CHEMISTRY

Chapter 9 GASES

Kevin Kolack, Ph.D.
The Cooper Union

HW problems: 5, 13, 15, 29, 35, 37, 53, 61, 75, 89, 101, 103

CH. 9 OUTLINE

9.1 Gas Pressure

9.2 Relating Pressure, Volume, Amount, and Temperature: The Ideal Gas Law

9.3 Stoichiometry of Gaseous Substances, Mixtures, and Reactions

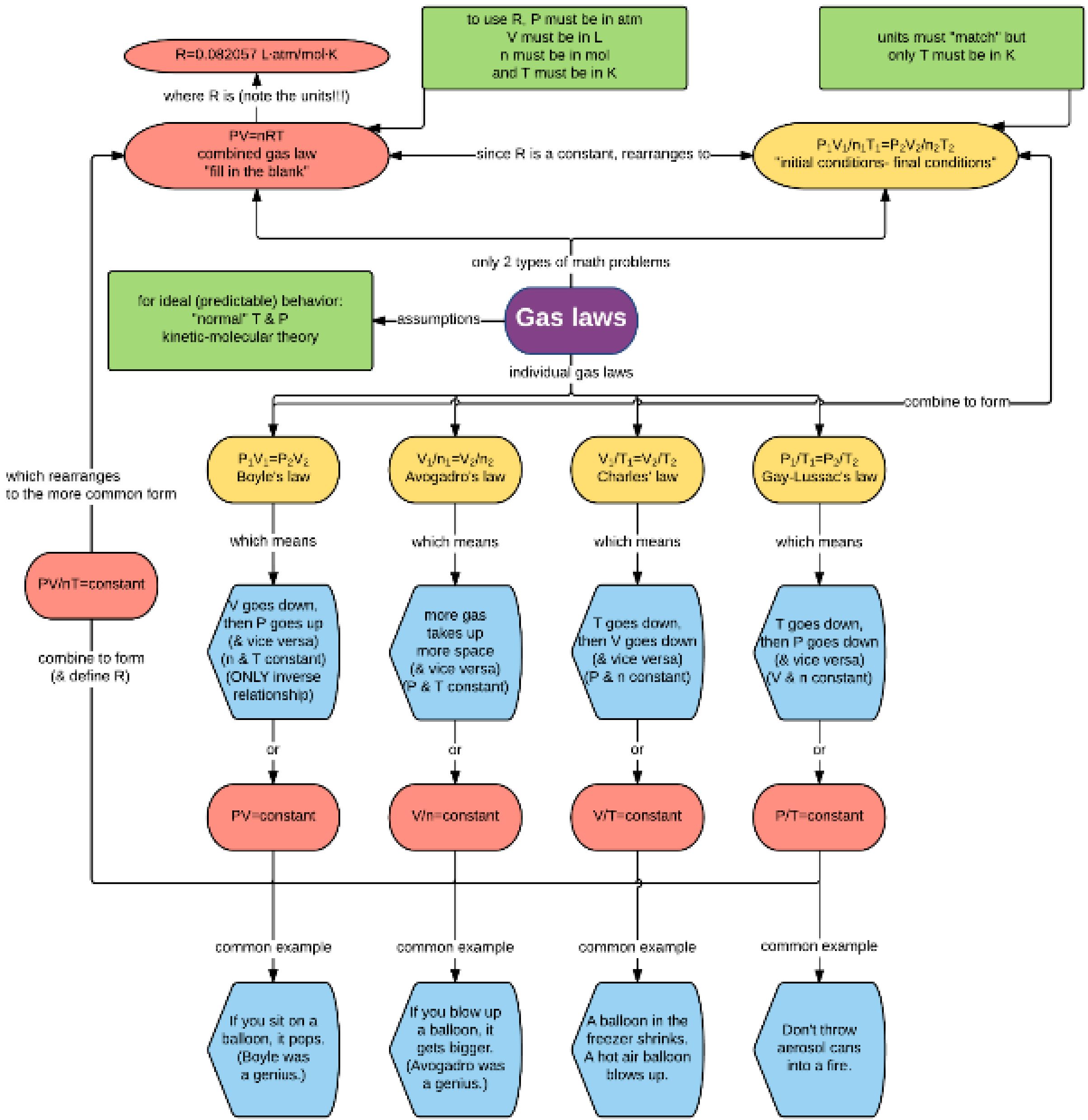
9.4 Effusion and Diffusion of Gases

9.5 The Kinetic-Molecular Theory

9.6 Non-Ideal Gas Behavior

CONCEPT MAP

openstax COLLEGE



Lucidchart.com

REVIEW:
PHYSICAL CHARACTERISTICS OF GASES

- Gases assume the volume and shape of their containers.
- Gases are the most compressible state of matter.
- Gases will mix evenly and completely when confined to the same container.
- Gases have much lower densities than liquids and solids.

Cl_2 gas

NO_2 gas

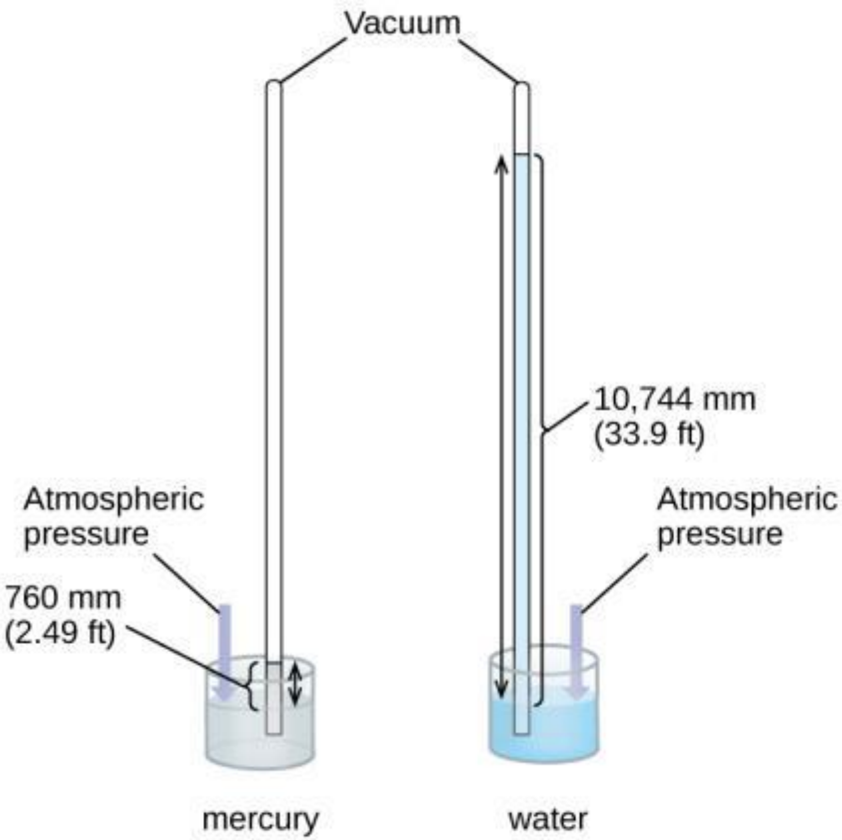
PRESSURE

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

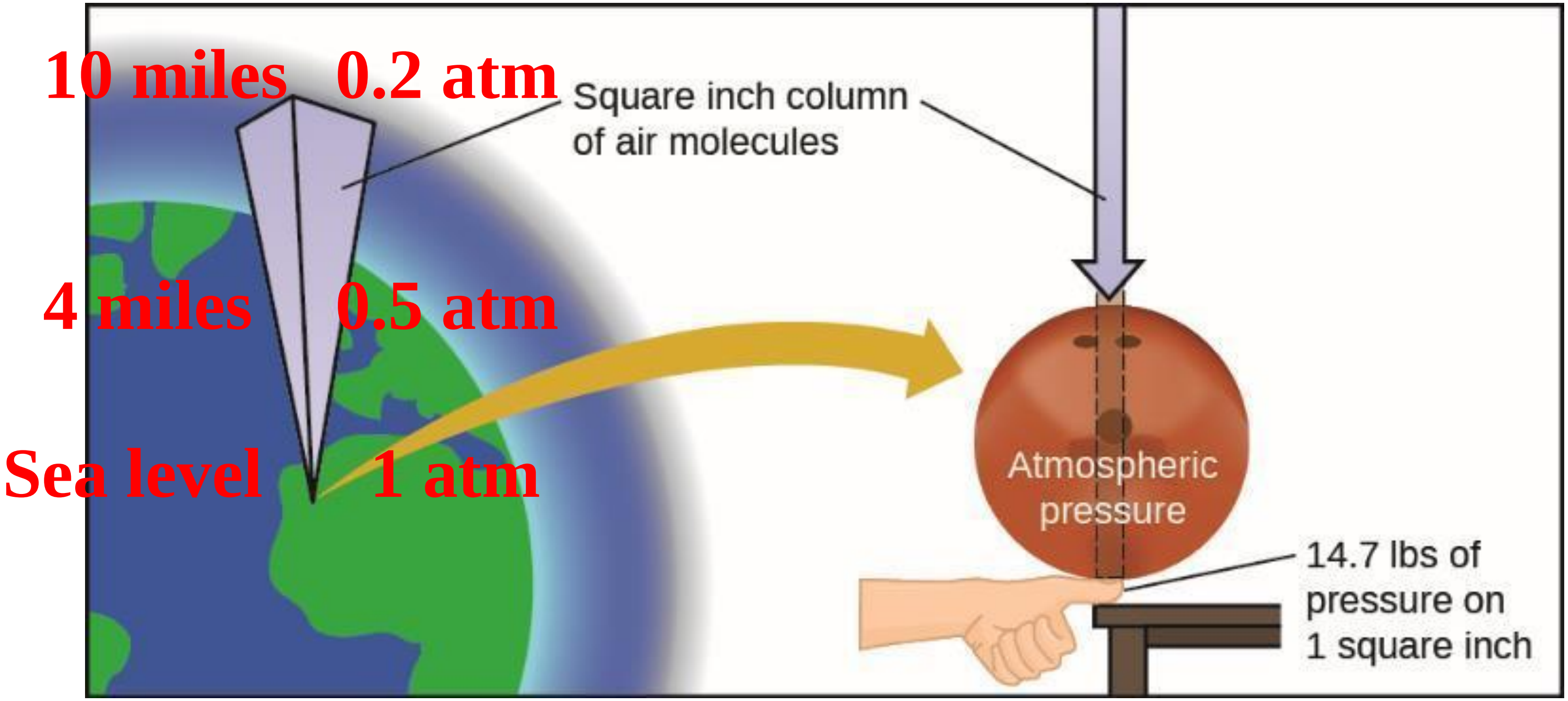
(force = mass x acceleration)

Units of pressure

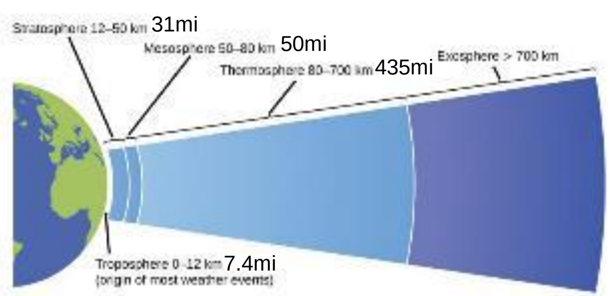
- 1 pascal (Pa) = 1 N/m^2 (SI unit)
- 1 atm = 760 mmHg = 760 torr
- 1 atm = 101,325 Pa
- 1 atm = 101.325 kPa
- 1 atm = 1.01325 bar
- 1 atm = 29.921” Hg
- 1 atm = 14.7 psi



PRESSURE VS. ALTITUDE (FIG 9.2)



(If you actually put a bowling ball on your thumb, the pressure experienced would be *twice* the usual pressure, and the sensation would be… unpleasant.)



AMUSING COMPARISONS

$$\text{pressure per elephant foot} = 14{,}000\,\frac{\text{lb}}{\text{elephant}} \times \frac{1\text{ elephant}}{4\text{ feet}} \times \frac{1\text{ foot}}{250\text{ in}^2} = 14\text{ lb/in}^2$$

$$\text{pressure per skate blade} = 120\,\frac{\text{lb}}{\text{skater}} \times \frac{1\text{ skater}}{2\text{ blades}} \times \frac{1\text{ blade}}{2\text{ in}^2} = 30\text{ lb/in}^2$$

$$\text{pressure per human foot} = 120\,\frac{\text{lb}}{\text{skater}} \times \frac{1\text{ skater}}{2\text{ feet}} \times \frac{1\text{ foot}}{30\text{ in}^2} = 2\text{ lb/in}^2$$

QUESTION

The pressure outside an airplane flying at high altitude falls considerably below standard atmospheric pressure. Therefore, the air inside the cabin must be pressurized to protect the passengers…

What is the pressure in atmospheres in the cabin if the barometer reading is 688 mmHg?

$$\text{pressure} = 688\ \cancel{\text{mmHg}} \times \frac{1\ \text{atm}}{760\ \cancel{\text{mmHg}}}$$

$$= 0.905\ \text{atm}$$

QUESTION

The atmospheric pressure in San Francisco on a certain day is 732 mmHg...

What is the pressure in kPa?

$$\text{pressure} = 732\ \cancel{\text{mmHg}} \times \frac{1.01325 \times 10^5\ \text{Pa}}{760\ \cancel{\text{mmHg}}}$$

$$= 9.76 \times 10^4\ \text{Pa}$$

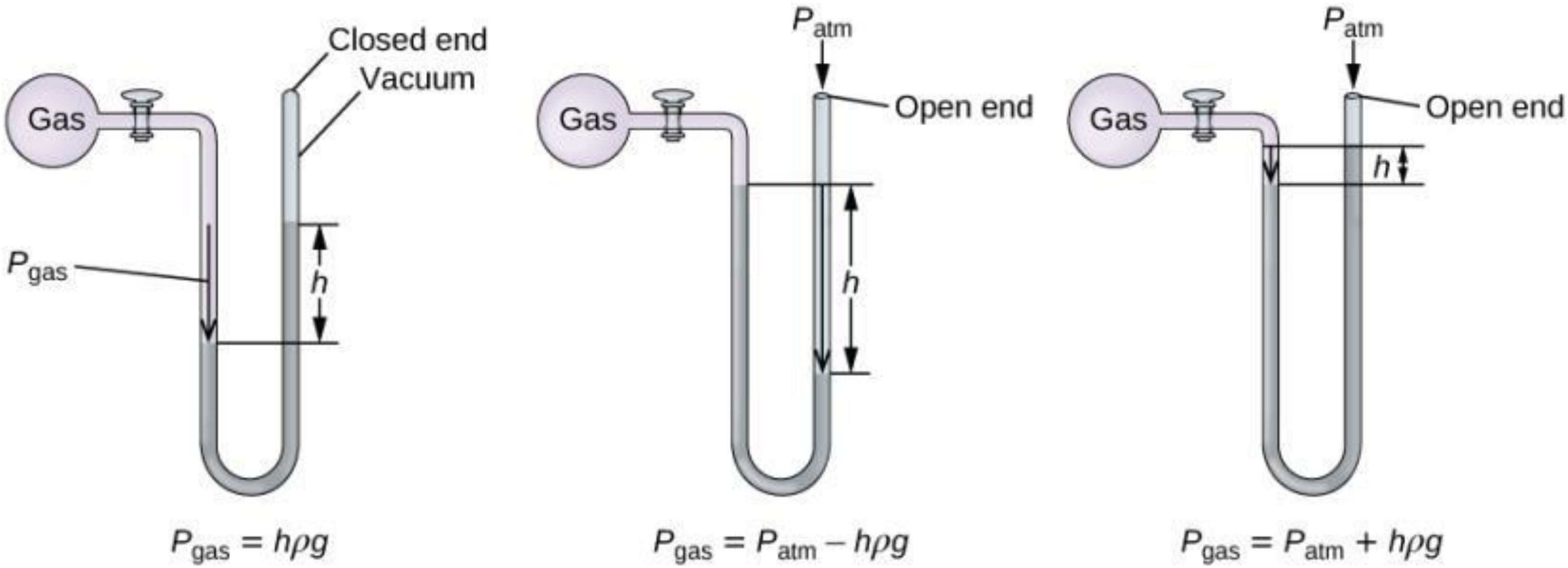
$$= 97.6\ \text{kPa}$$

INSTRUMENTS TO MEASURE GAS PRESSURE

Barometers measure atmospheric pressure.

Manometers measure (other) gas pressures.

closed open



A manometer can be used to measure the pressure of a gas. The (difference in) height between the liquid levels (h) is a measure of the pressure. Mercury is usually used because of its large density.

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PUNCH LINE: 2 TYPES OF GAS PROBLEMS

Fill in the blank

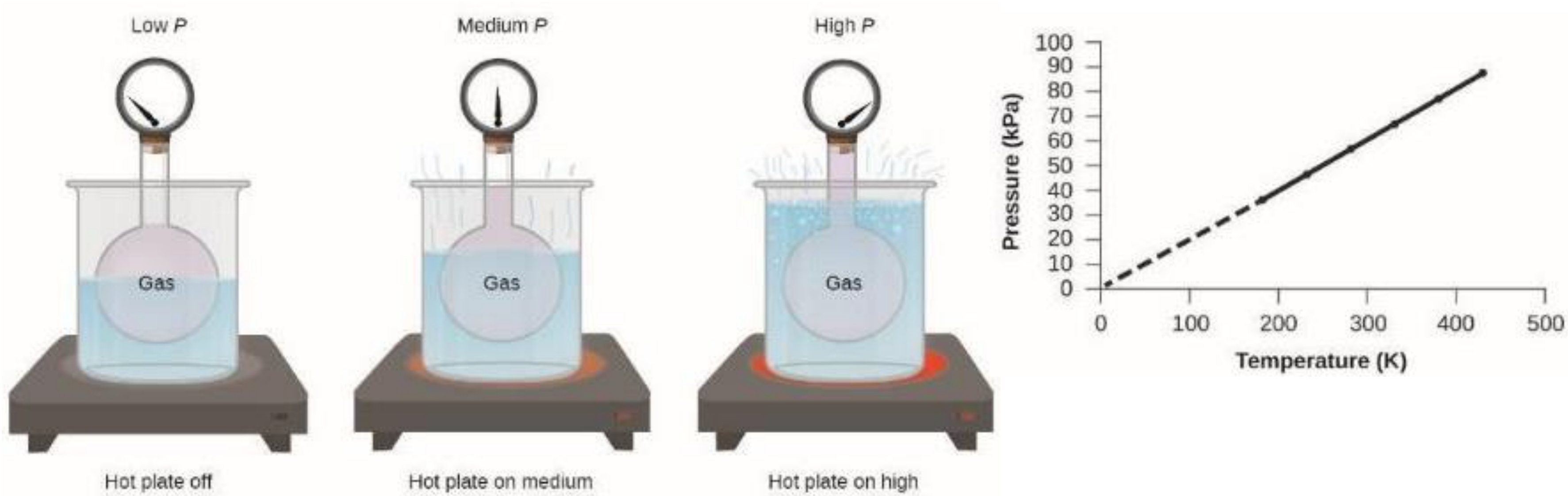
- PV=nRT
  - MUST use the correct units for R to work

Initial conditions – final conditions

- Since PV = nRT, then PV/nT = constant
- Therefore $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$
  - Note that at constant n and T, we have Boyle's law
  - At constant n and P, we have Charles' law
  - At constant P and T, we have Avogadro's law
  - Other than T, you can often get away with not converting units

AMONTON'S LAW (~1700)

P increases with T at constant V and n
(Gay-Lussac ~1800)



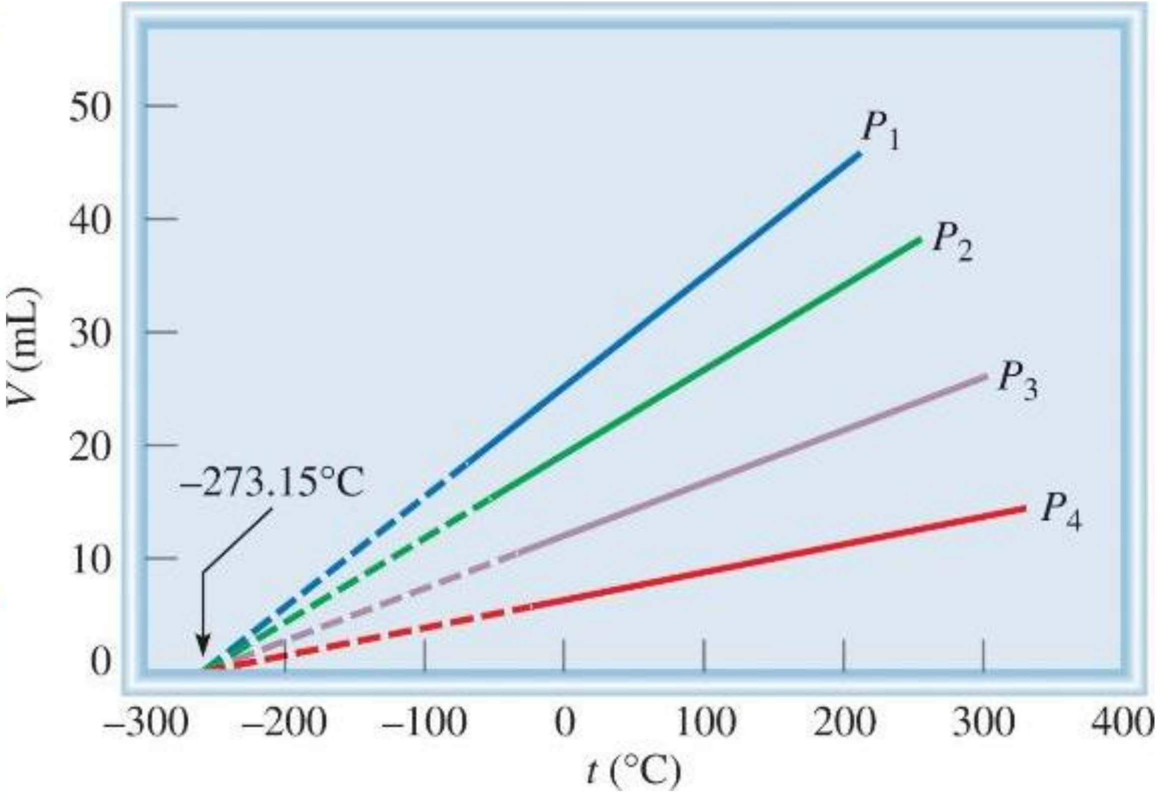
$P \alpha T$ *and thus* $P/T \alpha 1$

$P/T = \text{constant}$

$P_1/T_1 = P_2/T_2$

Temperature **must** be in Kelvin
(The origin of absolute 0.)

CHARLES' LAW



V increases with T at constant P and n

https://www.youtube.com/watch?v=ZvrJgGhnmJo

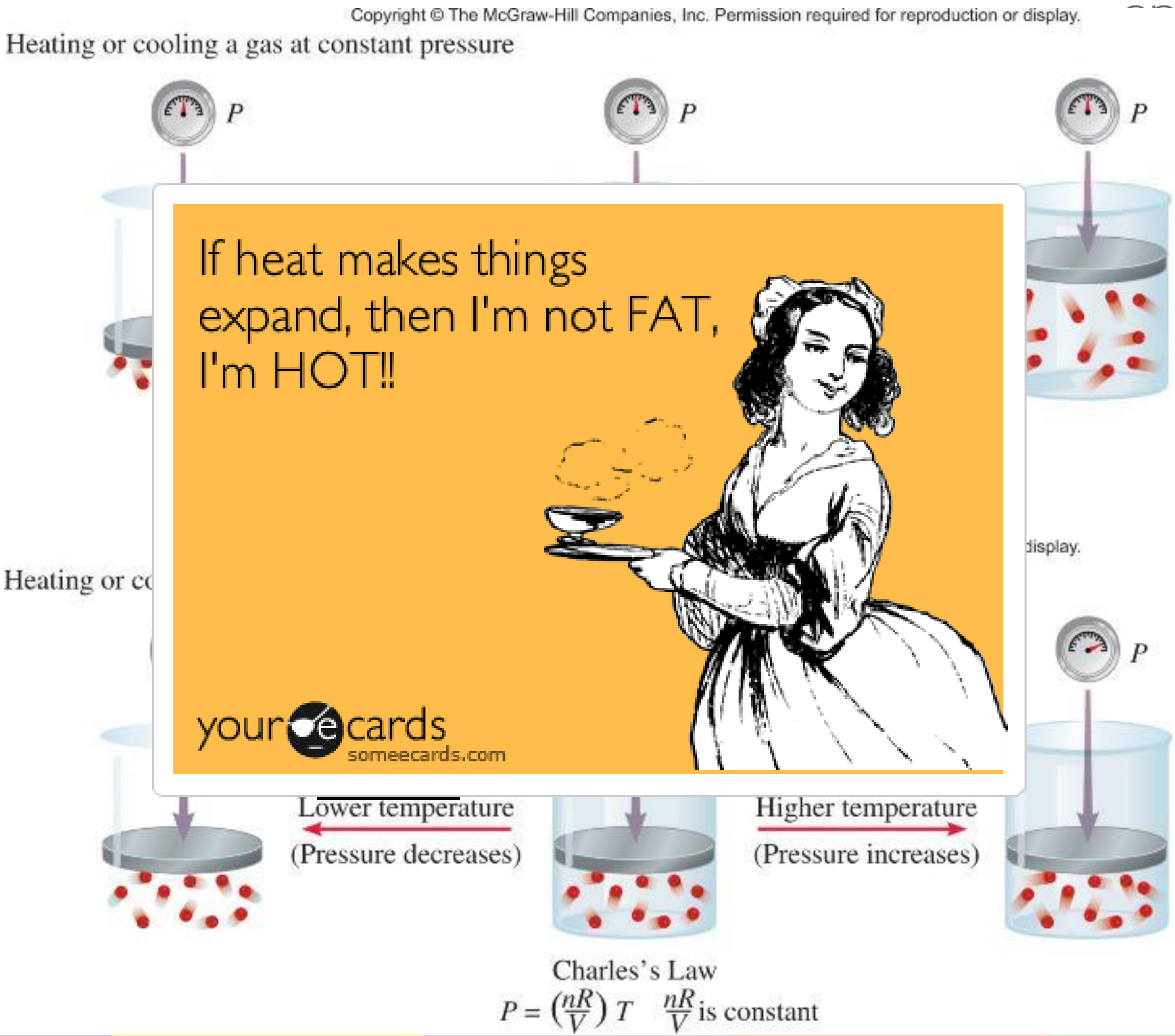
V α T *and thus* V/T α 1

V/T = constant

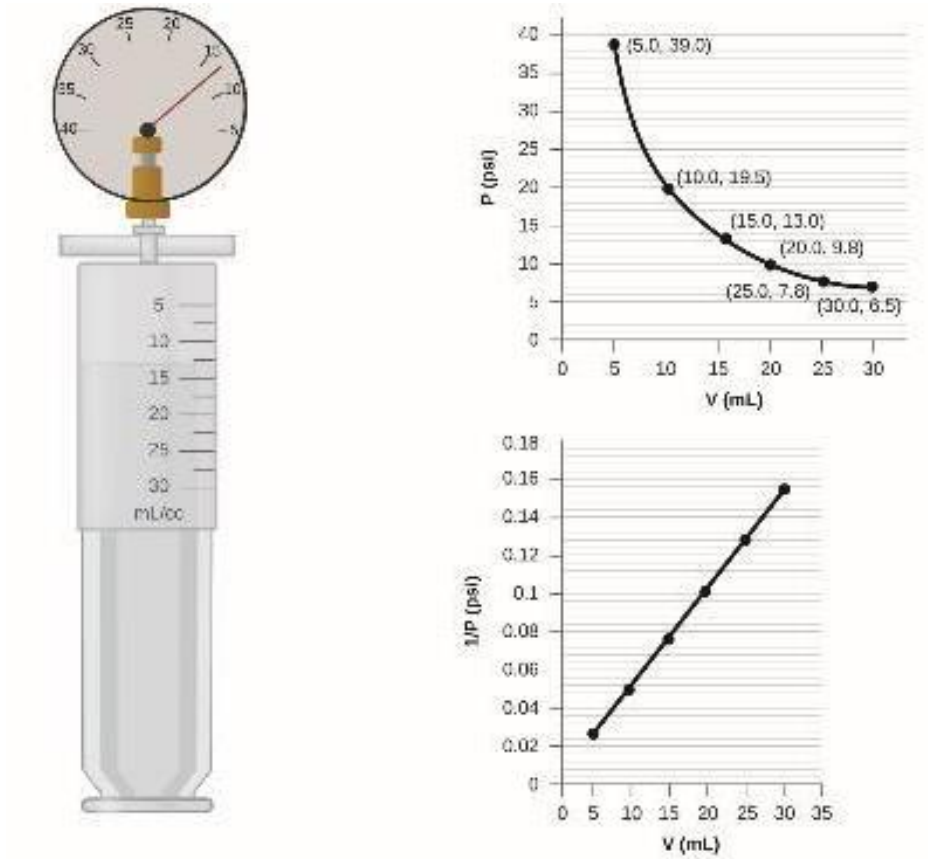
$V_1/T_1 = V_2/T_2$

Temperature **must** be in Kelvin (The origin of absolute 0.)

CHARLES' LAW (CONT'D)



BOYLE'S LAW



When a gas occupies a smaller volume, it exerts a higher pressure; when it occupies a larger volume, it exerts a lower pressure (assuming the amount of gas and the temperature do not change). Since *P* and *V* are inversely proportional, a graph of $\frac{1}{P}$ vs. *V* is linear.

$P\ \alpha\ 1/V$ *and thus* $PV\ \alpha\ 1$

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

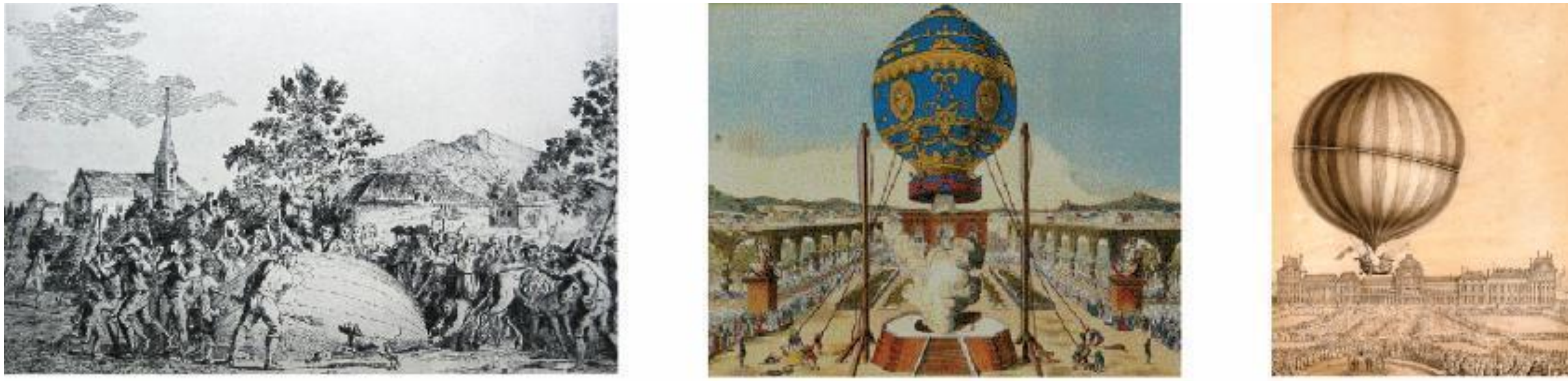
$P_1 \times V_1 = P_2 \times V_2$

http://openstaxcollege.org/l/16atmospressur1

(At constant temperature)

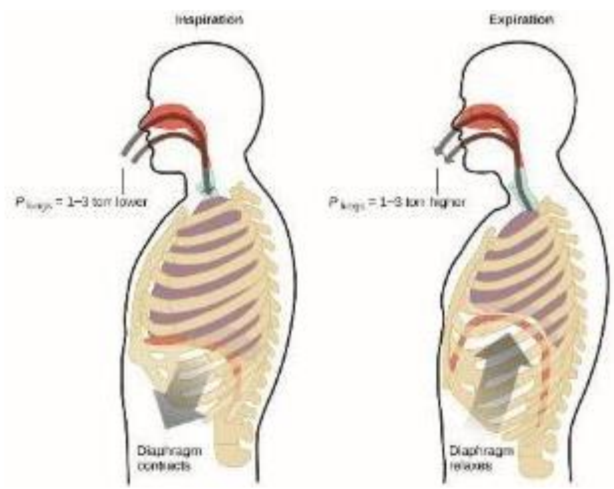
(And constant amount of gas)

FIGURE 9.9



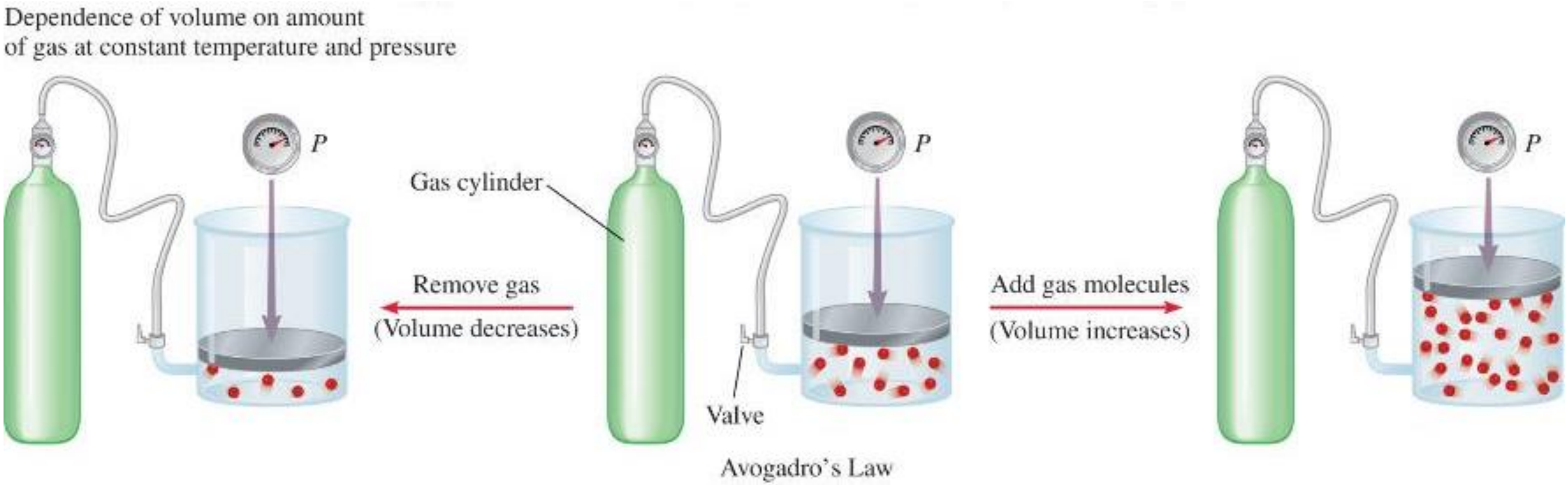
(a) (b) (c)

In 1783, the first (a) hydrogen-filled balloon flight, (b) manned hot air balloon flight, and (c) manned hydrogen-filled balloon flight occurred. When the hydrogen-filled balloon depicted in (a) landed, the frightened villagers of Gonesse reportedly destroyed it with pitchforks and knives. The launch of the latter was reportedly viewed by 400,000 people in Paris.



Breathing occurs because expanding and contracting lung volume creates small pressure differences between your lungs and your surroundings, causing air to be drawn into and forced out of your lungs.

AVOGADRO'S LAW



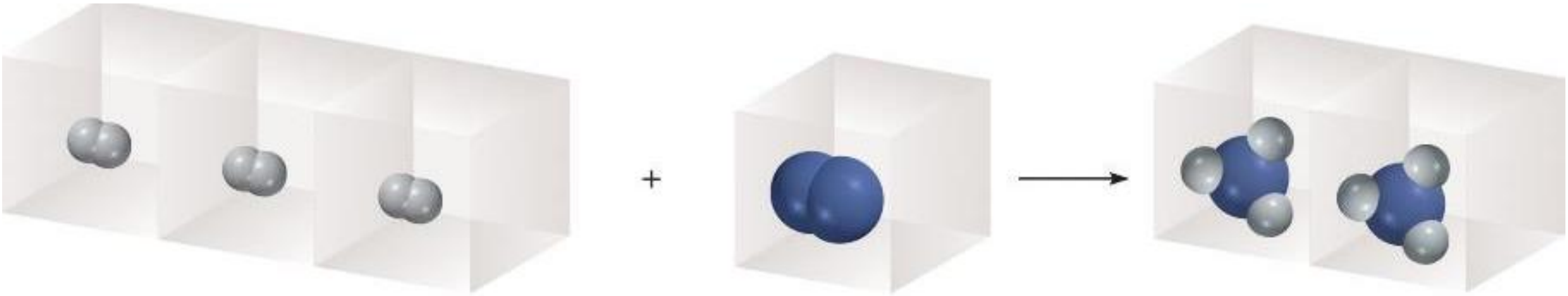
AVOGADRO'S LAW (CONT'D)

$V \alpha$ number of moles (n) *and thus* $V/n \; \alpha \; 1$

V/n = constant

$V_1 / n_1 = V_2 / n_2$

(At constant T)
(And constant P)



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

(COMBINED) IDEAL GAS LAW

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

Charles's law: $V \alpha T$ (at constant n and P)
Guy-Lussac's law: $P \alpha T$ (at constant n and V)

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

$PV \alpha nT$

PV = constant x nT = nRT
where R is the ideal gas constant

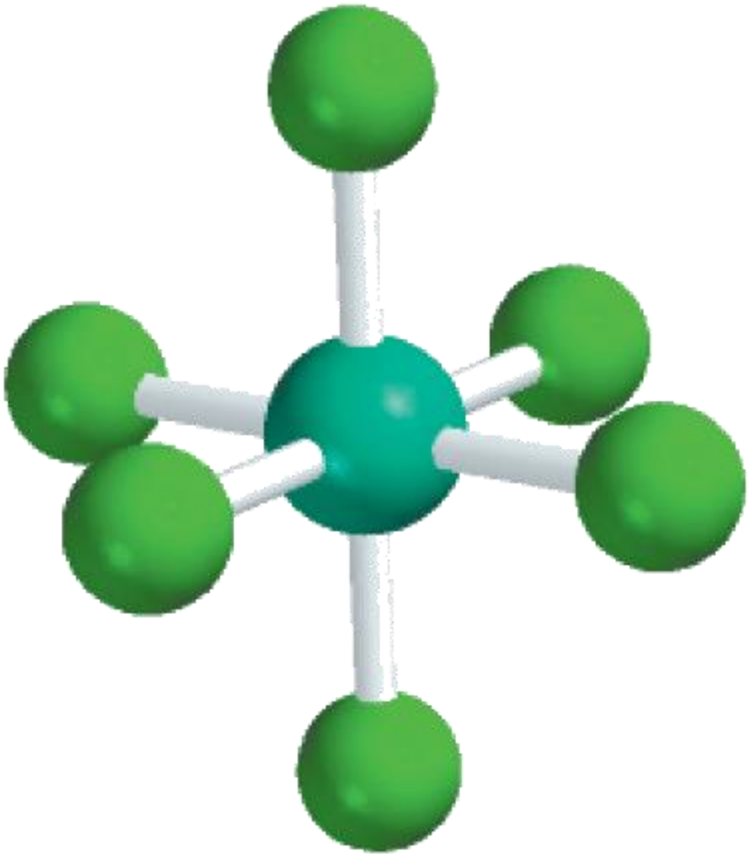
$$\boxed{PV = nRT}$$

QUESTION

Sulfur hexafluoride (SF_6) is a colorless and odorless gas.

Due to its lack of chemical reactivity, it is used as an insulator in electronic equipment.

Calculate the pressure (in atm) exerted by 1.82 moles of the gas in a 5.43 L steel vessel at 69.5°C.



SF_6

SOLUTION

Use the ideal gas law, converting the temperature to Kelvin.

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

$$= \frac{(1.82\ \text{mol})(0.0821\ \text{L}\cdot\text{atm/K}\cdot\text{mol})(69.5 + 273)\text{K}}{5.43\ \text{L}}$$

$$= 9.42\ \text{atm}$$

QUESTION 5.5

An inflated helium balloon with a volume of 0.55 L (not pictured) at sea level (1.0 atm) is allowed to rise to a height of 6.5 km, where the pressure is about 0.40 atm.

Assuming that the temperature remains constant, what is the final volume of the balloon?

A research He balloon.

SOLUTION

The amount of gas inside the balloon and its temperature remain constant, but both the pressure and the volume change. What gas law do you need?

$$\frac{P_1V_1}{\cancel{n_1}\cancel{T_1}} = \frac{P_2V_2}{\cancel{n_2}\cancel{T_2}}$$

which is Boyle's law:

$$P_1V_1 = P_2V_2$$

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

$$= 0.55\ \text{L} \times \frac{1.0\ \text{atm}}{0.40\ \text{atm}}$$

$$= 1.4\ \text{L}$$

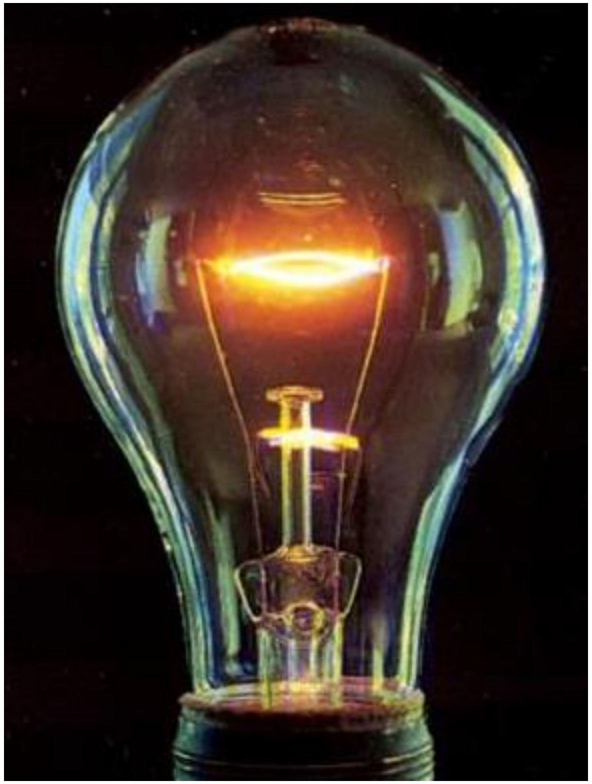
Check When pressure applied on the balloon is reduced (at constant temperature), the helium gas expands and the balloon's volume increases. The final volume is greater than the initial volume, so the answer is reasonable.

QUESTION 5.6

Argon is an inert gas used in light bulbs to retard the vaporization of the tungsten filament.

A certain light bulb containing argon at 1.20 atm and 18.0°C is heated to 85°C at constant V.

Calculate the final P inside the bulb (in atm).



SOLUTION

Strategy The temperature and pressure of argon change but the amount and volume of gas remain the same.

What equation would you use to solve for the final pressure?

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

which is Charles's law [see Equation (5.6)].

(What temperature unit should you use?)

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

$$= 1.20 \text{ atm} \times \frac{358 \text{ K}}{291 \text{ K}}$$

$$= 1.48 \text{ atm}$$

Check At constant volume, the pressure of a given amount of gas is directly proportional to its absolute temperature (ie- P goes up when T goes up). Therefore, the increase in pressure is reasonable.

SCUBA DIVING

760 mm Hg = 33 ft of water (density difference)

Thus:

Depth (ft)	Pressure (atm)
0	1
33	2
66	3

Since V↑ when P↓, small bubbles in the bloodstream become a big problem when surfacing.

QUESTION 5.7

A small bubble rises from the bottom of a lake, where the temperature and pressure are 8.0°C and 6.4 atm, to the water's surface, where the temperature is 25.0°C and the pressure is 1.0 atm.

Calculate the final volume (in mL) of the bubble if its initial volume was 2.1 mL. (Divers, beware!)

SOLUTION

According to the combined gas law:

$$\frac{P_1V_1}{\cancel{n_1}T_1} = \frac{P_2V_2}{\cancel{n_2}T_2}$$

We assume that the amount of air in the bubble remains constant, that is, $n_1 = n_2$ so that:

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

$T_1 = (8 + 273)\text{ K} = 281\text{ K}$

$T_2 = (25 + 273)\text{K} = 298\text{ K}$

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

$$= 2.1\text{ mL} \times \frac{6.4\text{ atm}}{1.0\text{ atm}} \times \frac{298\text{ K}}{281\text{ K}}$$

$$= 14\text{ mL}$$

STANDARD MOLAR VOLUME

The conditions 0 °C and 1 atm are called standard temperature and pressure (STP).

Experiments show that at STP, 1 mole of an ideal gas occupies 22.414 L.

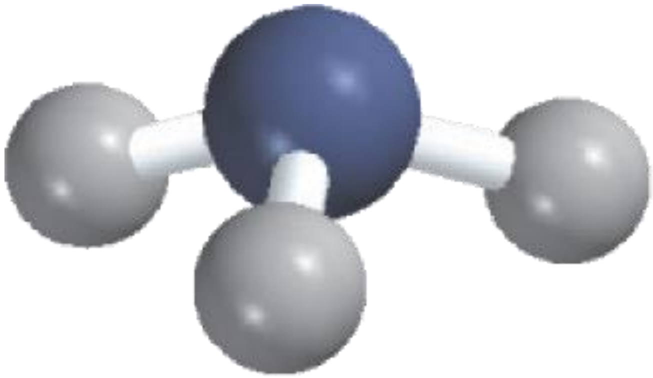
$$PV = nRT$$

$$R = \frac{PV}{nT} = \frac{(1\ \text{atm})(22.414\text{L})}{(1\ \text{mol})(273.15\ \text{K})}$$

$$R = 0.082057\ \text{L} \cdot \text{atm} / (\text{mol} \cdot \text{K})$$

QUESTION

Calculate the volume (in L) occupied by 7.40 g of NH3 at STP.



NH_3

Recognizing that 1 mole of an ideal gas occupies 22.41 L at STP and using the molar mass of NH3 (17.03 g), we write the sequence of conversions as:

$$\text{grams of } NH_3 \longrightarrow \text{moles of } NH_3 \longrightarrow \text{liters of } NH_3 \text{ at STP}$$

SOLUTION

So the volume of NH_3 is given by

$$V = 7.40\ \cancel{\text{g NH}_3} \times \frac{1\ \cancel{\text{mol NH}_3}}{17.03\ \cancel{\text{g NH}_3}} \times \frac{22.41\ \text{L}}{1\ \cancel{\text{mol NH}_3}}$$

$$= 9.74\ \text{L}$$

Alternately, the problem can also be solved by first converting 7.40 g of NH_3 to number of moles of NH_3, and then applying the ideal gas equation ($V = nRT/P$).

Make sense? Because 7.40 g of NH_3 is smaller than its molar mass, its volume at STP should be smaller than 22.41 L. Therefore, the answer is reasonable.

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GAS DENSITY, MW, & THE GAS LAWS

Since d = m/V and V=nRT/P and MW = m/n

$$d = \frac{m}{V} = \frac{Pm}{nRT} = \frac{P(MW)}{RT}$$

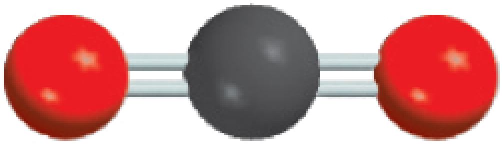
And thus, the molar mass (MW) of a gas is

$$MW = \frac{dRT}{P}$$

where d is the density of the gas in g/L

QUESTION

Calculate the density of carbon dioxide (CO_2) in grams per liter (g/L) at 0.990 atm and 55°C.



CO_2

$T = 273 + 55 = 328\ \text{K}$

$$d = \frac{P\mathscr{M}}{RT}$$

$$= \frac{(0.990\ \text{atm})\ (44.01\ \text{g/mol})}{(0.0821\ \text{L} \cdot \text{atm/K} \cdot \text{mol})\ (328\ \text{K})}$$

$$= 1.62\ \text{g/L}$$

Alternately (preferably),

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

$$= \frac{(1\ \text{mol})\ (0.0821\ \text{L} \cdot \text{atm/K} \cdot \text{mol})\ (328\ \text{K})}{0.990\ \text{atm}}$$

$$= 27.2\ \text{L}$$

$$d = \frac{44.01\ \text{g}}{27.2\ \text{L}} = 1.62\ \text{g/L}$$

QUESTION

Chemical analysis of a gaseous compound showed that it contained 33.0 percent silicon (Si) and 67.0 percent fluorine (F) by mass.

At 35°C, 0.210 L of the compound exerted a pressure of 1.70 atm.

If the mass of 0.210 L of the compound was 2.38 g, calculate the molecular formula of the compound.

SOLUTION

Assume that we have 100 g of the compound, so the percentages are converted to grams.

The number of moles of Si and F are given by

$$n_{\mathrm{Si}} = 33.0\ \cancel{\mathrm{g\ Si}} \times \frac{1\ \mathrm{mol\ Si}}{28.09\ \cancel{\mathrm{g\ Si}}} = 1.17\ \mathrm{mol\ Si}$$

$$n_{\mathrm{F}} = 67.0\ \cancel{\mathrm{g\ F}} \times \frac{1\ \mathrm{mol\ F}}{19.00\ \cancel{\mathrm{g\ F}}} = 3.53\ \mathrm{mol\ F}$$

Therefore, the empirical formula is $Si_{1.17}F_{3.53}$, or, dividing by the smaller subscript (1.17), we obtain SiF_3.

SOLUTION (CONT'D)

To calculate the molar mass of the compound, we need first to calculate the number of moles contained in 2.38 g of the compound. From the ideal gas equation

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

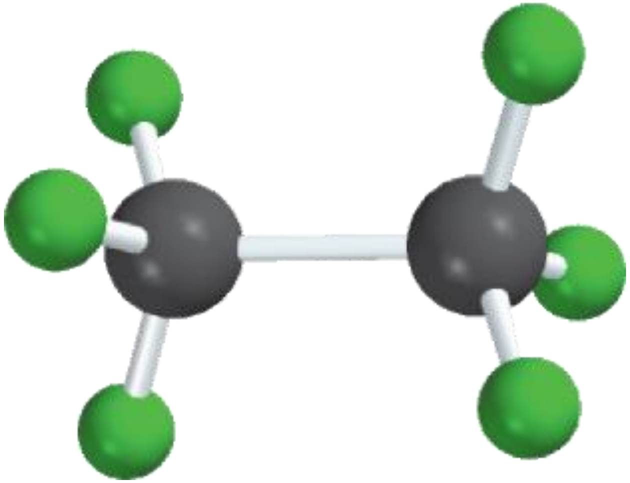
$$= \frac{(1.70 \text{ atm})(0.210 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(308 \text{ K})} = 0.0141 \text{ mol}$$

Because there are 2.38 g in 0.0141 mole of the compound, the mass in 1 mole, or the molar mass, is given by

$$\mathcal{M} = \frac{2.38 \text{ g}}{0.0141 \text{ mol}} = 169 \text{ g/mol}$$

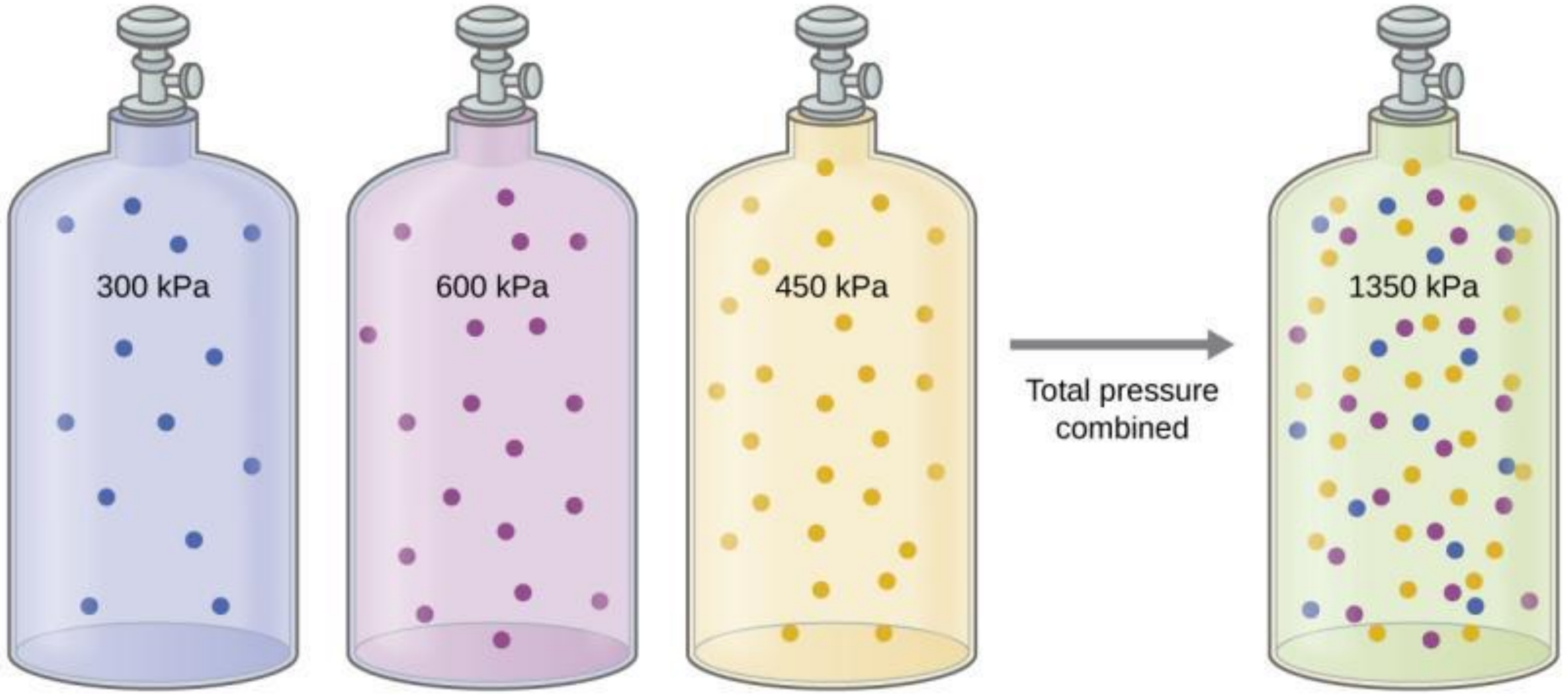
SOLUTION (CONT'D)

- The molar mass of the empirical formula SiF_3 is 85.09 g.
- Recall that the ratio (molar mass/empirical molar mass) is always an integer (169/85.09 ≈ 2).
- Therefore, the molecular formula of the compound must be $(SiF_3)_2$ or Si_2F_6.



DALTON'S LAW OF PARTIAL PRESSURES

At constant V and T, the total pressure of a mixture of ideal gases is equal to the sum of the partial pressures of the component gases.



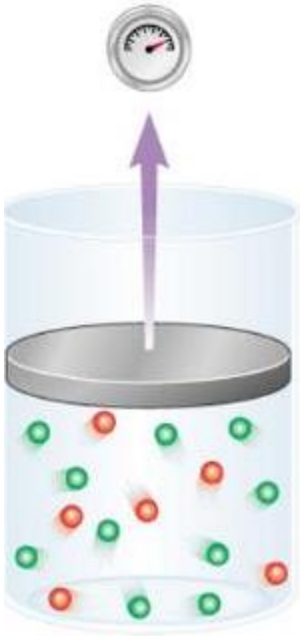
If equal-volume cylinders containing gas A at a pressure of 300 kPa, gas B at a pressure of 600 kPa, and gas C at a pressure of 450 kPa are all combined in the same-size cylinder, the total pressure of the mixture is 1350 kPa.

MOLE FRACTION AND PARTIAL P

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

$$P_B = \frac{n_B RT}{V}$$ n_B is the number of moles of B



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

$$P_A = X_A P_T \qquad P_B = X_B P_T$$

$$\boxed{P_i = X_i P_T}$$

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

QUESTION 5.14

A mixture of gases contains 4.46 moles of neon (Ne), 0.74 mole of argon (Ar), and 2.15 moles of xenon (Xe).

need to find

$$P_{\mathrm{Ne}} = X_{\mathrm{Ne}} P_{\mathrm{T}}$$

want to calculate

given

Calculate the partial pressures of the gases if the total pressure is 2.00 atm (at constant T).

$$X_{\mathrm{Ne}} = \frac{n_{\mathrm{Ne}}}{n_{\mathrm{Ne}} + n_{\mathrm{Ar}} + n_{\mathrm{Xe}}} = \frac{4.46 \text{ mol}}{4.46 \text{ mol} + 0.74 \text{ mol} + 2.15 \text{ mol}}$$

$$= 0.607$$

$$P_{\mathrm{Ne}} = X_{\mathrm{Ne}} P_{\mathrm{T}}$$
$$= 0.607 \times 2.00 \text{ atm}$$
$$= 1.21 \text{ atm}$$

$$P_{\mathrm{Ar}} = X_{\mathrm{Ar}} P_{\mathrm{T}}$$
$$= 0.10 \times 2.00 \text{ atm}$$
$$= 0.20 \text{ atm}$$

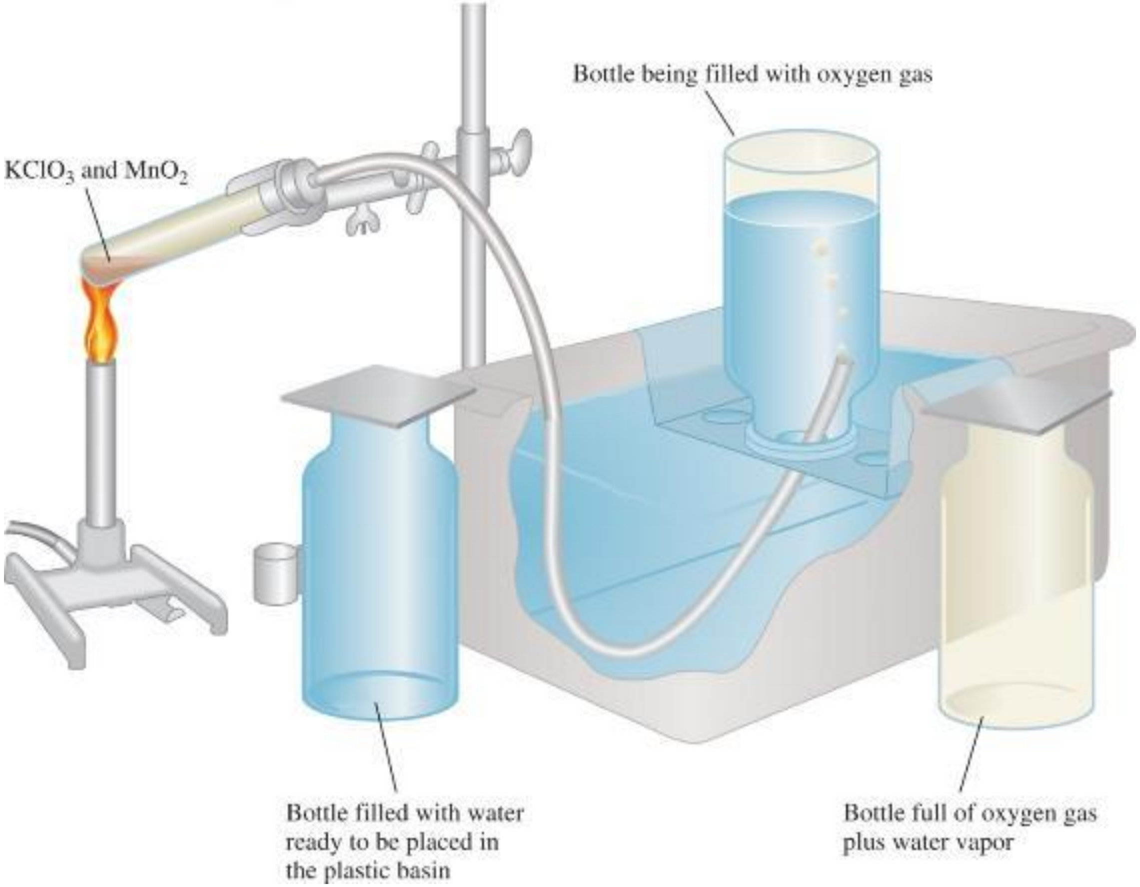
$$P_{\mathrm{Xe}} = X_{\mathrm{Xe}} P_{\mathrm{T}}$$

(or subtract)

$$= 0.293 \times 2.00 \text{ atm}$$
$$= 0.586 \text{ atm}$$

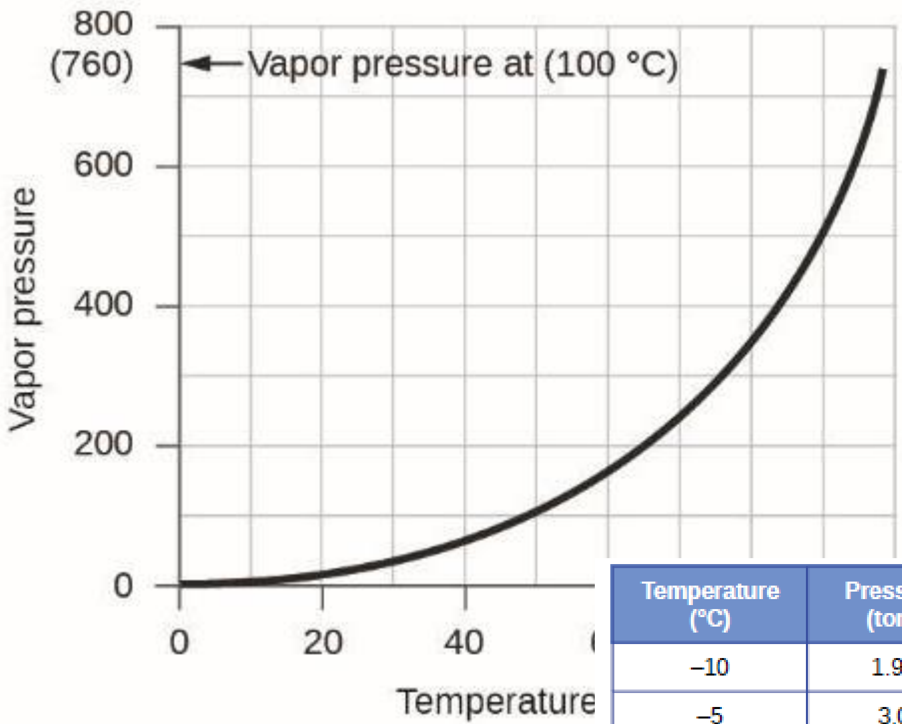
COLLECTING A GAS OVER WATER

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



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

VAPOR PRESSURE OF H_2O VS. T



Temperature (°C)	Pressure (torr)		Temperature (°C)	Pressure (torr)		Temperature (°C)	Pressure (torr)
–10	1.95		18	15.5		30	31.8
–5	3.0		19	16.5		35	42.2
–2	3.9		20	17.5		40	55.3
0	4.6		21	18.7		50	92.5
2	5.3		22	19.8		60	149.4
4	6.1		23	21.1		70	233.7
6	7.0		24	22.4		80	355.1
8	8.0		25	23.8		90	525.8
10	9.2		26	25.2		95	633.9
12	10.5		27	26.7		99	733.2
14	12.0		28	28.3		100.0	760.0
16	13.6		29	30.0		101.0	787.6

Table 9.2

QUESTION

- Oxygen gas generated by the decomposition of potassium chlorate is collected over water.
- 128 mL of gas is collected at 24°C and an atmospheric pressure of 762 mmHg.
- Calculate the mass (in grams) of oxygen gas obtained.
    - The pressure of the water vapor at 24°C is 22.4 mmHg.

$$PV = nRT = \frac{m}{\mathscr{M}}RT$$

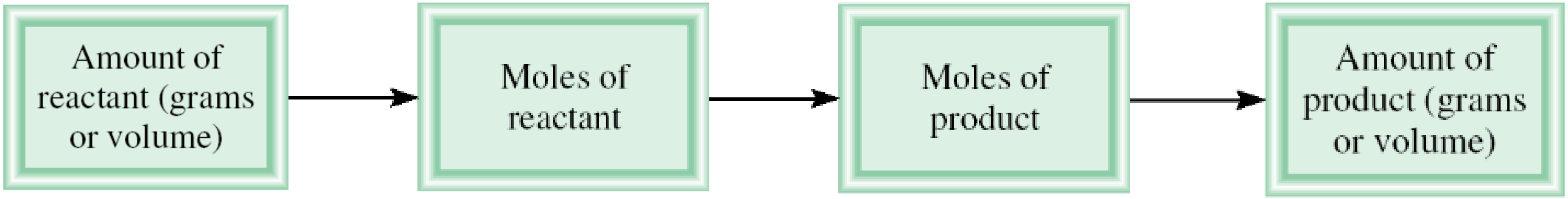
$$\begin{aligned} P_{O_2} &= P_T - P_{H_2O} \\ &= 762 \text{ mmHg} - 22.4 \text{ mmHg} \\ &= 740 \text{ mmHg} \end{aligned}$$

$$m = \frac{PV\mathscr{M}}{RT} = \frac{(740/760)\text{atm}(0.128 \text{ L})(32.00 \text{ g/mol})}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(273 + 24) \text{ K}}$$

$$= 0.164 \text{ g}$$

GAS STOICHIOMETRY

Given a balanced equation and the gas laws, it is possible to determine the amount (mass or volume) of any reactant required or product produced.



Chapter 3 mantra: grams to moles to moles to grams

Now we are able to use V instead of g to start!

QUESTION

Calculate the volume of O_2 (in liters) required for the complete combustion of 7.64 L of acetylene (C_2H_2) measured at the same temperature and pressure.

$$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(l)$$

$$\text{volume of } O_2 = 7.64\ \cancel{\text{L } C_2H_2} \times \frac{5\text{ L } O_2}{2\ \cancel{\text{L } C_2H_2}}$$

$$= 19.1\text{ L}$$

The reaction of calcium carbide (CaC_2) with water produces acetylene (C_2H_2), a flammable gas.

QUESTION

Sodium azide (NaN_3) is used in some automobile air bags. The impact of a collision triggers the decomposition of NaN_3 as follows:

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

The nitrogen gas produced quickly inflates the bag between the driver and the windshield and dashboard.

Calculate the volume of N_2 generated at 80.0°C and 823 mmHg by the decomposition of 60.0 g of NaN_3.

QUESTION 5.12 SOLUTION

First, calculate number of moles of N_2 produced by 60.0 g NaN_3 using the following sequence of conversions

$$\text{grams of NaN}_3 \longrightarrow \text{moles of NaN}_3 \longrightarrow \text{moles of N}_2$$

$$\text{moles of N}_2 = 60.0\ \cancel{\text{g NaN}_3} \times \frac{1\ \cancel{\text{mol NaN}_3}}{65.02\ \cancel{\text{g NaN}_3}} \times \frac{3\ \text{mol N}_2}{2\ \cancel{\text{mol NaN}_3}}$$

$$= 1.38\ \text{mol N}_2$$

The volume of 1.38 moles of N_2 can be obtained by using the ideal gas equation:

$$V = \frac{nRT}{P} = \frac{(1.38\ \text{mol})\,(0.0821\ \text{L} \cdot \text{atm/K} \cdot \text{mol})\,(80 + 273\ \text{K})}{(823/760)\ \text{atm}}$$

$$= 36.9\ \text{L}$$

QUESTION

Aqueous lithium hydroxide solution is used to purify air in spacecrafts and submarines because it absorbs carbon dioxide, which is an end product of metabolism, according to the equation

$$2LiOH(aq) + CO_2(g) \longrightarrow Li_2CO_3(aq) + H_2O(l)$$

The pressure of carbon dioxide inside the cabin of a submarine having a volume of 2.4×10^5 L is 7.9 $\times 10^{-3}$ atm at 312 K. A solution of lithium hydroxide (LiOH) of negligible volume is introduced into the cabin. Eventually the pressure of CO_2 falls to 1.2×10^{-4} atm. How many grams of lithium carbonate are formed by this process?

SOLUTION

The drop in CO_2 pressure is

$$(7.9 \times 10^{-3}\ \text{atm}) - (1.2 \times 10^{-4}\ \text{atm})\ \text{or}\ 7.8 \times 10^{-3}\ \text{atm}$$

Therefore, the number of moles of CO_2 reacted is given by

$$\Delta n = 7.8 \times 10^{-3}\ \text{atm} \times \frac{2.4 \times 10^{5}\ \text{L}}{(0.0821\ \text{L} \cdot \text{atm/K} \cdot \text{mol})\ (312\ \text{K})}$$

$$= 73\ \text{mol}$$

From the chemical equation, the ratio of carbon dioxide to lithium carbonate is

$$1\ \text{mol}\ CO_2 \ :\ 1\ \text{mol}\ Li_2CO_3$$

so the amount of Li_2CO_3 formed is also 73 moles.

$$\text{mass of}\ Li_2CO_3\ \text{formed} = 73\ \cancel{\text{mol}\ Li_2CO_3} \times \frac{73.89\ \text{g}\ Li_2CO_3}{1\ \cancel{\text{mol}\ Li_2CO_3}}$$

$$= 5.4 \times 10^{3}\ \text{g}\ Li_2CO_3$$

CH. 9 OUTLINE

9.1 Gas Pressure

9.2 Relating Pressure, Volume, Amount, and Temperature: The Ideal Gas Law

9.3 Stoichiometry of Gaseous Substances, Mixtures, and Reactions

9.4 Effusion and Diffusion of Gases

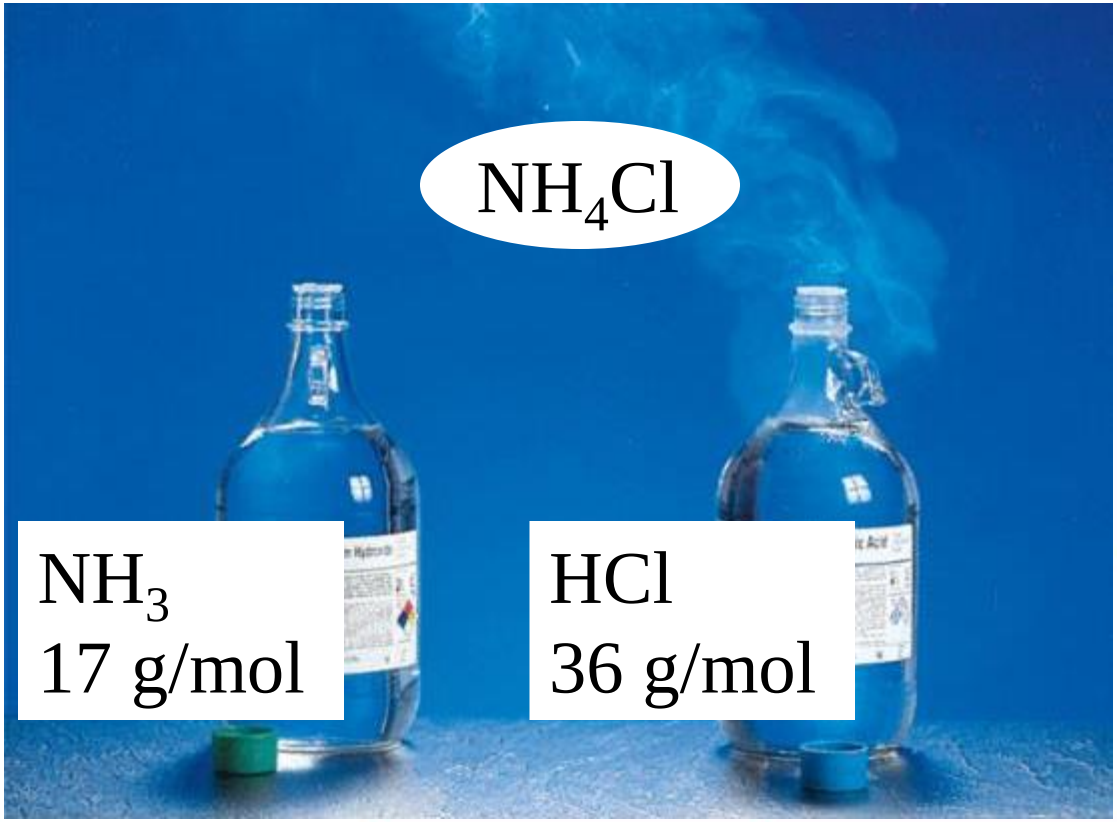
9.5 The Kinetic-Molecular Theory

9.6 Non-Ideal Gas Behavior

DIFFUSION

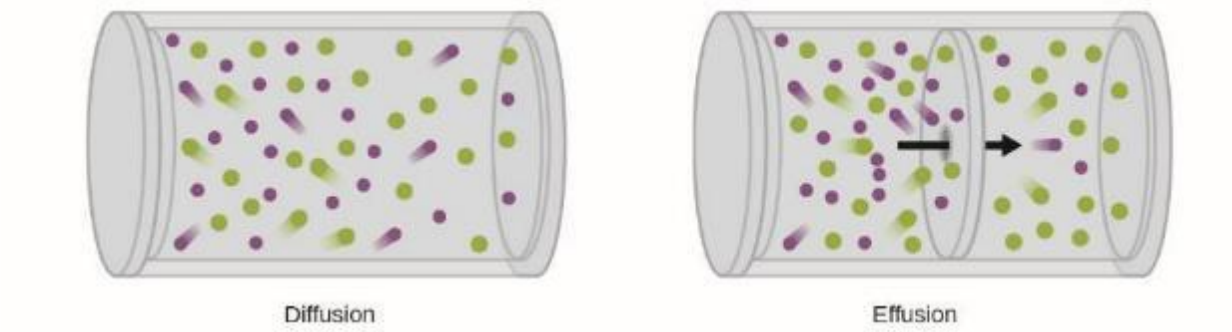
Gas diffusion is the gradual mixing of molecules of one gas with molecules of another by virtue of their kinetic properties.

Lighter is faster.



EFFUSION

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



$$\frac{r_1}{r_2} = \frac{t_2}{t_1} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$

Graham's law of effusion: *The rate of effusion of a gas is inversely proportional to the square root of the mass of its particles*.

EXAMPLE 9.20

Example 9.20

Applying Graham's Law to Rates of Effusion

Calculate the ratio of the rate of effusion of hydrogen to the rate of effusion of oxygen.

Solution

From Graham's law, we have:

$$\frac{\text{rate of effusion of \ ydrogen}}{\text{rate of effusion of \ xygen}} = \frac{\sqrt{1.43\ \cancel{\text{g L}^{-1}}}}{\sqrt{0.0899\ \cancel{\text{g L}^{-1}}}} = \frac{1.20}{0.300} = \frac{4}{1}$$

Using molar masses:

$$\frac{\text{rate of effusion of \ ydrogen}}{\text{rate of effusion of \ xygen}} = \frac{\sqrt{32\ \cancel{\text{g mol}^{-1}}}}{\sqrt{2\ \cancel{\text{g mol}^{-1}}}} = \frac{\sqrt{16}}{\sqrt{1}} = \frac{4}{1}$$

Hydrogen effuses four times as rapidly as oxygen.

Check Your Learning

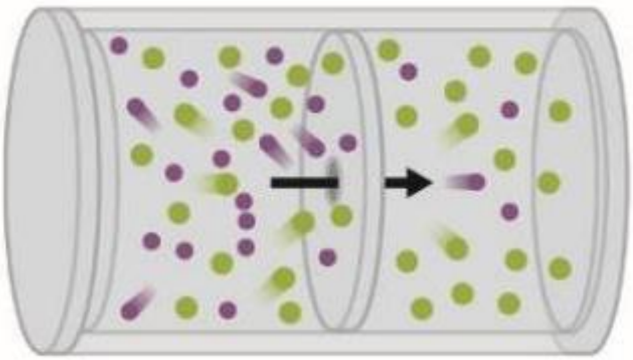
At a particular pressure and temperature, nitrogen gas effuses at the rate of 79 mL/s. Using the same apparatus at the same temperature and pressure, at what rate will sulfur dioxide effuse?

QUESTION

A flammable gas made up only of carbon and hydrogen is found to effuse through a porous barrier in 1.50 min.

Under the same conditions of temperature and pressure, it takes an equal volume of bromine vapor 4.73 min to effuse through the same barrier.

Calculate the molar mass of the unknown gas, and suggest what this gas might be.



Effusion

SOLUTION

From the molar mass of Br_2, we write

$$\frac{1.50\text{ min}}{4.73\text{ min}} = \sqrt{\frac{\mathscr{M}}{159.8\text{ g/mol}}}$$

Where $\mathscr{M}$ is the molar mass of the unknown gas. Solving for $\mathscr{M}$ we obtain

$$\mathscr{M} = \left(\frac{1.50\text{ min}}{4.73\text{ min}}\right)^2 \times 159.8\text{ g/mol}$$
$$= 16.1\text{ g/mol}$$

Because the molar mass of carbon is 12.01 g and that of hydrogen is 1.008 g, the gas is methane (CH_4).

EXAMPLE

Use of Diffusion for Nuclear Energy Applications: Uranium Enrichment

Gaseous diffusion has been used to produce enriched uranium for use in nuclear power plants and weapons. Naturally occurring uranium contains only 0.72% of ^{235}U, the kind of uranium that is "fissile," that is, capable of sustaining a nuclear fission chain reaction. Nuclear reactors require fuel that is 2–5% ^{235}U, and nuclear bombs need even higher concentrations. One way to enrich uranium to the desired levels is to take advantage of Graham's law. In a gaseous diffusion enrichment that is volatile enough to work) is slowly pum contain porous barriers with microscopic openi of the barrier is not evacuated. The $^{235}UF_6$ mol barrier a little faster than the heavier $^{238}UF_6$ mol enriched in $^{235}UF_6$ and the residual gas is slightl $^{235}UF_6$ and $^{238}UF_6$ only about 0.4% enrichment many diffusers in a sequence of stages (called a

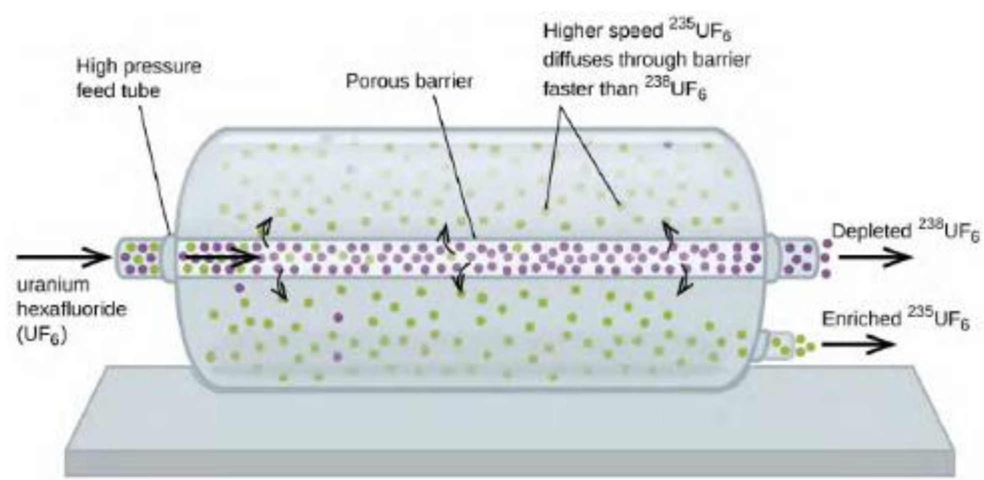


Figure 9.30 In a diffuser, gaseous UF_6 is pumped through a porous barrier, which partially separates $^{235}UF_6$ from $^{238}UF_6$ The UF_6 must pass through many large diffuser units to achieve sufficient enrichment in ^{235}U.

The large scale separation of gaseous $^{235}UF_6$ from $^{238}UF_6$ was first done during the World War II, at the atomic energy installation in Oak Ridge, Tennessee, as part of the Manhattan Project (the development of the first atomic bomb). Although the theory is simple, this required surmounting many daunting technical challenges to make it work in practice. The barrier must have tiny, uniform holes (about 10^{-6} cm in diameter) and be porous enough to produce high flow rates. All materials (the barrier, tubing, surface coatings, lubricants, and gaskets) need to be able to contain, but not react with, the highly reactive and corrosive UF_6.

Because gaseous diffusion plants require very large amounts of energy (to compress the gas to the high pressures required and drive it through the diffuser cascade, to remove the heat produced during compression, and so on), it is now being replaced by gas centrifuge technology, which requires far less energy. A current hot political issue is how to deny this technology to Iran, to prevent it from producing enough enriched uranium for them to use to make nuclear weapons.

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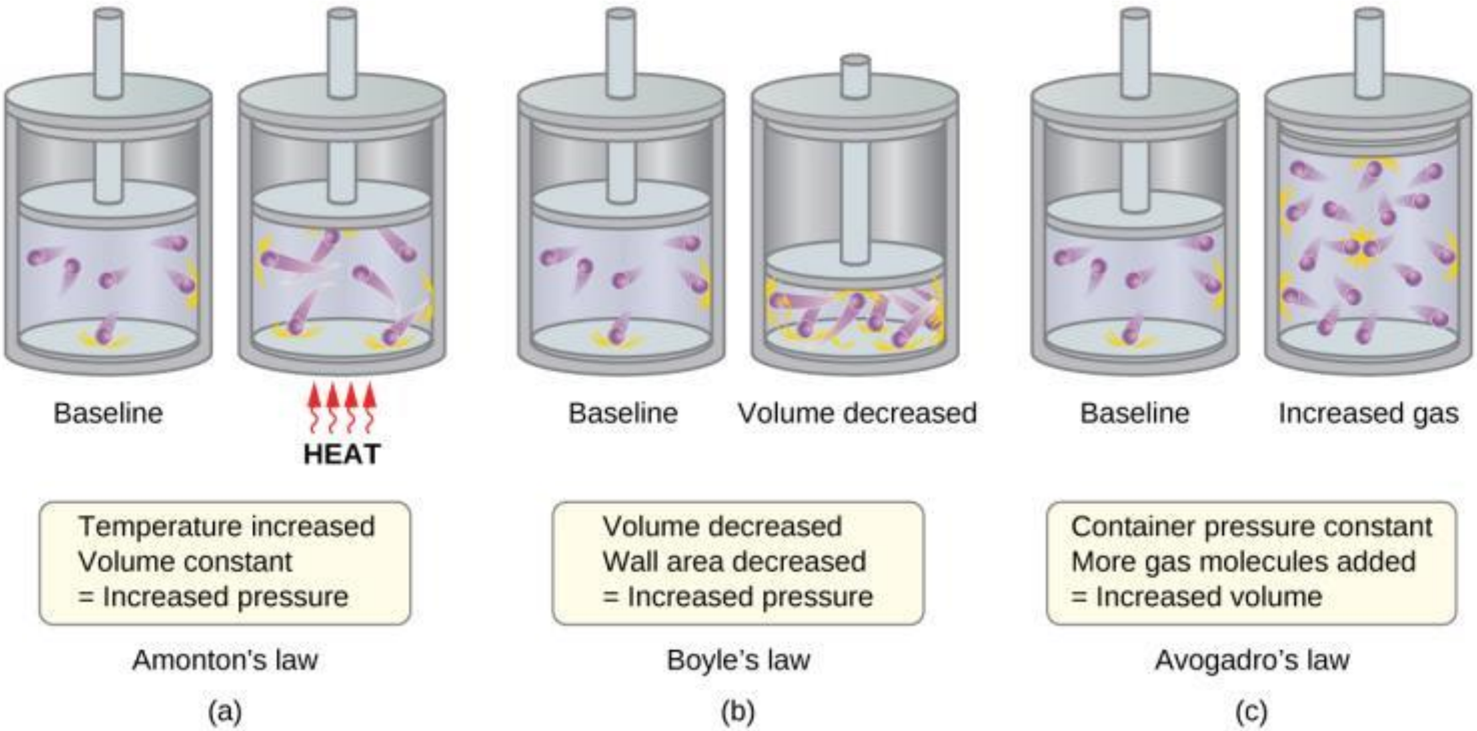
KINETIC-MOLECULAR THEORY (KMT)

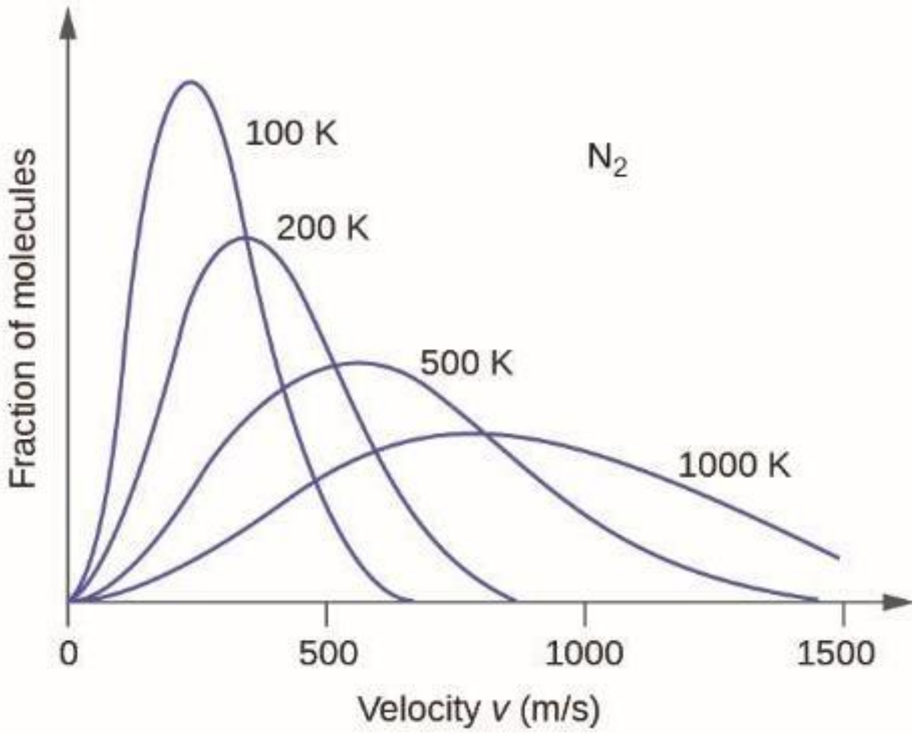
1. Gases are composed of particles that are in continuous motion, travelling in straight lines and changing direction only when they collide with other particles or with the walls of a container.
2. The particles composing the gas are negligibly small compared to the distances between them. (“Gases are point masses.” They possess mass but have negligible volume.)
3. The pressure exerted by a gas in a container results from collisions between the gas molecules and the container walls.
4. Gas molecules exert no attractive or repulsive forces on each other or the container walls; therefore, their collisions are perfectly elastic (do not involve a loss of energy).
5. The average kinetic energy of the gas molecules is proportional to the kelvin temperature of the gas. (“Hotter is faster.”)
    - KE is energy of motion, but is NOT the same as speed...

$$KE = \tfrac{1}{2} m v^2$$

KMT EXPLAINS THE GAS LAWS



MOLECULAR SPEEDS



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

$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{MW}}}$$

The distribution of speeds of five different gases at the same temperature

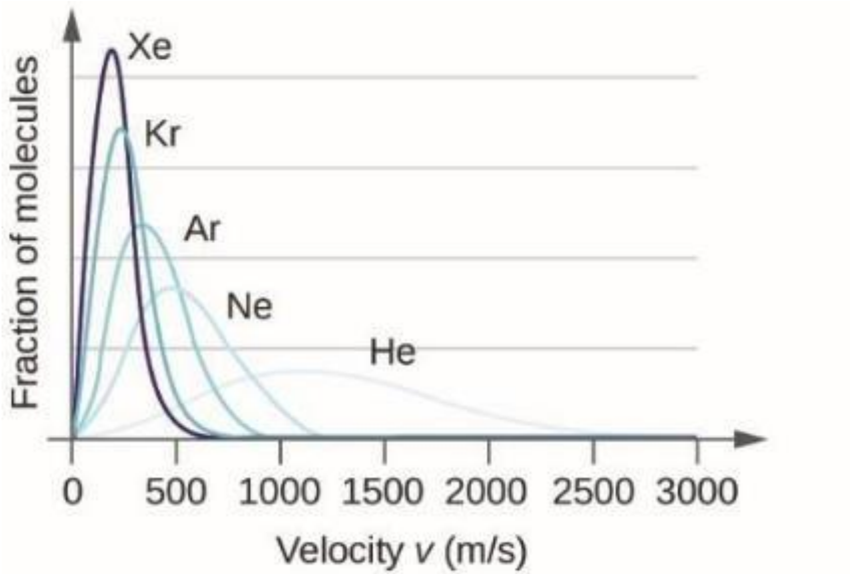
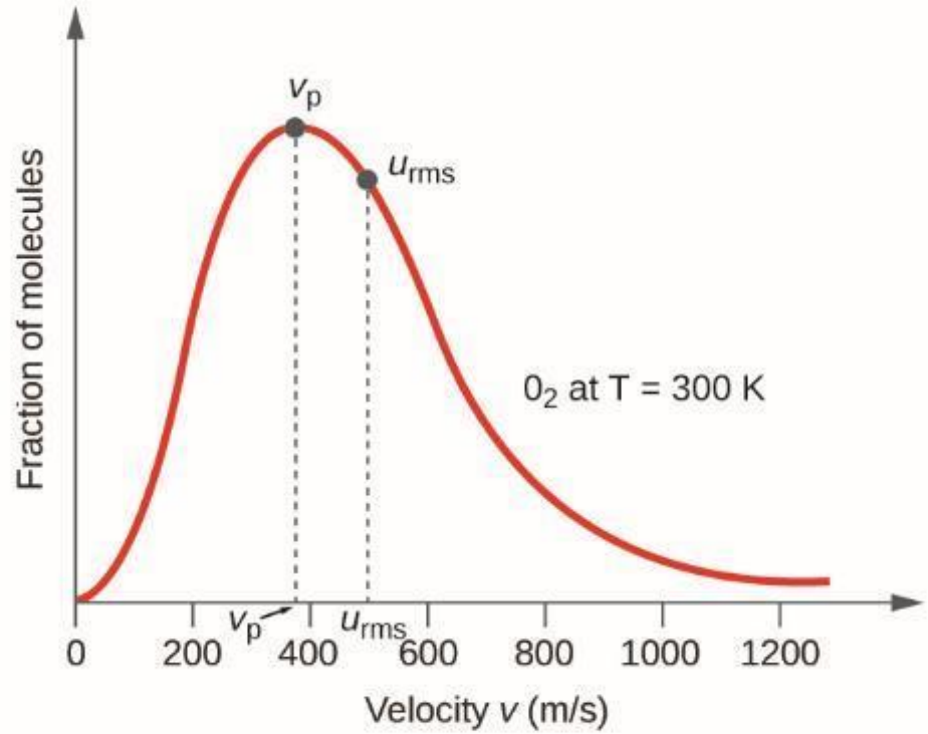


FIGURE 9.32



The molecular speed distribution for oxygen gas at 300 K is shown here. Very few molecules move at either very low or very high speeds. The number of molecules with intermediate speeds increases rapidly up to a maximum, which is the most probable speed, then drops off rapidly. Note that the most probable speed, v_p, is a little less than 400 m/s, while the root mean square speed, u_{rms}, is closer to 500 m/s.

CHEMISTRY 151 - KOLACK

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Chapter 5

Section 5.8

CH. 9 OUTLINE

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9.1 Gas Pressure

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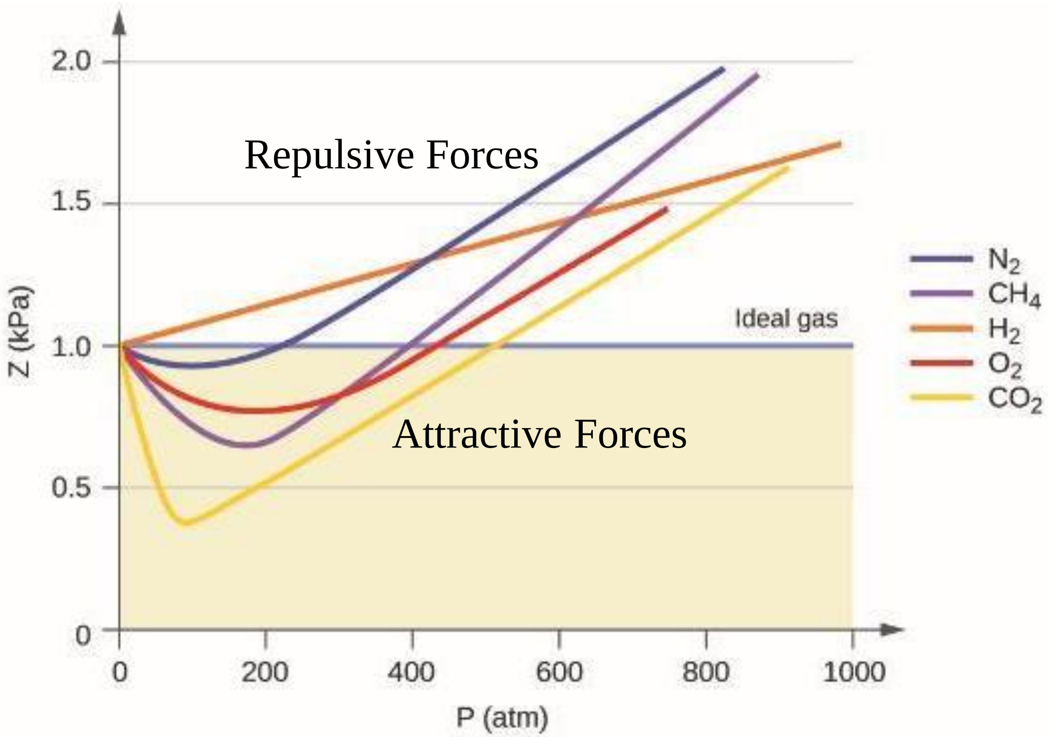
DEVIATIONS FROM IDEAL BEHAVIOR

<u>1 mole of ideal gas</u>

$$PV = nRT$$

$$n = \frac{PV}{RT} = 1.0$$

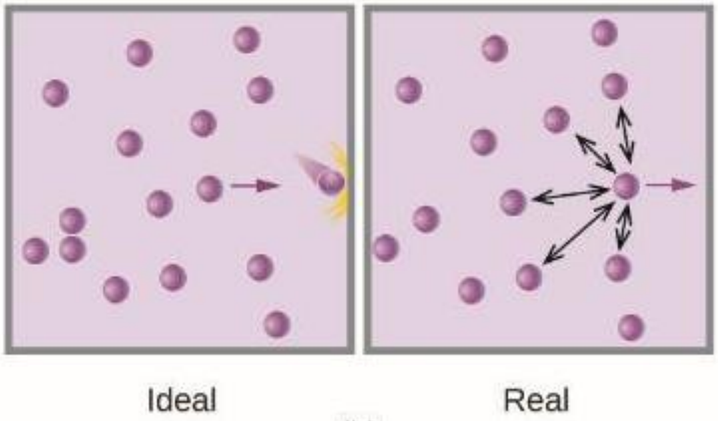
$$Z = \frac{PV}{RT} \text{ measured}$$



A graph of the compressibility factor (Z) vs. pressure shows that gases can exhibit significant deviations from the behavior predicted by the ideal gas law.

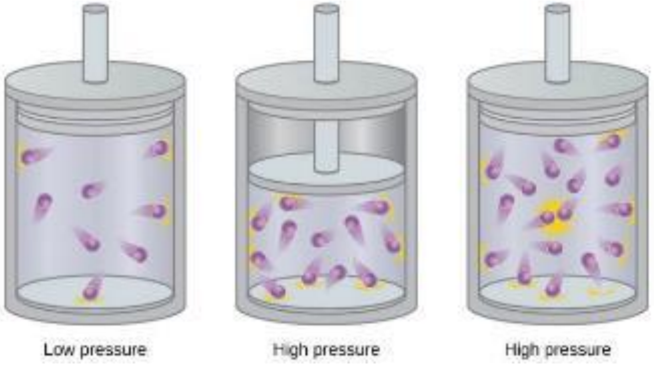
DEVIATIONS (CONT'D)

What would be the effect of intermolecular (attractive) forces on the pressure exerted by a gas?



When would this happen?

Also, at high P (or small V), the volume of the particles begins to matter.



THE REAL GAS EQUATION

Van der Waals equation

nonideal gas

$$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$$

Correction for molecular attraction

Correction for volume of molecules

van der Waals Constants of Some Common Gases

Gas	a $\left(\frac{\text{atm} \cdot \text{L}^2}{\text{mol}^2}\right)$	b $\left(\frac{\text{L}}{\text{mol}}\right)$
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

QUESTION

Given that 3.50 moles of NH_3 occupy 5.20 L at 47°C, calculate the pressure of the gas (in atm) using

(a) the ideal gas equation, and

(b) the van der Waals equation.

SOLUTION

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

$$= \frac{(3.50\ \text{mol})\ (0.0821\ \text{L} \cdot \text{atm/K} \cdot \text{mol})\ (320\ \text{K})}{5.20\ \text{L}}$$

$$= 17.7\ \text{atm}$$

$$\frac{an^2}{V^2} = \frac{(4.17\ \text{atm} \cdot \text{L}^2/\text{mol}^2)\ (3.50\ \text{mol})^2}{(5.20\ \text{L})^2} = 1.89\ \text{atm}$$

$$nb = (3.50\ \text{mol})\ (0.0371\ \text{L/mol}) = 0.130\ \text{L}$$

$$(P + 1.89\ \text{atm})\ (5.20\ \text{L} - 0.130\ \text{L}) = (3.50\ \text{mol})\ (0.0821\ \text{L} \cdot \text{atm/K} \cdot \text{mol})\ (320\ \text{K})$$

$$P = 16.2\ \text{atm}$$

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HW problems:

5, 13, 15, 29, 35, 37, 53, 61, 75, 89, 101, 103