Your LARC and PEER Tutors

PEER TUTORING

- Please take advantage of the free tutoring provided by the General Chemistry Peer Tutoring in room RH517. Your tutors are **Ani Orujyan**, **Anais Panossian**, and **Joseph Bui**. Here is their tutoring schedule.

LARC (watch a 2 min YouTube video if you are new to LARC)

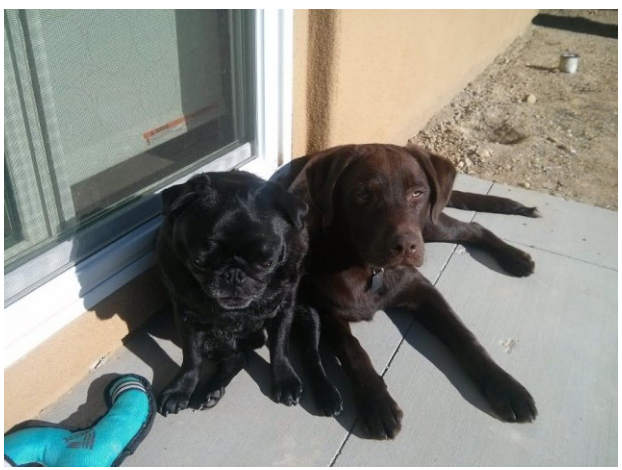
- LARC tutorials: two one-hour sessions per week ($110/qtr).
- The LARC tutor for this class is **Cally Chung**. Use this link to find her tutorial schedule.
- The job of LARC tutors is to help you study more efficiently and to help you understand and retain the course material.
- For the schedule please go to LARC Website. Enroll through WebReg for the appropriate LARC Tutorial.
- If you have any questions email larc@uci.edu or stop by the front desk in Rowland Hall 284.

Your Instructor

Random/boring facts about your Instructor

My dad was a ranger in Yosemite, where we lived each summer.

I'm a huge fan of world cup cross-country skiing.



I have two dogs, Odin and Bean.

Your Instructor's Research

Ultrafine Aerosol Laboratory
http://sites.uci.edu/uagroup/

Limonene ($C_{10}H_{16}$) + Ozone (O_3) $\rightarrow$ Gaseous Products

Gaseous Products $\xrightarrow{\text{nucleation}}$ New Particles

Exothermic reaction (Chapter 5) involving two gases (Chapter 10)
Condensation of gaseous compounds into liquids (Chapters 11,12)

Why should we care about particle formation?

Particles are a massive worldwide health threat

Beijing, Jan. 1, 2017
(video: Chas Pope)

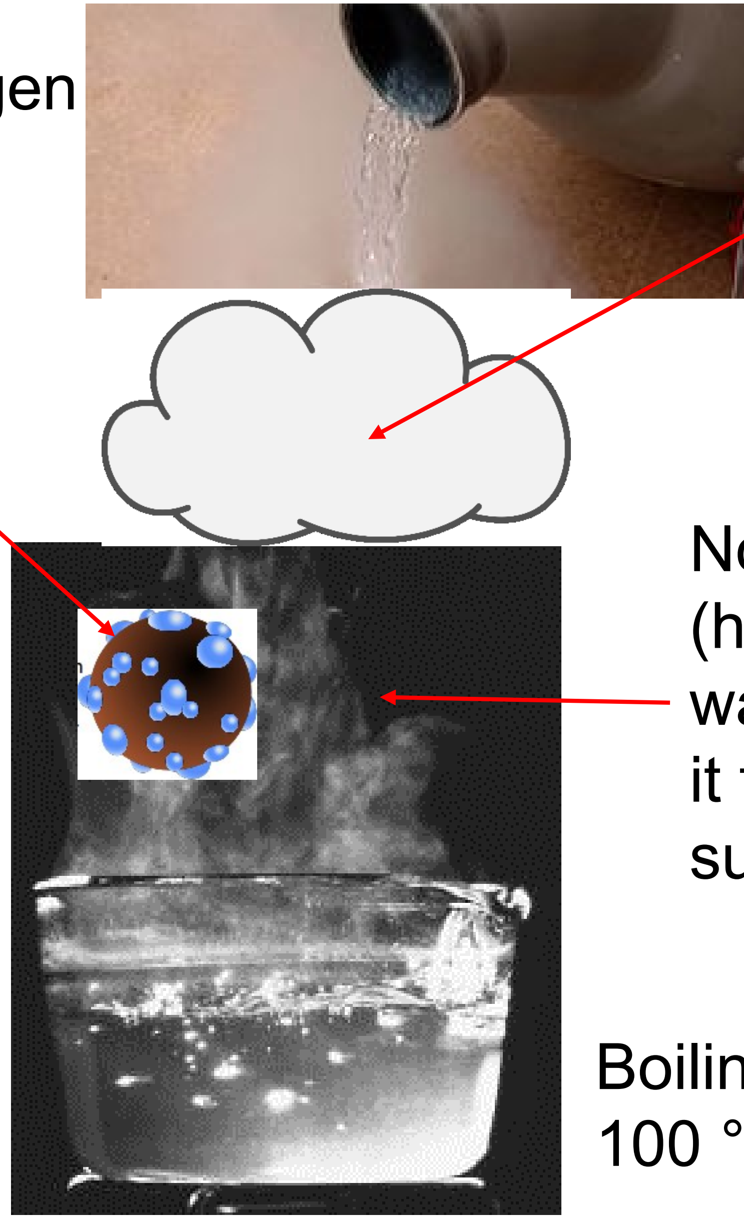
Particles can affect climate by making and modifying clouds

(video: NASA)

Cloud Formation Demo

Liquid Nitrogen
-196 °C

Water vapor
collects on
room particles



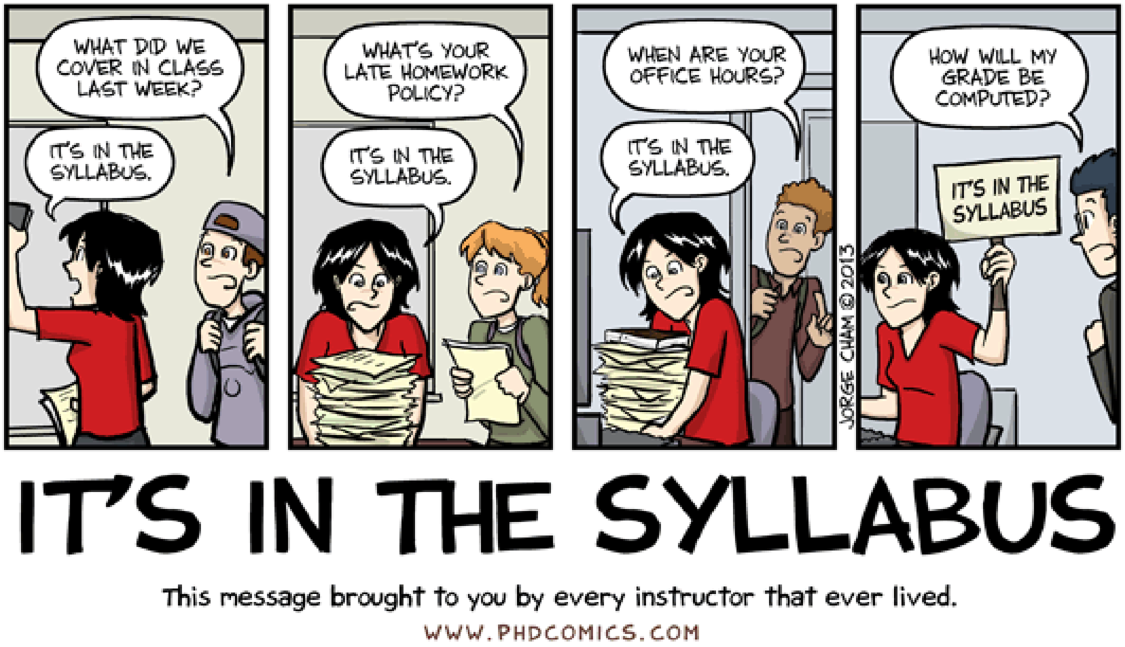
Cool air
drives equilibrium
towards
condensation

Non-equilibrium
(high) concentration
water vapor causes
it to condense on
surfaces (particles).

Boiling water
100 °C

Course Information

- Read all the e-mail announcements carefully (copied to the website).
- There is a piazza group for this course if you want to post a question to everyone about homework, exams, logistics, etc.
- All 3 exams are **COMPREHENSIVE** and **REQUIRED**. Please check the exam schedule on the course website ASAP and switch sections if you have a conflict with one of the exam dates.
- Prof. Smith or TAs cannot help you with your enrollment issues. Please go to the chemistry graduate office with all such issues.



Before asking the instructors and TAs for any info, please read the course information on the syllabus section of the website or ask your classmates (e.g., using the piazza group).

Homework

[Modified Mastering Chemistry]

The registration instructions can be found in the syllabus section of the course website. Course ID: smith66181

IMPORTANT

- *If you can't solve a given problem **SKIP IT**; **never press the "give up" button**! The homework submission windows will re-open during finals week to allow studying and wrapping up the few(!) problems you could not get to.*
- *Each module takes 5-7 hours to complete. Start early!!!*
- *Solve the problems yourself, do not copy the solutions from others. Otherwise you will not be prepared to the exams!!!*

Grading

	Homework	Midterm 1	Midterm 2	Final
Percent weight	14%	18%	28%	40%

- For each of the above activities you will receive a percentage score.
- *Your final percentage score is calculated using the weights shown above:*
  0.14 x *HW% + 0.18* x *MT1% + 0.28* x *MT2% + 0.4* x *FinalEx%*
- Your **final grade** will be obtained using a curve.
- For fairness, the same curve will be used for all sections of Chemistry 1B this quarter.
- You will have an idea of your grade since the class average (mean of the curve) will lie **approximately** at the border between a B- and a C+.
- I will provide your current points along with the class mean after every exam.

Common Final

- All Chem 1B students will have a Common Final on Sunday March 18th, 10:30 am – 12:30 pm.
- Students who have a conflict should fill out the makeup exam form in a timely manner (by week 7).
- That form and other info are available on http://sites.uci.edu/chemcommonfinal/.

Chemistry Common Final

Chemistry Common Final.

Chemistry Common Final.

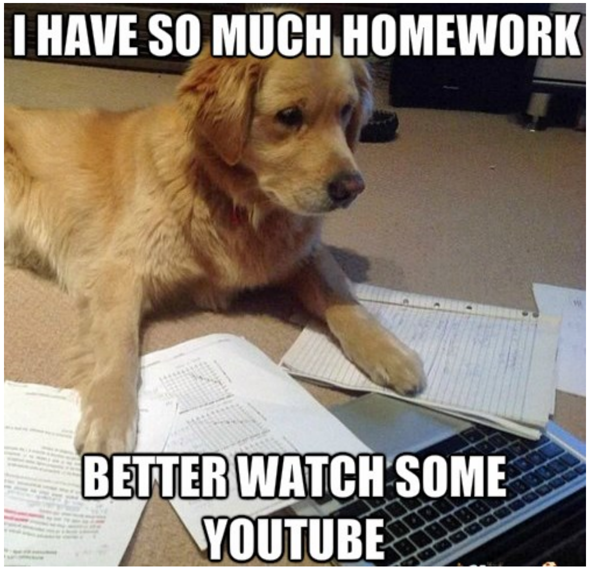
Chem 1A, 1B and 1C Common Final Exam

The UCI Chemistry Department offers a single Common Final exam for all students taking each chemistry class in the general chemistry series (1A, 1B, 1C). We are using this Common Final exam for several reasons, including:

- The common final helps ensure that all students are taught and are responsible for material at the same level of rigor and are similarly prepared for subsequent math courses.
- The Common Final helps unify grading across different sections of courses, ensuring a fairer grading system for all students, not dependent on particular section enrollment.
- The Common Final helps the chemistry department assess and redesign the chemistry courses to help aid student learning.

How To Pass This Class

1. **Read the book** both before and after attending a lecture
2. Attend every lecture and **take notes by hand** even though lecture notes are posted
3. Work in **groups** to know your level relative to your peers, to get help, and to learn while helping others
4. Take advantage of **tutors**
5. Take advantage of **office hours** (schedules on website)
6. Solve as many **practice problems** in lectures, discussions, and tutoring sessions as you can
7. **Do not procrastinate !!!!!!!!!** Give yourself enough time for homework (at least 6 hours per module).
8. Eliminate distractions (Facebook, Twitter, YouTube).



Respect Your Classmates!

1. Please do not use cell phones or computers for surfing the internet during lectures. It is disrespectful to your instructor may be very distracting to students behind you.
2. If you must use the computers for educational purposes (such as note taking) please try to sit in the last row or around the edges so that students behind you are not too distracted.
3. Turn your cell phones off so that they do not ring.
4. Please do not start packing your belongings until you are dismissed (a few minutes before the official lecture ending time). Making noise by packing in the last minute may be annoying to students around you.
5. Please refrain from noisy activities such as talking, chewing, slurping, burping, etc.
6. If you are sick and contagious please consider skipping a lecture to avoid spreading the disease and to get better.
7. If you are right-handed avoid taking left-handed seats.

Chem 1B Topics

Chemistry: Structure and Properties, *Nivaldo J. Tro; Pearson Education Inc. 2015, 2017.*

Week*	Chapter	Topic
1-2	*10 (11)*	*Gases*
3-4	*11 (12)*	*Liquids, Solids, Intermolecular Forces*
5	*12 (13)*	*Phase Diagrams & Crystalline Solids*
6-7	*13 (14)*	*Properties of Solutions*
8-9	*9 (10)*	*Thermochemistry*
10	*18 (19)*	*Free Energy & Thermodynamics*

** We will be adjusting the exact schedule as we go along…*

Gases

Elements that exist as gases at 25°C and 1 atm

H_2, N_2, O_2, F_2, Cl_2, He, Ne, Ar, Kr, Xe, Rn

1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg							

Compounds that exist as gases at 25°C and 1 atm

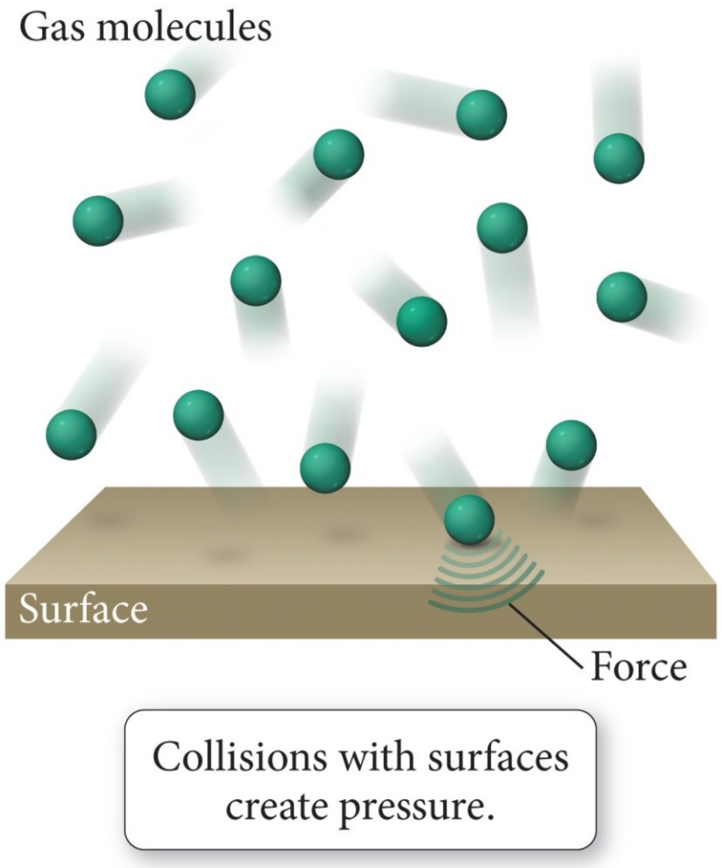
CO, CO_2, SO_2, N_2O, NO, NO_2, NO_3, CH_4, C_2H_6, …

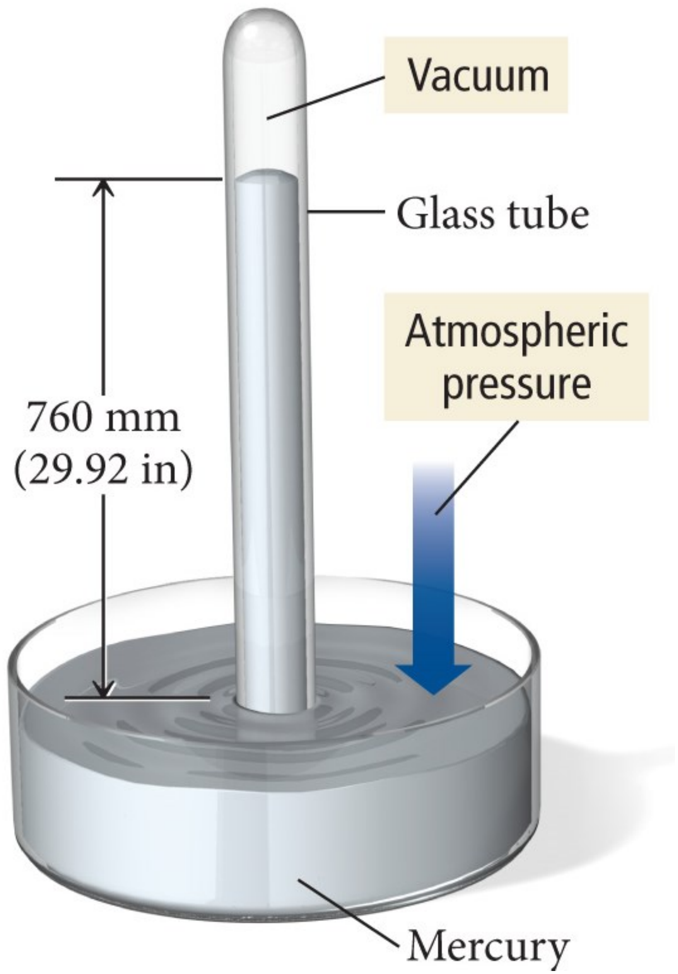
Physical Characteristics of Gases

- Gases assume the volume and shape of their containers
- Gases are the most compressible state of matter
- Two gases mix evenly when confined to the same container
- Gases have much lower densities than liquids and solids and have to be compressed for storage or transport

Gas carrier – designed to transport liquefied natural gas (LNG)

Gas Pressure





$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

1 Pascal (Pa) = 1 N/m^2

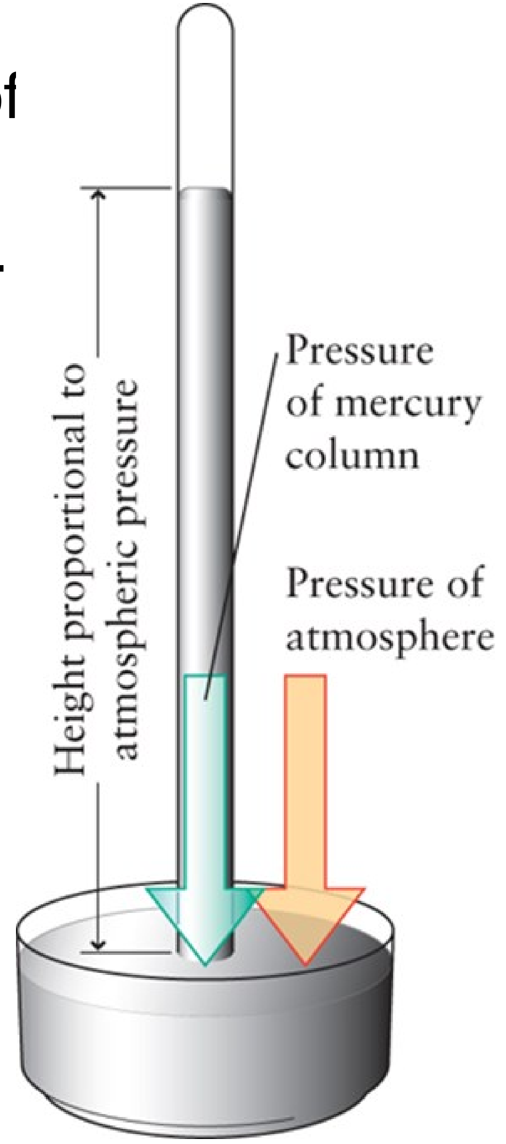
1 atm = 101,325 Pa

1 atm = 760 mmHg = 760 torr

Barometer

Evangelista Torricelli, 17th century, a student of Galileo, made the first barometer – an instrument for measuring atmospheric pressure.

A glass tube of liquid mercury is inverted into a beaker. The mercury falls until its pressure matches the atmosphere's pressure.



$$P = \frac{Force}{Area} = \frac{mass \cdot g}{Area}$$

g = gravity constant

$$mass = d \cdot h \cdot Area$$

d = density of mercury

$$P = d \cdot g \cdot h$$

h = height of mercury column

Why Mercury? (Sample Problem)

The height of the column of mercury in a barometer is 760.0$\underline{0}$ mm. What is the atmospheric pressure in pascals? If mercury were replaced by water, what would the height of the column be? d_{Hg} = 13,595 kg·m^{-3} , d_{water} = 998.0 kg·m^{-3}, g = 9.80665 m·s^{-2}.

Pressure and Weather

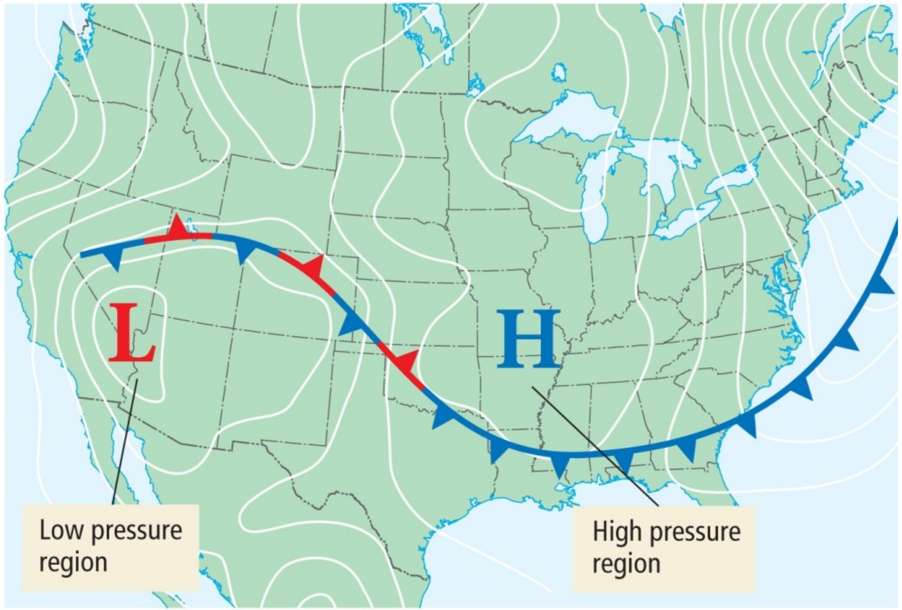
Meteorologists commonly report pressure in two units:

- In mbar: 1 mbar = 10^2 Pa; 1 atm = 1013.25 mbar
- In inches of Hg (only in US): 1 atm = 29.92 in Hg

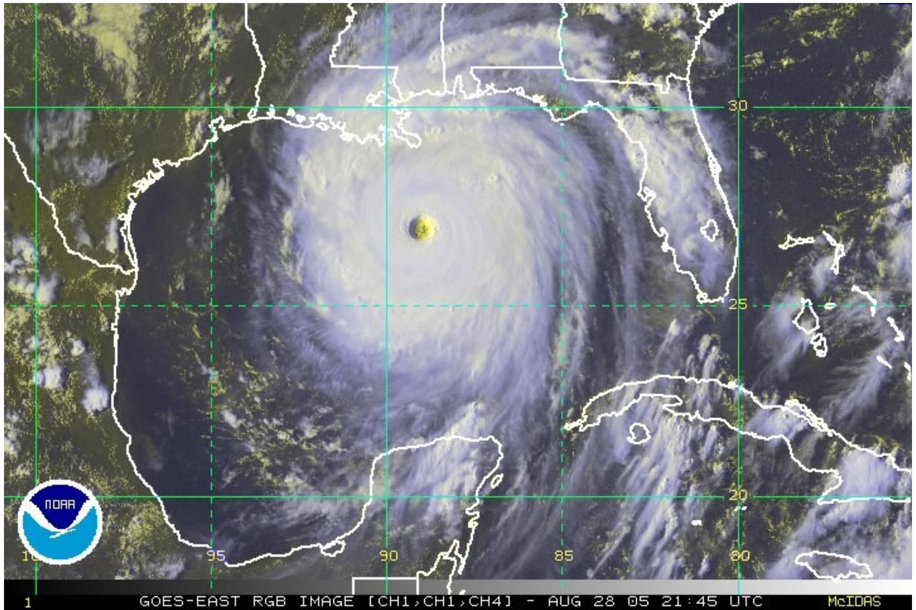
High-pressure area (anticyclone) – surface pressure is greater than the surroundings

Low-pressure area (cyclone) – surface pressure is lower than the surroundings. Can develop into powerful storm.

Typical barometric map



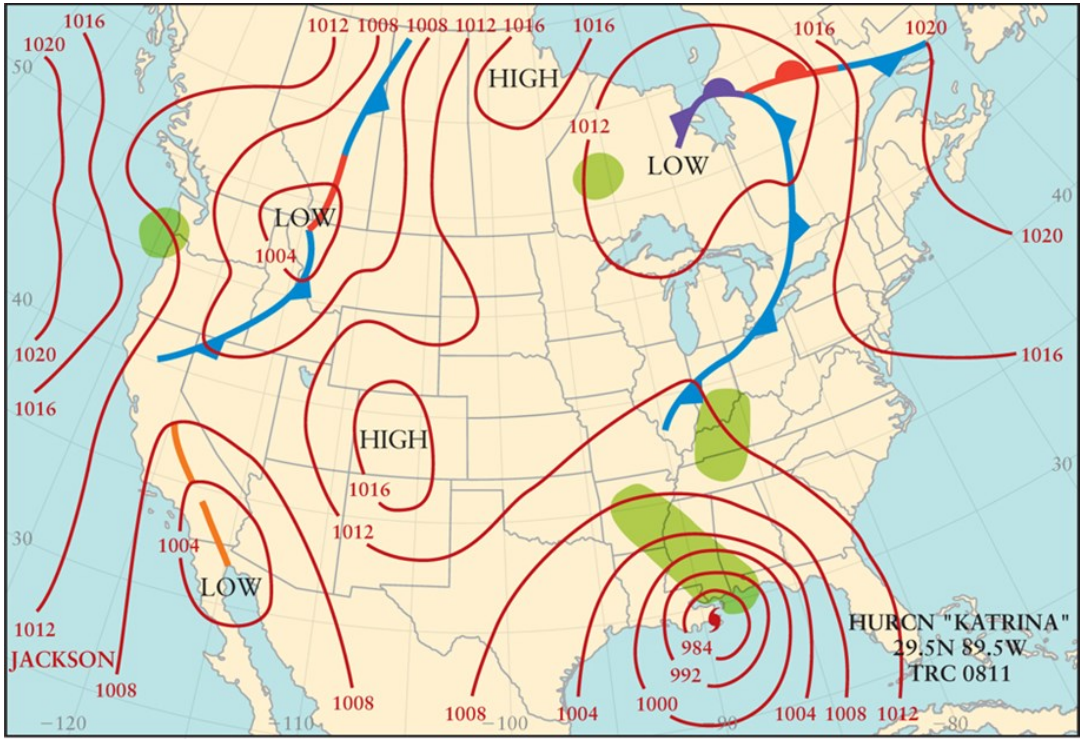
Hurricane Katrina



Pressure and Weather

Units of Pressure

This is a weather map of North America during Hurricane Katrina in 2005. Isobars show atmospheric pressure. The pressure at the center of the hurricane had fallen to as low as 902 mbar. What is this pressure in atm?



1 bar = 10^5 Pa

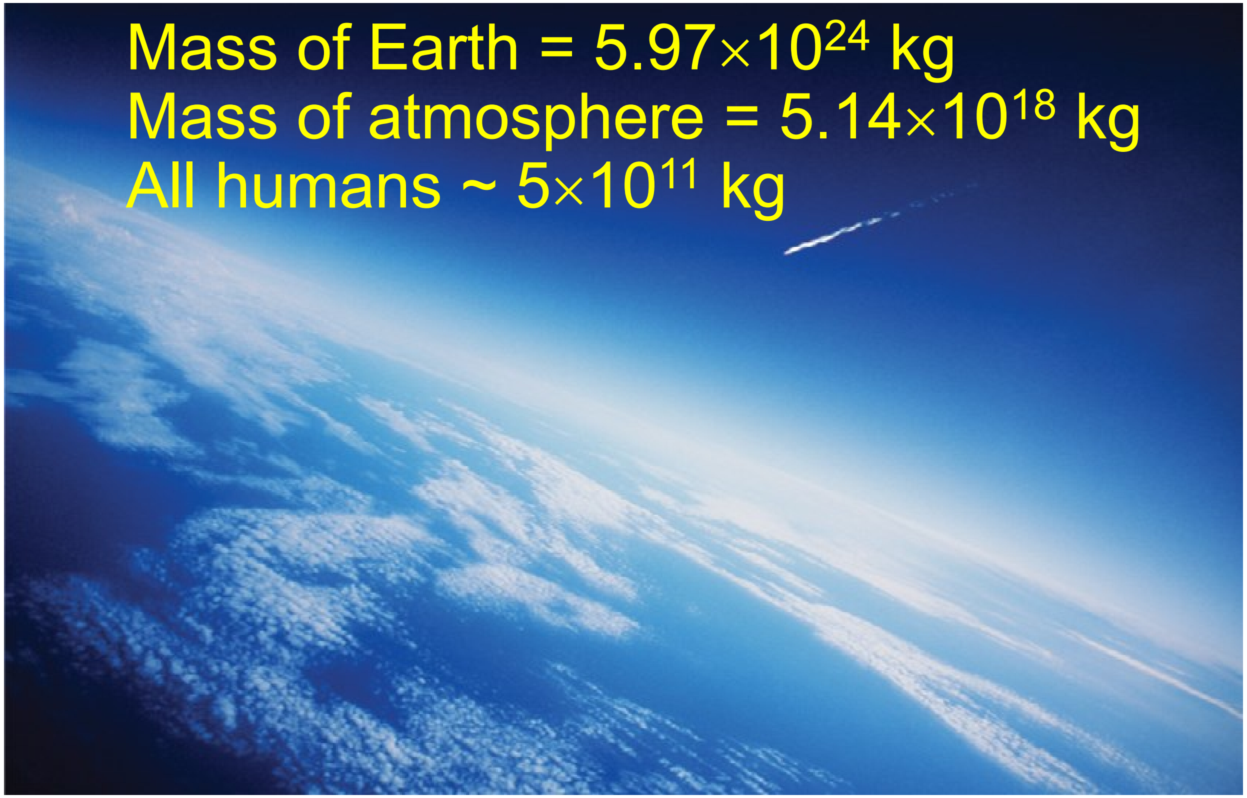
1 mbar = 10^2 Pa

1 atm = 1.01325×10^5 Pa

$$\frac{1 \text{ atm}}{1.01325 \times 10^5 \text{ Pa}} \times \frac{10^2 \text{ Pa}}{1 \text{mbar}} \times \frac{902 \text{ mbar}}{1} = 0.890 \text{ atm}$$

Pressure at Higher Altitudes

The pressure drops roughly ***exponentially*** with altitude. Half of the atmospheric mass lies within 5.5 km (3.4 mi) above our heads. 99.9% mass is contained below the stratopause (50 km).



Height	Pressure
Surface	1 atm
5 km	0.5 atm
8 km	0.3 atm
16 km	0.1 atm
31 km	0.01 atm

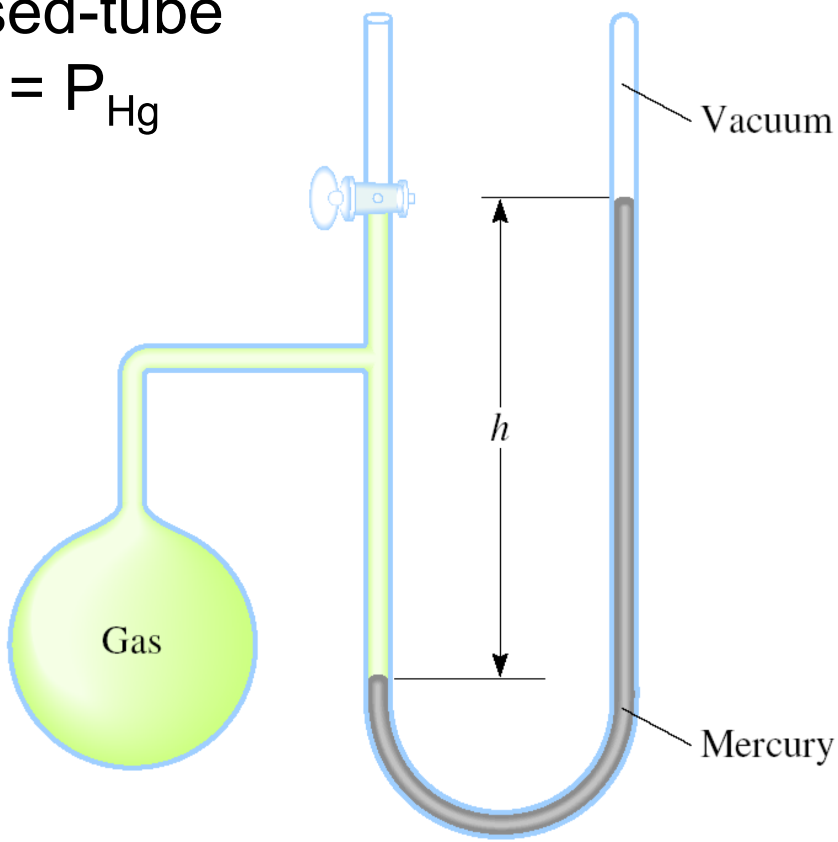
The atmosphere is very thin and fragile! If Earth were the size of a basketball, the atmosphere would only be 1 mm thick!

Manometer

Measures gas pressure from a height of a liquid needed to counterbalance the pressure in the container.

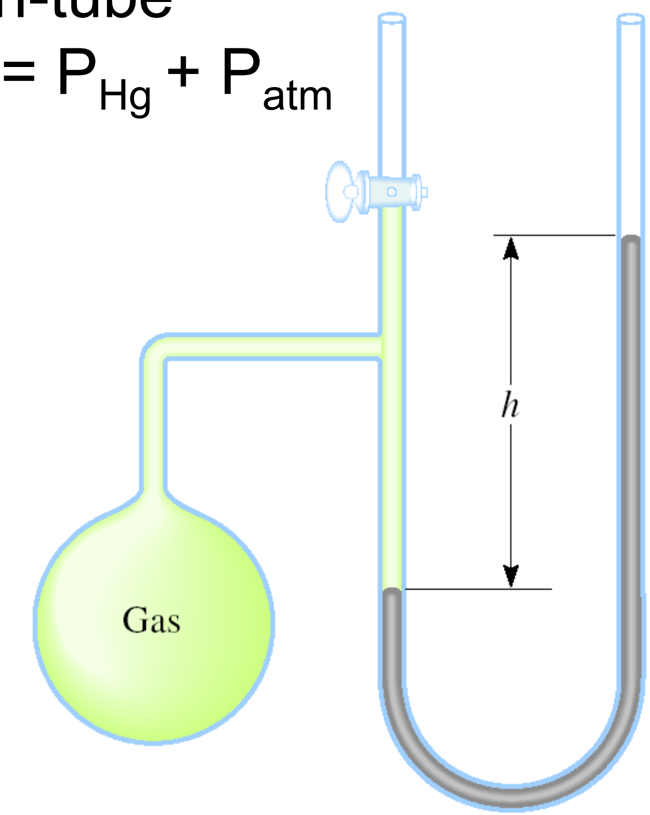
Closed-tube

$P_{gas} = P_{Hg}$



Open-tube

$P_{gas} = P_{Hg} + P_{atm}$

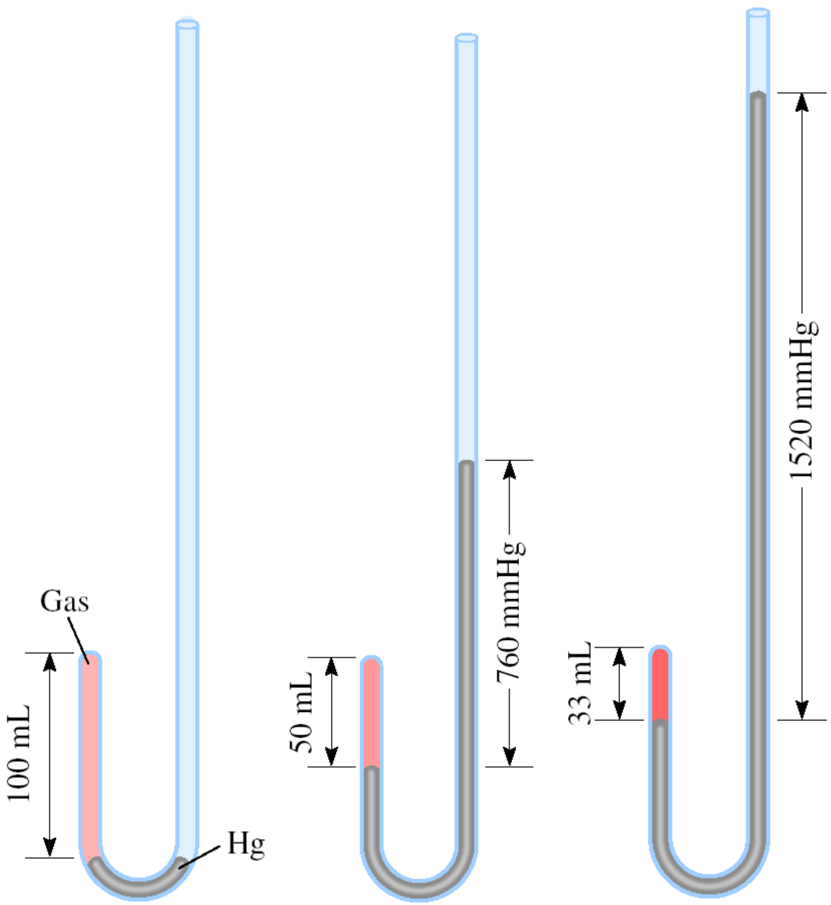


Gas pressure balanced by the weight of the mercury column

Gas pressure balanced by the Hg column AND atmospheric pressure

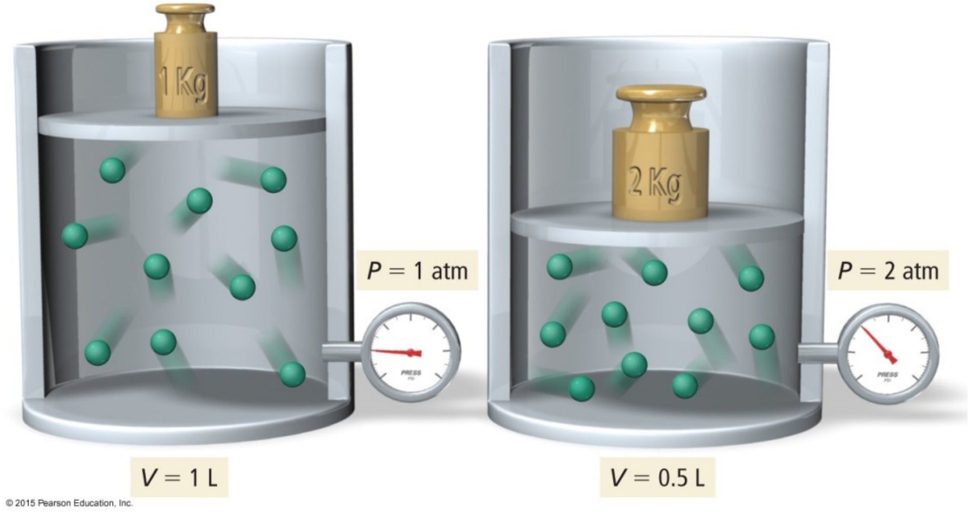
Boyle's Law

Robert Boyle studied the relationship between pressure and volume of a gas. He took a J-shaped tube, with the short end sealed, and poured different amounts of mercury into the tube.



As *P* (h) increases

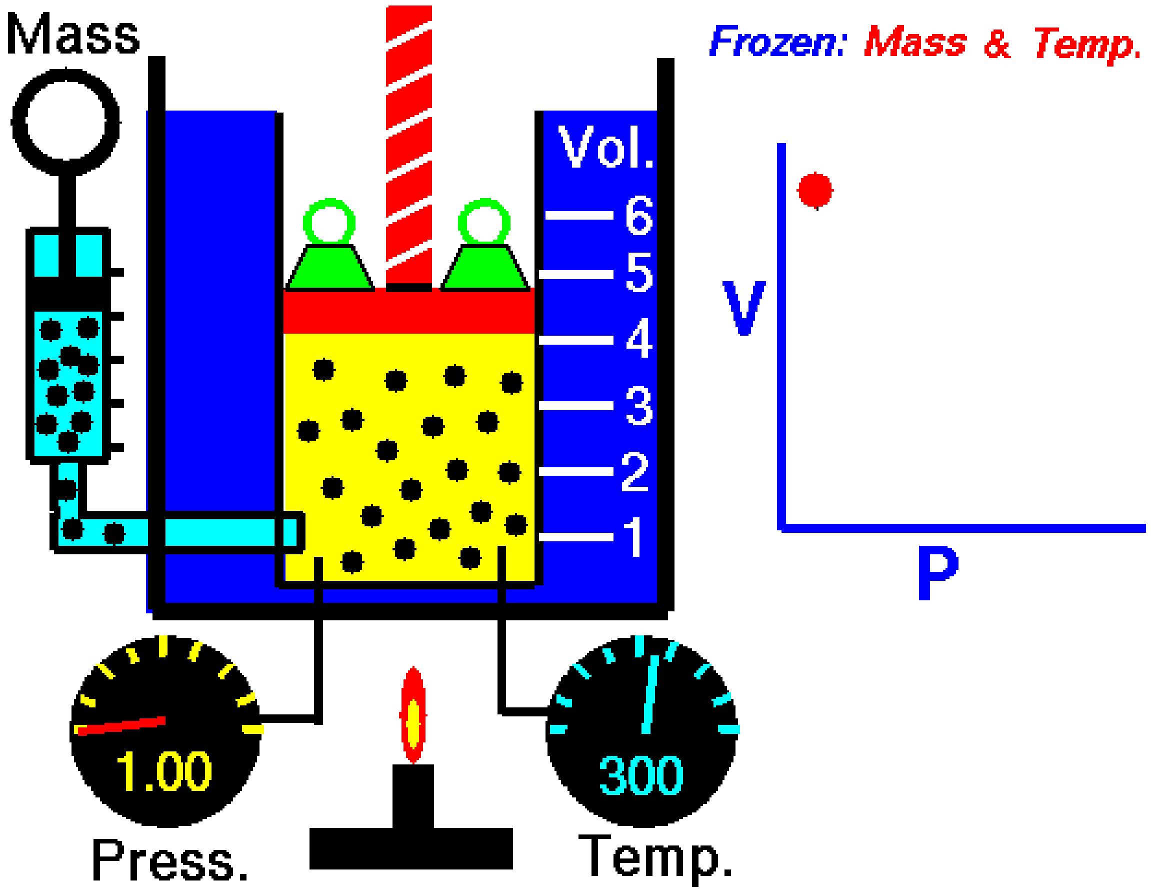
***V* decreases**



***Discussion**: what happens on a molecular level?*

Boyle's Law

At constant temperature and constant amount of gas:



∝ means "proportional to"

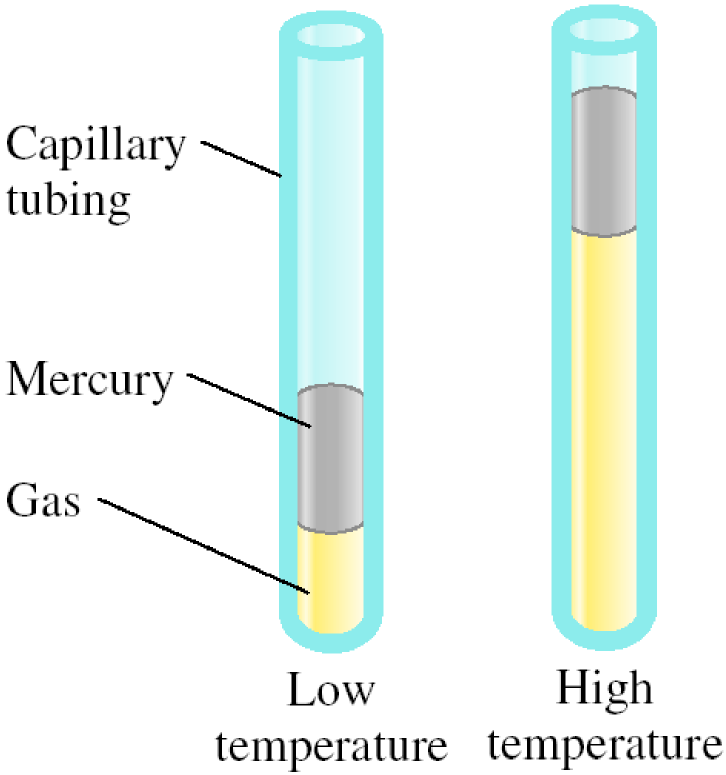
$$P \propto 1/V$$

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

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

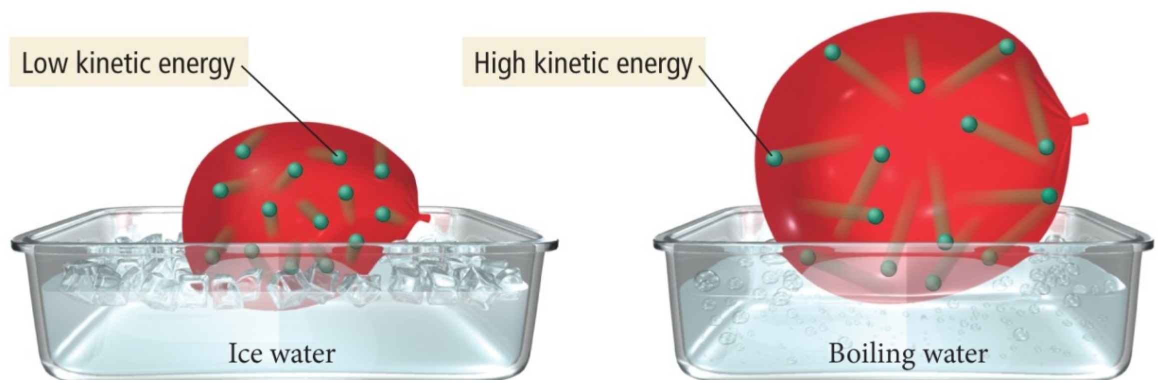
Charles's Law

Jacques Charles (an inventor and balloonist) studied the relationship between temperature and volume of a gas. He observed that volume increases linearly with temperature.



As *T* increases

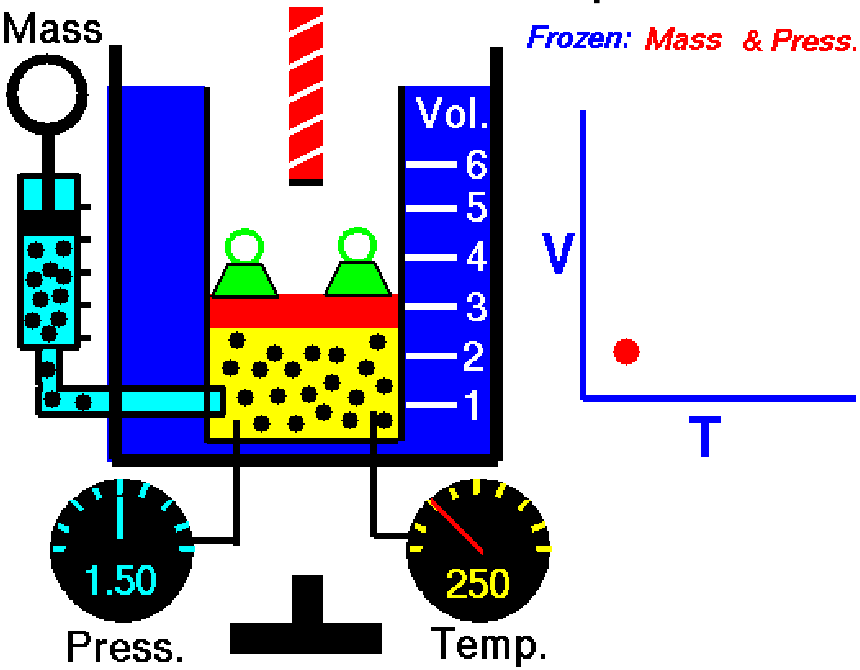
_V_ increases

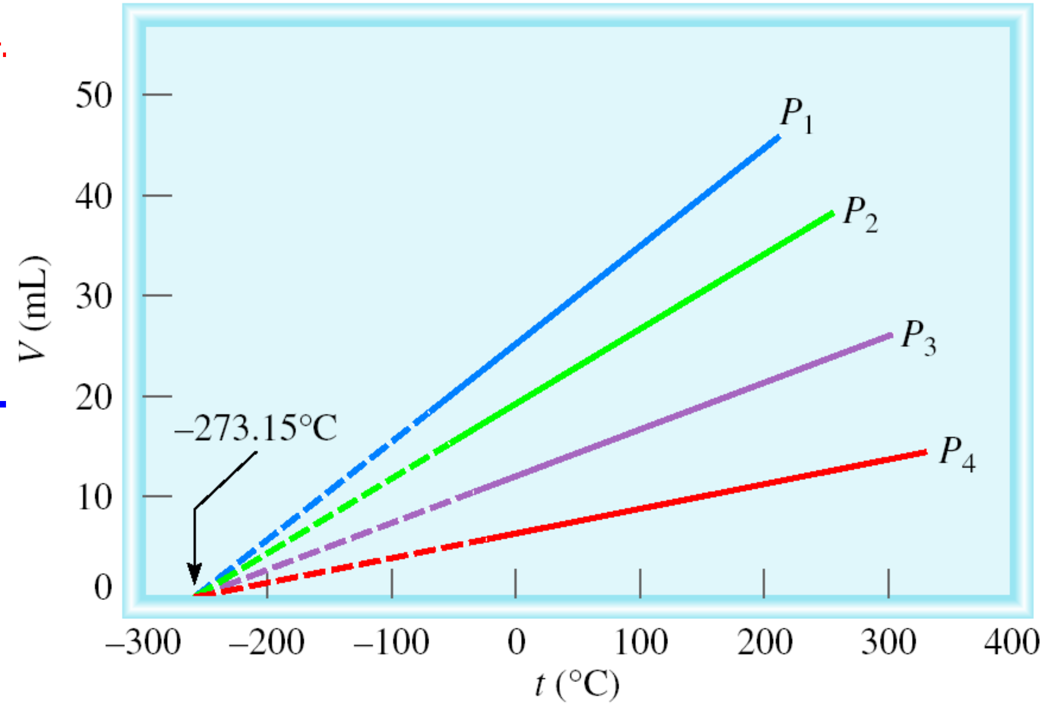


***Discussion**: what happens on a molecular level?*

Charles's Law

At constant external pressure





$V \propto T$

V = constant x T

$V_1/T_1 = V_2/T_2$

Temperature **must** be in Kelvin

T (K) = T (^{0}C) + 273.15

Sample Problem

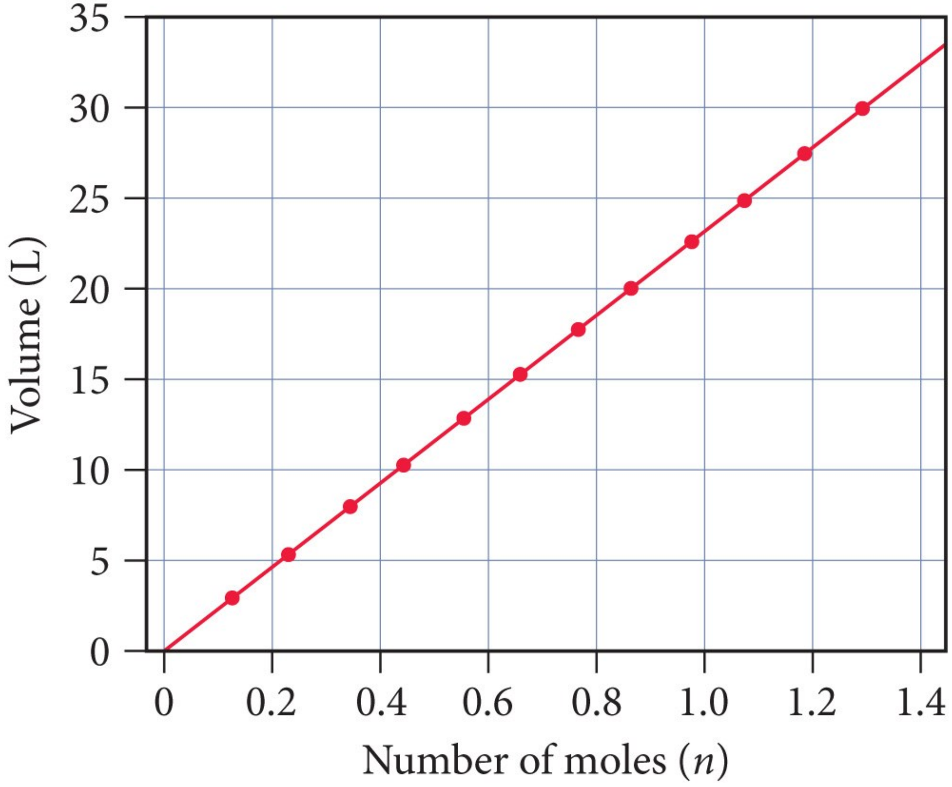
A sample of carbon monoxide gas occupies 3.20 L at 125 ^{0}C. At what temperature will the gas occupy a volume of 1.54 L if the pressure remains constant?

Sample Problem

A gas occupies a volume of 946 mL at a pressure of 726 mmHg. What is the pressure of the gas (in mmHg) if the volume is reduced at constant temperature to 154 mL?

Avogadro's Law

Amedeo Avogadro observed a linear relationship between the amount of gas and the volume it occupies at constant temperature and pressure



$V \propto$ number of moles (n)

V = constant x n

$V_1 / n_1 = V_2 / n_2$

The nature of the gas doesn't matter (ideal gas approximation).

Ideal Gas Law

Boyle's law: $V \propto \frac{1}{P}$ (at constant n and T)

Charles' law: $V \propto T$ (at constant n and P)

Avogadro's law: $V \propto n$ (at constant P and T)

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

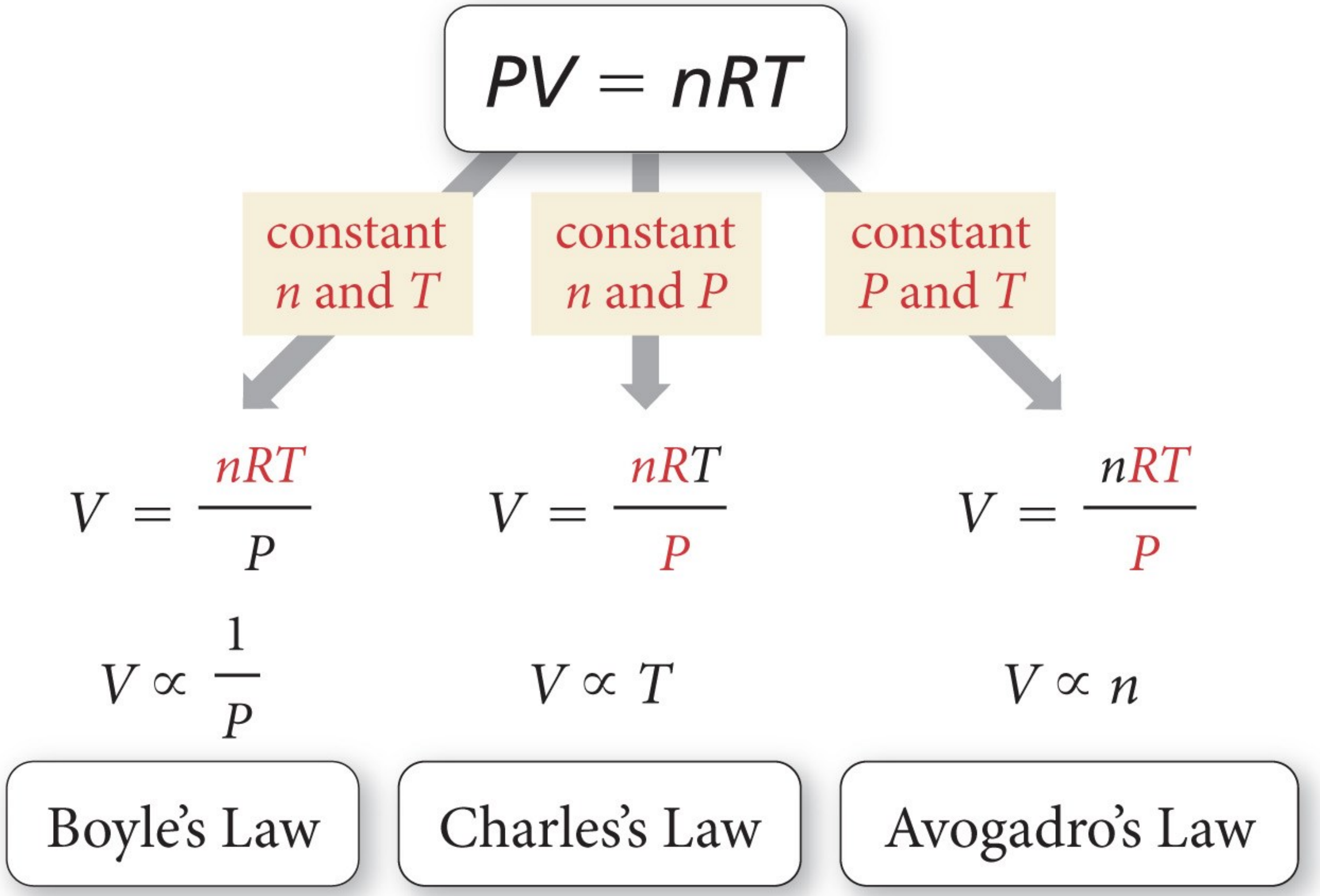
$$V = \text{constant} \times \frac{nT}{P} = R\frac{nT}{P}$$

$$\boxed{PV = nRT}$$

R is the **gas constant**

- $R = 0.0820573 \text{ L atm K}^{-1} \text{ mol}^{-1}$
- $R = 8.31445 \text{ m}^3 \text{ Pa K}^{-1} \text{ mol}^{-1}$

Ideal Gas Law



The other gas laws are obtained from the ideal gas law if two of the variables (V, P, T, n) are kept constant.

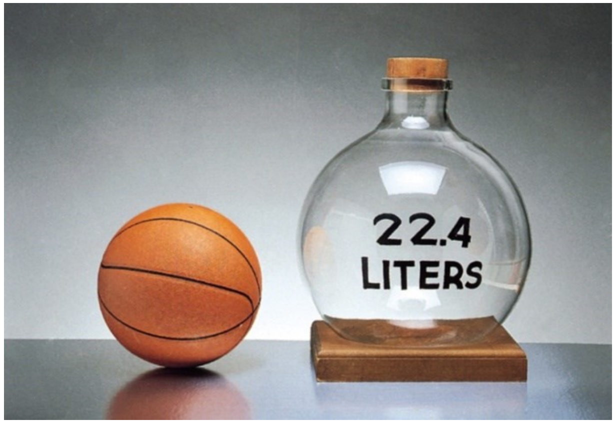
Standard Conditions

Standard Temperature and Pressure (STP) conditions

- Defined to permit inter-comparison of measurements
- Standard pressure = 1 atm
- Standard temperature = 273.15 K = 0 °C
- Standard amount = 1 mol

Standard molar volume

- The volume occupied by one mole of a substance at STP.
- $V = \dfrac{1\ \text{mol} \times 0.082057\ \ \text{atm}\cdot\text{L}\cdot\text{mol}^{-1}\cdot\text{K}^{-1} \times 273.15\ \text{K}}{1.0\ \text{atm}} = 22.41\ L$



Also used: Standard Ambient Temperature and Pressure (SATP) conditions

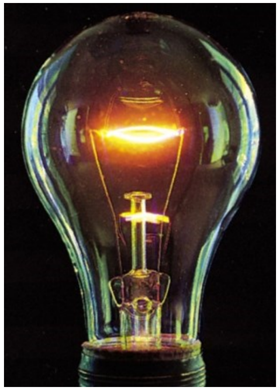
- Standard ambient pressure = 1 atm
- Standard ambient temperature = 298.15 K = 25 °C
- This would give a molar volume of 24.47 L

Sample Problem

What is the volume (in liters) occupied by 49.8 g of HCl at STP?

Sample Problem

Argon is an inert gas used in lightbulbs to retard the vaporization of the filament. A certain lightbulb containing argon at 1.20 atm and 18 ^{0}C is heated to 85 ^{0}C at constant volume. What is the final pressure of argon in the lightbulb (in atm)?



Ideal Gas Law Applications

Density (d) Calculations

$$d = \frac{m}{V} = \frac{n\mathcal{M}}{V} = \frac{P\mathcal{M}}{RT}$$

m is the mass of the gas; *V* is the gas volume

$\mathcal{M}$ is the molar mass of the gas

Molar Mass ($\mathcal{M}$) of a Gaseous Substance

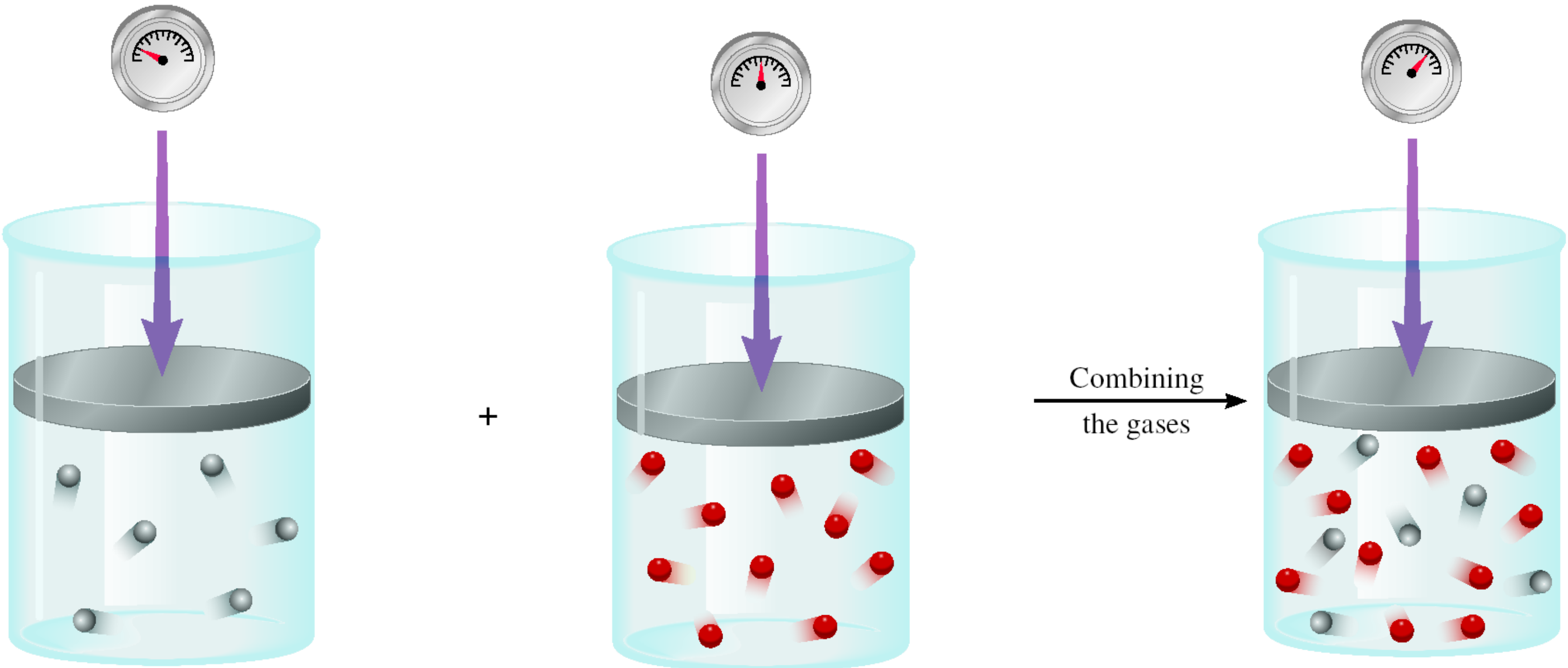
$$\mathcal{M} = \frac{dRT}{P}$$

Sample Problem

A 2.10-L vessel contains 4.65 g of a gas at 1.00 atm and 27.0 ^{0}C. What is the molar mass of the gas?

Dalton's Law of Partial Pressures

V and *T* are **constant**



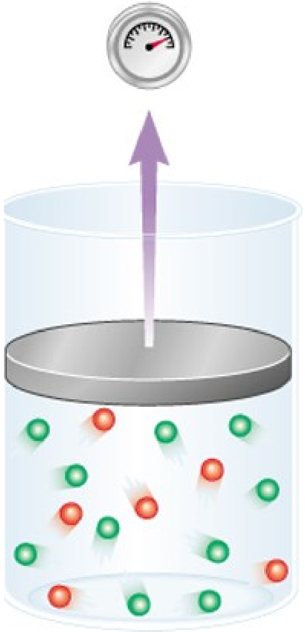
P_1 P_2 $P_T = P_1 + P_2$

P_1 and P_2 are called ***partial pressures***

P_T is the total pressure of the mixture

Partial Pressures and Mole Fractions

Consider a case in which two gases, A and B, are in a container of volume V.



$$P_A = \frac{n_A RT}{V}$$

n_A is the number of moles of A
P_A is the partial pressure of A

$$P_B = \frac{n_B RT}{V}$$

n_B is the number of moles of B
P_B is the partial pressure of B

$$P_T = P_A + P_B \qquad X_A = \frac{n_A}{n_A + n_B} \qquad X_B = \frac{n_B}{n_A + n_B}$$

$$\textbf{\textit{mole fraction}}\ (X_i) = \frac{n_i}{n_T}$$

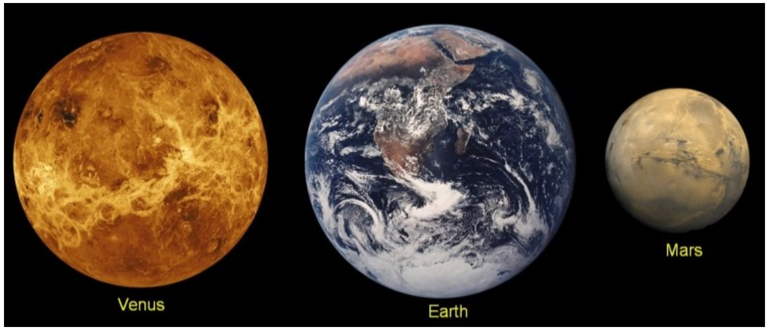
$$P_A = X_A P_T \qquad P_B = X_B P_T \qquad \boxed{P_i = X_i P_T}$$

Sample Problem

A sample of natural gas contains 8.24 moles of CH_4, 0.421 moles of C_2H_6, and 0.116 moles of C_3H_8. If the total pressure of the gases is 1.37 atm, what is the partial pressure of propane (C_3H_8)?

Planetary Atmospheres

Atmospheric Chemistry and Global Change, NCAR, 1999

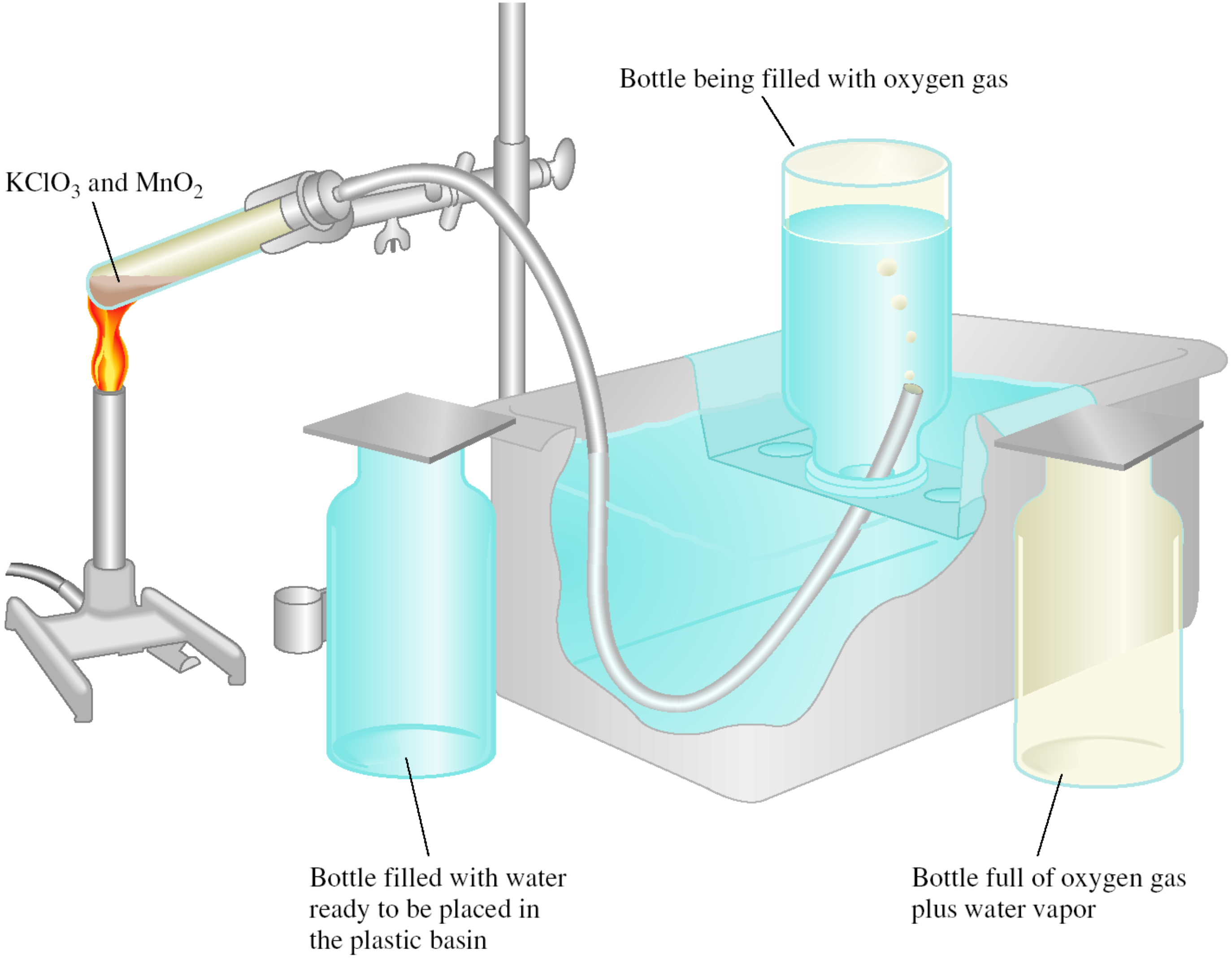


Characteristic	*Venus*	*Earth*	*Mars*
Total mass (10^{27} g)	5	6	0.6
Radius (km)	6049	6371	3390
Atmospheric mass (ratio)	100	1	0.06
Distance from Sun (10^6 km)	108	150	228
Solar constant (W m^{-2})[a]	2613	1367	589
Albedo (%)	75	30	15
Cloud cover (%)	100	50	Variable
Effective radiative (°C) temperature	–39	–18	–56
Surface temperature (°C)	427	15	–53
Greenhouse warming (°C)	466	33	3
N_2 (%)	<2	78	<2.5
O_2 (%)	<1 ppmv	21	<0.25
CO_2 (%)	98	0.035	>96
H_2O (range %)	$1 \times 10^{-4} - 0.3$	$3 \times 10^{-4} - 4$	<0.001
SO_2 (fraction)	150 ppmv	<1 ppbv	Nil
Cloud composition	H_2SO_4	H_2O	Dust, H_2O, CO_2

Fractions (X) are often expressed:

- In percent:
  % = X×100
- In parts per million by volume (ppmv):
  ppmv = X×10⁶

Collecting a Gas over Water

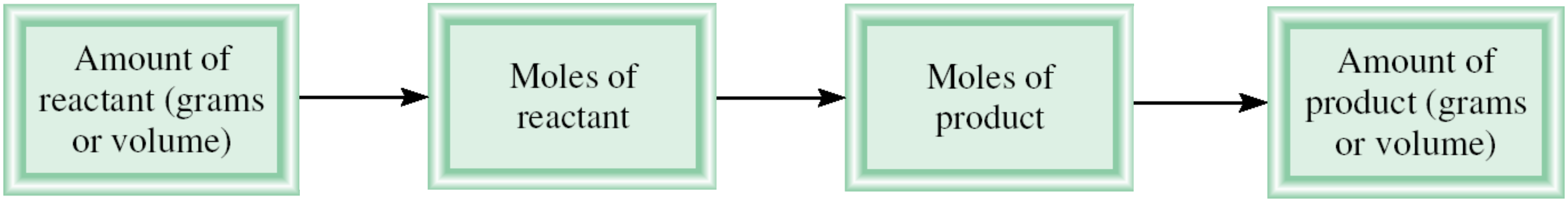


$$2KClO_3\ (s) \longrightarrow 2KCl\ (s) + 3O_2\ (g)$$

$$P_T = P_{O_2} + P_{H_2O}$$

Temperature (°C)	Water Vapor Pressure (mmHg)
0	4.58
5	6.54
10	9.21
15	12.79
20	17.54
25	23.76
30	31.82
35	42.18
40	55.32
45	71.88
50	92.51
55	118.04
60	149.38
65	187.54
70	233.7
75	289.1
80	355.1
85	433.6
90	525.76
95	633.90
100	760.00

Gas Stoichiometry



When gases are at STP, use 1 mol = 22.41 L

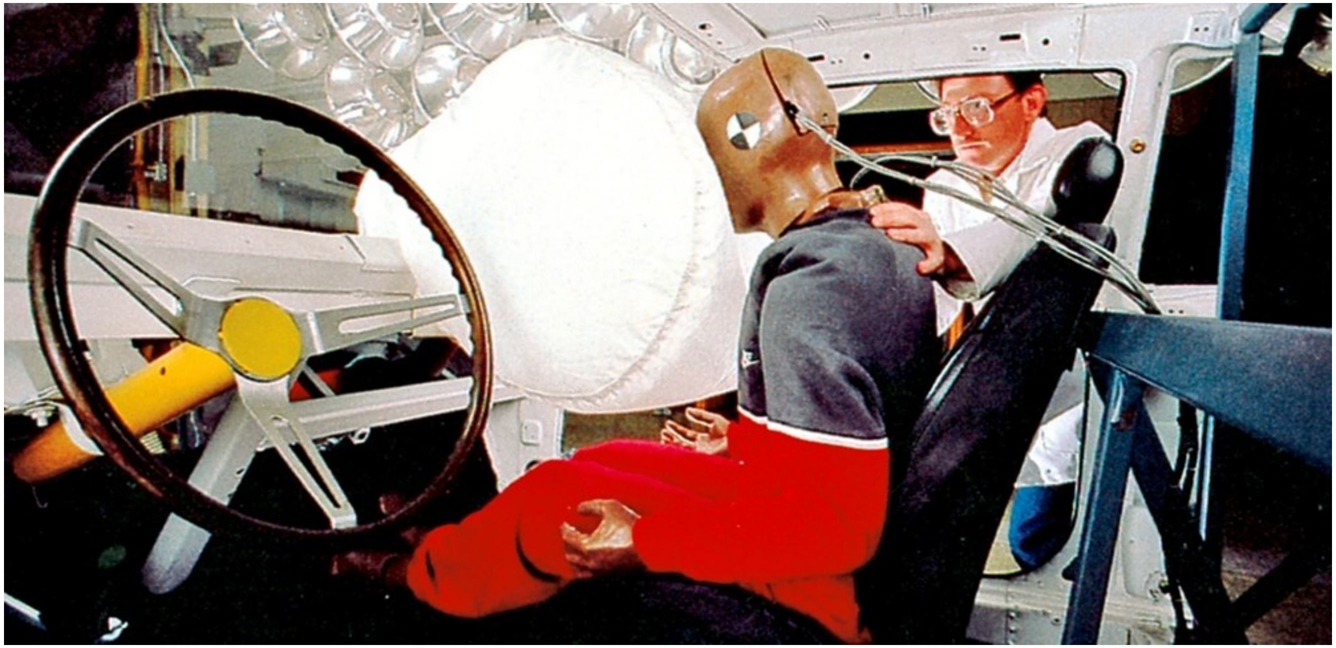
When gases are at SATP, use 1 mol = 24.47 L

At other conditions use the ideal gas law

The pressures here could also be partial pressures

Gas Stoichiometry

Sodium azide, NaN_3, forms a large volume of nitrogen gas in a reaction triggered electrically in air bags.



$$2\ NaN_3\ (s) \rightarrow\ 2\ Na\ (s) + 3\ N_2\ (g)$$

$$NaN_3\ = 65.01\ g{\cdot}mol^{-1}$$

2 moles of NaN_3 (130.02 g) would then give 3 moles of N_2 gas (about 3×24.5 = 73 L at SATP)

Gas Stoichiometry Calculations

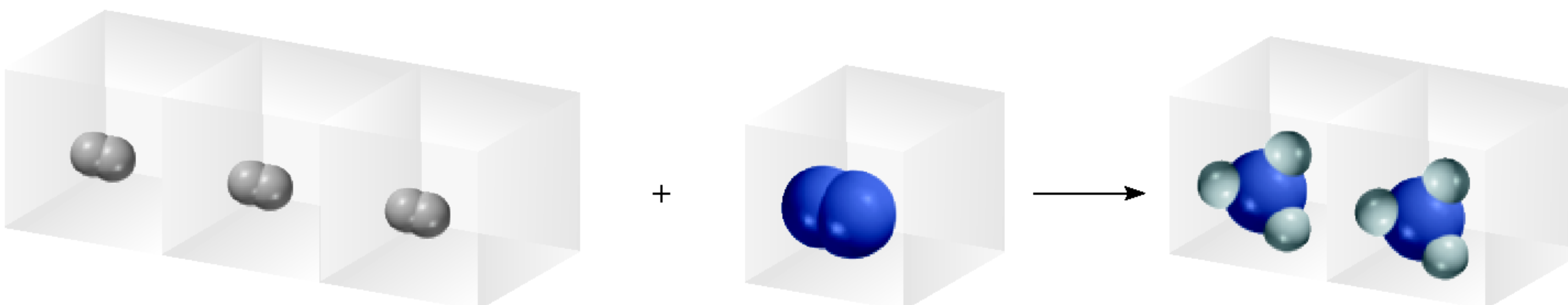
How many liters of gaseous NH_3 can be obtained from 1 liter of gaseous hydrogen reacting with excess nitrogen at constant temperature and pressure?

$$3H_2 + N_2 \rightarrow 2NH_3$$

$$3 \text{ mole } H_2 \rightarrow 2 \text{ mole } NH_3$$

At constant *T* and *P*

$$1 \text{ liter } H_2 \rightarrow 2/3 \text{ liters } NH_3$$



$3H_2(g)$	+	$N_2(g)$	$\longrightarrow$	$2NH_3(g)$
3 molecules	+	1 molecule	$\longrightarrow$	2 molecules
3 moles	+	1 mole	$\longrightarrow$	2 moles
3 volumes	+	1 volume	$\longrightarrow$	2 volumes

Sample Problem

What is the volume of CO_2 produced at 37 °C and 1.00 atm when 5.60 g of glucose are used up in the reaction:

$$C_6H_{12}O_6\ (s) + 6O_2\ (g) \rightarrow 6CO_2\ (g) + 6H_2O\ (l)$$

Water-Candle Experiment

Paraffin is composed of a mix of hydrocarbons with the general formula: C_nH_{2n+2}

This is liquid because the water condenses on the glass jar.

One such compound is hentriacontane:

$$\mathbf{C_{31}H_{64\,(s)} + 47\ O_{2\,(g)} = 31\ CO_{2\,(g)} + 32\ H_2O_{\,(l)}}$$

This is solid because it’s in the candle

$$47 \text{ mole } O_{2\,(g)} \quad \rightarrow \quad 31 \text{ mole } CO_{2\,(g)}$$

At constant T and P (Avogadro’s Law):

$$1 \text{ liter } O_2 \quad \rightarrow \quad 31/47 \text{ or } \sim 2/3 \text{ liter } CO_{2\,(g)}$$

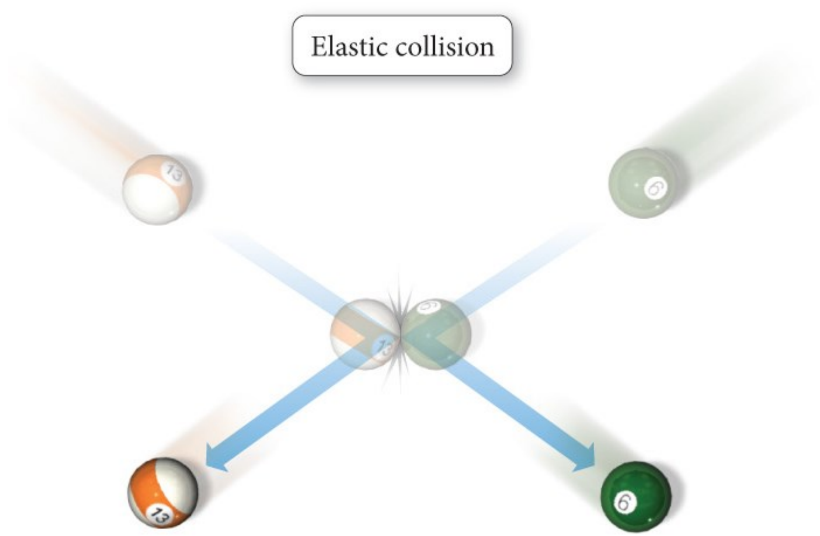
If this were done with pure $O_{2\,(g)}$, volume would decrease by 1/3

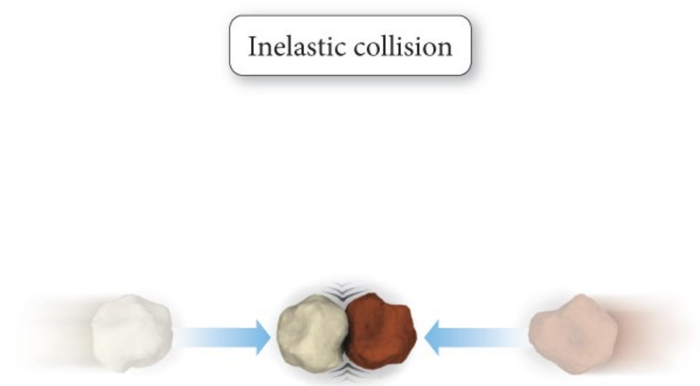
Since air is made of only 21% $O_{2\,(g)}$ volume decreases by ~7%

$$1/3 \times 0.21 = 0.07$$

Molecular Theory of Gases

- Gas is modeled as a collection of small particles in a large volume with lots of empty space between the particles
- Particles are constantly moving in random directions and with different speeds
- Particles undergo elastic collisions (like billiard balls) – they exchange energy but there is no overall loss of kinetic energy





Temperature & Molecular Speeds

Temperature is ***defined*** in molecular theory of gases as a measure of the average kinetic energy of particles

$$\langle E_{particle} \rangle = \frac{m\langle \upsilon^2 \rangle}{2} = \frac{3}{2}kT$$

Average kinetic energy of one particle

k = 1.3806503 × 10^{-23} J K^{-1} Boltzmann constant

$$\langle E_{molar} \rangle = N_A \langle E_{particle} \rangle = \frac{3}{2}RT$$

Average kinetic energy of one mole of particles

$R = k \times N_A$ Gas constant (8.314 J·mol^{-1}·K^{-1})

Root-mean-square speed is defined as (where $\mathcal{M}$ = molecular weight):

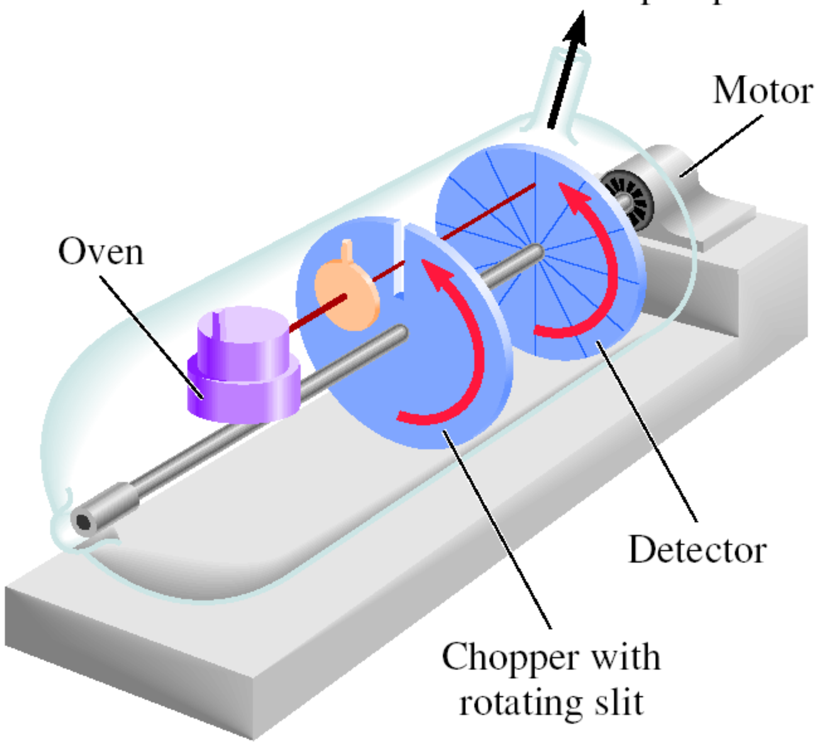
$$\upsilon_{rms} = \sqrt{\langle \upsilon^2 \rangle} = \sqrt{\frac{3kT}{m}} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

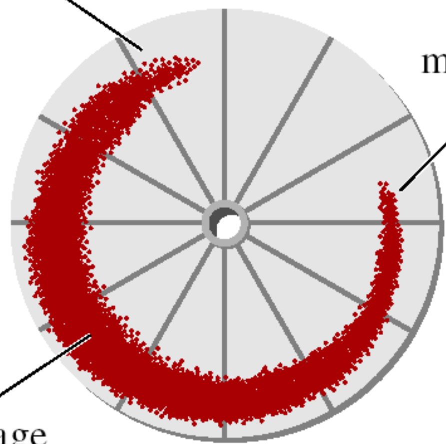
Sample Problem

Calculate the root mean square velocity of oxygen molecules at 25 °C.

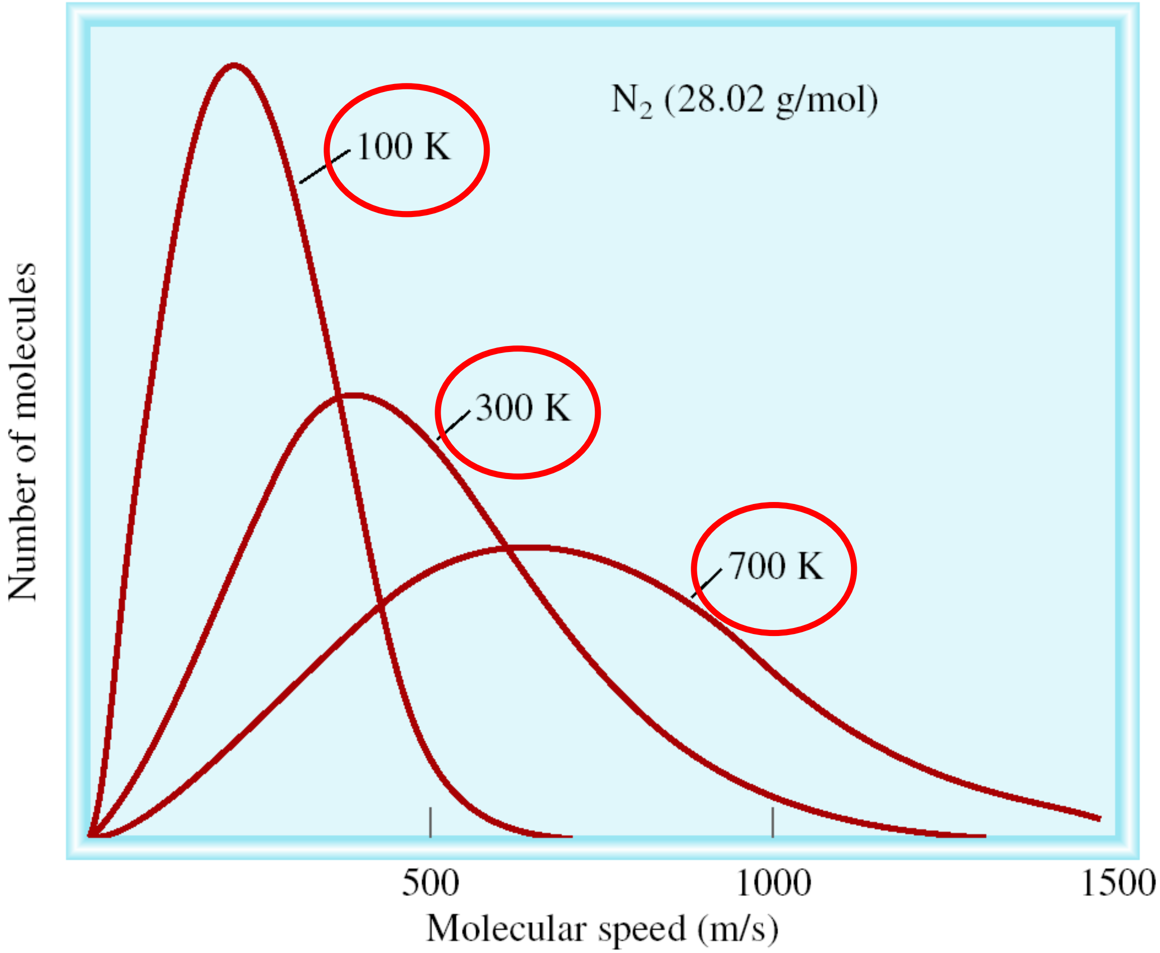
Distribution of Speeds in a Gas

Molecules do not have the same speed! The molecular speed distribution can be measured experimentally:





Distribution of Speeds in a Gas

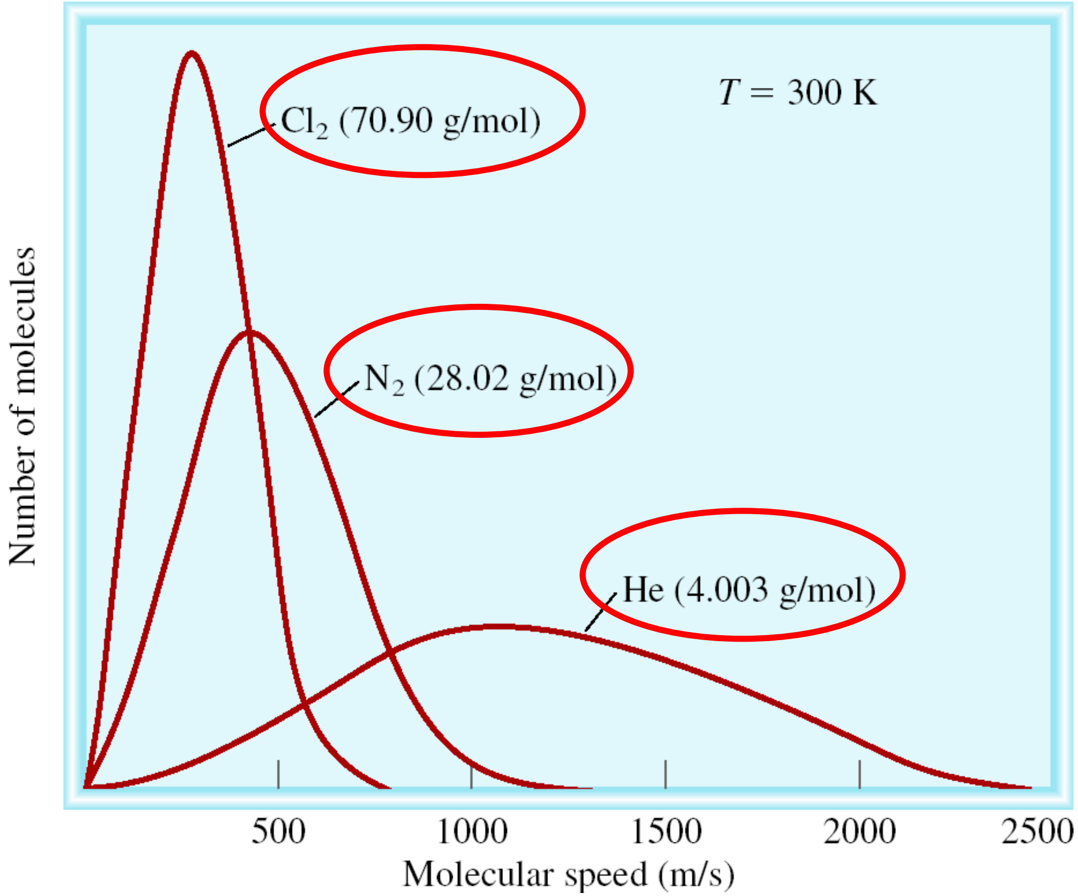


- The speed distribution "spreads out" at higher temperatures
- Average speed increases with temperature

$$\upsilon_{rms} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

The distribution of speeds for nitrogen gas molecules at three different temperatures

Distribution of Speeds in a Gas



The distribution of speeds of three different gases at the **same temperature**

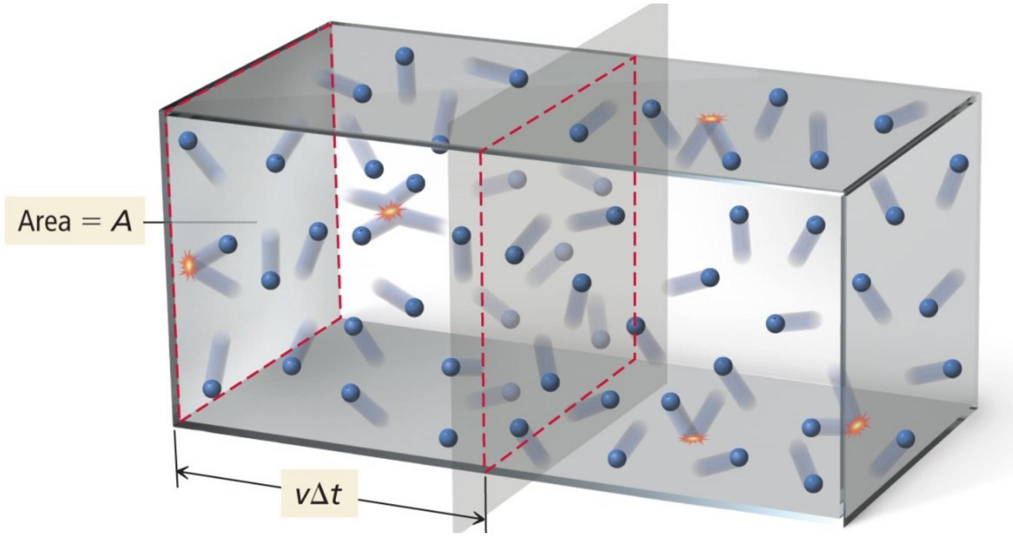
- The average energy is ***the same*** because it only depends on temperature

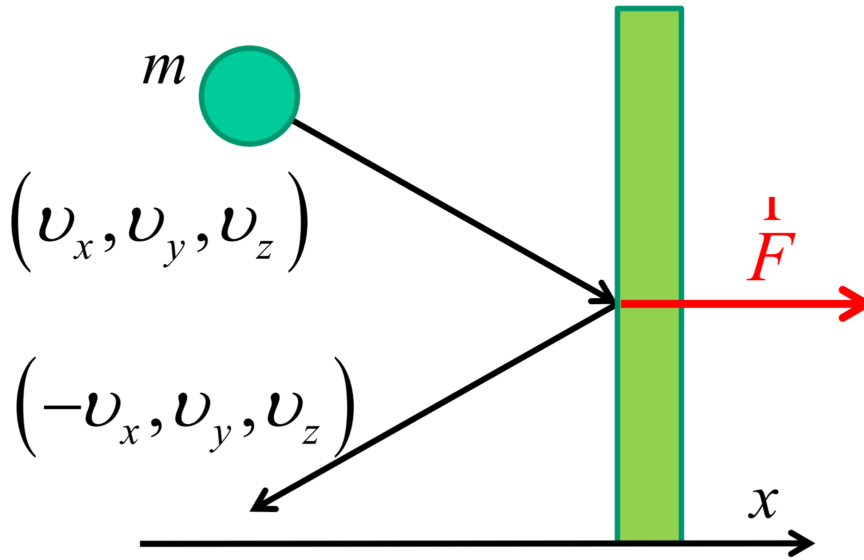
$$\langle E \rangle = \frac{3}{2} RT$$

- Lighter molecules have to move faster to have the same kinetic energy as the heavier ones

$$\upsilon_{rms} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

Molecular View





N particles in volume ***V***

$$\vec{F} = m\frac{\Delta \vec{\upsilon}}{\Delta t}$$

Force exerted by a ***single collision***

$$\Delta \upsilon_x = \upsilon_x - (-\upsilon_x) = 2\upsilon_x$$

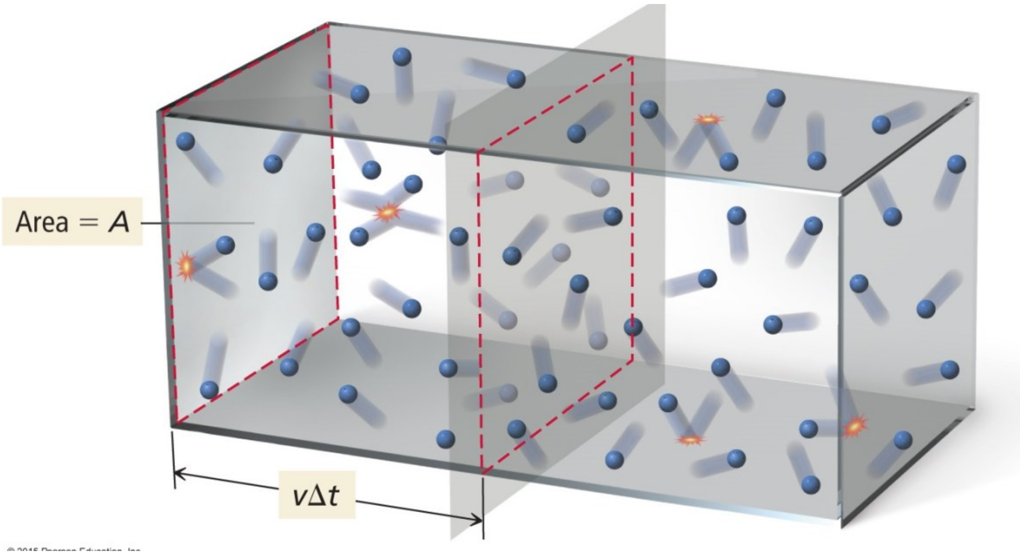
$$\Delta \upsilon_y = \Delta \upsilon_z = 0$$

Collision with the wall is elastic – the molecule bounces off with the same speed like a billiard ball

$$|F_x| = m\frac{2|\upsilon_x|}{\Delta t}$$

Force exerted by a ***single collision***

Molecular View



Number of collisions during time Δt = number of particles within distance $\upsilon_x \Delta t$ from the wall AND moving towards the wall

$$N_{collisions} = \frac{1}{2} \times \underbrace{|\upsilon_x| \Delta t \times Area}_{} \times \frac{N}{V}$$

Half of the particles move towards the wall

Effective volume occupied by these particles

Particle number density (N particles in volume V)

$$F_{total} = |F_x| \times N_{collisions}$$

Total force on the wall

Molecular View

$$F_{total} = |F_x| \times N_{collisions} = m\frac{2|\upsilon_x|}{\Delta t} \times \frac{1}{2} \times |\upsilon_x| \Delta t \times Area \times \frac{N}{V}$$

$$F_{total} = m(\upsilon_x)^2 \times Area \times \frac{N}{V}$$

Total force on the wall

$$P = \frac{F_{total}}{Area} = m(\upsilon_x)^2 \times \frac{N}{V}$$

Pressure is the force normalized by the area

$$E_x = \frac{m(\upsilon_x)^2}{2} = \frac{kT}{2}$$

Recall that the average kinetic energy is related to temperature

$$P = kT\frac{N}{V}$$

This is the ideal gas law!

$$N = n \times N_A \qquad R = k \times N_A$$

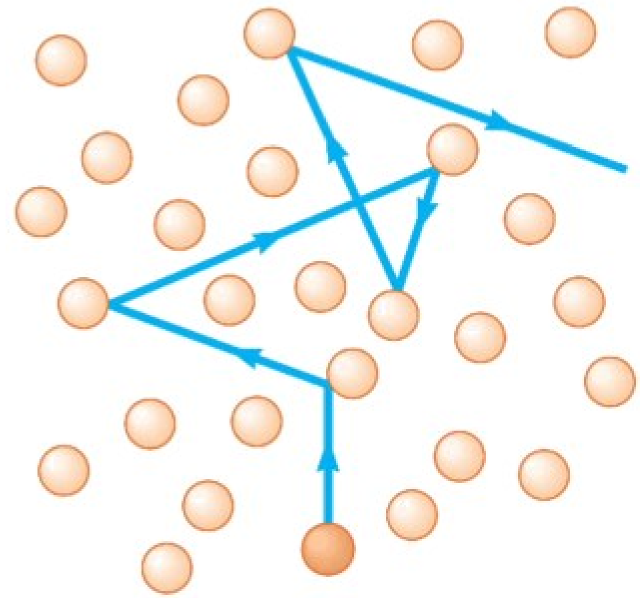
Converting to molar quantities

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

The ideal gas law!

Mean Free Path

Molecules in a gas travel in straight lines until they collide with another molecule or the container wall.



The average distance a molecule travels between collisions is called the ***mean free path***.

Mean free path is inversely proportional to the pressure, $\lambda \propto P^{-1}$

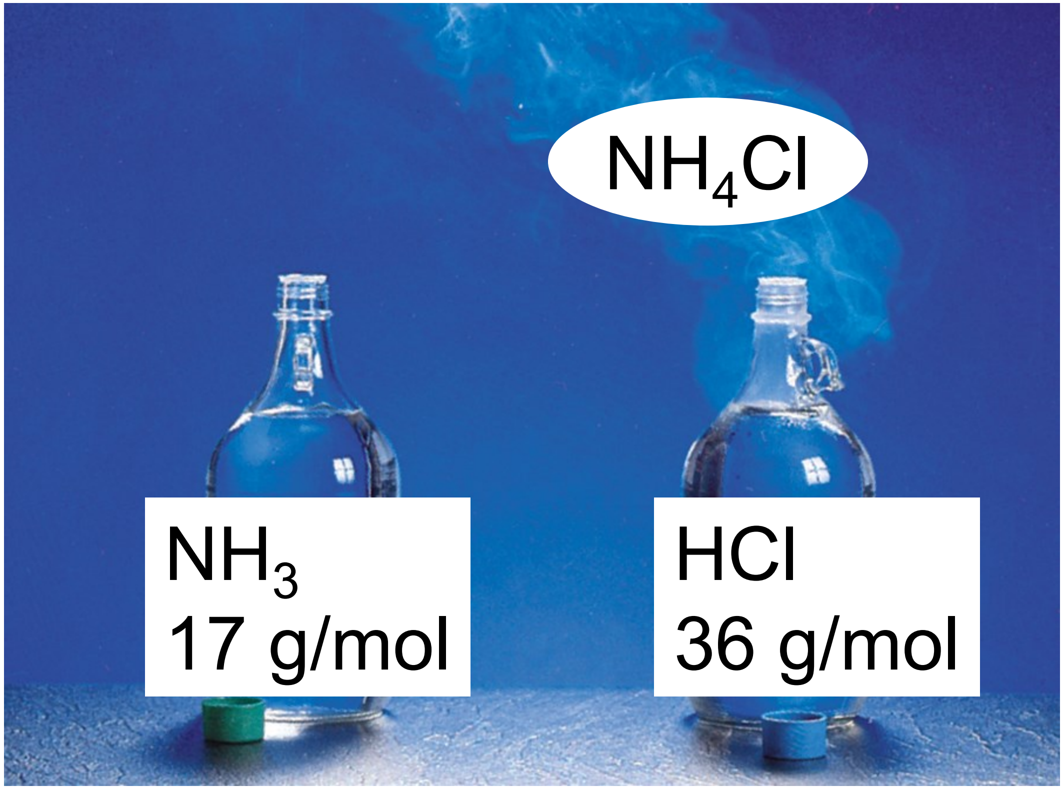
Environment	Pressure	Mean free path of N_2	For comparison:
Sea-level	1 atm	60 nm	Van der Waals size of N_2 molecule = 0.16 nm
Rough vacuum	10^{-3} atm	60 μm	Diameter of a human hair = 100 μm
High vacuum	10^{-6} atm	60 mm	One inch = 25.4 mm
Ultrahigh vacuum	10^{-10} atm	600 m	The tallest building (Burj Khalifa) = 828 m
Interstellar space	10^{-16} atm	6×10^8 m	Average Earth-moon distance = 3.8×10^8 m

Diffusion

Gas diffusion is the gradual mixing of molecules of one gas with molecules of another by virtue of their kinetic properties.

Molecular path

$$\frac{\text{rate}_1}{\text{rate}_2} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$

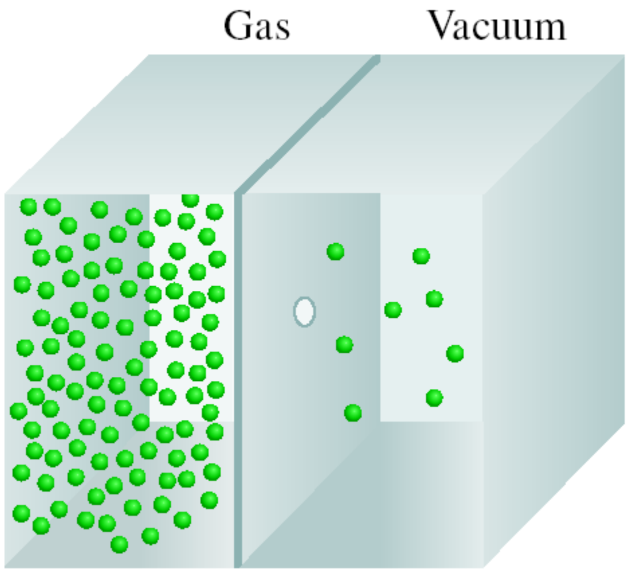


Average rates (and distances travelled) by two molecules

Diffusion Video

Effusion

Gas effusion is the is the process by which gas under pressure escapes from one compartment of a container to another by passing through a small opening.



$$\frac{\text{rate}_1}{\text{rate}_2} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$

The ratio of effusion rates can be used to estimate molecular weights of unknown gases

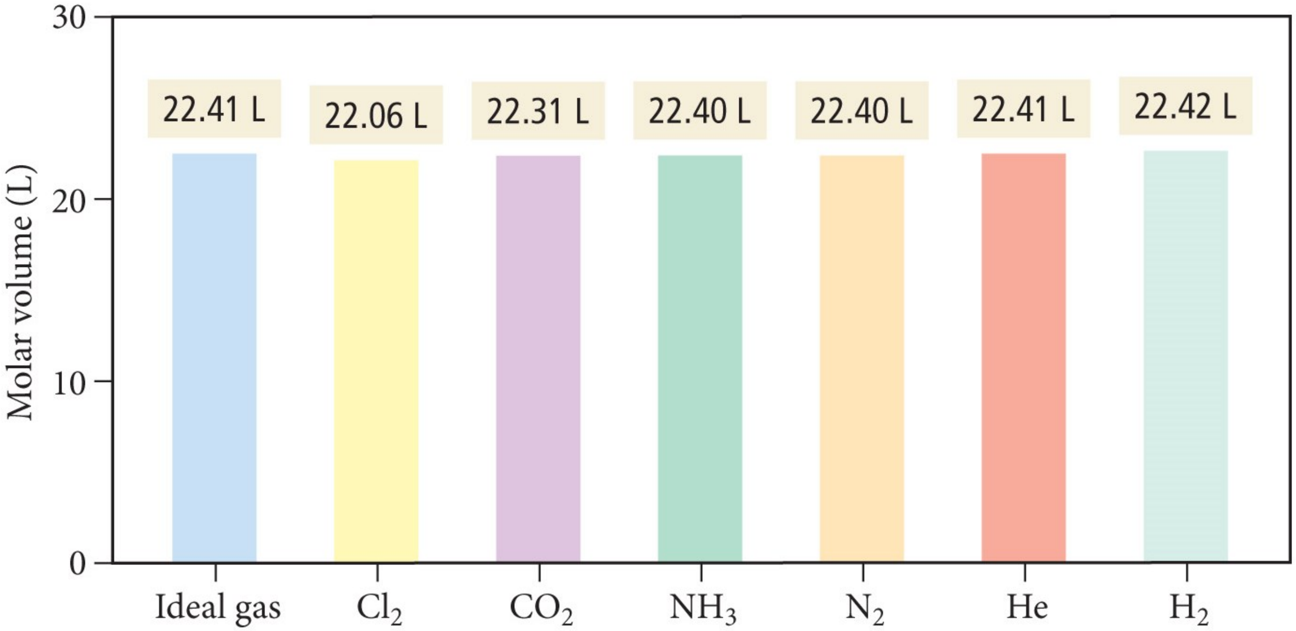
Sample Problem

Nickel forms a gaseous compound of the formula $Ni(CO)_x$ What is the value of *x* given that under the same conditions methane (CH_4) effuses 3.3 times faster than the compound?

Deviations from Ideal Behavior

At STP, gases are well described by the ideal gas approximation:

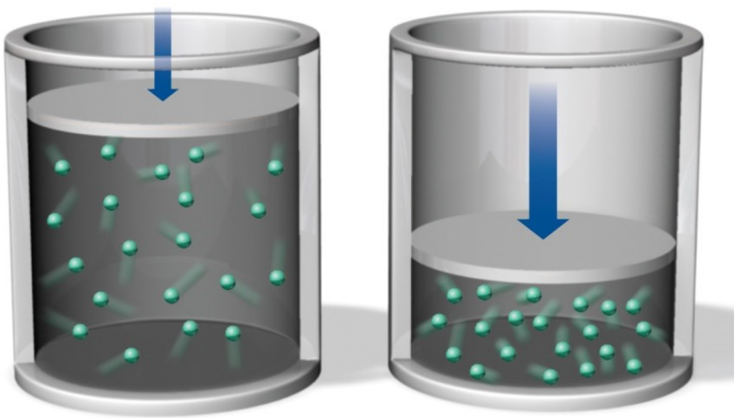
1. No attractions between gas molecules; and
2. Gas molecules do not take up space.



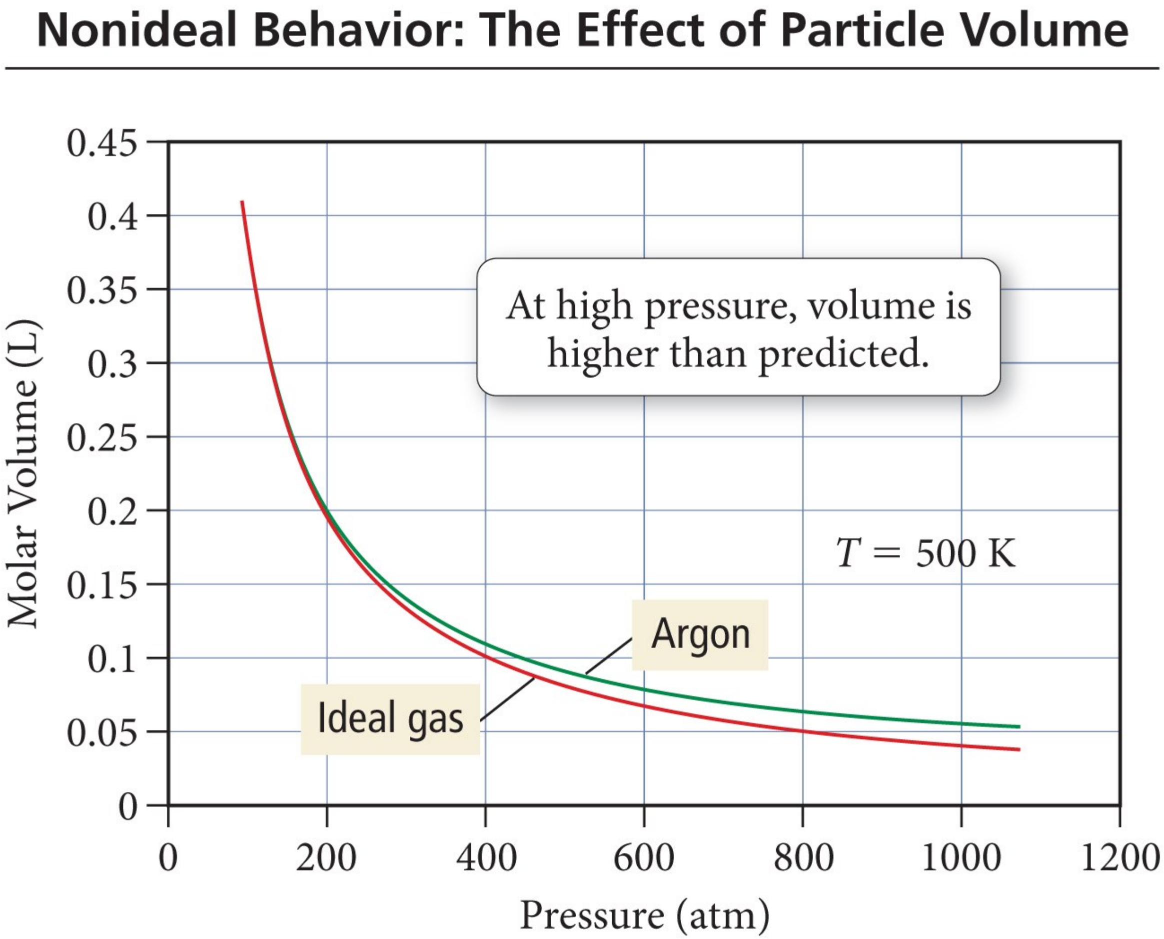
Molar volume at STP is very close to the ideal gas limit of 22.4 L

However, at ***low temperatures*** and ***high pressures***, these assumptions break down. Indeed, gases can be condensed into liquids if they are cooled or compressed (Chapter 13).

Deviations from Ideal Behavior

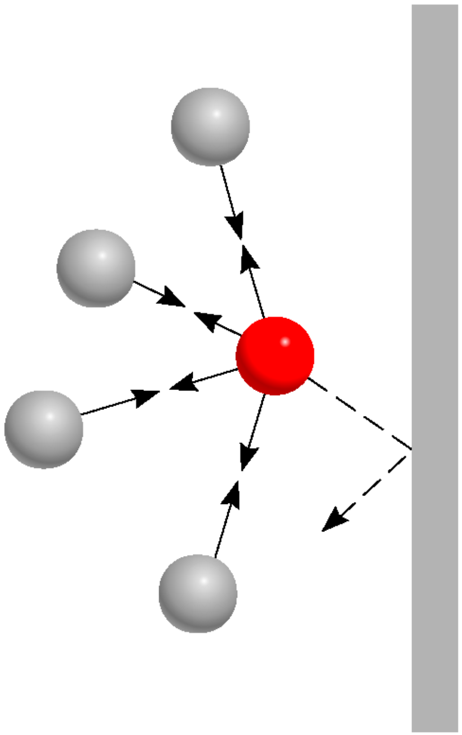


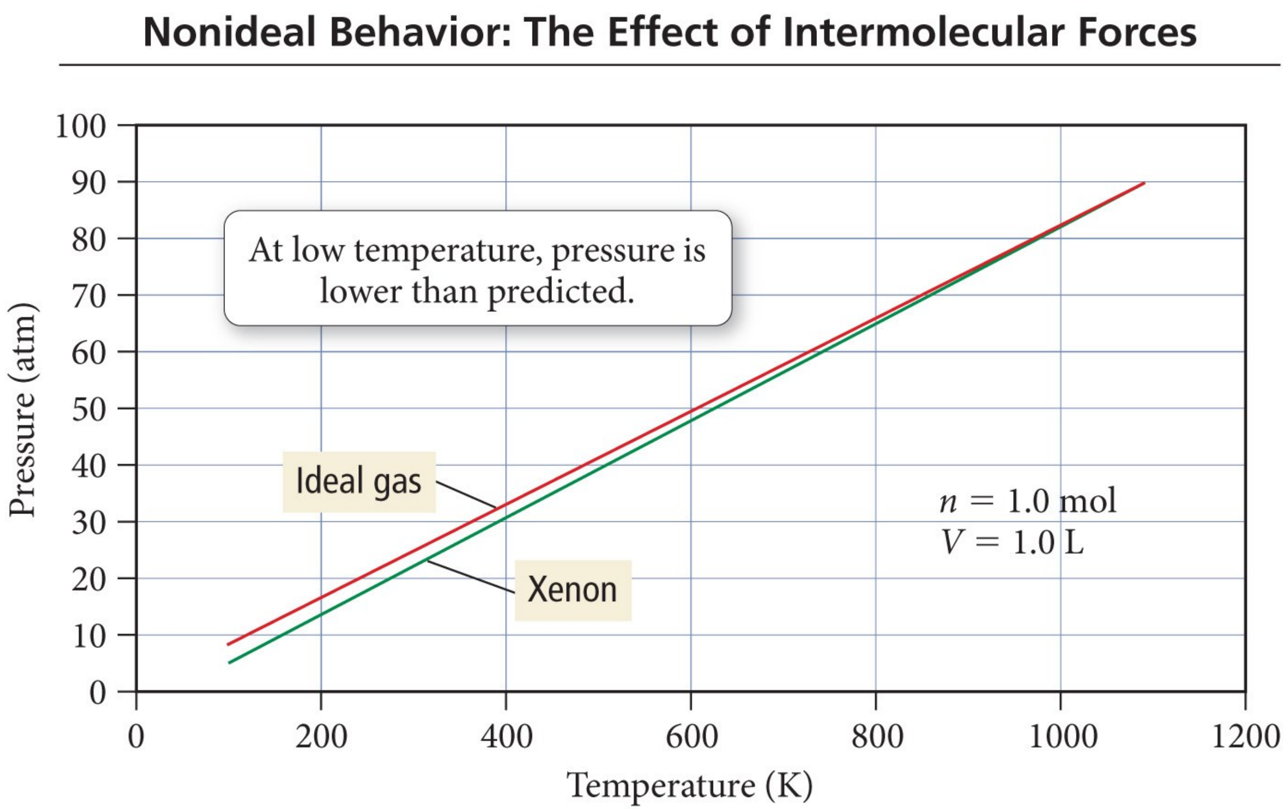
Because real molecules take up space, the molar volume of a real gas is larger than predicted by the ideal gas law at high pressures.



Deviations from Ideal Behavior

Because real molecules attract each other when they are close by, it reduces the pressure exerted on the walls of the container, especially at lower temperatures when molecules move slower.





Van der Waals Equation

To account for these deviations from ideality, van der Waals proposed two corrections to the ideal gas equation

$$PV = nRT$$ The ideal gas law

$$\left[P + a\left(\frac{n}{V}\right)^2\right] \times [V - nb] = nRT$$ Van der Waals equation

Correction for the reduction of the pressure due to attractive intermolecular forces. Larger parameter ***a*** corresponds to stronger intermolecular forces

Correction for the finite volume of the particles. Larger parameter ***b*** corresponds to bulkier molecules.

Deviations from Ideal Behavior

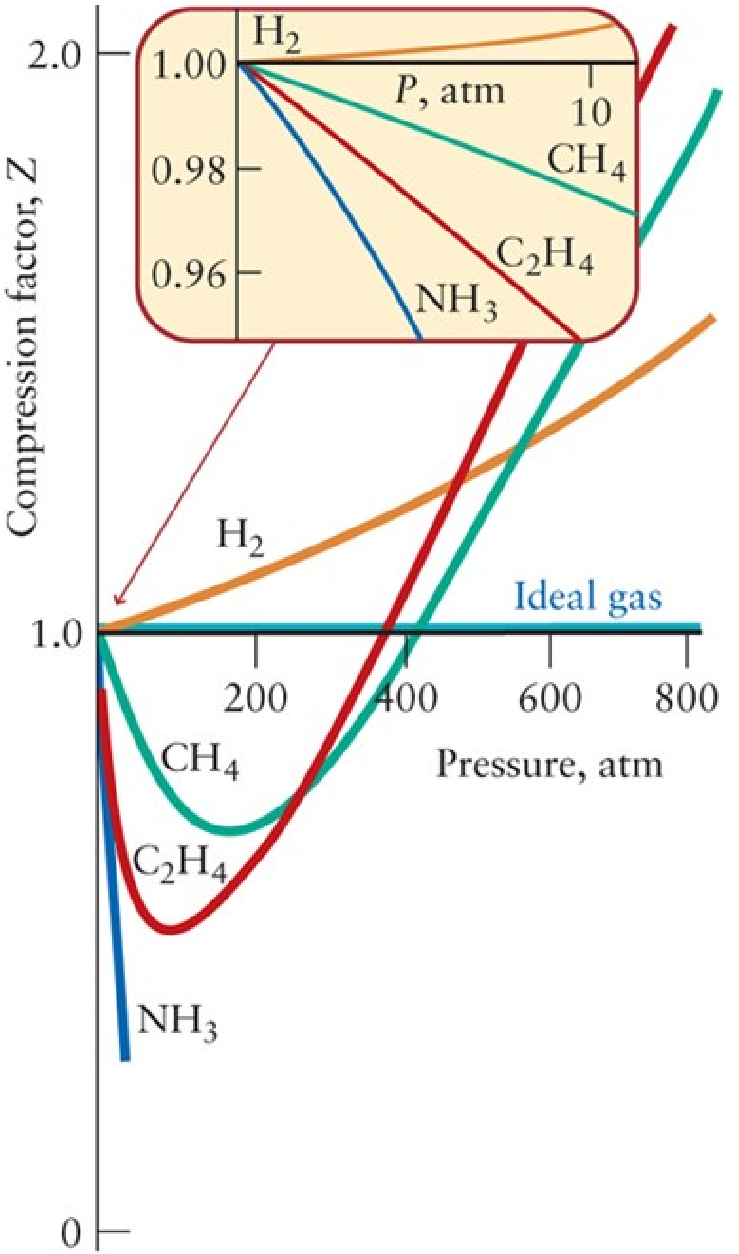
One measure of non-ideality is the compression factor, *Z*

$$Z = \frac{V_{molar}}{V_{molar}^{ideal}}$$

Z = 1 under ideal conditions.

For most gases, at low pressures the **attractive forces** (the effect of the ***a***-parameter) are dominant and Z < 1.

At high pressures, **repulsive forces** (effect of the ***b***-parameter) become dominant and *Z* > 1 for all gases.



Deviations from Ideal Behavior

Based on this table, molecule of which gases has the largest deviations from the ideal gas law?

Van der Waals equation

$$\underbrace{\left(P + \frac{an^2}{V^2}\right)}_{\text{corrected pressure}}\underbrace{(V - nb)}_{\text{corrected volume}} = nRT$$

Carbon tetrachloride (CCl_4) has the largest values of the a-parameter and b-parameter, so it is likely to deviate from the ideal gas equation the most.

van der Waals Constants of Some Common Gases

Gas	a $\left(\frac{atm \cdot L^2}{mol^2}\right)$	b $\left(\frac{L}{mol}\right)$
He	0.034	0.0237
Ne	0.211	0.0171
Ar	1.34	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0266
H_2	0.244	0.0266
N_2	1.39	0.0391
O_2	1.36	0.0318
Cl_2	6.49	0.0562
CO_2	3.59	0.0427
CH_4	2.25	0.0428
CCl_4	20.4	0.138
NH_3	4.17	0.0371
H_2O	5.46	0.0305

Sample Problem

10.11

The properties of carbon dioxide gas are well known in the bottled beverage industry. In an industrial process, a tank of volume 100. L at 20.0°C contains 50.0 mol CO_2. Estimate the pressure in the tank.

$a = 3.640\ L^2{\cdot}atm{\cdot}mol^{-2}$ and $b = 4.267 \times 10^{-2}\ L{\cdot}mol^{-1}$