motion No energy loss collisions



Compress



Mix



Fill

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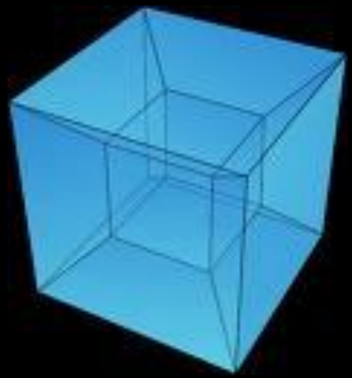


V

Gas behavior is described in terms of:

T

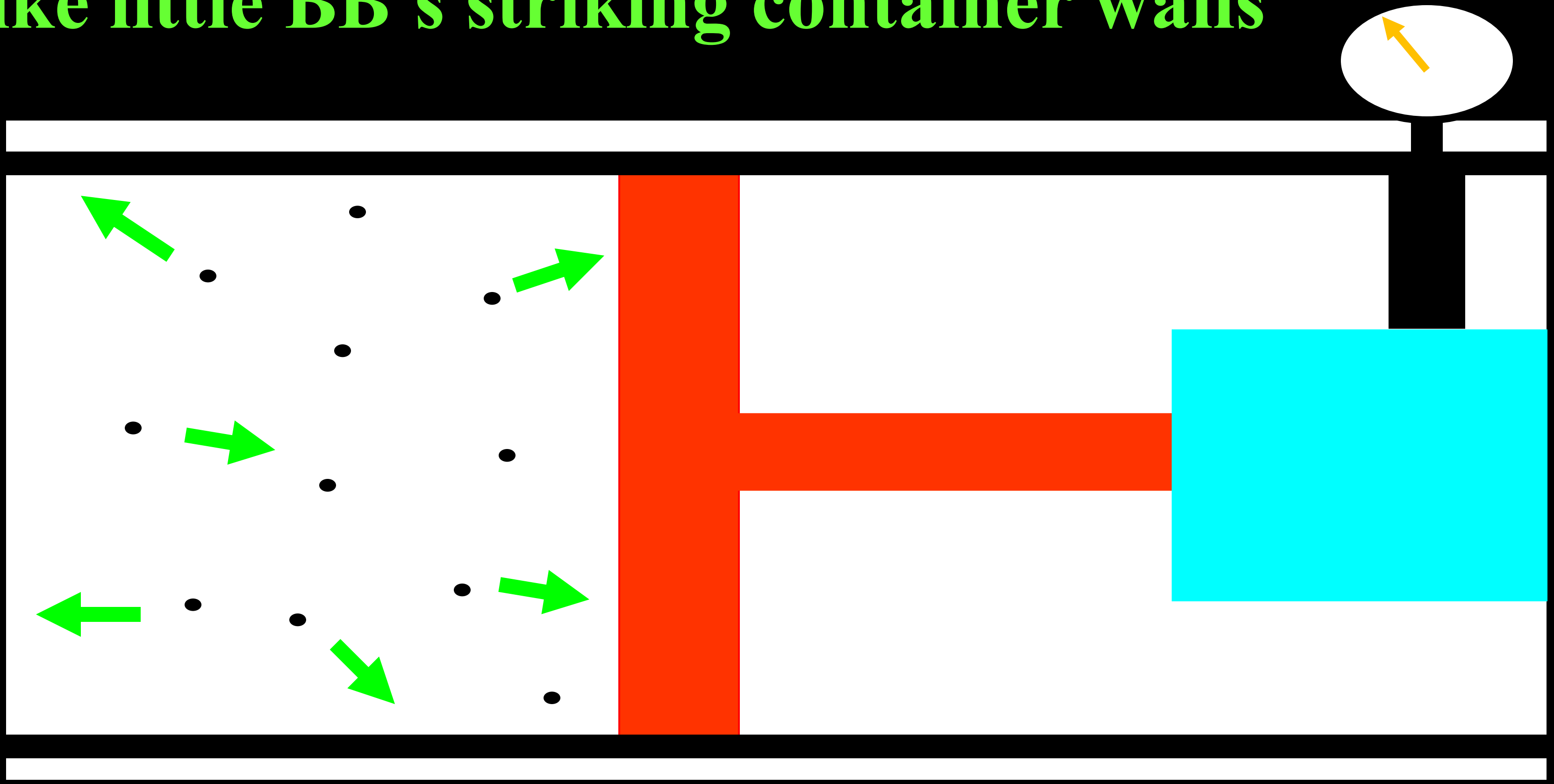
Volume	**(V)**
Pressure	**(P)**
Temperature	**(T)**
Quantity (moles)	**(n)**
# molecules	

P

n

Pressure = molecular impact on container walls

Like little BB's striking container walls



Pressure = force/area

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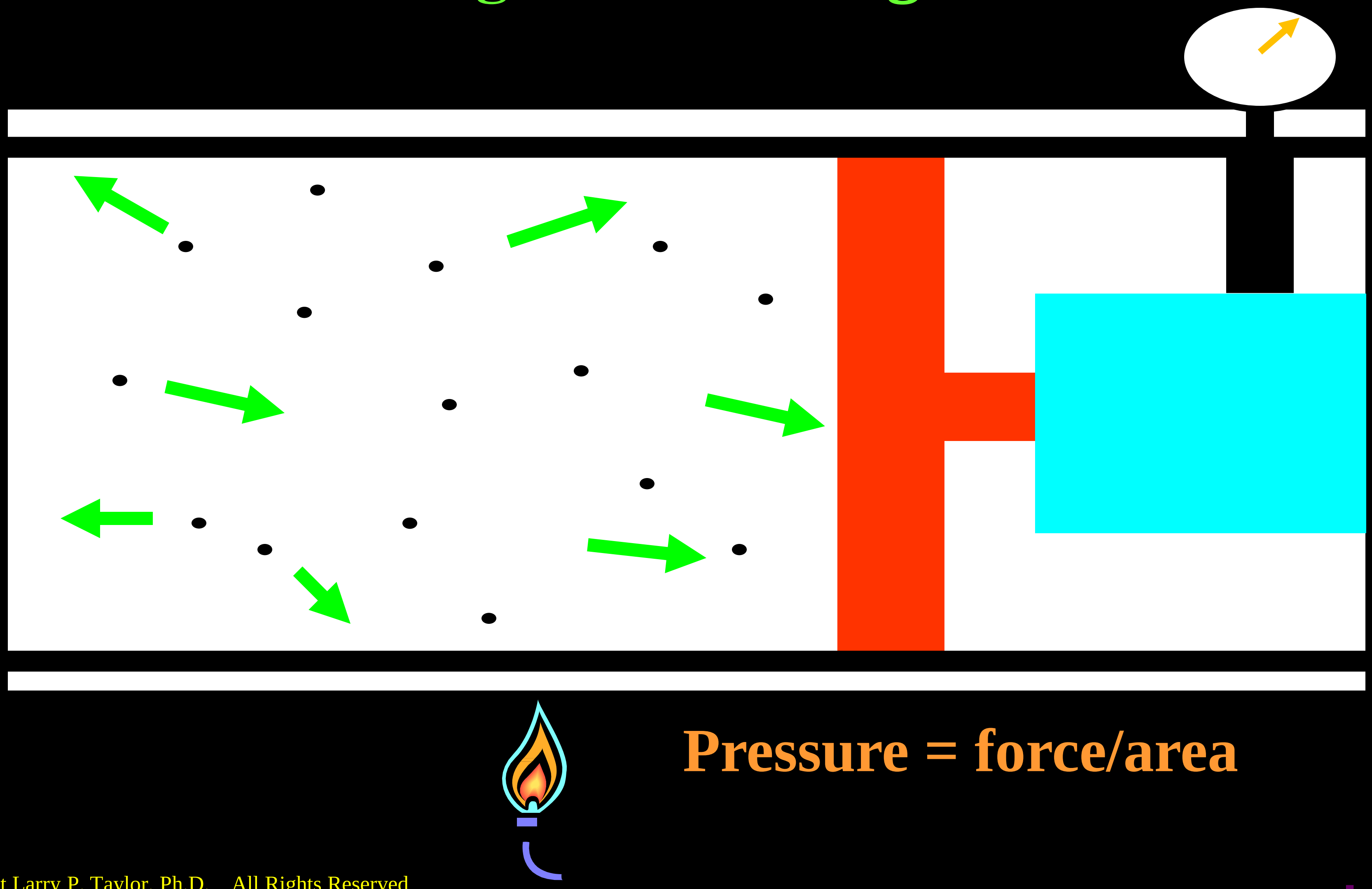
More impacts / time → Increased Pressure

Like faster moving BB's striking container walls



Pressure = force/area

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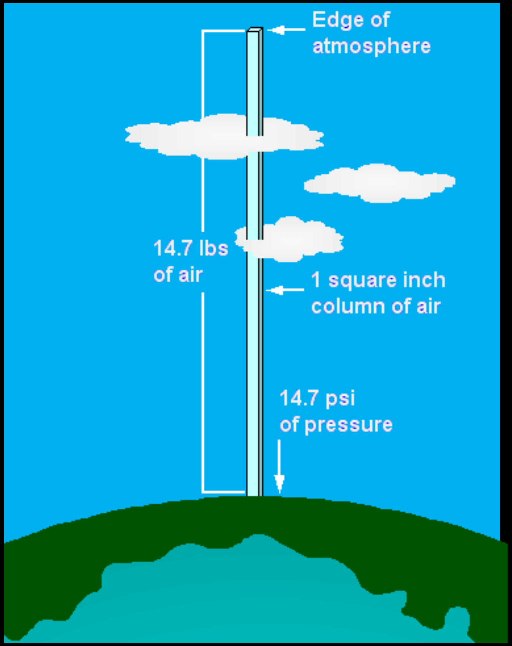
Pressure = Force per Unit Area



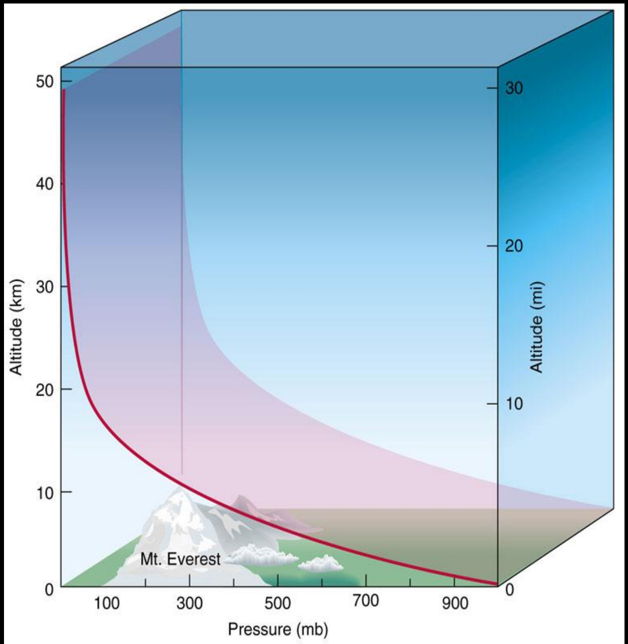
UNITS:

Related to atmosphere

- mm Hg (torr)
- inches Hg
- inches H_2O
- bar (or millibar)
- atm

Related to force

- psi (pounds / in^2)
- Pa (Pascals: newtons / m^2)



Atmospheric pressure = weight of atmosphere

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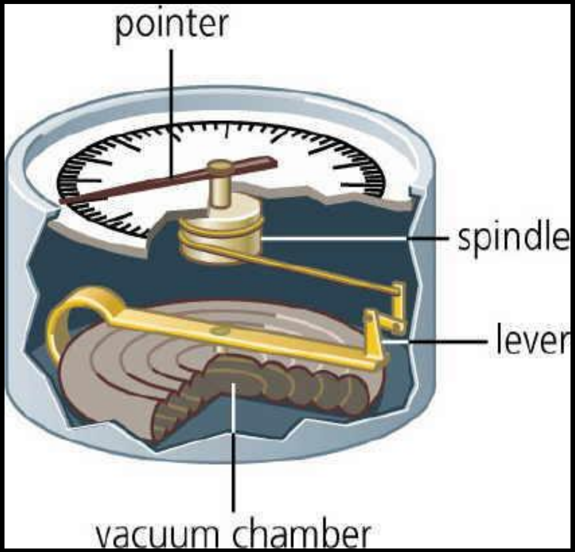
One atmosphere (atm) of Pressure



1 atm equals:

760 millimeters of mercury
760 torr
29.92 inches of mercury
101.3 kilopascals (kPa)
1.01325 bars
14.7 lbs/in^2 (psi)
33 feet of sea water (fsw)
34 feet of fresh water (ffw)
10 meters of sea water (msw)
10.4 meters of fresh water (mfw)

Barometer measures weight of atmosphere

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Converting Pressure Measurements

NOAA reports barometric pressure in mm of mercury.
What is atmospheric pressure in inches of mercury?

$$760.0 \text{ mm Hg} \times \frac{1 \text{ m}}{1000 \text{ mm}} \times \frac{100 \text{ cm}}{1 \text{ m}} \times \frac{1 \text{ in}}{2.54 \text{ cm}} = 29.92 \text{ in Hg}$$

What is the atmospheric pressure in bar?

$$760.0 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} \times \frac{1.01325 \text{ bar}}{1 \text{ atm}} = 1.013 \text{ bar}$$

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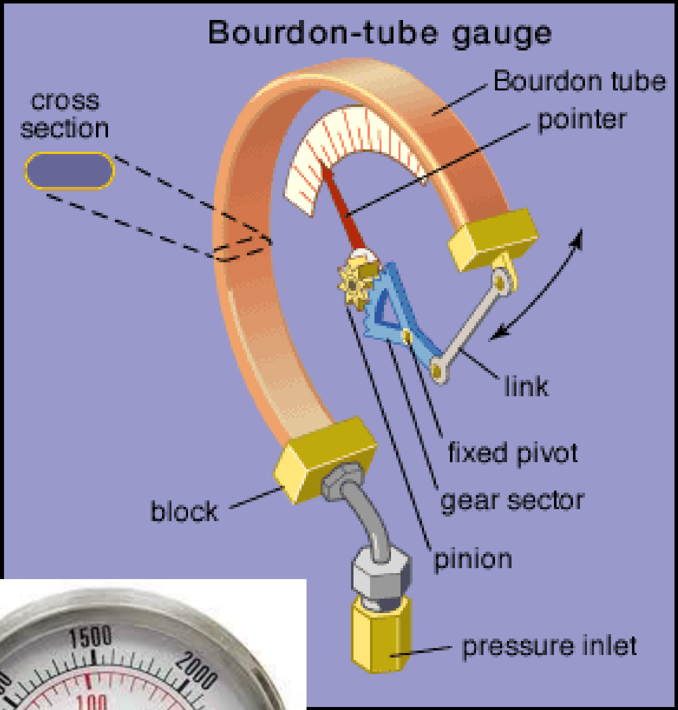
Converting Pressure Measurements

Americans commonly use psig for cylinder pressures
Others use units of bar (100 kiloPascals)



From psig to bar

$$1500\text{ psi} \times \frac{1\text{ atm}}{14.7\text{ psi}} \times \frac{1.01325\text{ bar}}{1\text{ atm}} = 103\text{ bar}$$

From bar to psig

$$100\text{ bar} \times \frac{1\text{ atm}}{1.01325\text{ bar}} \times \frac{14.7\text{ psig}}{1\text{ atm}} = 1451\text{ psig}$$

Let the units drive the solution

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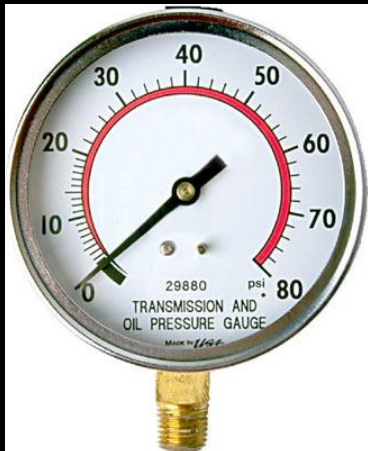
Gauge vs. Absolute Pressure



Gauges have a zero point

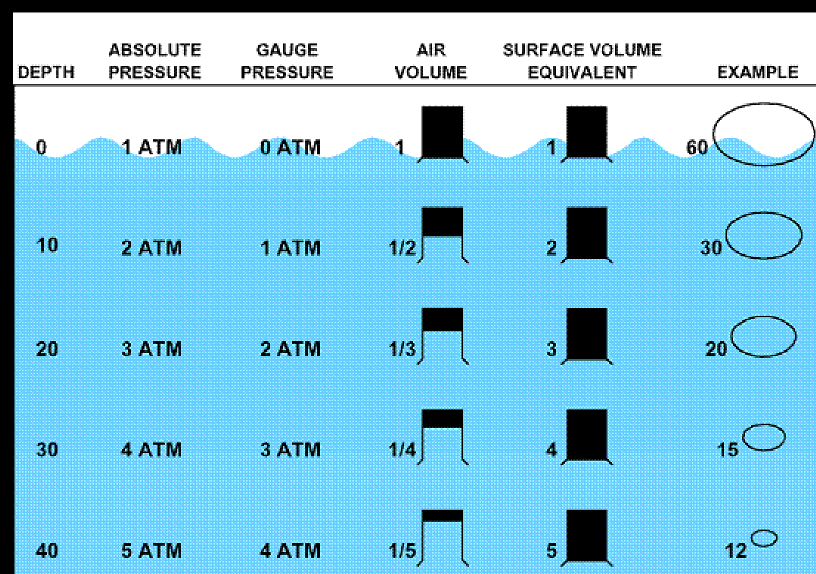
Gauge zero point = 1 atmospheric pressure

Absolute Pressure

Absolute pressure = pressure of gauge + atmospheric pressure

Represents the actual (total) pressure on the system

Gases respond to the absolute pressure



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Absolute Pressure

Total pressure on system: gauge pressure + atmospheric pressure

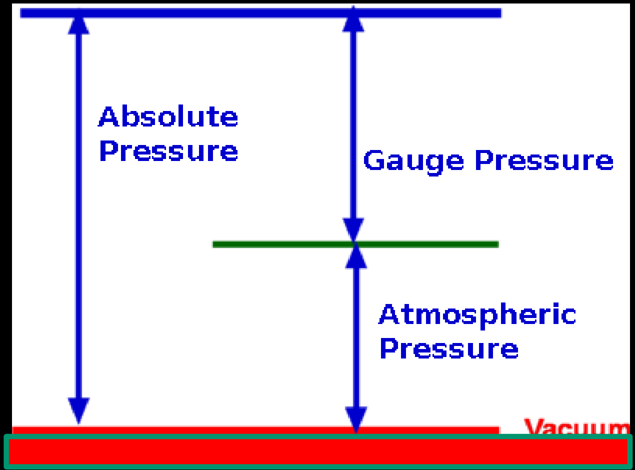
$$P_t = P_g + P_a$$

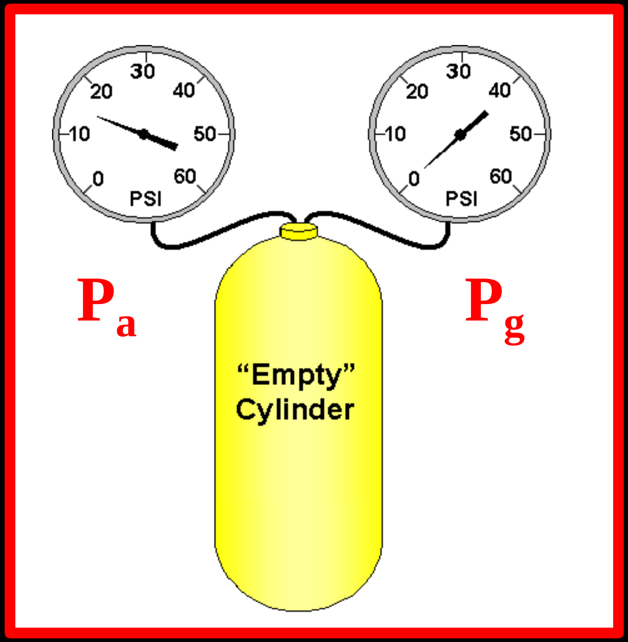
For absolute pressure:
Need to add 1 atm

SPG Pressure
Depth Gauge

Use Appropriate Units

Gauges calibrated:
fsw or msw





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Converting Depth To Pressure

Converting depth sea water (fsw) to absolute pressure in atmospheres (ata):
33 fsw of depth represents 1 atm of pressure (33 fsw / 1 atm)

$$\frac{(D\ fsw + 33\ fsw)}{33\ fsw / atm} = P\ ata$$

For a depth of 33 fsw

$$\frac{(33\ fsw + 33\ fsw)}{33\ fsw / atm} = 2.0\ ata$$

For a depth of 99 fsw

$$\frac{(99\ fsw + 33\ fsw)}{33\ fsw / atm} = 4.0\ ata$$

Pressure on Diver:
Weight of atmosphere +
Weight of water above diver

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Converting Depth To Pressure

It is critical to keep consistent units when solving problems

For example:

Assume an American trained diver on vacation rents a depth gauge calibrated in meters. The diver dives with a guide to 40 m.

If the American diver only remembers

ata = Depth/33 + 1 (without units)

the following diver calculation error is possible:

ata = 40/33 + 1 = 2.21

(Actual units the diver is using in {})

ata = (40 {m} / 33 fsw/atm) + 1 atm = 2.21 {m-atm / fsw}

SHOULD BE:

ata = (40 m / 10 msw/atm) + 1 atm = 5 ata

Let the units drive the solution

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Converting Pressure To Depth

Converting absolute pressure (ata) to depth sea water (fsw)
33 fsw of depth represents 1 atm of pressure (33 fsw / 1 atm)

D fsw = (Pata x 33 fsw/atm) - 33 fsw

For 5 atm of absolute pressure

D = (5 atm x 33 fsw /atm) - 33 fsw = 132 fsw

For 3 ata (ata = atm + 1 atm)

D = (3 atm x 33 fsw/atm) - 33 fsw = 66 fsw

For 2 ata

D = (2 atm x 33 fsw/atm) - 33 fsw = 33 fsw

For 1.4 ata

D = (1.4 atm x 33 fsw/atm) - 33 fsw = 13.2 fsw

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Temperature

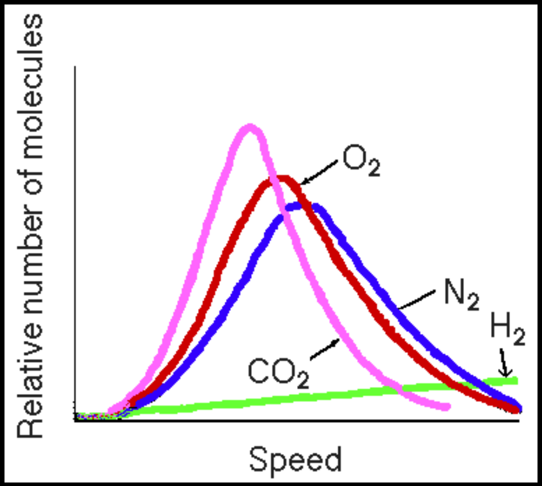
The *average* kinetic energy (K.E.) of molecules

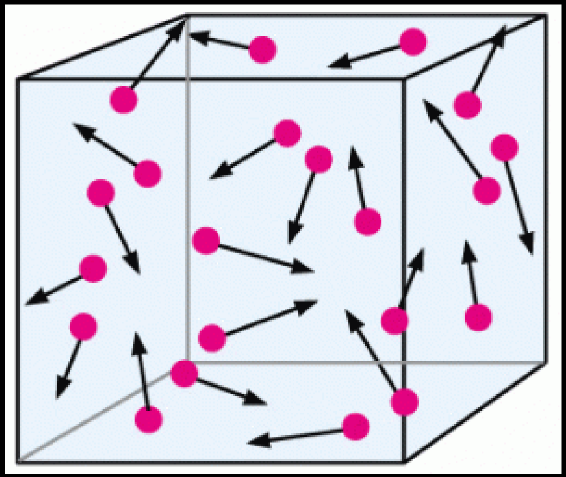
K.E. = energy of motion

= 1/2 mass(velocity)^2

As velocity (speed) slows, K.E. decreases & temperature falls

At K.E. = 0, v = 0 (Absolute zero … no molecular movement)

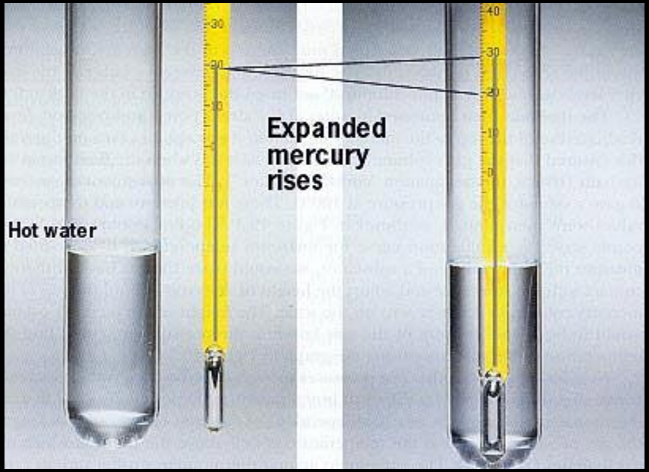




At constant temperature, larger molecules move slower

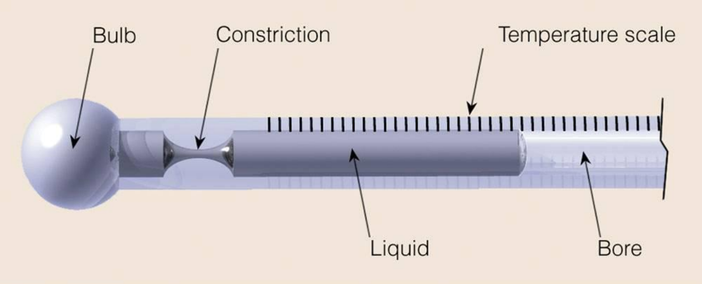
Thermometer

Device for measuring temperature



As liquid warms, it expands &
Amount of expansion is quantified

Clinical (Maximum Temperature):
Restriction Prevents Liquid Return to Bulb



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Temperature Scales

Fahrenheit (^{o}F)

Daniel Fahrenheit

Dutch scientist

1724

Placed mercury in closed glass column

Used three points to define scale:

- **0 = coldest that could be reached with water, ice, sea salt slush**
- **32 = water/ice**
- **96 = arm pit temperature (used bisection to create scale)**

Ice/ Water Slush	Freezing Water	Human Arm Pit	Boiling Water
			
0 ^{o}F	32 ^{o}F	96 ^{o}F	212 ^{o}F

water: freezing and boiling points 180 units apart

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Temperature Scales

Celsius (Centigrade, ^{o}C) — **Anders Celsius**

Swedish Astronomer — 1742

International scientific scale

Used two points to define scale:

100 = freezing point of water

0 = boiling point of water

Freezing Water

100 ^{o}C

Boiling Water

0 ^{o}C

water: freezing and boiling points 100 units apart

Carolus Linnaeus — 1744

Swedish Botanist

Reversed scale to set Freezing Point at 0 Degrees

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"Absolute" Temperature Scales

William Thompson (Lord Kelvin)   1848

Proposed a scale based on absolute zero as zero point

Uses the Centigrade (1/273 gas volume change) degree

Makes all temperatures have positive value

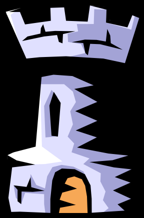
K = ^{o}C + 273 ^{o}C = K - 273

William Rankine   1859

Proposed a scale based on absolute zero as zero point

Uses the Fahrenheit degree

^{o}R = ^{o}F + 459 ^{o}F = ^{o}R - 459



Gas laws must use absolute temperatures

Science uses Kelvin

American engineering (especially heating / cooling) uses Rankine

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Temperature Conversions

Zero point difference:
32 degrees

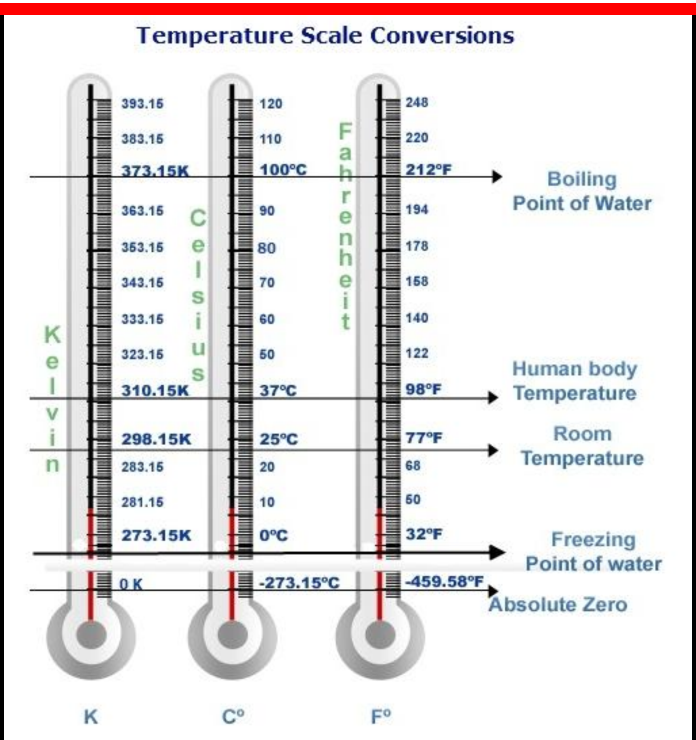
$$^{o}F = 9/5\ ^{o}C + 32$$

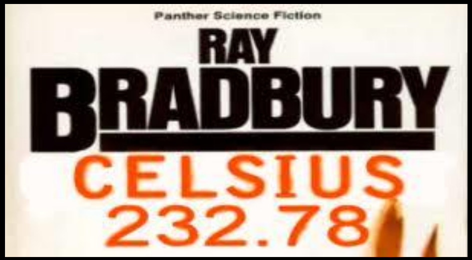
$$^{o}C = 5/9\ (^{o}F - 32)$$

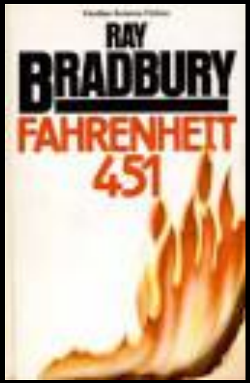
180 /100 = 9/5

K = ºC + 273 (Kelvin, Absolute Celsius Scale)

By convention, there is no º symbol for degrees kelvin







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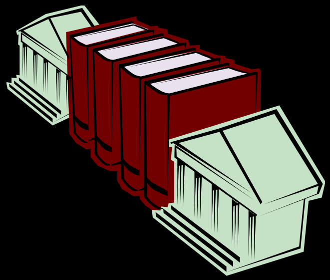
Gas Behavior (Equations or Laws)

Must use absolute temperature and pressure

"Ideal" Gases



Described by Kinetic Theory of Gases

Behavior predictable by "Ideal Gas Laws"

Valid at low pressures & high temperatures

Not valid at compressed gas cylinder pressures

Need more complex "Real" Gas Equations

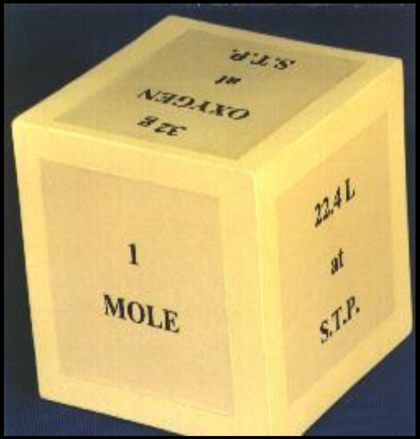
STP

Standard Temperature & Pressure

Standard Temperature = 0° C (273 K)

Standard Pressure = 1 atm (760 torr)

At STP:



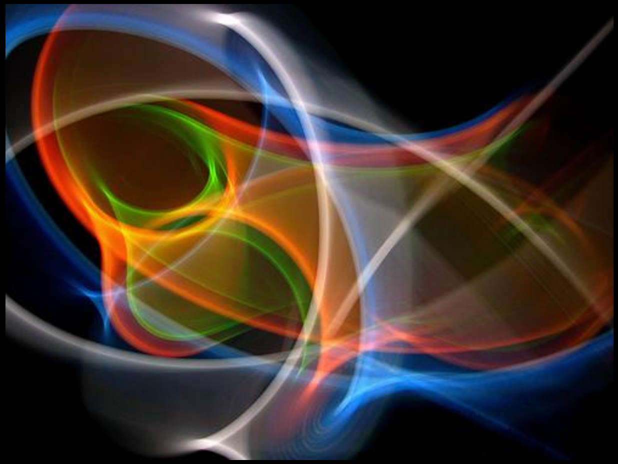
22.4 L = 1 mole of any gas

22.4 L weighs Molar Mass

22.4 L contains 6.02 x 10^{23} molecules

STP

Allows comparisons to different locations and conditions



Gas Laws

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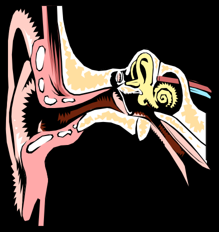
Rigid and Flexible Gas Containers

Walls: rigid
Volume: constant
Rupture when internal pressure exceeds container strength
Example: compressed gas cylinder

Walls: flexible
Volume: constant if internal & surroundings pressures equal
Volume: changes if internal & surroundings pressures unequal
Rupture when internal pressure exceeds container strength
Examples: balloon, internal air spaces (lungs, ears, sinus, gut)



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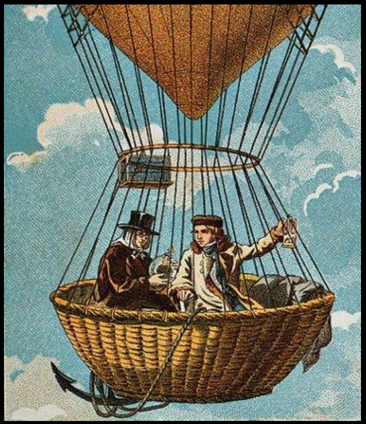
Joseph Louis Guy-Lussac



French chemist
Student of Jacques Charles
Studied Gases In Chemical Reactions

Pressure - Temperature relationship (1809)

Maybe called Charles's Law or Charles's Law #2
Sometimes called Amonton's Law
(Proposed relationship, but lacked technology to prove)
But,
Guy-Lussac was first to experimentally document P-T relation

His observations – a primary source of absolute temperature scale

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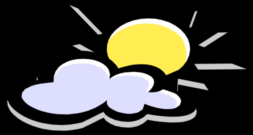
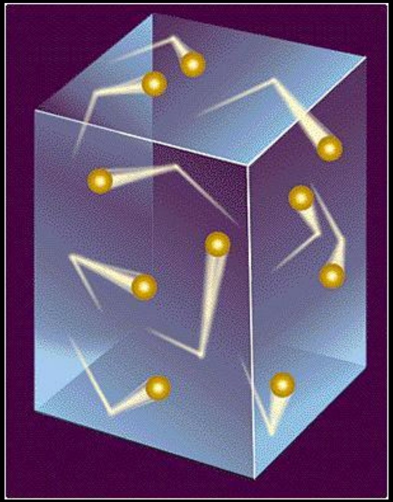
Guy-Lussac's Law



Heat energy increases molecular motion.

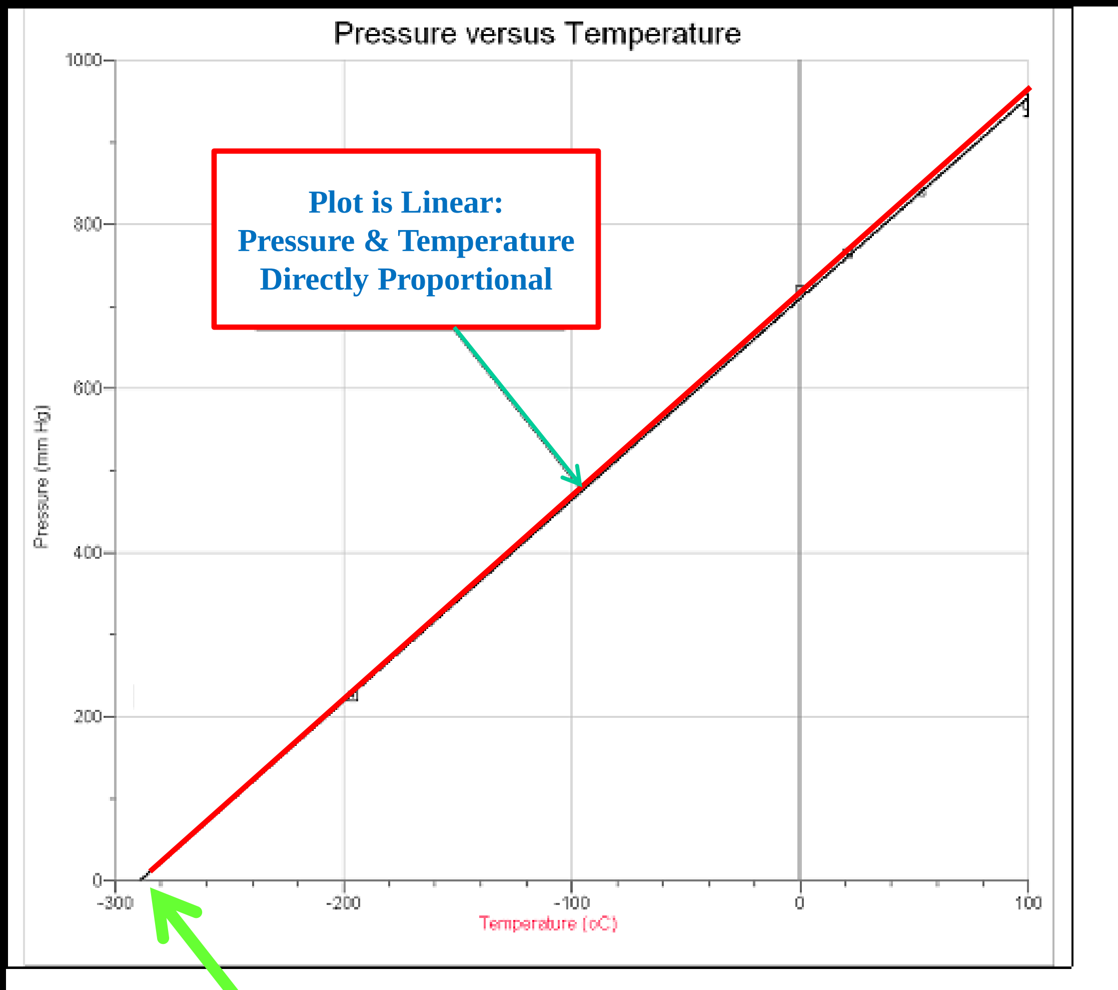
Volume of rigid cylinder cannot increase; the pressure increases

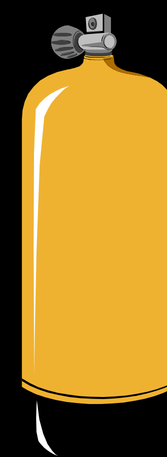
At constant *volume*, in a RIGID container:
pressure is directly proportional to the absolute temperature



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

Guy-Lussac's Law





Absolute Zero (-273.16 ºC)

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A scuba cylinder contains 3000 psig at 78 ºF. It is left in the trunk of a car on a hot summer day. If the temperature of the trunk is 115 ºF, what will be the gauge pressure of the cylinder?

Using Guy-Lussac's Law: $P_1 / T_1 = P_2 / T_2$

Remember P & T must in absolute values

Since P value given in psi, 1 atm = 14.7 psi

psiga = 3000 psi + 14.7 psi = 3014 psia

Since T given in ºF, use Rankin Absolute Scale

T_1 = 78 ºF + 460 = 538 ºR

T_2 = 115 ºF + 460 = 575 ºR

Substituting in Guy Lussac's Law

3014.7 psia / 538 ºR = P_2 / 575 ºR

P_2 = 3222 psia

Converting to gauge pressure

3222 psia - 14.7 psi = 3207 psig

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A cylinder at 25 ^oC has a gauge pressure of 200 bar. Predict the gauge pressure at 42 ^oC .

Using Guy-Lussac's Law: $P_1 \ / \ T_1 \ = \ P_2 \ / \ T_2$

Remember P & T must in absolute values

Since P value given in bar, 1 atm = 1.01 bar

psiga = 200 bar + 1.01 bar = 201.01 bar

Since T given in ^oC, use Kelvin Absolute Scale

$$T_1 = 25\ ^oC + 273 = 298\ K$$

$$T_2 = 42\ ^oC + 273 = 315\ K$$

Substituting in Guy Lussac's Law

$$201\ bar \ / \ 298\ K = P_2 \ / \ 315\ K$$

$$P_2 = 212.5\ bar$$

Converting to gauge pressure

$$212.5\ bar – 1.01\ bar = 212\ bar$$

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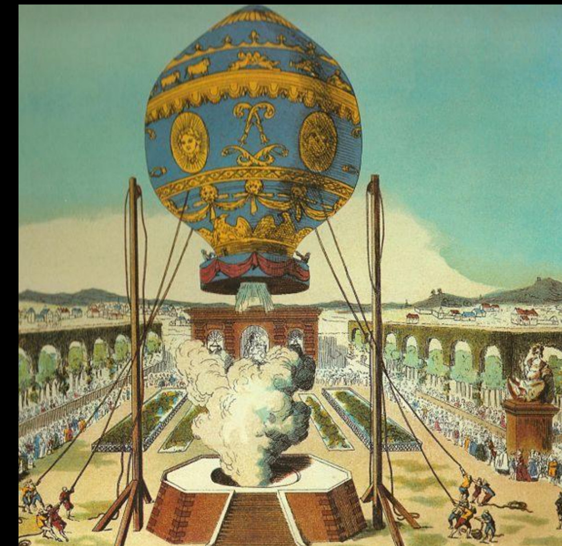
Jacques Charles

French chemist

Scientific Advisor to Montgolfier brothers



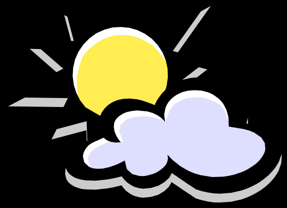
Volume - Temperature Relationship (1787)

1783 – First hot air balloon

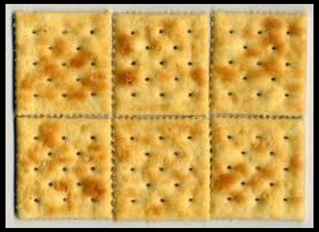
Sack cloth and paper with 1800 buttons

Redesigned the way hot-air balloons were built:

- **Silk instead of paper construction**
- **Hydrogen instead of hot air**
- **Valve line**
- **Wicker basket passenger compartment**



Charles' Law



Heat energy increases molecular motion.

Volume of flexible container increases

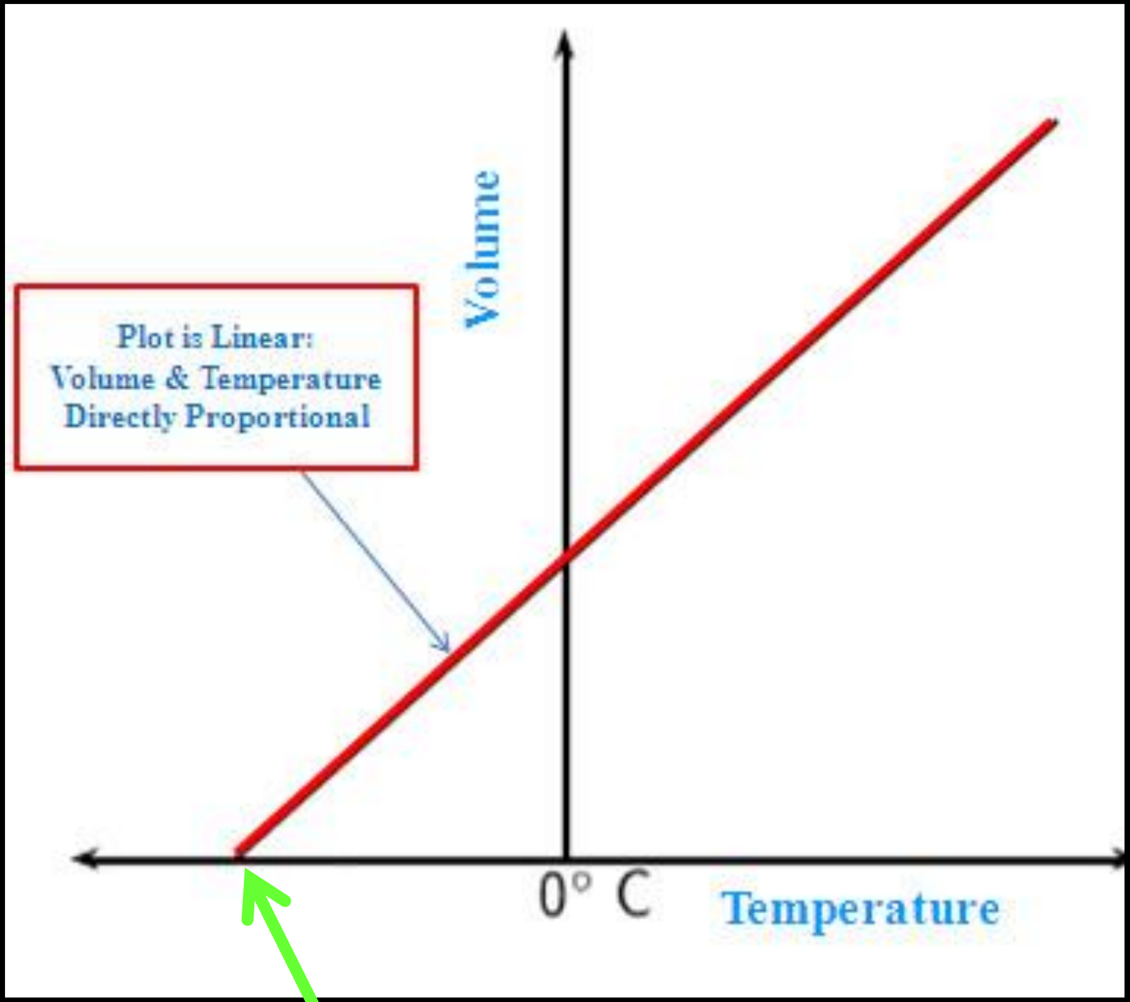
At constant *pressure*, in a FLEXIBLE container volume is directly proportional to the absolute temperature

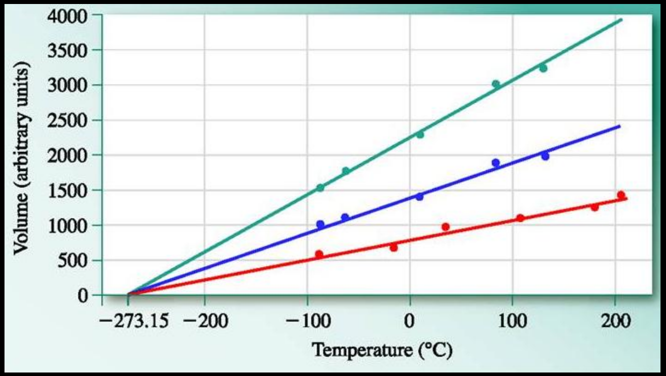
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

If T = negative, volume = negative (not realistic)

Need temperature to be positive

Charles' Law





Absolute Zero (-273.16 ºC)

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If a scuba cylinder is capable of delivering 40 ft^3 of air to a diver at 78 ^oF, how much air is available at 55 ^oF?

Using Charles' Law: $V_1 \;/\; T_1 \;=\; V_2 \;/\; T_2$

Since T given in ^oF, use Rankin Absolute Scale

$T_1 = 78\ ^oF + 460 = 538\ ^oR$

$T_2 = 55\ ^oF + 460 = 515\ ^oR$

Substituting into Charles' Law

$40\ ft^3 \;/\; 538\ ^oR \;=\; V_2 \;/\; 515\ ^oR$

Solving:

$V_2 = 38.3\ ft^3$

The temperature 55 ^oF is typically the temperature of the first thermocline of a fresh water lake. Charles' law explains why divers have less air available to them in colder water.

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A scuba cylinder delivers 1000 L of air at 25 ^oC. If this cylinder is used at 18 ^oC. how much air will be available to the diver?

Using Charles' Law: $V_1 \ / \ T_1 \quad = \quad V_2 \ / \ T_2$

Since T given in ^oC, use Kelvin Absolute Scale

$T_1 = 25\ ^oC + 273 \ = 298\ K$

$T_2 = 18\ ^oC + 273 \ = 291\ K$

Substituting into Charles' Law

$1000\ L \ / \ 298\ K \quad = \quad V_2 \ / \ 291\ K$

Solving:

$V_2 \quad = \quad 977\ L$

When descending below the thermocline, the decrease in BCD volume changes buoyancy and divers need to add gas to compensate

Thermoclines are often visualized by silt particles resting on the colder, more dense water

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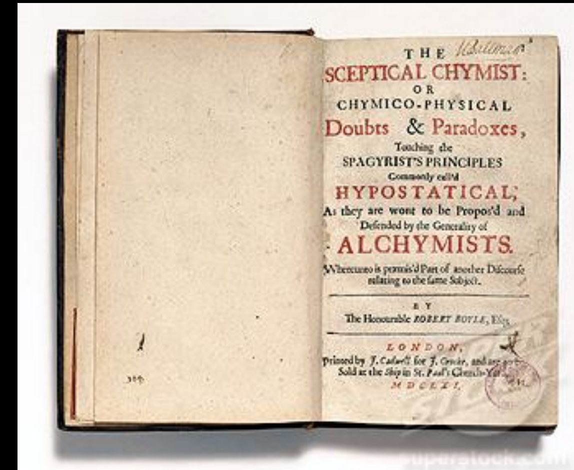
Robert Boyle

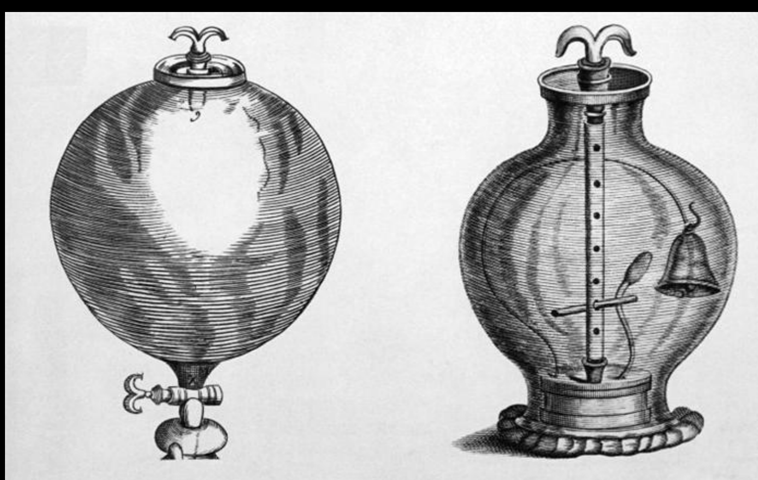
Irish Alchemist
Father of modern chemistry
Founder of Royal Society



Pressure - Volume relationship (1660)

New Experiments: Phsico-Mechanical Touching the spring of air and their effects (1660)

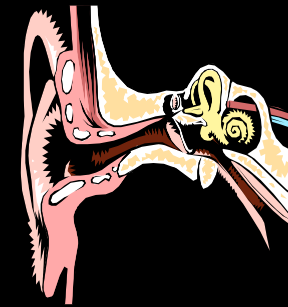
The Sceptical Chymst (Air, Earth, Fire, & Water not elements) (1661)



In an evacuated chamber
Observed bubble in snake's eye
Reduced Pressure Changes Physiology
Bell produced no sound
Air needed to carry sound

LPT

Boyles's Law



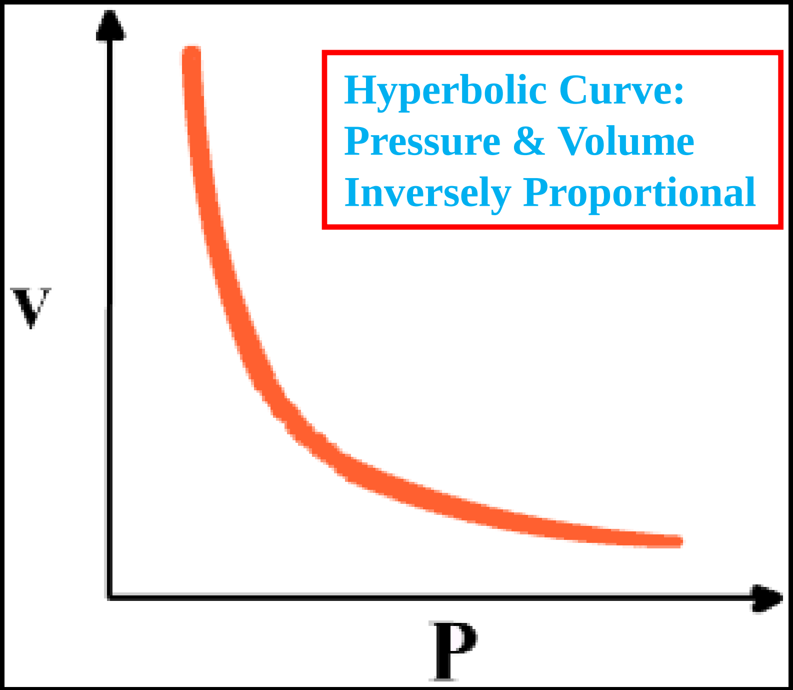
At constant *temperature*,
the volume of a flexible container
depends upon the surrounding pressure

At constant *temperature*, in a FLEXIBLE container volume is indirectly proportional to the absolute pressure

$$P_1 V_1 = P_2 V_2$$

LPT

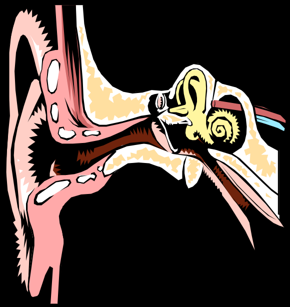
Boyles's Law



Greatest volume change:
~12 feet to surface

Means greatest risk to tissue:
shallow water

Explains:



- **Ear Discomfort while ascending / descending**
- **Grandpa's knee forecasting weather**
- **Changes in all gas volumes with altitude / depth**
- **Changes in pressure with altitude / depth**

What is the physical volume (in ft^3) of an aluminum "80"?

An aluminum "80" delivers 80 ft^3 at one atmosphere pressure when filled to 3000 psig.

Using Boyle's Law: $P_1 V_1 = P_2 V_2$

Since P value given in psi, 1 atm = 14.7 psi

psiga = 3000 psi + 14.7 psi = 3014.7 psia

Substituting into Boyle's Law:

$$(14.7 \text{ psia})\ (80 \text{ ft}^3) = (3014.7 \text{ psia})\ V_2$$

Solving:

$$V_2 = 0.39 \text{ ft}^3$$

This physical volume represents how much water the cylinder would hold if the valve were removed and the cylinder were filled with water. This is the "water capacity" of a scuba cylinder.

LPT

A scuba cylinder is rated at 2400 L with a pressure of 200 bar. What is the physical volume (water capacity) of the cylinder.?

Using Boyle's Law: $P_1 V_1 = P_2 V_2$

Since P value given in bar, 1 atm = 1 bar

psiga = 200 bar + 1 bar = 201 bar

Substituting into Boyle's Law:

$(1 \text{ bar})(2400 \text{ L}) = (201) V_2$

Solving:

$V_2 = 11.9 \text{ L}$

LPT

Determine the volume of air from an “80” ft^3 cylinder that will be available to the diver at 33, 66, 99 and 132 fsw

Using Boyles’ Law: $P_1 V_1 = P_2 V_2$

For 33 feet:

$(1\ ata)(80\ ft^3) = (2\ ata)\ V_2$

$V_2 = 40\ ft^3$

For 66 feet:

$(1\ ata)(80\ ft^3) = (3\ ata)\ V_2$

$V_2 = 26.7\ ft^3$

For 99 feet:

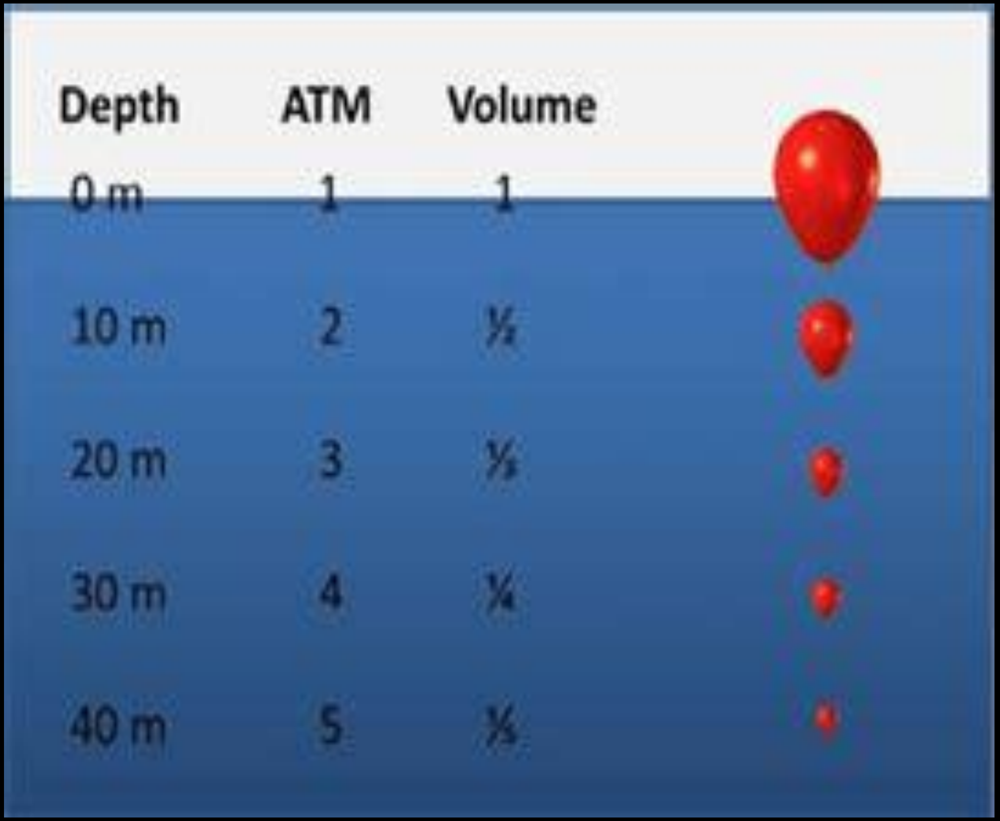
$(1\ ata)(80\ ft^3) = (4\ ata)\ V_2$

$V2 = 20\ ft^3$

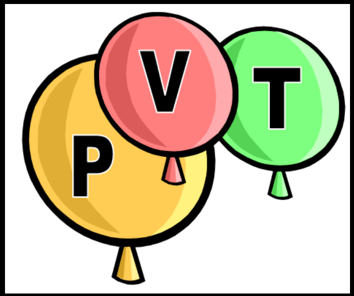
For 132 feet:

$(1\ ata)(80\ ft^3) = (5\ ata)\ V_2$

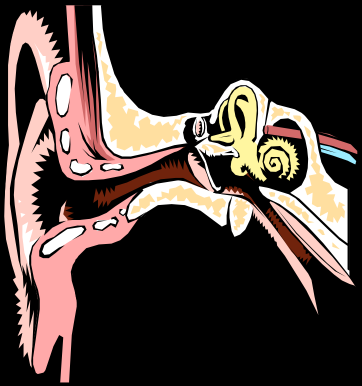
$V2 = 16\ ft^3$



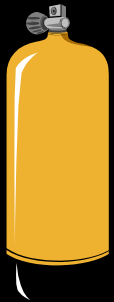
LPT



General Gas Law

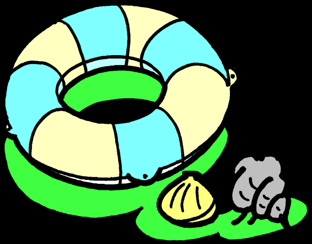


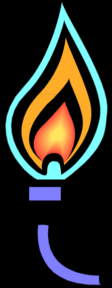
$$\frac{p_1 v_1}{t_1} = \frac{p_2 v_2}{t_2}$$



Units need to be same on both sides of =

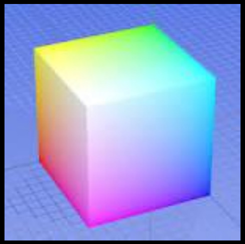
P & T must be in absolute measure





LPT

General Gas Law


$$\frac{p_1 v_1}{t_1} = \frac{p_2 v_2}{t_2}$$

If P constant:

$$\frac{v_1}{t_1} = \frac{v_2}{t_2}$$

Charles

If V constant:

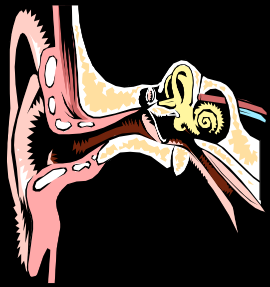
$$\frac{p_1}{t_1} = \frac{p_2}{t_2}$$

Guy-Lussac

If T constant:

$$p_1 v_1 = p_2 v_2$$

Boyle



LPT

Universal Gas Law Problem Matrix

A standard method for determining Gas Law Solutions Facilitates Solving

Set up the universal gas law table of values

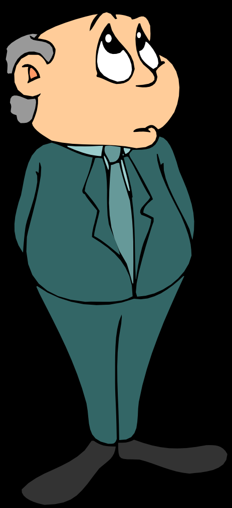
	Pressure (units)	→	Pressure (absolute)	Volume (units)	Temperature (°C)	→	Temperature (K)
Initial						+273	
Final						+273	

Fill in table values from the problem description

Place problem table values into general gas law

Solve

Proportional Thinking

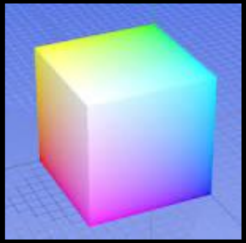


$$\frac{pv}{t} = k$$

$$\frac{p_1 v_1}{t_1} = k = \frac{p_2 v_2}{t_2}$$

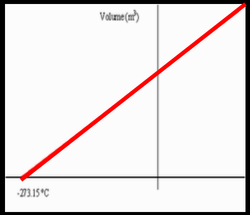
LPT

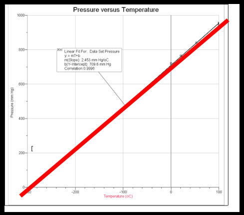
Proportional Thinking

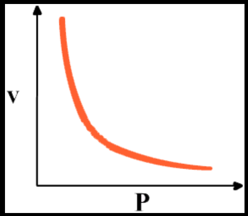


$$\frac{p\,v}{t} = k$$

If P constant:	If V constant:	If T constant:
$\frac{v}{t} = k$	$\frac{p}{t} = k$	$pv = k$
Charles' Law	Guy-Lussac's Law	Boyle's Law
Direct Proportion	Direct Proportion	Inverse Proportion

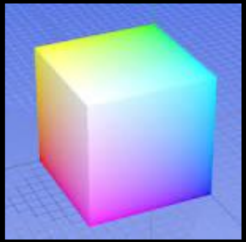






LPT

Proportional Thinking



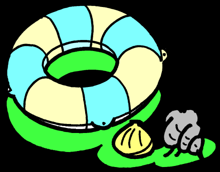
$$\frac{p\,v}{t} = k$$

Variables change to keep k constant

If P constant:

$$\frac{\uparrow v}{\uparrow t} = k$$

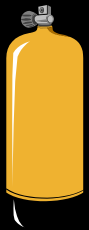
v and t change
(increase or decrease)
in same direction



If V constant:

$$\frac{\uparrow p}{\uparrow t} = k$$

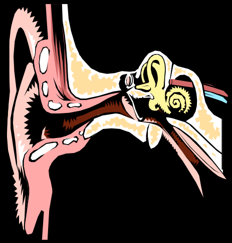
p and t change
(increase or decrease)
in same direction



If T constant:

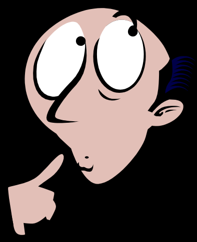
$$\underset{\uparrow}{p}\,\underset{\downarrow}{v} = k$$

p and v change
(increase or decrease)
in opposite direction



LPT

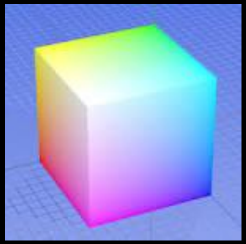
Proportional Thinking: Word problems



At constant volume, if temperature decreases, pressure ________

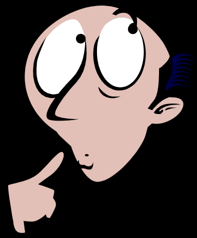
At constant pressure, if temperature increases, volume ________

At constant temperature, if pressure increases, volume ________



LPT

Proportional Thinking: Word problems



At constant volume, if temperature decreases, pressure $\boxed{\underline{\text{decreases}}}$

$$\frac{p}{t} = k \longrightarrow \frac{p}{t} = k$$

P & T Move same direction

At constant pressure, if temperature increases, volume $\boxed{\underline{\text{increases}}}$

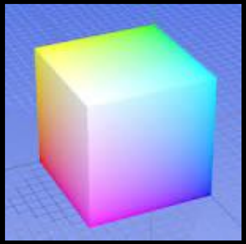
$$\frac{v}{t} = k \longrightarrow \frac{v}{t} = k$$

V & T Move same direction

At constant temperature, if pressure increases, volume $\boxed{\underline{\text{decreases}}}$

$$p\,v = k \longrightarrow p\,v = k$$

P & V Move opposite direction

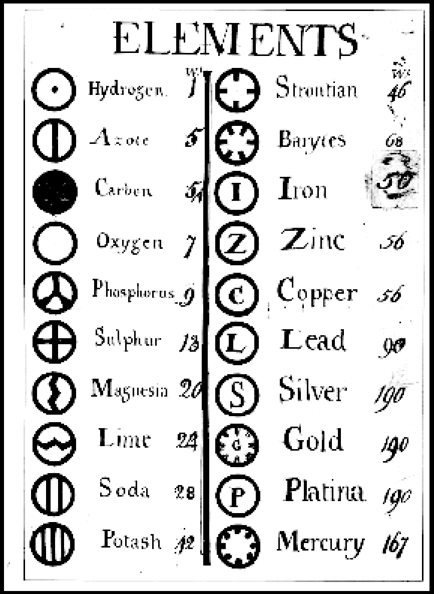


LPT

John Dalton

School teacher with contributions to:

- Atomic Theory
- Understanding Color Blindness
- Studies on Gas Behavior



Dalton's Law of Partial Pressure (1803)

$$P_{total} = P_1 + P_2 + P_3 + \dots P_n$$

For a mixture of ideal gases,
total pressure = sum of the partial pressures of gases present

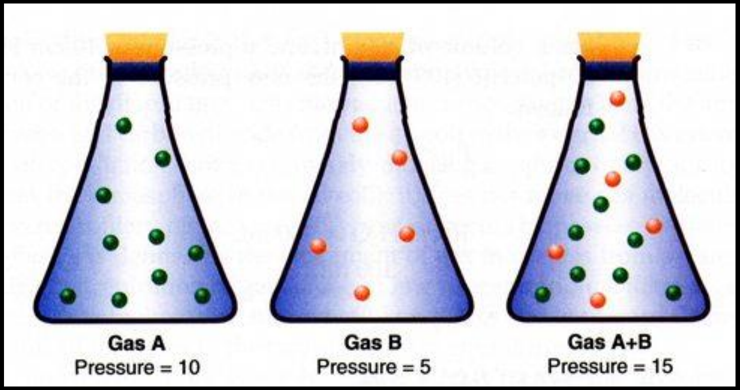
LPT

Dalton's Law: Partial Pressures

Dalton's law: In a mixture of gases, the total pressure is the sum of the partial pressures of the individual components

$$P = P_1 + P_2 + P_3 + \ldots + P_n$$

The partial pressure of a gas is the product of the fraction of that gas times the total pressure



$$P_g = F_g \times P_{total}$$

Where

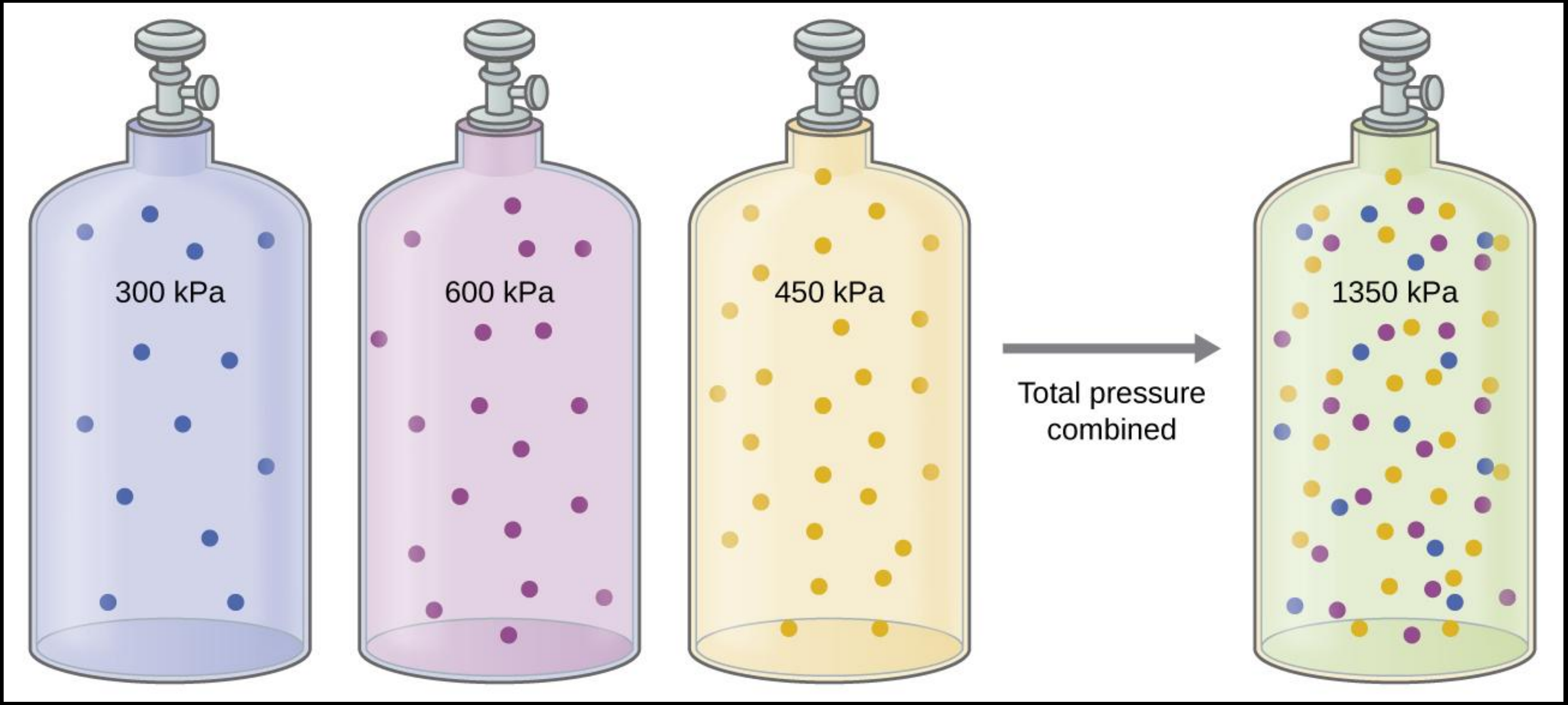
P_g = partial pressure of the component gas

F_g = fraction of the component gas in the mixture

P_{total} = the total pressure of the gas mixture

LPT

Dalton’s Law: Partial Pressures

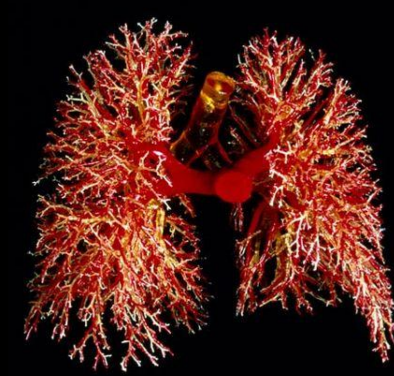


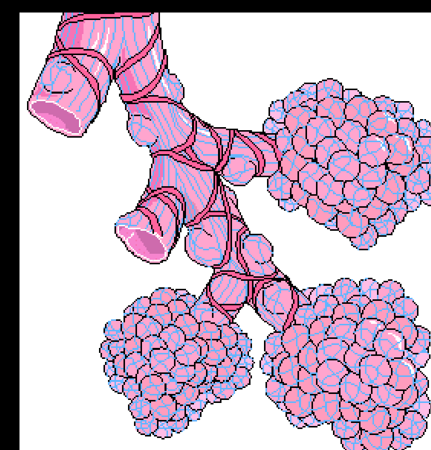
Total pressure is always the sum of component gas pressures

Dalton's Law: Partial Pressures

Pressure in alveolar spaces immediately equilibrates with blood

	Inspired air	Alveolar air
H_2O	Variable	47 mmHg
CO_2	000.3 mmHg	40 mmHg
O_2	159 mmHg	105 mmHg
N_2	601 mmHg	568 mmHg
Total pressure	760 mmHg	760 mmHg





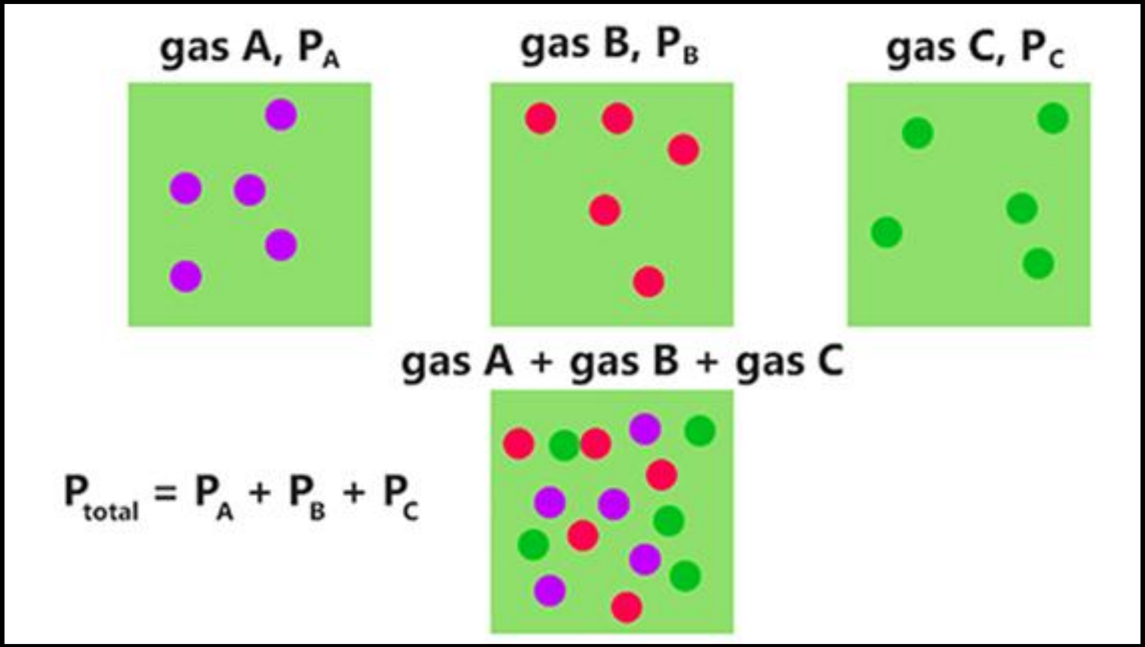
A mixture of gases at 760 torr contains 55.0 % N_2, 25.0 % O_2, and 20.0 % CO_2 by volume. What is the partial pressure of each gas?

N_2: 55.0/100 x 760 torr = 418 torr

O_2: 25.0/100 x 760 torr = 150 torr

<u>CO_2: 20.0/100 x 760 torr =152 torr</u>

Total (check) = 760 torr



LPT

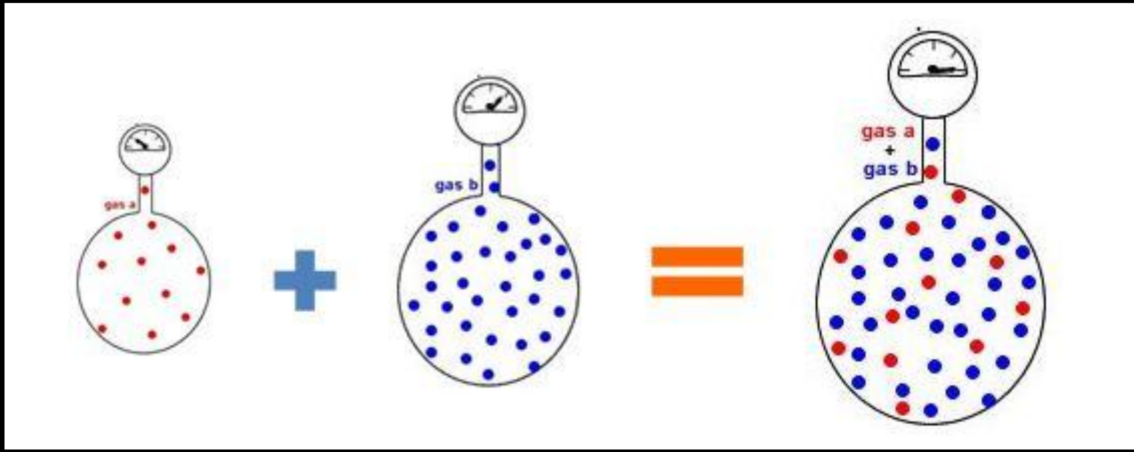
A 200 mL flask contains O_2 at 220 torr and a 300 mL flask contains N_2 at 100 torr. The flasks are connected and the gasses are allowed to completely fill the system. There is no temperature change. What is the partial pressure of each gas and the total pressure?

The final volume is 200 mL + 300 mL = 500 mL

O_2: 220 torr (200 / 500) = 88 torr

N_2: 100 torr (300 / 500) = 60 torr

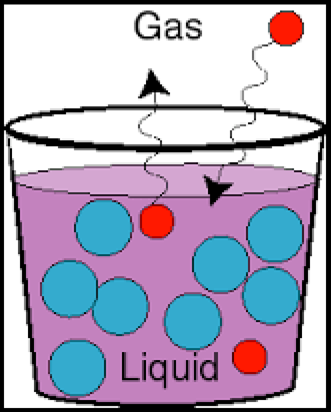
Total: 60 torr + 88 torr = 148 torr



LPT

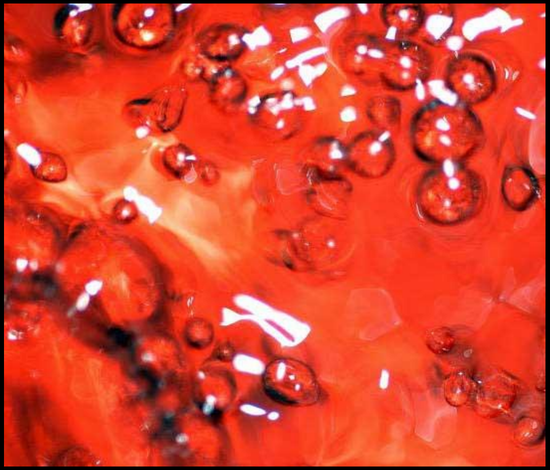
William Henry

British chemist
Solubility of gases
Composition of HCl and NH_3
Disinfecting powers of heat



Gas in liquid solubility: Henry's Law (1803)

Determined solubility of gases in liquids a function of:
- Partial pressure of the gas
- Temperature of the system
- Characteristics of the liquid



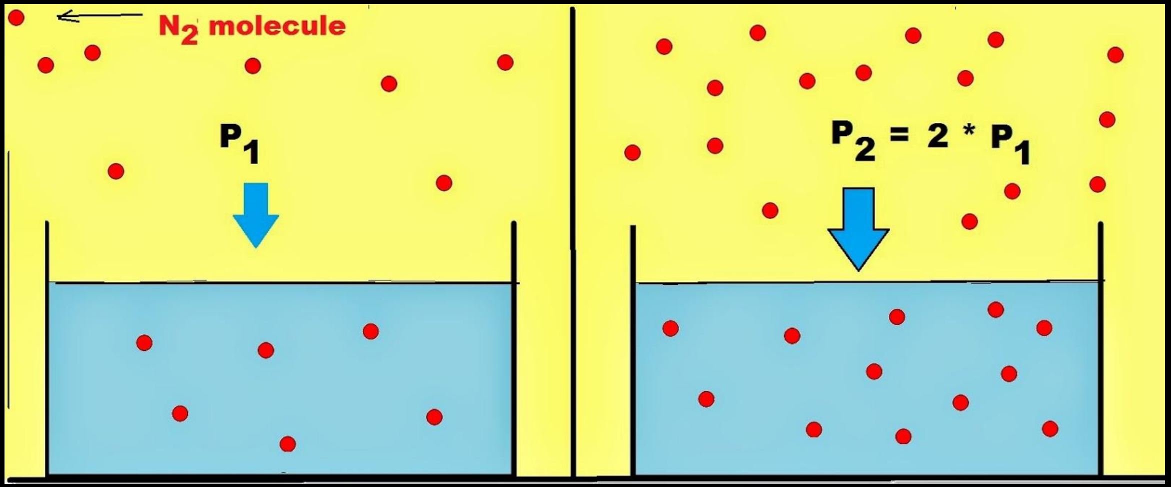
Very important when environmental pressure changes (alters gasses dissolved in the body)

Henry's Law

The amount of any given gas that will dissolve in a liquid at a given temperature is a function of the partial pressure of the gas that is in contact with the liquid and the solubility coefficient of the gas in the particular liquid

$$S_g = K_H \times P_g$$

S_g solubility of the gas
K_h liquid solubility constant
P_g Partial pressure of the gas

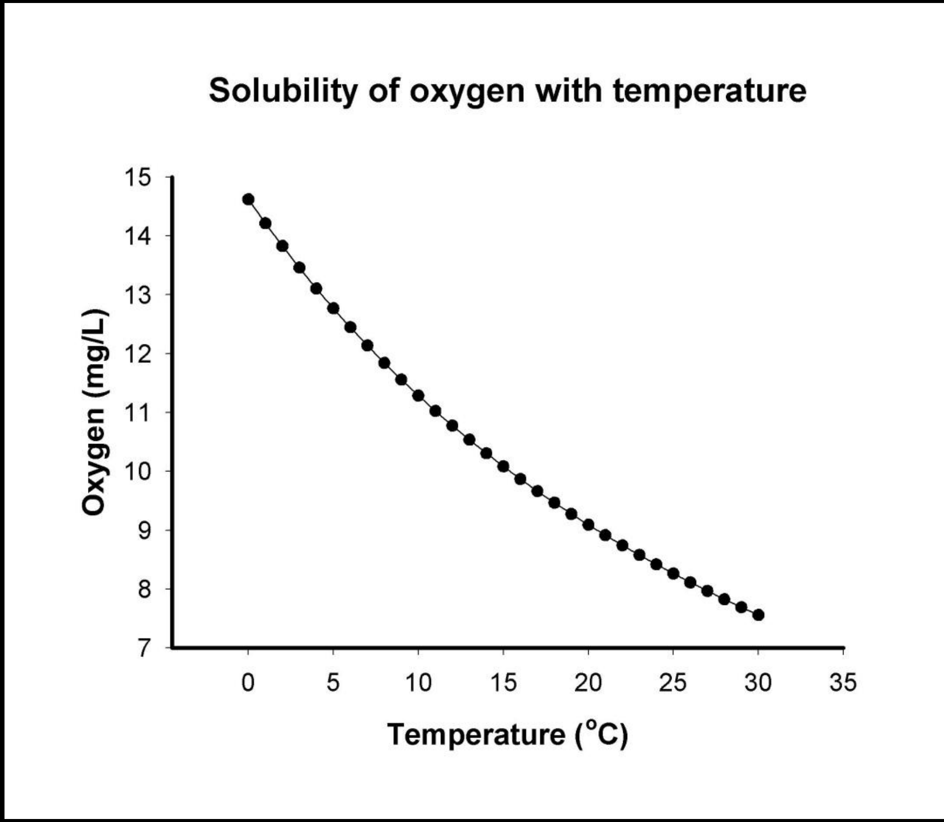


Increase in pressure → increase in solubility
Decrease in pressure → decrease in solubility

LPT

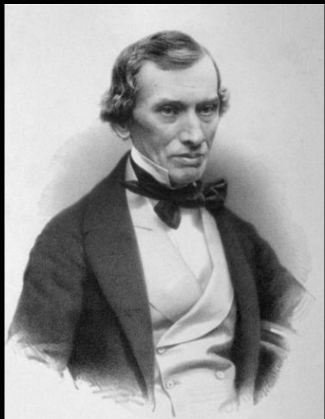
Henry's Law

Gas solubility changes with temperature



Colder water (Great Lakes): Divers carry additional gas loads

LPT

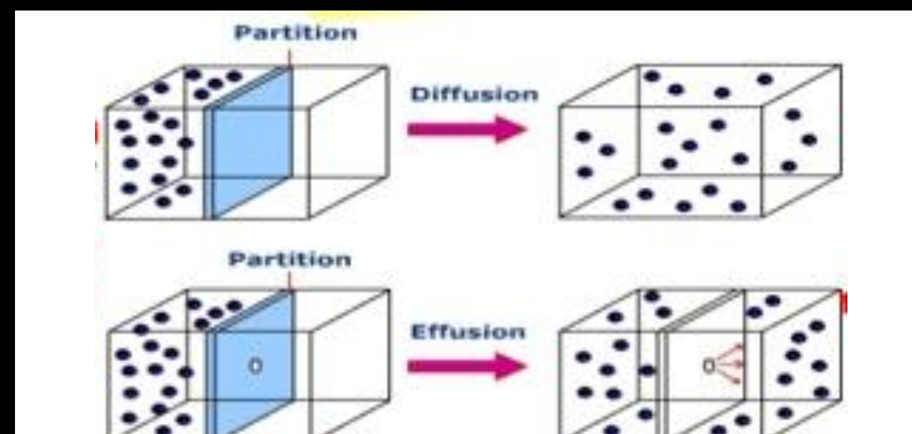


Graham's Law

The speed of gas diffusion (gas mixing due to kinetic energy) or effusion (gas moving through a tiny opening) is inversely proportional to the square root of their molar masses.

This is commonly written for effusion:

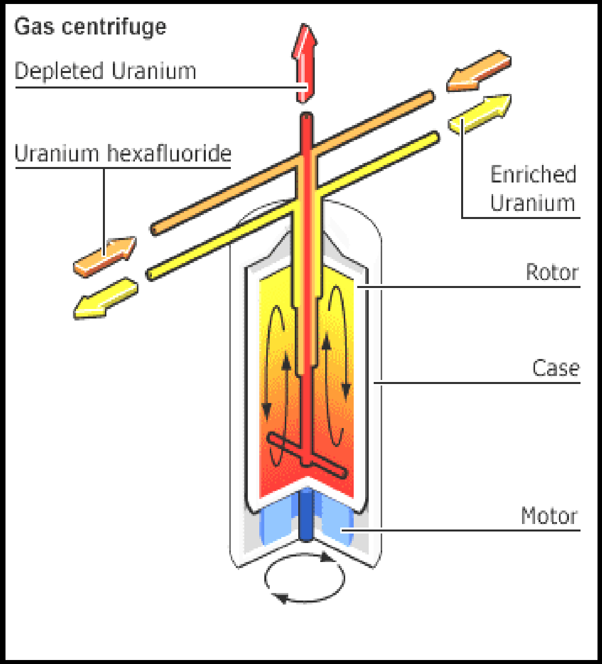
$$\frac{Rate_1}{Rate_2} = \sqrt{\frac{M_2}{M_1}}$$

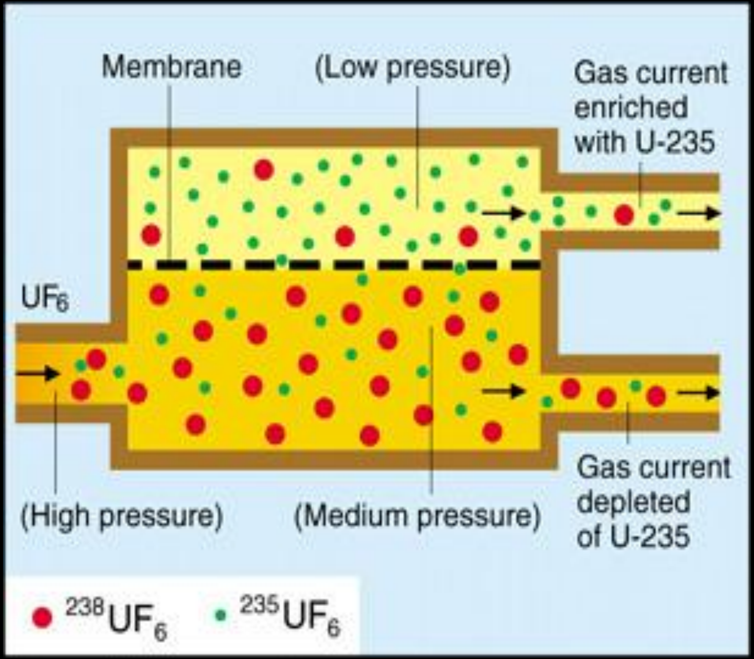


Graham's Law

Useful for:

- separation of gases of different densities
- separation of isotopes
  - prime means for producing nuclear material
- determining molar mass of unknown material





LPT

Compare the relative rates of effusion of H_2 and CO through a fine pinhole.

Molar masses:

CO = 28.0

H_2 = 2.00

$$\frac{\text{Rate } H_2}{\text{Rate CO}} = \sqrt{\frac{28.0}{2.00}} = 3.74$$

Compare the relative rates of effusion of H_2 and O_2 through a fine pinhole.

Molar masses:

O_2 = 32.0

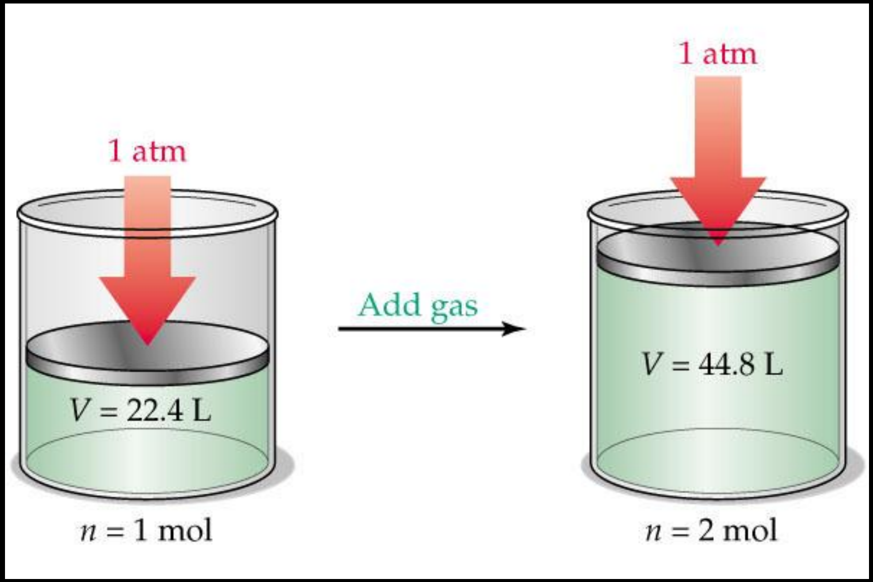
H_2 = 2.0

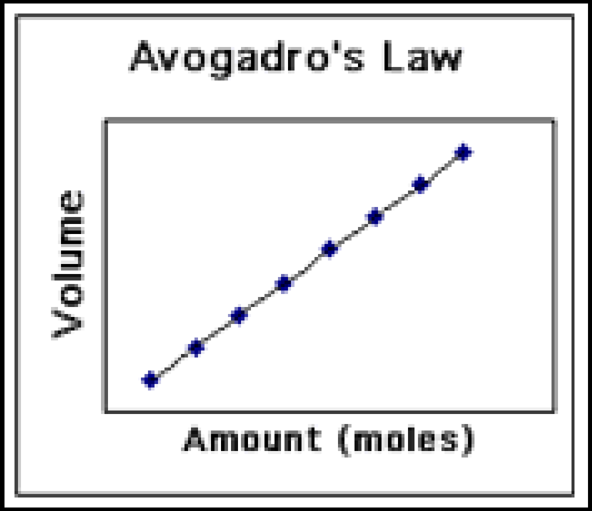
$$\frac{\text{Rate } H_2}{\text{Rate } O_2} = \sqrt{\frac{32.0}{2.0}} = 4.0$$

LPT

Avogadro's Law (1811)

More Moles (Molecules) means more molecular collisions

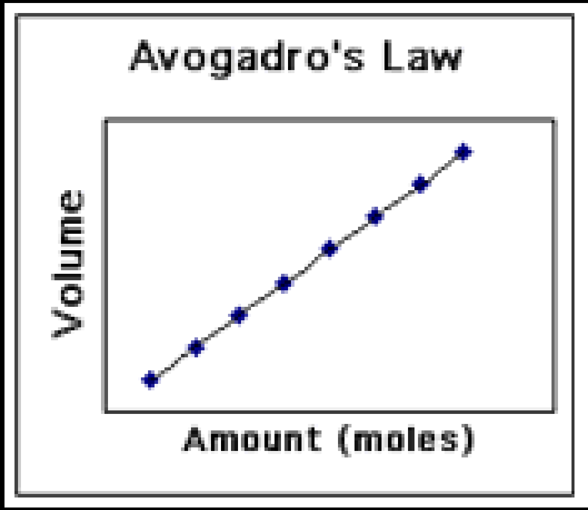




Plot is Linear:
Volume & # moles
Directly Proportional

LPT

Avagadro's Law



At constant *temperature and pressure,*
volume is directly proportional to the number of moles present

$$\frac{v_1}{n_1} = \frac{v_2}{n_2}$$

LPT

1.00 mole of gas occupies 1.45L. If the quantity of gas is increased to 2.50 moles, what is the new volume of gas?

	Pressure (torr)	Volume (L)	# moles
Initial	constant	1.45	1.00
Final	constant	?	2.50

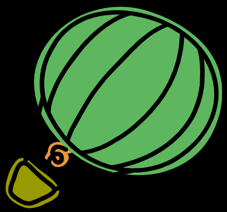
Pressure constant, use Avagadro's Law

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

$$\frac{1.45\ \text{L}}{1.00\ \text{m}} = \frac{V_2}{2.50\ \text{m}} \longrightarrow V_2 = \frac{(1.45\ \text{L})\,(2.50\ \text{m})}{1.00\ \text{m}}$$

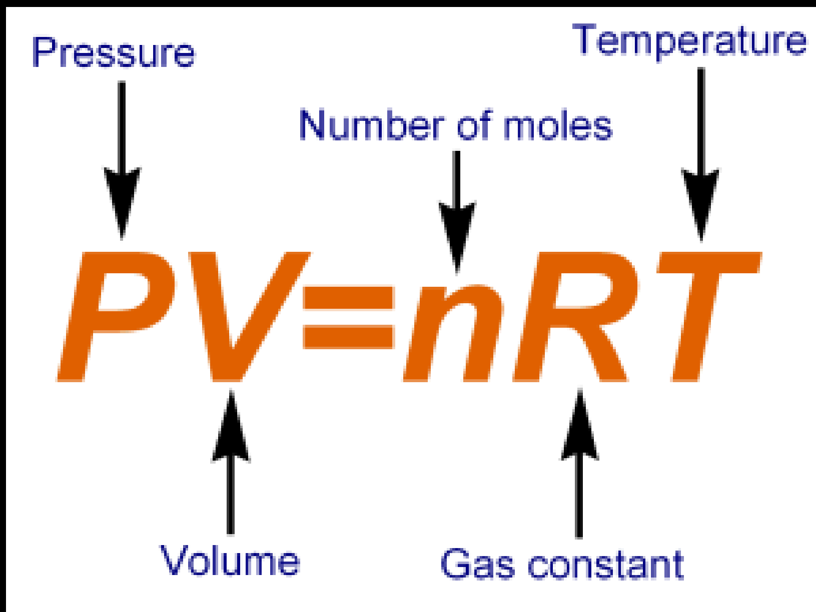
$$V_2 = 3.625\ \text{L} \longrightarrow 3.63\ \text{L}$$

LPT



Ideal Gas Law

$$PV = kT$$



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

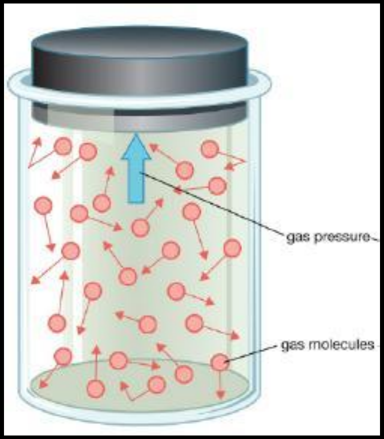
R makes "numbers work"

For:

P in atm; V in L; T in K

R = universal gas constant: 0.08206 L · atm / K · mol

Ideal Gas Law



PV = nRT

Based on kinetic theory of gases

Primary use is single point determination

Assumes:

- Volume of individual gas molecules is negligible
- There is no attraction between individual molecules

Strongest correlation between calculated and measured:

- Low pressure
- High temperature
- Monatomic gases

Ideal gas law assumes vast distances between gas molecules

Typically not discussed in scuba classes

Idea equation examples follow (for enlightenment)

LPT

Calculate the pressure of a 2.50 mole sample of a gas in a 5.50 L container at 27 °C.

Only one condition given → use Ideal Gas Law:

$$PV = nRT$$

$$P\ (5.50\text{ L}) = (2.50\text{ mole})\ (0.08206\text{ L-atm/K-mol})\ (27 + 273\text{ K})$$

$$P = \frac{(2.50\text{ mole})\ (0.08206\text{ L-atm/K-mol})\ (300\text{ K})}{(5.50\text{ L})}$$

$$P = 11.19\text{ atm} \rightarrow 11.2\text{ atm}$$

LPT

How many moles of a gas occupy 2.67 L at STP?

Only one condition given → use Ideal Gas Law:
STP: Standard Temperature & Pressure → 1 atm and 273 K

$$PV = nRT$$

$$(1.00 \text{ atm})\ (2.67 \text{ L}) = n\ (0.08206 \text{ L-atm/K-mol})\ (273 \text{ K})$$

$$n = \frac{(1.00 \text{ atm})\ (2.67 \text{ L})}{(0.08206 \text{ L-atm/K-mol})\ (273 \text{ K})}$$

$n = 0.119184$ → 0.119 mole

LPT

Real (Non-Ideal) Gases

Ideal gases (because of decreased distance between molecules) often will differ in behavior from ideal gas equations. Modifications of ideal laws to account for molecular volumes and attractions are termed Real or non-ideal situations.

Van der Waals' Real Gas Equation

Modifies ideal gas law to account for molecular volume and attractions

New Volume term: V-nb (n = moles; b= specific constant for each gas)

New Pressure term: $P + an^2 / V^2$ (a= specific constant for each gas)

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

Values of a and b available in gas tables

Complex calculations best done by calculator / computer

Details of solution on next slide from Gas Law Primer
http://www-personal.umich.edu/~lpt/primer.htm

LPT

Assume : 80 ft^3 scuba cylinder with water capacity of 0.4 ft^3 (11.3 L)
This corresponds to ~ 100 moles of air

Using the Ideal gas equation at 25 ^{o}C, the pressure would be:

$$P = \frac{(100 \text{ moles})(0.0821 \text{ L-ata/deg K moles})(298 \text{ K})}{11.3 \text{ L}}$$

P = 216.5 ata (This corresponds to 3182.5 psia or 3168 psig)

Using Van DerWaal's for compressed air:

$$P = \frac{(100 \text{ mole})(0.0821 \text{ L-ata/K moles})(298 \text{ K})}{(11.3 \text{ l} - 0.036 \text{ L/mole } (100 \text{ mole}))} - \frac{(100 \text{ mole})^2 \ (1.33 \text{ L}^2\text{-atm/mole}^2)}{(11.3 \text{ L})^2}$$

Solving:

P = 317.738 ata - 104.159 ata

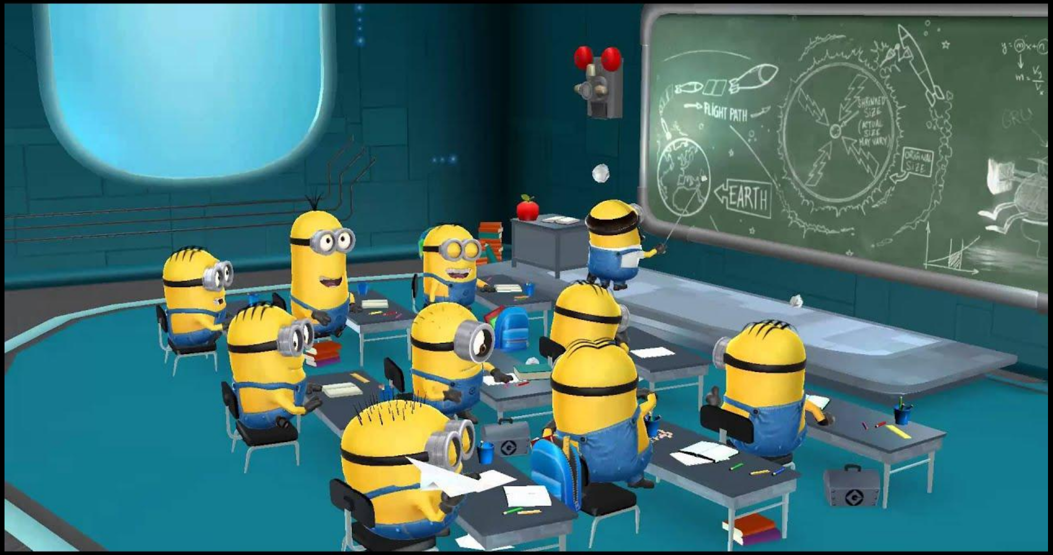
P = 213.58 ata (This corresponds to 3139 psia or 3124 psig)

For most scuba applications, this difference is not significant, but Becomes critical when mixing gases

LPT

Extra Example Problems
Using Gas Table For Setup

From CEM 101
All pressures in class assumed absolute



LPT

A sample of oxygen occupies a volume of 1240 mL at temperature of 45° C. What is the volume of this gas sample if the temperature is raised to 85° C?

	Pressure	Volume (mL)	Temperature (ºC)	→	Temperature (K)
Initial	constant	1240	45	+ 273	318
Final	constant	?	85	+ 273	358

Pressure constant, use Charles' Law

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

$$\frac{1240 \text{ mL}}{318 \text{ K}} = \frac{V_2}{358 \text{ K}} \longrightarrow \frac{(1240 \text{ mL})(358 \text{ K})}{318 \text{ K}} = V_2$$

$$V_2 = 1395.97 \text{ mL} \rightarrow 1400 \text{ mL} \quad (1.40 \times 10^3 \text{ mL})$$

LPT

Calculate the volume a gas will occupy at 15 °C if the gas has a volume of 830. mL at 42 °C.

	Pressure	Volume (mL)	Temperature (ºC)	→	Temperature (K)
Initial	constant	830	42	+ 273	315
Final	constant	?	15	+ 273	288

Pressure constant, use Charles' Law

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

$$\frac{830\ \text{mL}}{315\ \text{K}} = \frac{V_2}{288\ \text{K}} \longrightarrow \frac{(830\ \text{mL})(288\ \text{K})}{315\ \text{K}} = V_2$$

$$V_2 = 758.857\ \text{mL} \rightarrow 759\ \text{mL}$$

Calculate the final temperature in °C of a gas initially at 39 °C whose volume changes from 348 ml to 657 mL. The pressure remains constant.

	Pressure	Volume (mL)	Temperature (ºC)	→	Temperature (K)
Initial	constant	348	39	+ 273	312
Final	constant	657	?	+ 273	?

Pressure constant, use Charles' Law

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

$$\frac{348 \text{ mL}}{312 \text{ K}} = \frac{657 \text{ mL}}{T_2} \longrightarrow \frac{(657 \text{ mL})(312 \text{ K})}{(348 \text{ mL})} = T_2$$

$$T_2 = 589.038 \rightarrow 589 \text{ K}$$

$$T_2 = 589 \text{ K} - 273 = 316 \text{ ºC}$$

A sample of oxygen has a pressure of 1420. mm Hg at a temperature of 75 °C. What is the pressure of this gas sample if temperature is lowered to 19° C?

	Pressure (torr)*	Volume	Temperature (ºC)	→	Temperature (K)
Initial	1420	constant	75	+ 273	348
Final	?	constant	19	+ 273	292

Volume constant, use Guy-Lussac's Law

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

$$\frac{1420\ \text{torr}}{348\ \text{K}} = \frac{P_2}{292\ \text{K}} \longrightarrow \frac{(1420\ \text{torr})\,(292\ \text{K})}{348\ \text{K}} = P_2$$

$P_2 =$ 1191.49 torr → 1190 torr → 1190 mm Hg

*** 1 mm Hg = 1 torr**

LPT

Calculate the pressure a gas will exert at 65 °C if the gas has a pressure of 830. torr at 52 °C.

	Pressure (torr)	Volume	Temperature (°C)	→	Temperature (K)
Initial	830	constant	52	+ 273	325
Final	?	constant	65	+ 273	338

Volume constant, use Guy-Lussac's Law

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

$$\frac{830\text{ torr}}{325\text{ K}} = \frac{P_2}{338\text{ K}} \longrightarrow \frac{(830\text{ torr})(338\text{ K})}{325\text{ K}} = P_2$$

P_2 = 863.2 torr → 863 torr

LPT

A sample of nitrogen has a pressure of 1420. torr at a temperature of 75 °C. What is the ° C temperature of this gas if the pressure is lowered to 258 torr?

	Pressure (torr)	Volume	Temperature (°C)	→	Temperature (K)
Initial	1420	constant	75	+ 273	348
Final	258	constant	?	+ 273	?

Volume constant, use Guy-Lussac's Law

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

$$\frac{1420 \text{ torr}}{348 \text{ K}} = \frac{258 \text{ torr}}{T_2} \longrightarrow \frac{(258 \text{ torr})(348 \text{ K})}{(1420 \text{ torr})} = T_2$$

$$T_2 = 63.2262 \text{ K} \rightarrow T_2 = 63.2 \text{ K}$$

$$T_2 = 63.2 \text{ K} - 273 = -210\ °\text{C}$$

LPT

At 723 mm Hg a gas has a volume of 294 mL. What is the new volume of this gas if the pressure is changed to 585 mm Hg?

	Pressure (mm Hg)	Volume (mL)	Temperature (°C)	Temperature (K)
Initial	723	294	constant	constant
Final	585	?	constant	constant

Temperature constant, use Boyle's Law

$$P_1 V_1 = P_2 V_2$$

$$(723 \text{ mm Hg})\ (294 \text{ mL}) = (585 \text{ mm Hg})\ V_2$$

$$\frac{(723 \text{ mm Hg})\ (294 \text{ mL})}{(585 \text{ mm Hg})} = V_2$$

$$V_2 = 363.354 \text{ mL} \rightarrow 363 \text{ mL}$$

LPT

At 723 torr a gas has a volume of 294 mL. What is the new pressure of this gas if the volume is changed to 1256 mL?

	Pressure (torr)	Volume (mL)	Temperature (°C)	Temperature (K)
Initial	723	294	constant	constant
Final	?	1256	constant	constant

Temperature constant, use Boyle's Law

$$P_1 V_1 = P_2 V_2$$

$$(723 \text{ torr})\,(294 \text{ mL}) = (1256 \text{ mL})\, P_2$$

$$\frac{(723 \text{ torr})\,(294 \text{ mL})}{(1256 \text{ mL})} = P_2$$

$P_2 = 169.237 \rightarrow$ 169 torr

LPT

A sample of neon with a volume of 825 mL at a temperature of 37 °C and a pressure of 600. torr is heated to a temperature of 68 °C and a pressure of 940. mm Hg. What is the new volume of the gas?

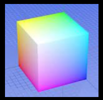
	Pressure (torr)	Volume (mL)	Temperature (ºC)	→	Temperature (K)
Initial	600	825	37	+ 273	310
Final	940	?	68	+ 273	341

Pressure, Volume & Temperature change, use General Gas Law

$$\frac{p_1 v_1}{t_1} = \frac{p_2 v_2}{t_2}$$

$$\frac{(600\ \text{torr})\ (825\ \text{mL})}{310\ \text{K}} = \frac{(940\ \text{torr})\ v_2}{341\ \text{K}} \rightarrow \frac{(600\ \text{torr})\ (825\ \text{mL})\ (341\ \text{K})}{(310\ \text{K})\ (940\ \text{torr})} = v_2$$

$$V_2 = 579.255\ \text{mL} \rightarrow 579\ \text{mL}$$


LPT

A sample of argon with a volume of 4.37 L at a temperature of 58 °C and a pressure of 725 torr is cooled to a temperature of 22 °C and a pressure of 615 mm Hg. What is the new volume of the gas?

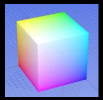
	Pressure (torr)	Volume (L)	Temperature (°C)	→	Temperature (K)
Initial	725	4.37	58	+ 273	331
Final	615	?	22	+ 273	295

Pressure, Volume & Temperature change, use General Gas Law

$$\frac{p_1 v_1}{t_1} = \frac{p_2 v_2}{t_2}$$

$$\frac{(725\ \text{torr})\,(4.37\ \text{L})}{331\ \text{K}} = \frac{(615\ \text{torr})\,v_2}{295\ \text{K}} \rightarrow \frac{(725\ \text{torr})\,(4.37\ \text{L})\,(295\ \text{K})}{(331\ \text{K})\,(615\ \text{torr})} = v_2$$

$$V_2 = 4.59133\ \text{L} \rightarrow 4.59\ \text{L}$$


LPT

A sample of nitrogen with a volume of 14.7 L at a temperature of 95 °C and a pressure of 485 torr is brought to STP. What is the new volume?

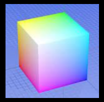
	Pressure (torr)	Volume (L)	Temperature (°C)	→	Temperature (K)
Initial	485	14.7	95	+ 273	368
Final	760	?	0	+ 273	273

Pressure, Volume & Temperature change, use General Gas Law

$$\frac{p_1 v_1}{t_1} = \frac{p_2 v_2}{t_2}$$

$$\frac{(485 \text{ torr})(14.7 \text{ L})}{368 \text{ K}} = \frac{(760 \text{ torr})\, v_2}{273 \text{ K}} \rightarrow \frac{(485 \text{ torr})(14.7 \text{ L})(273 \text{ K})}{(368 \text{ K})(760 \text{ torr})} = v_2$$

$$V_2 = 6.95922 \text{ L} \rightarrow 6.96 \text{ L}$$


LPT

A sample of neon at STP has a volume of 286 L. What is the pressure in atmospheres if the temperature is changed to 95 °C at a new volume of 26.5 L?

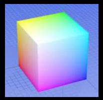
	Pressure (ata)*	Volume (L)	Temperature (°C)	→	Temperature (K)
Initial	1.00	286	0	+ 273	273
Final	?	26.5	95	+ 273	368

Pressure, Volume & Temperature change, use General Gas Law

$$\frac{p_1 v_1}{t_1} = \frac{p_2 v_2}{t_2}$$

$$\frac{(1.00 \text{ ata})(286 \text{ L})}{273 \text{ K}} = \frac{P_2(26.5 \text{ L})}{368 \text{ K}} \rightarrow \frac{(1.00 \text{ ata})(286 \text{ L})(368 \text{ K})}{(273 \text{ K})(26.5 \text{ L})} = P_2$$

$$P_2 = 14.5481 \text{ ata} \rightarrow 14.5 \text{ ata}$$


* ata = atmospheres absolute

LPT

A sample of xenon with a volume of 825 mL at a temperature of 37 °C and a pressure of 600. torr is changed to a pressure of 940. mm Hg at a volume of 628 mL. What is the temperature in °C of the gas?

	Pressure (torr)	Volume (mL)	Temperature (°C)	→	Temperature (K)
Initial	600	825	37	+ 273	310
Final	940	628	?	+ 273	?

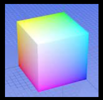
Pressure, Volume & Temperature change, use General Gas Law

$$\frac{p_1 v_1}{t_1} = \frac{p_2 v_2}{t_2}$$

$$\frac{(600 \text{ torr})(825 \text{ mL})}{310 \text{ K}} = \frac{(940 \text{ torr})(628 \text{ mL})}{T_2} \rightarrow \frac{(940 \text{ torr})(628\text{mL})(310 \text{ K})}{(600 \text{ torr})(825 \text{ mL})} = T_2$$

$$T_2 = 369.695 \text{ K} \rightarrow 370 \text{ K}$$

$$T_2 = 370 \text{ K} - 273 = 97\ ^{o}\text{C}$$


Solving Gas Law Problems:
Merely a Matter of Paying Attention to Details!

(As far as we know, picture is public domain)

LPT