

Recitation, Jan 18

- **Topics:**
 - Ideal Gas Law: Applications
 - Kinetic Molecular Theory
 - Diffusion and Effusion
 - Real Gases
- **Do first:** Q. 2, 3 (sufficient to just solve for partial pressure of Ar), 5 and 7; do Q. 4 if time permits
 - Q. 1 – similar to experiment 12 prelab/report
 - Q. 4 – **key point: note vapor pressure of water**
 - See sample exercise 10.12; will go over
 - Q. 6 – similar to Q. 5 in terms of the theory
 - See sample exercise 10.15
- Handout includes: solutions to Q. 1, 4, and 6
- $R = 0.08206 \text{ L atm/mol K} = 8.314 \text{ J/(mol.K)}$

Q. 1

Chemical analysis of a gaseous compound showed that it contained 33.0% Si and 67.0% F by mass. At 35.00°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.

1. Determine the empirical formula, assuming 100g of material (exactly)

$$\text{number of moles of Si} = \frac{33.0\text{g}}{28.09\text{g/mol}} = 1.175\text{mol} \div 1.175 = 1.000$$

$$\text{number of moles of F} = \frac{67.0\text{g}}{19.00\text{g/mol}} = 3.526\text{mol} \div 1.175 = 3.001$$

Therefore, the empirical formula is SiF_3

2. Find molecular weight using ideal gas law

$$\text{molecular weight} = \frac{mRT}{PV} = \frac{(2.38\text{g}) \left(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (308\text{K})}{(1.70\text{atm})(0.210\text{L})} = 169\text{g/mol}$$

Can show that the empirical formula weight for SiF_3 is 85.09 g

$$169/85.09 = 1.99 \sim 2$$

Therefore, the molecular formula is Si_2F_6

Factor of 2 from the division \rightarrow increase in coefficients by a factor of 2

Q. 2

From the previous problem, and noting that the density $d = m/V$

$$\text{molecular weight} = \frac{dRT}{P} = \frac{(7.71\text{g/L}) \left(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (309\text{K})}{(2.88\text{atm})} = 67.9\text{g/mol}$$

Most chlorine oxides are of the form ClO_n – we will assume this formula.

Note in particular that it can't have more than one Cl since Cl_2 has a formula weight of 71 g/mol which is greater than that measured

Mass of oxygen per mole = $67.9 - 35.5 = 32.4$ g

Therefore, number of moles of oxygen $n = 32.4 / 16.00 = 2.03 \sim 2$

Therefore, we have ClO_2

Q. 3

Recall Dalton's Law of Partial Pressures

$$P_i = \chi_i P_T$$

Partial pressure of component i

Total pressure of system

Mole fraction of component i

$$P_{\text{Ne}} = \frac{4.46}{7.35} \cdot 2.00 \text{ atm} = 1.21 \text{ atm}$$

$$P_{\text{Ar}} = \frac{0.74}{7.35} \cdot 2.00 \text{ atm} = 0.20 \text{ atm}$$

$$P_{\text{Xe}} = \frac{2.15}{7.35} \cdot 2.00 \text{ atm} = 0.585 \text{ atm}$$

Check: sum of all pressures = 2.00 atm = total pressure

Gas Mixtures & Partial Pressures

Collecting Gases over Water

- It is common to synthesize gases and collect them by displacing a volume of water.
- To calculate the amount of gas produced, we need to correct for the partial pressure of the water:

$$P_{\text{total}} = P_{\text{gas}} + P_{\text{water}}$$

Vapor pressure of water



Gas Mixtures & Partial Pressures

Example: Q. 4

The partial pressure of the gas collected is

$$P_{\text{gas}} = 762 - 22.4 = 750 \text{ mmHg} = \frac{750}{760} = 0.987 \text{ atm}$$

Therefore, using the Ideal Gas Equation

$$n = \frac{PV}{RT} = \frac{(0.987 \text{ atm})(0.128 \text{ L})}{\left(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(297.2 \text{ K})} = 5.06 \times 10^{-3} \text{ mol}$$

Therefore, as the **molecular weight = mass/number of moles (important to remember!)**

$$(5.06 \times 10^{-3} \text{ mol}) \times \left(32.00 \frac{\text{g}}{\text{mol}}\right) = 0.162 \text{ g}$$

Q. 5

The root mean square speed is

$$u = \sqrt{\frac{3RT}{M}}$$

$R = 8.314 \text{ J}/(\text{mol}\cdot\text{K})$

Temperature in Kelvins

Molecular weight in **kilograms/mole**

Therefore

$$u = \sqrt{\frac{3(8.315\text{JK}^{-1}\text{mol}^{-1})(298\text{K})}{(4.003 \times 10^{-3}\text{kg}\cdot\text{mol}^{-1})}} = 1.36 \times 10^3 \text{ms}^{-1}$$

Q. 6

Graham's Law of Molecular Effusion

$$rate \propto \frac{1}{\sqrt{M}}$$

Hence

$$\frac{rate_{\text{unknown}}}{rate_{\text{Br}_2}} = \sqrt{\frac{M_{\text{Br}_2}}{M_{\text{unknown}}}}$$

Therefore

$$M_{\text{unknown}} = M_{\text{Br}_2} \cdot \left(\frac{rate_{\text{Br}_2}}{rate_{\text{unknown}}} \right)^2 = 159.8 \cdot \left(\frac{(1/4.73)}{(1/1.50)} \right)^2 = 16.1 \text{g/mol}$$

Q. 7

(a) Ideal Gas Law

$$P = \frac{nRT}{V} = \frac{(3.50\text{mol}) \left(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (320.2\text{K})}{(5.20\text{L})} = 17.7\text{atm}$$

(b) Van der Waals equation

$$P = \frac{nRT}{V - nb} - \frac{n^2 a}{V^2} = \frac{(3.50\text{mol}) \left(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right) (320.2\text{K})}{(5.20\text{L}) - (3.50\text{mol})(0.0371\text{L/mol})} - \frac{(3.50\text{mol})^2 \left(4.170 \frac{\text{L}^2\text{atm}}{\text{mol}^2} \right)}{(5.20\text{L})^2} = 16.2\text{atm}$$