

# 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.

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

<b>Section 006 (Lipson)</b>	<b>Room</b>
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<sub>2</sub> in 2.00 g of air in a 1 L container at 20°C?

Useful information: mole % air = 21.0% O<sub>2</sub>, 78.0% N<sub>2</sub>, 1.00% Ar; average molar mass of air = 29.0 g mol<sup>-1</sup>

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

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

Since 
$$P_{\text{O}_2} = X_{\text{O}_2} P_{\text{TOT}} \quad \text{and} \quad X_{\text{O}_2} = \frac{21.0}{100}$$

$$P_{\text{O}_2} = \left(\frac{21}{100}\right)(1.68) = 0.353 \text{ bar} = 35.3 \text{ kPa}$$

**Example:** A steel flask of volume 23.0 L contains a gas mixture consisting of 0.475 g H<sub>2</sub> 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<sub>2</sub> = n = 0.475 g / 2.02 g mol<sup>-1</sup> = 0.235 mol

$$\therefore P_{\text{H}_2} = \frac{nRT}{V} = \frac{(0.235 \text{ mol})(8.314 \text{ kPa L mol}^{-1} \text{ K}^{-1})(302 \text{ K})}{(23.0 \text{ L})}$$

$$= 2.56 \text{ kPa}$$

$$\therefore P_{\text{He}} = P_{\text{TOT}} - P_{\text{H}_2} = 100 - 2.56 = 97.44 \text{ kPa}$$

**Example:** A gas mixture at 27 °C contains methane (CH<sub>4</sub>) at 70.0 kPa and ethane (C<sub>2</sub>H<sub>6</sub>) at 35.0 kPa. What is the average molecular mass of the gas mixture?

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

$$\therefore X_{\text{CH}_4} = \frac{70.0}{105.0} = 0.667 \quad \text{and} \quad X_{\text{C}_2\text{H}_6} = 1 - 0.667 = 0.333$$

$$\begin{aligned} \bar{M} &= X_{\text{CH}_4} M(\text{CH}_4) + X_{\text{C}_2\text{H}_6} M(\text{C}_2\text{H}_6) \\ &= (0.667)(16.05) + (0.333)(30.08) \end{aligned}$$

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

### Example:

The density of a mixture of  $\text{N}_2$  gas and  $\text{Cl}_2$  gas is  $1.58 \text{ g L}^{-1}$  at  $25^\circ\text{C}$  and 1 bar total pressure. What is the partial pressure of  $\text{N}_2$ ?

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

$$\text{where } \bar{M} = X_{\text{N}_2} M_{\text{N}_2} + X_{\text{Cl}_2} M_{\text{Cl}_2}$$

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

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

In this problem, since  $X_{\text{N}_2} + X_{\text{Cl}_2} = 1$   $\therefore X_{\text{Cl}_2} = 1 - X_{\text{N}_2}$

$$\therefore \bar{M} = X_{\text{N}_2} M_{\text{N}_2} + (1 - X_{\text{N}_2}) M_{\text{Cl}_2}$$

$$\begin{aligned} \therefore 39.1 &= X_{\text{N}_2} (28.0) + (1 - X_{\text{N}_2}) (70.8) \\ &= 70.8 - 42.8 X_{\text{N}_2} \end{aligned}$$



Solve for  $X_{N_2} = 0.74$

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

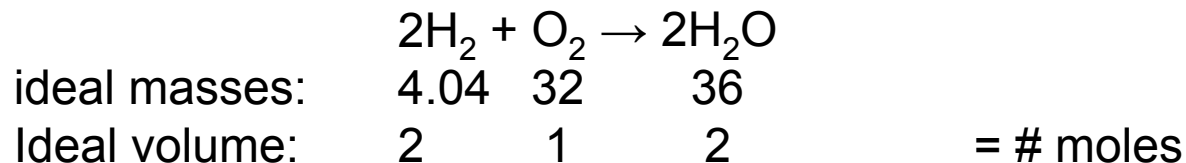
# Gases and Stoichiometry

Volume measurements can be converted to mass measurements for chemical reactions

## Example:

What mass of H<sub>2</sub>O can be produced by the complete reaction of 1.68 L O<sub>2</sub> and 2.90 L H<sub>2</sub>, all at 25°C and 1 bar?

Begin with a balanced equation:



**Why?**  $V = n \left( \frac{RT}{P} \right) = nk$  at fixed T, P k = constant

Determine the limiting reagent L.R.

Ideal volume ratio:  $\frac{\text{H}_2}{\text{O}_2} = \frac{2}{1}$       Actual:  $\frac{\text{H}_2}{\text{O}_2} = \frac{2.90}{1.68} = \frac{1.73}{1}$

∴ L. R. =  $\text{H}_2$

∴ mass of  $\text{H}_2$  consumed:  $w = \frac{\text{MPV}}{\text{RT}} = \frac{(2.02)(1)(2.90)}{(0.0821)(298)}$   
= 0.237 g  $\text{H}_2$

Now calculate the mass of  $\text{H}_2\text{O}$  produced:

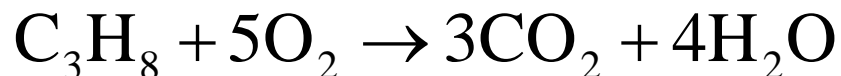
mass  $\frac{\text{H}_2}{\text{H}_2\text{O}}$       Ideal:  $\frac{4.04}{36.0}$       Actual:  $\frac{0.237}{x}$

$x = 2.11 \text{ g H}_2\text{O}$

**Example:** A 2.00 L vessel contains 1.40 g of propane (C<sub>3</sub>H<sub>8</sub>) 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:



$$M(\text{C}_3\text{H}_8) = 44.11 \text{ g mol}^{-1} \quad \therefore \quad n_{\text{C}_3\text{H}_8} = \frac{1.40}{44.11} = 0.032 \text{ mol}$$

$$\text{By stoichiometry: } n_{\text{O}_2} = 5 \times n_{\text{C}_3\text{H}_8} = 0.159 \text{ mol}$$

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

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

$$= 238.2 \text{ kPa}$$

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

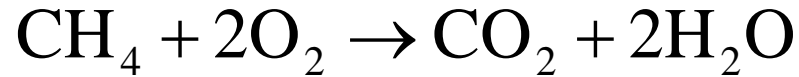
By stoichiometry: 0.032 mol of propane produces  $3 \times 0.032 = 0.096$  mol  $\text{CO}_2$   
and  $4 \times 0.032 = 0.128$  mol  $\text{H}_2\text{O}$ . The total number of moles,  $n_{\text{TOT}}$ , = 0.224 mol.

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

$$= 838.1 \text{ kPa}$$

Another limiting reagent **example**: A mixture contains 50 kPa methane (CH<sub>4</sub>) and 80 kPa O<sub>2</sub>. The mixture combusts. What pressure of CO<sub>2</sub> and H<sub>2</sub>O are produced if measured at the original temperature? What remains unreacted?

First, as always, write a balance reaction:



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

For complete combustion, by stoichiometry, 50 kPa of CH<sub>4</sub> requires 100 kPa of O<sub>2</sub>. Therefore O<sub>2</sub> is the L. R.

Therefore, 80 kPa will produce 40 kPa of CO<sub>2</sub> and 80 kPa of H<sub>2</sub>O.

More so, 80 kPa of O<sub>2</sub> will react with 40 kPa of CH<sub>4</sub>.  
Therefore 50 – 40 = 10 kPa of CH<sub>4</sub> will remain unreacted.