Announcements

Mid Term Exam

When and Where

• Saturday, November 3, 7:00 – 9:00 PM. Since we will start setting up about half an hour prior to the exam, please remain outside of the rooms, as they cannot be used for studying while we set up.

• Assigned rooms are based on the lecture section of registration and the **last five digits of the student number**. You ***must*** write in the assigned room, as space is limited. Those who arrive at the incorrect room will be asked to leave.

Section 003 (Lipson)	Room
04620 – 14942	SH 2355
15176 – 21129	SH 3315
21169 – 27343	SH 3317
27350 – 40575	TH 3101
40576 – 99554	TH 3102

Section 006 (Lipson)	Room
01806 – 15307	HSB 236
15315 – 27190	HSB 240
27250 – 98264	HSB 35

TH = Thames Hall
HSB = Health Sciences Building
SH = Sommerville House

Exam Content

- Format: 25 multiple-choice questions of equal value. There is no penalty for wrong answers, so it is to your advantage to make an informed guess if you cannot answer a question.

- Two questions will be derived from Experiments A and B (Synthesis and Acid/Base Titration, respectively). These questions will be based on lab theory or procedure. Essentially, if you have used a procedure, calculation, apparatus, or reagent, it is expected that you know what it does, how it works, why you used it, etc. No lab questions will be based directly on your own data, so the lab reports are not needed to study for the test.

- The remaining 23 questions will be derived from material covered in class and in the tutorial manual. The *approximate* question breakdown is provided below

Topic	Questions
Fundamentals and Stoichiometry	10
Strong Acids and Bases	8
Redox	5

Full details at: https://instruct.uwo.ca/chemistry/020/

Example *(dealing with partial pressures)*:

What is the partial pressure of O_2 in 2.00 g of air in a 1 L container at 20°C?

Useful information: mole % air = 21.0% O_2, 78.0% N_2, 1.00% Ar; average molar mass of air = 29.0 g mol^{-1}

Total air pressure calculated from: $$P_{TOT} = \frac{n_{TOT}RT}{V}$$

∴ P_{TOT} = (2.00 g/29.0 g mol^{-1})(0.0831 L-bar K^{-1}mol^{-1})(293 K)/(1.00 L) = 1.68 bar

Since $P_{O_2} = X_{O_2} P_{TOT}$ and $X_{O_2} = \frac{21.0}{100}$

$$P_{O_2} = \left(\frac{21}{100}\right)(1.68) = \boxed{0.353 \text{ bar} = 35.3 \text{ kPa}}$$

Example: A steel flask of volume 23.0 L contains a gas mixture consisting of 0.475 g H_2 and an unknown amount of He. If the total pressure of the mixture is 100 kPa at 29 °C, what is the partial pressure of He in the mixture?

number of moles of H_2 = n = 0.475 g/2.02 g mol^{-1} = 0.235 mol

$$\therefore P_{H_2} = \frac{nRT}{V} = \frac{(0.235\,\text{mol})(8.314\,\text{kPa L mol}^{-1}\,\text{K}^{-1})(302\,\text{K})}{(23.0\,\text{L})}$$

= 2.56 kPa

$$\therefore P_{He} = P_{TOT} - P_{H_2} = 100 - 2.56 = \boxed{97.44\,\text{kPa}}$$

Example: A gas mixture at 27 ^oC contains methane (CH_4) at 70.0 kPa and ethane (C_2H_6) at 35.0 kPa. What is the average molecular mass of the gas mixture?

$$P_{TOT} = 70.0 + 35.0 = 105.0 \text{ kPa}$$

$$\therefore X_{CH_4} = \frac{70.0}{105.0} = 0.667 \quad \text{and} \quad X_{C_2H_6} = 1 - 0.667 = 0.333$$

$$\bar{M} = X_{CH_4} M(CH_4) + X_{C_2H_6} M(C_2H_6)$$

$$=(0.667)(16.05) + (0.333)(30.08)$$

$$= \boxed{20.72 \text{ g mol}^{-1}}$$

Example:

The density of a mixture of N_2 gas and Cl_2 gas is 1.58 g L^{-1} at 25^{o}C and 1 bar total pressure. What is the partial pressure of N_2?

In this problem the correct form of PV = nRT to use is: $P(\overline{M}) = dRT$

where $\overline{M} = X_{N_2} M_{N_2} + X_{Cl_2} M_{Cl_2}$

In a mixture with n components: $\overline{M} = X_1 M_1 + X_2 M_2 + \cdots + X_n M_n$

$$\therefore \quad \overline{M} = \frac{dRT}{P}$$ = (1.58)(0.0831)(298)/(1.0) = 39.1 g mol^{-1}

In this problem, since $X_{N_2} + X_{Cl_2} = 1 \quad \therefore \quad X_{Cl_2} = 1 - X_{N_2}$

$$\therefore \quad \overline{M} = X_{N_2} M_{N_2} + (1 - X_{N_2}) M_{Cl_2}$$

$$\therefore \quad 39.1 = X_{N_2}(28.0) + (1 - X_{N_2})(70.8)$$

$$= 70.8 - 42.8 X_{N_2}$$

Solve for $X_{N_2} = 0.74$

$\therefore \quad P_{N_2} = X_{N_2} P_{TOT}$ = (0.74)(1.0 bar) = 0.74 bar

Gases and Stoichiometry

Volume measurements can be converted to mass measurements for chemical reactions

Example:

What mass of H_2O can be produced by the complete reaction of 1.68 L O_2 and 2.90 L H_2, all at 25^oC and 1 bar?

Begin with a balanced equation:

	$2H_2$ +	O_2 →	$2H_2O$	
ideal masses:	4.04	32	36	
Ideal volume:	2	1	2	= # moles

Why?

$$V = n\left(\frac{RT}{P}\right) = nk$$

at fixed T, P k = constant

Determine the limiting reagent L.R.

Ideal volume ratio: $\frac{H_2}{O_2} = \frac{2}{1}$ Actual: $\frac{H_2}{O_2} = \frac{2.90}{1.68} = \frac{1.73}{1}$

$\therefore$ L. R. = H_2

$\therefore$ mass of H_2 consumed: $w = \frac{MPV}{RT} = \frac{(2.02)(1)(2.90)}{(0.0821)(298)}$

= 0.237 g H_2

Now calculate the mass of H_2O produced:

mass $\frac{H_2}{H_2O}$ Ideal: $\frac{4.04}{36.0}$ Actual: $\frac{0.237}{x}$

$\boxed{x = 2.11 \text{ g } H_2O}$

Example: A 2.00 L vessel contains 1.40 g of propane (C_3H_8) at 300 K. Exactly enough oxygen is added for a complete reaction forming water and carbon dioxide.

a) What is the initial pressure at 300 K before the reaction?

First write a balanced equation for this combustion reaction:

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

$M(C_3H_8) = 44.11 \text{ g mol}^{-1}$ $\therefore$ $$n_{C_3H_8} = \frac{1.40}{44.11} = 0.032\,\text{mol}$$

By stoichiometry: $n_{O_2} = 5 \text{ x } n_{C_3H_8} = 0.159\,\text{mol}$

$n_{TOT} = 0.032 + 0.159 = 0.191 \text{ mol}$

$$\therefore P_{TOT} = \frac{n_{TOT}RT}{V} = \frac{(0.191\,\text{mol})(8.314\,\text{kPa L mol}^{-1}\,\text{K}^{-1})(300\,\text{K})}{(2.00\,\text{L})}$$

$$\boxed{= 238.2 \text{ kPa}}$$

b) After the reaction is complete, the temperature has risen to 900 K. What is the final pressure?

By stoichiometry: 0.032 mol of propane produces 3 x 0.032 = 0.096 mol CO_2 and 4 x 0.032 = 0.128 mol H_2O. The total number of moles, n_{TOT}, = 0.224 mol.

$$\therefore P_{final} = \frac{n_{TOT}RT}{V} = \frac{(0.224\,\text{mol})(8.314\,\text{kPa L mol}^{-1}\text{K}^{-1})(900\,\text{K})}{(2.00\,\text{L})}$$

$$\boxed{= 838.1\ \text{kPa}}$$

Another limiting reagent **example:** A mixture contains 50 kPa methane (CH_4) and 80 kPa O_2. The mixture combusts. What pressure of CO_2 and H_2O are produced if measured at the original temperature? What remains unreacted?

First, as always, write a balance reaction:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

At fixed V and T, P is proportional to n.

For complete combustion, by stoichiometry, 50 kPa of CH_4 requires 100 kPa of O_2. Therefore O_2 is the L. R.

Therefore, 80 kPa will produce 40 kPa of CO_2 and 80 kPa of H_2O.

More so, 80 kPa of O_2 will react with 40 kPa of CH_4.
Therefore 50 – 40 = 10 kPa of CH_4 will remain unreacted.