Gases and Their Properties
Chapter 9

Chemistry 222
Professor Michael Russell

MAR *Last update: 4/29/24*

Importance of Gases

Airbags fill with N_2 gas in an accident.

Gas is generated by the decomposition of sodium azide, NaN_3.

$2\ NaN_{3(s)}$ ---> $2\ Na_{(s)}$ + $3\ N_{2(g)}$

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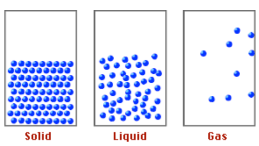
THREE STATES OF MATTER



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General Properties of Gases

There is a lot of "free" space in a gas.

Gases can be expanded infinitely.

Gases occupy containers uniformly and completely.

Gases diffuse and mix rapidly.

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Properties of Gases

Gas properties can be modeled using math. Model depends on:

- V = volume of the gas (L)
- T = temperature (K)
- n = amount (moles)
- P = pressure (atm)

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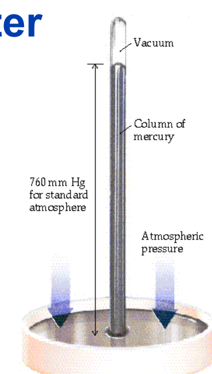
The Barometer

Pressure of air is measured with a BAROMETER (developed by Torricelli in 1643)

Hg rises in tube via atmosphere (pushing up), opposed by gravity (pulling down)

Barometer calibrated for column width, pool width, depth, Hg density, etc.



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Pressure Units

Column height measures P of atmosphere *(atm)*

1 standard atm

= 760 mm Hg

= 76 cm Hg

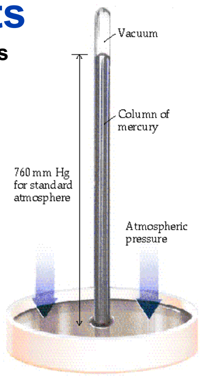
= 760 torr *(torr = mm Hg)*

= 1.013 bar = 1013 mbar

= 29.9 inches Hg

= about *34 feet* of water

SI unit is PASCAL, Pa, where 1 atm = 101.325 kPa *(1 mbar = 1hPa)*



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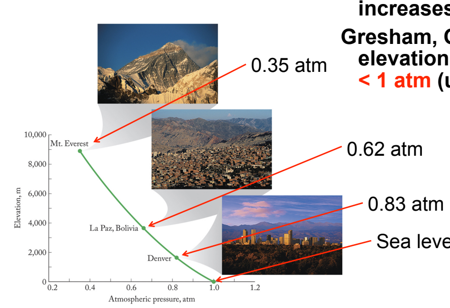
Pressure

Pressure about 1.0 atm at sea level

Pressure decreases as elevation increases

Gresham, OR elevation: 301 feet, < 1 atm (usually)



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Boyle's Law

If n and T are constant, then

PV = (nRT) = k

This means, for example, that P goes up as V goes down, *or:*

$P_1V_1 = P_2V_2$

Robert Boyle (1627-1691). Son of Early of Cork, Ireland.

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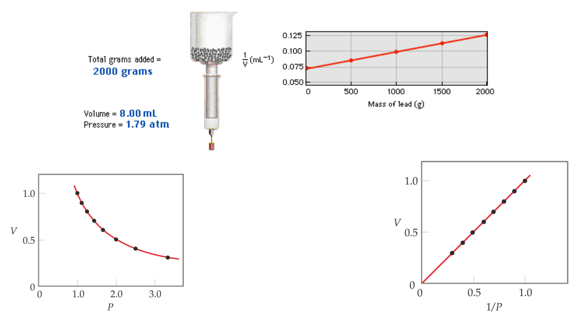
Boyle's Law

Boyle's law states that the pressure and volume of a gas are inversely related



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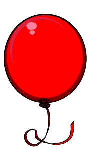


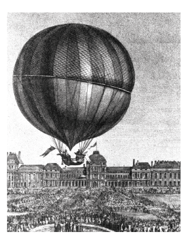
Charles's Law

If n and P are constant, then

V = (nR/P)T = kT

V and T are directly related, *or:*

$V_1 / T_1 = V_2 / T_2$



Jacques Charles (1746-1823). Isolated boron and studied gases. Balloonist.

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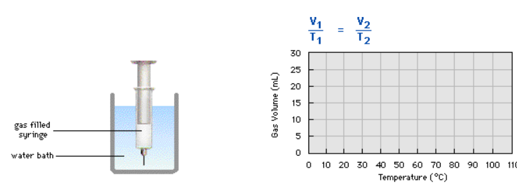
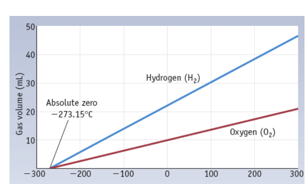


Charles's Law



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Charles's Law

Balloons immersed in liquid N_2 (at -196 ˚C) will shrink as the air cools (and is liquefied).

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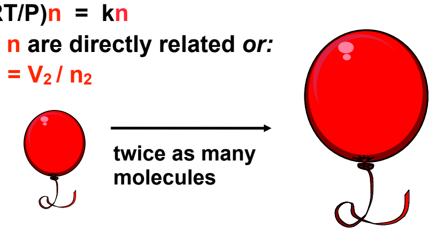
Avogadro's Hypothesis

Equal volumes of gases at the same T and P have the same number of molecules.

$V = (RT/P)n = kn$

V and n are directly related *or:*

$V_1 / n_1 = V_2 / n_2$



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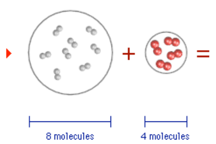
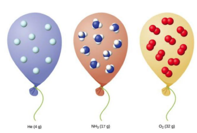
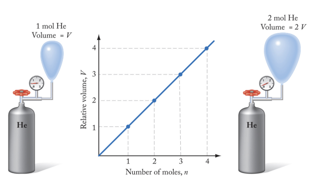
Avogadro's Hypothesis





$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$

The gases in this experiment are all measured at the same T and P.



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IDEAL GAS LAW

P V = n R T

Brings together gas properties.

Can be derived from experiment and theory.

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IDEAL GAS LAW

P V = n R T

The constant of proportionality is known as *R*, the gas constant.

***Memorize* R! *Always* use 0.082057!**

We will also use 8.3145 later...

Units	Numerical Value
L-atm/mol-K	0.082057
J/mol-K*	8.3145
cal/mol-K	1.987
m^3-Pa/mol-K*	8.3145
L-torr/mol-K	62.36

*SI unit

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Using PV = nRT

How much N_2 is req'd to fill a small room with a volume of 960. cubic feet ($2.70 * 10^4$ L) to P = 745 mm Hg at 25 °C?

R = 0.082057 L•atm/K•mol

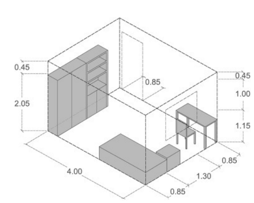
Solution

1. Get all data into proper units

$V = 2.70 * 10^4$ L

T = 25 °C + 273 = 298 K

P = 745 mm Hg (1 atm/760 mm Hg) = 0.980 atm



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Using PV = nRT

How much N_2 is req'd to fill a small room with a volume of 960. cubic feet ($2.70 * 10^4$ L) to P = 745 mm Hg at 25 ^oC?

R = 0.082057 L•atm/K•mol

Solution

2. Now calc. n = PV / RT

$$n = \frac{(0.980 \text{ atm})(2.70 \times 10^4 \text{ L})}{(0.082057 \text{ L} \bullet \text{atm/K} \bullet \text{mol})(298 \text{ K})}$$

n = 1.08×10^3 mol (30.3 kg of N_2)

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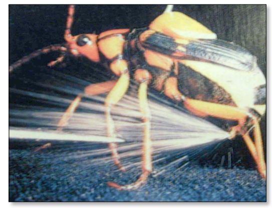
Gases and Stoichiometry

2 H_2O_2(liq) ---> 2 H_2O(g) + O_2(g)

Decompose 1.1 g of H_2O_2 in a flask with a volume of 2.50 L. What is the pressure of O_2 at 25 ^oC? Of H_2O?



Bombardier beetle uses decomposition of hydrogen peroxide to defend itself.

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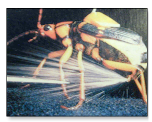
Gases and Stoichiometry



2 H_2O_2(liq) ---> 2 H_2O(g) + O_2(g)

Decompose 1.1 g of H_2O_2 in a flask with a volume of 2.50 L. What is the pressure of O_2 at 25 ^oC? Of H_2O?

Solution

Strategy:

- Calculate moles of H_2O_2 and then moles of O_2 and H_2O.
- Finally, calc. P from n, R, T, and V.

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Gases and Stoichiometry

2 H_2O_2(liq) ---> 2 H_2O(g) + O_2(g)

Decompose 1.1 g of H_2O_2 in a flask with a volume of 2.50 L. What is the pressure of O_2 at 25 ^oC? Of H_2O?

Solution

$$1.1 \text{ g } H_2O_2 \bullet \frac{1 \text{ mol}}{34.0 \text{ g}} = 0.032 \text{ mol}$$

$$0.032 \text{ mol } H_2O_2 \bullet \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O_2} = 0.016 \text{ mol } O_2$$

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Gases and Stoichiometry

2 H_2O_2(liq) ---> 2 H_2O(g) + O_2(g)

Decompose 1.1 g of H_2O_2 in a flask with a volume of 2.50 L. What is the pressure of O_2 at 25 ^oC? Of H_2O?

Solution

P of O_2 = nRT/V

$$= \frac{(0.016 \text{ mol})(0.082057 \text{ L} \bullet \text{atm/K} \bullet \text{mol})(298 \text{ K})}{2.50 \text{ L}}$$

P of O_2 = 0.16 atm

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Gases and Stoichiometry

2 H_2O_2(liq) ---> 2 H_2O(g) + O_2(g)

Solution

What is P of H_2O? Could calculate as above. But *recall* Avogadro's hypothesis.

V α n at same T and P, *and*

P α n at same T and V

There are 2 times as many moles of H_2O as moles of O_2. P is proportional to n. Therefore, P of H_2O is twice that of O_2.

P of H_2O = 0.32 atm

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Dalton's Law of Partial Pressures

$2\ H_2O_2(liq) \longrightarrow 2\ H_2O(g) + O_2(g)$

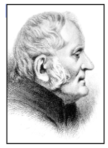
0.32 atm   0.16 atm

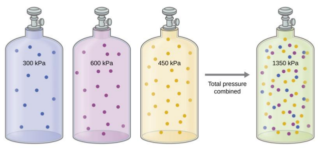
P_{total} in gas mixture = $P_A + P_B + \ldots$

So:

$P_{total} = P(H_2O) + P(O_2) = 0.48$ atm

Dalton's Law: total P equals sum of PARTIAL pressures.





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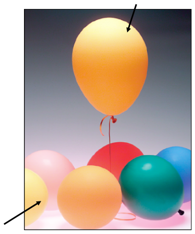
GAS DENSITY

Density of a gas is proportional to the molar mass

$SF_6(g)$, 146.1 g mol^{-1}

$Br_2(g)$, 159.8 g mol^{-1}



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GAS DENSITY

PV = nRT

$$\frac{n}{V} = \frac{P}{RT}$$

$$\frac{m}{M \bullet V} = \frac{P}{RT}$$

where M = molar mass

$$d = \frac{m}{V} = \frac{PM}{RT}$$

or **PM = dRT** *(evening dirt equation)*

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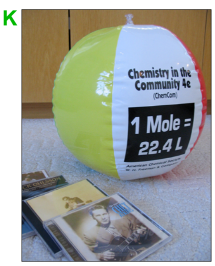
Standard Temperature and Pressure (STP)

A common reference point used in applications using gases

- **Standard Temperature = 273.15 K**
- **Standard Pressure = 1.000 atm**

and if 1.00 mol of gas used,

- **Standard Volume = 22.4 L**



1.00 mol of an ideal gas occupies 22.4 L at 273 K and 1.00 atm of pressure!

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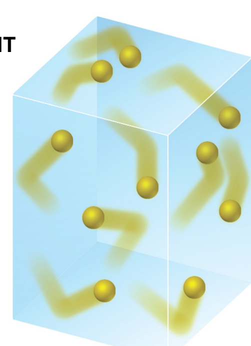
KINETIC MOLECULAR THEORY (KMT)

Theory used to explain gas laws. KMT assumptions are

- **Gases consist of molecules in constant, random motion.**
- **P arises from collisions with container walls.**
- **No attractive or repulsive forces between molecules. Collisions elastic.**
- **Volume of molecules is negligible.**

see Principal Assumptions of KMT Handout



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Kinetic Molecular Theory

We assume molecules of **mass** (**m, kg/mol**) are in **motion** (**velocity, v, m/s**), so they have **kinetic energy** (**KE, J**).

Molecules at the same **temperature** (**T, K**) also have the same kinetic energy, so:

$$KE = \tfrac{1}{2}mv^2 = \tfrac{3}{2}RT$$

*Note: this R = 8.3145 J/mol*K ("energy R")*

At the same T, all gases have the same average KE.
As T goes up, KE also increases - and so does speed.

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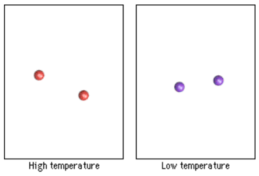
Kinetic Molecular Theory

At the same T, all gases have the same average KE.

As T goes up, KE also increases - and so does speed.



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Kinetic Molecular Theory

Expressed by Maxwell's equation

$$\underbrace{\sqrt{\overline{u^2}}}_{\text{root mean square speed}} = \sqrt{\frac{3RT}{M}}$$

where u is the speed and M is the molar mass.

- speed INCREASES with increasing T
- speed DECREASES with increasing M

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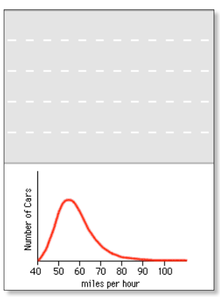
*Use R = 8.3145 * 10^{-3} kJ mol^{-1} K^{-1}, MM = kg mol^{-1}*



Distribution of Gas Molecule Speeds

What is an "average" speed?

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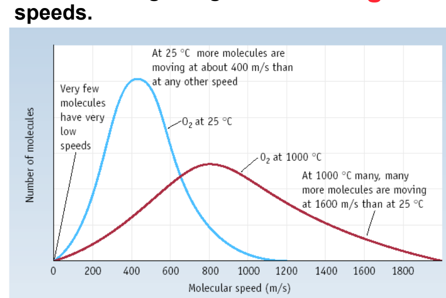
Velocity of Gas Molecules

Molecules of a given gas have a range of speeds.



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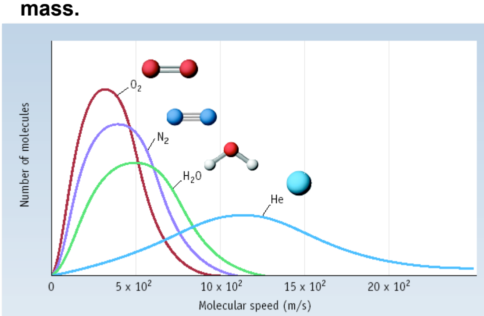
Velocity of Gas Molecules

Average velocity decreases with increasing mass.



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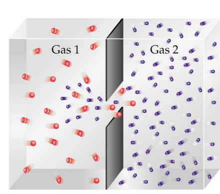
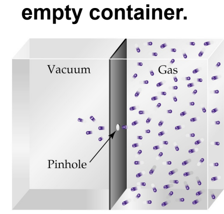
GAS DIFFUSION AND EFFUSION

diffusion is the gradual mixing of molecules of different gases.

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





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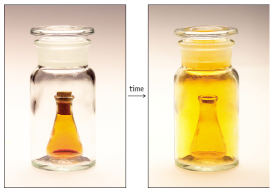
Molecules effuse through holes in a rubber balloon, for example, at a rate (= moles/time) that is

- proportional to T
- inversely proportional to M.

Therefore, He effuses more rapidly than O_2 at same T.

GAS DIFFUSION AND EFFUSION



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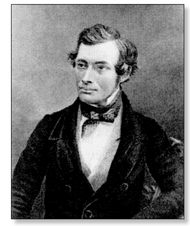
GAS DIFFUSION AND EFFUSION

Graham's law governs effusion and diffusion of gas molecules.

$$\frac{\text{Rate for A}}{\text{Rate for B}} = \sqrt{\frac{\text{M of B}}{\text{M of A}}}$$

Rate of effusion is inversely proportional to its molar mass.



Thomas Graham, 1805-1869. Professor in Glasgow and London.

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Gas Diffusion

relation of mass to rate of diffusion

Gaseous diffusion of $NH_3(g)$ and $HCl(g)$

- HCl and NH_3 diffuse from opposite ends of tube.
- Gases meet to form NH_4Cl
- HCl heavier than NH_3
- Therefore, NH_4Cl forms closer to HCl end of tube.

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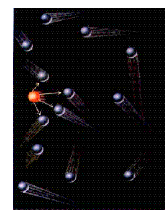
Deviations from Ideal Gas Law

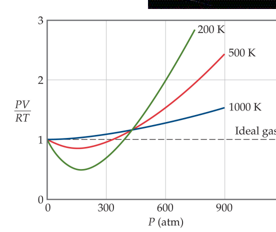


Real molecules have volume.

There are intermolecular forces.

Otherwise a gas could not become a liquid.

High Pressure and Low Temperature conditions show greatest deviation



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Deviations from Ideal Gas Law

Account for volume of molecules and intermolecular forces with VAN DER WAAL'S EQUATION.

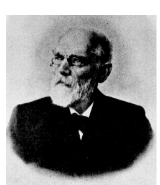
Measured P

Measured V = V(ideal)

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

intermol. forces

vol. correction



J. van der Waals, 1837-1923, Professor of Physics, Amsterdam. Nobel Prize 1910.

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Deviations from Ideal Gas Law

Cl_2 gas has a = 6.49, b = 0.0562

For 8.0 mol Cl_2 in a 4.0 L tank at 27 °C.

P (ideal) = nRT/V = 49.3 atm

P (van der Waals) = 29.5 atm

van der Waals Constants for Gas Molecules

Substance	a (L^2-atm/mol^2)	b (L/mol)
He	0.0341	0.02370
Ne	0.211	0.0171
Ar	1.34	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0510
H_2	0.244	0.0266
N_2	1.39	0.0391
O_2	1.36	0.0318
Cl_2	6.49	0.0562
H_2O	5.46	0.0305
CH_4	2.25	0.0428
CO_2	3.59	0.0427
CCl_4	20.4	0.1383

Measured P

Measured V = V(ideal)

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

intermol. forces

vol. correction

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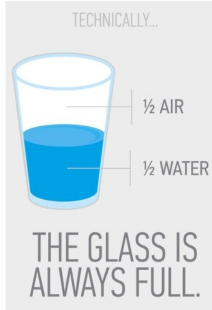
End of Chapter 9



See:

- *Chapter Nine Study Guide*
- *Chapter Nine Concept Guide*
- *Important Equations (following this slide)*
- *End of Chapter Problems (following this slide)*

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Important Equations, Constants, and Handouts from this Chapter:

- **know how to use the gas laws, desired units for the gas law, STP uses, Dalton's Law of Partial Pressure, etc.**
- **understand pressure**
- **know how to use gases in stoichiometry problems**
- **know how the KMT (Kinetic Molecular Theory) describes gases**

- **$PV = nRT$**
- **$PM = dRT$**
- **mole = 6.022×10^{23}**
- **760 mm Hg = 1 atm**
- **1013 mbar = 1 atm**
- *metric prefixes* **(m, k,** *etc.***)**
- **STP = 1 atm, 273.15 K**

$R = 0.082057$ L atm mol^{-1} K^{-1} *(the "gas R")*
$R = 8.3145$ J mol^{-1} K^{-1} *(the "energy R")*

$$KE = {}^1/_2mv^2 = {}^3/_2RT$$

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End of Chapter Problems: ***Test Yourself***

1. A sample of nitrogen gas has a pressure of 67.5 mm Hg in a 500. mL flask. What is the pressure of this gas sample when it is transferred to a 125 mL flask at the same temperature?
2. You have 3.5 L of NO at a temperature of 22.0 °C. What volume would the NO occupy at 37 °C? (Assume the pressure is constant.)
3. An automobile cylinder has a volume of 400. cm^3. The engine takes in air at a pressure of 1.00 atm and a temperature of 15 °C and compresses the air to a volume of 50.0 cm^3 at 77 °C. What is the final pressure of the gas in the cylinder?
4. A 1.25 g sample of CO_2 is contained in a 750. mL flask at 22.5 °C. What is the pressure of the gas?
5. A gaseous organofluorine compound has a density of 0.355 g/L at 17 °C and 189 mm Hg. What is the molar mass of the compound?
6. Sodium azide, the explosive compound in automobile air bags, decomposes according to the following equation:
   $2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$
   What mass of sodium azide is required to provide the nitrogen needed to inflate a 75.0 L bag to a pressure of 1.3 atm at 25 °C?

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End of Chapter Problems: ***Answers***

1. 270. mm Hg
2. 3.7 L
3. 9.72 atm
4. 0.919 atm
5. 34.0 g/mol
6. 170 g

Be sure to view practice problem set #3 and self quizzes for ***nomenclature*** *examples and practice*

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