Chapter 5
Gases - 4
Gas Stoichiometry

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Stoichiometry in Gases

Amounts of gaseous reactants and products can be calculated by utilizing

- The ideal gas law to relate moles to T, P and V.
- Moles can be related to mass by the molar mass
- The coefficients in the balanced equation to relate moles of reactants and products

Solved Problem:

When a 2.0-L bottle of concentrated HCl was spilled, 1.2 kg of $CaCO_3$ was required to neutralize the spill. What volume of CO_2 was released by the neutralization at 735 mmHg and 20.°C?

First, write the balanced chemical equation:

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

Second, calculate the moles of CO_2 produced:
Molar mass of $CaCO_3$ = 100.09 g/mol

$$1.2 \times 10^3 \text{ g CaCO}_3 \frac{1 \text{ mol CaCO}_3}{100.09 \text{ g CaCO}_3} \frac{1 \text{ mol CO}_2}{1 \text{ mol CaCO}_3} = 11.99 \text{ mol}$$

n = 11.99 mol
P = 735 mmHg
= 0.967 atm
T = 20° C = 293 K

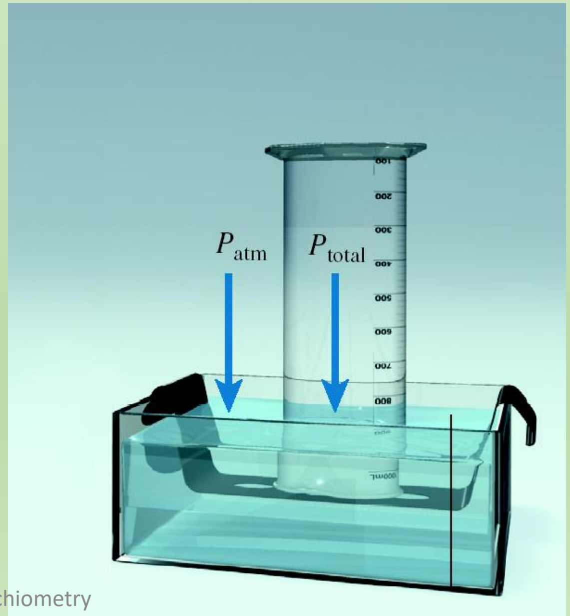
$$V = \frac{nRT}{P}$$

$$V = \frac{(11.99 \text{ mol})\left(0.08206 \frac{\text{L} \bullet \text{atm}}{\text{mol} \bullet \text{K}}\right)(293 \text{ K})}{(0.967 \text{ atm})} = 2.98 \times 10^2 \text{ L}$$

(3 significant figures)

Collecting Gas Over Water

- Gases are often collected over water. The result is a mixture of the gas and water vapor.
- The total pressure is equal to the sum of the gas pressure and the vapor pressure of water.
- The partial pressure of water depends only on temperature and is known (Table 5.6).
- The pressure of the gas can then be found using Dalton's law of partial pressures.



Vapor Pressure of Water at Various Temperatures*

Temperature (°C)	Pressure (mmHg)
0	4.6
10	9.2
15	12.8
17	14.5
19	16.5
21	18.7
23	21.1
25	23.8
27	26.7
30	31.8
40	55.3
60	149.4
80	355.1
100	760.0

$$P = P_{H_2} + P_{H_2O}$$

$$P_{H_2} = P - P_{H_2O}$$

$$P_{H_2} = 769 \text{ mmHg} - 16.5 \text{ mmHg}$$

$$P_{H_2} = 752.5 \text{ mmHg}$$

$$P_{H_2} = 753 \text{ mmHg}$$

(no decimal places)

Solved Problem:

You prepare nitrogen gas by heating ammonium nitrite:

$$NH_4NO_2(s) \rightarrow N_2(g) + 2H_2O(l)$$

If you collected the nitrogen over water at 23°C and 727 mmHg, how many liters of gas would you obtain from 5.68 g NH_4NO_2?

P = 727 mmHg

P_{vapor} = 21.1 mmHg

P_{gas} = 706 mmHg

T = 23° C = 296 K

Molar mass NH_4NO_2 = 64.06 g/mol

$$5.68\ \text{g}\ NH_4NO_2 \frac{1\ \text{mol}\ NH_4NO_2}{64.04\ \text{g}\ NH_4NO_2} \frac{1\ \text{mol}\ N_2}{1\ \text{mol}\ NH_4NO_2} = 0.08887\ \text{mol}\ N_2$$

n = 0.0887 mol

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

$$V = \frac{(0.0887\ \text{mol})\left(0.08206\ \frac{\text{L} \bullet \text{atm}}{\text{mol} \bullet \text{K}}\right)(296\ \text{K})}{\left(706\ \text{mmHg}\ \frac{1\ \text{atm}}{760\ \text{mmHg}}\right)} = 2.32\ \text{L of}\ N_2$$

(3 significant figures)

Solved Problem:

Oxygen was produced and collected over water at 22ºC and a pressure of 754 torr.

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

325 mL of gas were collected and the vapor pressure of water at 22ºC is 21 torr. Calculate the number of moles of O_2 and the mass of $KClO_3$ decomposed.

$$P_{\text{total}} = P_{O_2} + P_{H_2O} = P_{O_2} + 21\text{ torr} = 754\text{ torr}$$

$$P_{O_2} = 754\text{ torr} - 21\text{ torr} = 733\text{ torr} = 733/760\text{ atm}$$

$$V = 325\text{ mL} = 0.325\text{ L}$$

$$T = 22ºC + 273 = 295\text{ K}$$

$$n_{O_2} = \frac{\left(\frac{733}{760}\text{ atm}\right)(0.325\text{ L})}{\left(0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(295\text{ K})} = 1.29\times10^{-2}\text{ mol O}_2$$

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

$$1.29\times10^{-2}\text{ mol O}_2\left(\frac{2\text{ mol KClO}_3}{3\text{ mol O}_2}\right)\left(\frac{122.6\text{ g KClO}_3}{1\text{ mol KClO}_3}\right) =$$

$$= 1.06\text{ g KClO}_3$$

Speed of Gas

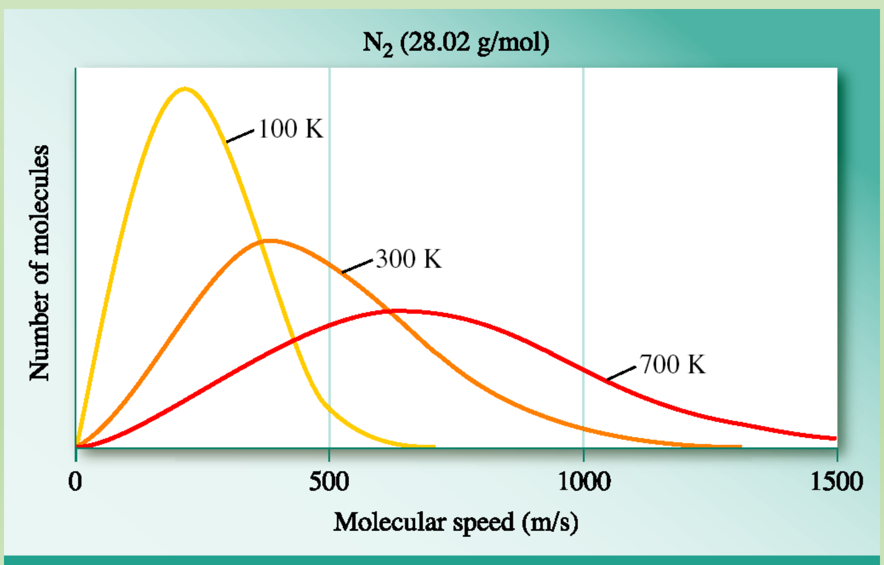
- Root mean square (rms) speed (u_{rms})

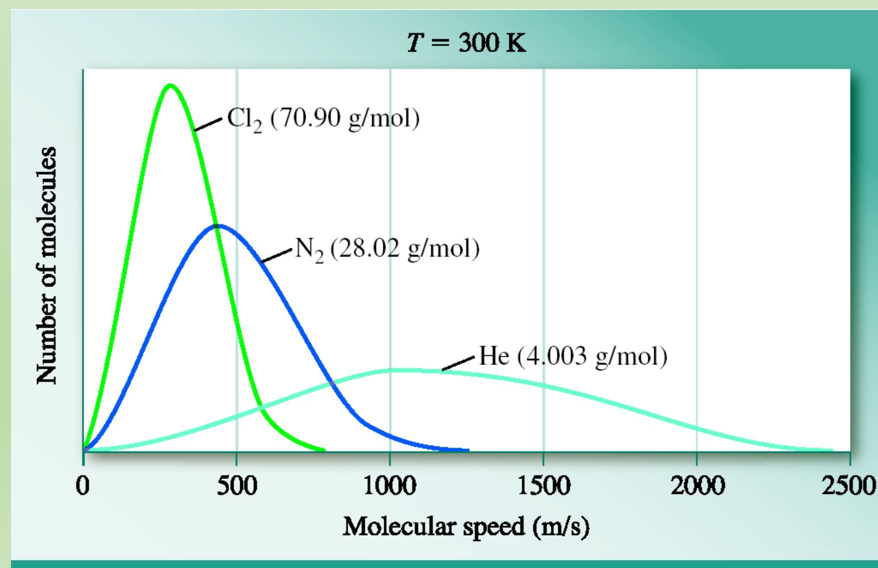
$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}}$$

- For two gases (1 and 2)

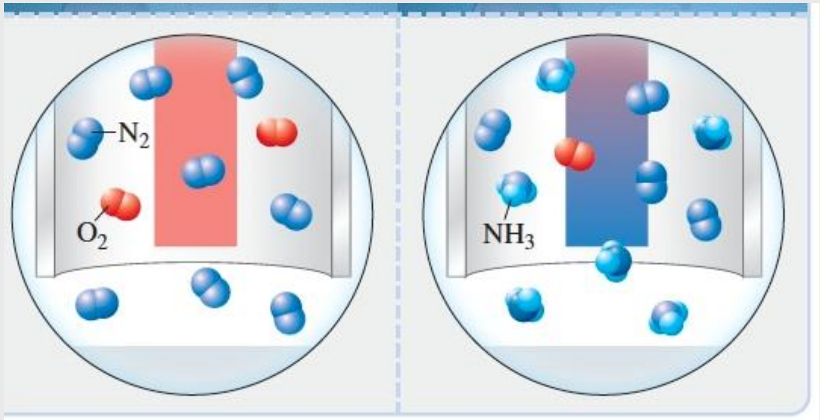
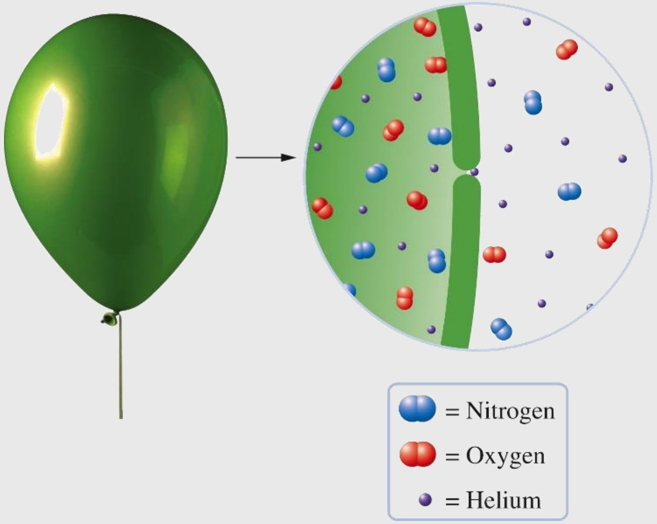
$$\frac{u_{\text{rms}}(1)}{u_{\text{rms}}(2)} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$$

- Effect of Temperature on Molecular Speed (1st graph)
- Effect of Molar Mass on Molecular Speed (2nd graph)



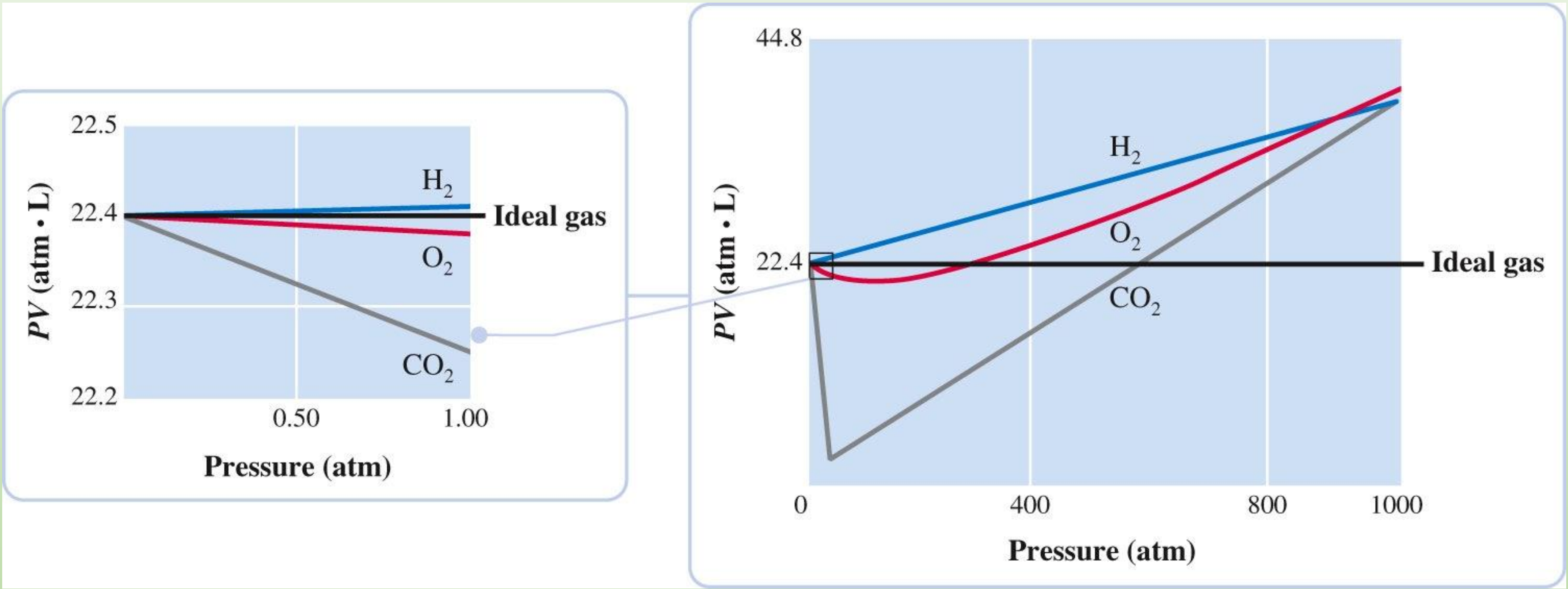


Diffusion and Effusion

Diffusion	Effusion
The process whereby a gas spreads out through another gas to occupy the space uniformly. Below NH_3 diffuses through air. The indicator paper tracks its progress.	The process by which a gas flows through a small hole in a container. A pinprick in a balloon is one example of effusion.
	

Real Gases

At high pressure the relationship between pressure and volume does not follow Boyle's law. This is illustrated on the graph below.



At high pressure, some assumptions of the kinetic theory no longer hold true. At high pressure:

1. the volume of the gas molecule is not negligible.
2. the intermolecular forces are not negligible.

Van der Waal's Equation

An equation that is similar to the ideal gas law, but which includes two constants, *a* and *b*, to account for deviations from ideal behavior.

The term V becomes ($V - nb$) to account for the space between molecules.

The term P becomes ($P + n^2a/V^2$) to account for attraction/repulsion between molecules.

Values for *a* and *b* are different for different gases and can be found in Table 5.7

- The ideal gas law $PV = nRT$

becomes van der Waal's equation

a and *b* have specific values for each gas

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

Key Points

- Gas stoichiometry
- Collecting gas over water
- Gas mixtures
    - Molecular speed
    - Diffusion and effusion
- Deviation from ideal behavior
    - Factors causing deviation
    - Van der Waal's equation