What are 4 Assumptions of Ideal Gases

- Gases move with high Kinetic Energy (Temp is proportional to its KE)
- Gas molecules have negligible volume
- Gas molecules exert no attraction
- Gas molecules exhibit elastic collisions (do not lose energy)

What are three cases in which ideal gases begin to deviate to be Real Gases

- Low Temperature
- High Pressure
- High Molar Mass
Sketch Your Prediction and Give the Relationship

**Boyle's Law**

\[ P_1 V_1 = P_2 V_2 \]

**Volume**

Oct 24-8:12 AM

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Sketch Your Prediction and Give the Relationship

**Charles's Law**

\[ \frac{V_1}{T_1} = \frac{V_2}{T_2} \]

**Volume**

**Temp (in K)**

Oct 24-8:12 AM

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Sketch Your Prediction and Give the Relationship

**Gay-Lussac's Law**

\[ \frac{P_1}{T_1} = \frac{P_2}{T_2} \]

**Pressure**

**Temp (in K)**

Oct 24-8:12 AM
1) The most plentiful gas in the earth's atmosphere.
2) A 1 mole sample of this gas occupying 1 liter will have the greatest density.
3) At the same temperature, this gas will have the greatest speed.
4) The molecules of this nonpolar gas contain polar bonds.
5) The molecules of this gas contain 1 sigma and 2 pi bonds.

What will happen to the gas pressure if:

(A) The volume is halved.
(B) The volume is quadrupled.
(C) The absolute temperature is doubled.
(D) The volume is doubled and the absolute temperature is halved.
(E) The number of moles of the gas is doubled.

A container with a volume of 500 mL has 56g of nitrogen gas (N₂) at 0°C.

(A) How many moles of nitrogen gas are there?
(B) What is the gas pressure at this temp and volume?
(C) If the volume doubles, what is the pressure?
(D) What volume will this amount of gas be at standard temperature and pressure?
Hydrogen gas is placed in a container with oxygen gas. If 250 mL of hydrogen gas at a pressure of 0.80 atm and 25°C reacts with excess oxygen gas:

(A) How many moles of water will be produced?
(B) What mass of water will be produced?
(C) How many moles of oxygen gas will react with this hydrogen gas?

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(A) What is the total pressure in the container after they mix?
(B) What is the partial pressure of Ar gas?
(C) What is the partial pressure of Kr gas?

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(A) What is the partial pressure of He?
(B) What is the partial pressure of Ne?
2. A rigid and sealed cylinder contains 0.250 mol of helium gas, 0.500 mol of neon gas, and 0.150 mol of nitrogen gas. The total pressure inside the cylinder is 9.0 atm. Helium has a partial pressure of:
   (A) 0.250 atm  
   (B) 0.500 atm  
   (C) 2.60 atm  
   (D) 2.50 atm  
   (E) 5.00 atm

7. An ideal gas is heated from 20°C to 75°C in a sealed container at constant volume. Which of the following statements are true?
   I. The partial pressure of the gas increases.
   II. The average distance between the gas particles decreases.
   III. The average kinetic energy of the gas particles increases.
   (A) I only  
   (B) I and II only  
   (C) I, II, and III  
   (D) I and III only  
   (E) III only

1. A rigid container holds nitrogen gas at 3.70 atm and 25°C. Additional nitrogen is added until the pressure inside the tank reaches 5.00 atm. The tank remains at 25°C throughout this process. Which of the following statements are true?
   I. The volume of the gas increased.
   II. The average speed of the nitrogen molecules increased.
   III. The partial pressure of the nitrogen gas increased.
   (A) I only  
   (B) II only  
   (C) III only  
   (D) II and III only  
   (E) I, II, and III

3. A rigid and sealed cylinder contains 0.300 mol of helium gas, 0.500 mol of neon gas, and 0.400 mol of nitrogen gas. The total pressure inside the cylinder is 12.00 atm. Ne has a partial pressure of:
   (A) 0.400 atm  
   (B) 4.00 atm  
   (C) 3.00 atm  
   (D) 2.50 atm  
   (E) 5.00 atm

Order in increasing rate of effusion

C\textsubscript{2}H\textsubscript{6}    H\textsubscript{2}    Cl\textsubscript{2}    Ne
1) Which two gases would be the most difficult to separate from each other?

\[ \text{C}_2\text{H}_4 \quad \text{He} \quad \text{F}_2 \quad \text{N}_2 \quad \text{CO}_2 \]

2) A gas mixture with a total pressure of 2.0 atm is made up of He and Ne gases. There is three times as much He as there is Ne. What is the partial pressure of He?

3) \( \text{H}_2 \) gas is collected over water at 29°C. The atm pressure is 760mmHg and the vapor pressure of \( \text{H}_2\text{O} \) is 30mmHg at this temp. What is the partial pressure of the \( \text{H}_2 \) gas?

1) At what two conditions will ideal gases deviate?

2) A gas at 273K and 1.5 atm has a density of 1.87 g L\(^{-1}\). What is this gas’ molar mass?

3) A gas sample contains 0.4 moles of He, 0.4 moles of Ne, and 0.8 moles of Ar. The total pressure is 2.4 atm. What is the partial pressure of each of the three gases?

C\(_3\)H\(_8\)  
25°C  
1.2 atm  
1.5 L  

CO\(_2\)  
25°C  
1.5 atm  
1.5 L  

Compare:  
(a) # of molecules  
(b) density  
(c) avg. Kinetic Energy  
(d) rms speed  
(e) rate of effusion
Kinetic Energy deals with _______

Effusion deals with _______

*rms speed* deals with _______ & _______

---

A different compound, which has the empirical formula C\textsubscript{3}H\textsubscript{8}Br, has a vapor density of 6.00 g L\textsuperscript{-1} at 373 K and 3005 atm. Using these data, determine the following.

(i) The molar mass of the compound

(ii) The molecular formula of the compound

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2005 Form B
A rigid 5.00 L cylinder contains 24.5 g of \( \text{N}_2(g) \) and 20.0 g of \( \text{O}_2(g) \).

(a) Calculate the total pressure, in atm, of the gas mixture in the cylinder at 298 K.

(b) The temperature of the gas mixture in the cylinder is decreased to 280 K. Calculate each of the following:
   (i) The mole fraction of \( \text{N}_2(g) \) in the cylinder
   (ii) The partial pressure, in atm, of \( \text{N}_2(g) \) in the cylinder

(c) If the cylinder develops a pinhole-sized leak and some of the gaseous mixture escapes, would the ratio of \( \text{N}_2(g) \) to \( \text{O}_2(g) \) in the cylinder increase, decrease, or remain the same? Justify your answer.

A different rigid 5.00 L cylinder contains 0.176 mol of \( \text{NO}(g) \) at 298 K. A 0.476 mol sample of \( \text{O}_2(g) \) is added to the cylinder, where a reaction occurs to produce \( \text{NO}_2(g) \).

(d) Write the balanced equation for the reaction.

(e) Calculate the total pressure, in atm, in the cylinder at 298 K after the reaction is complete.

Cumulative MC Problems
Finding R by Collecting a gas over water
Using magnesium with hydrochloric acid, will collect a gas over water to calculate R.

1) Write a balanced net ionic reaction.
2) What 4 things do you need to measure?
3) What is one constant you need from me?
4) What are 3 calculations you must do to find R? Setup those calculations.

Nov 3-8:09 AM

3 cm of Mg
1.8 g = 100 cm
20 mL of 6.0 M HCl and fill the grad cylinder with water.
Invert and measure the amount of gas produced.
**Calculate R.
**Find % error.
**What could have attributed to error?

Nov 3-8:11 AM

Goal: To design an experiment to determine the factors that affect the rate of a reaction.

Materials: CaCO₃ (solid) and HCl (aq)

WHITEBOARD

Factors to Investigate:

Develop a Systematic Procedure of Testing:

Nov 27-7:41 AM
Your responses to the rest of the questions in this part of the examination will be graded on the basis of the accuracy and relevance of the information cited. Explanations should be clear and well organized. Examples and equations may be included in your responses where appropriate. Specific answers are preferable to broad, diffuse responses.

Answer BOTH Question 5 below AND Question 6 printed on the next page. Both of these questions will be graded. The Section II score weighting for these questions is 30 percent (15 percent each).

5. A student performs an experiment to determine the molar mass of an unknown gas. A small amount of the pure gas is released from a pressurized container and collected in a graduated tube over water at room temperature, as shown in the diagram above. The collection tube containing the gas is allowed to stand for several minutes, and its depth is adjusted until the water levels inside and outside the tube are the same. Assume that:

- the gas is not appreciably soluble in water
- the gas collected in the graduated tube and the water are in thermal equilibrium
- a barometer, a thermometer, an analytical balance, and a table of the equilibrium vapor pressure of water at various temperatures are also available.

(a) Write the equation(s) needed to calculate the molar mass of the gas.

(b) List the measurements that must be made in order to calculate the molar mass of the gas.

(c) Explain the purpose of equalizing the water levels inside and outside the gas collection tube.

(d) The student determines the molar mass of the gas to be 64 g mol⁻¹. Write the expression (set-up) for calculating the percent error in the experimental value, assuming that the unknown gas is butane (molar mass 58 g mol⁻¹). Calculations are not required.

(e) If the student fails to use information from the table of the equilibrium vapor pressures of water in the calculation, the calculated value for the molar mass of the unknown gas will be smaller than the actual value. Explain.
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Question 5

8 Points

(a) \( PV = nRT \) AND \( n = \frac{m}{M} \), OR molar mass = \( \frac{mRT}{PV} \), OR \( M = \frac{DRT}{P} \)  \( 1 \) pt

(b) temperature, atmospheric pressure, volume of the gas, and mass of gas (mass of pressurized container before and after releasing the gas)  \( 3 \) pts

**Note:** 1 point earned for any two of the above, 2 points earned for any three of them.

“The mass of the gas” is acceptable as a “measurement” for the 1st or 2nd point. Extraneous measurements (e.g., density, volume of liquid, etc.) are ignored. To earn 3rd point, “mass of pressurized container before and after releasing the gas”, or “change in mass of container” must be indicated.

(c) to equalize internal pressure with room pressure (atmospheric pressure), or the pressure(s) will be the same.  \( 1 \) pt

(d) \( \% \) error = \( \frac{(64 - 58) g}{58 g} \times 100 \% \) (or \( \frac{6}{58} \times 100 \%, \) or \( \frac{6}{58} \))  \( 1 \) pt

**Note:** No points earned for generic response (e.g., \( \frac{(\text{expt.} - \text{theor.})}{\text{theor.}} \times 100 \)), or for \( \frac{6}{64} \times 100 \% \). No penalty if “\( \times 100 \% \)” is absent or if value (10\%) is not calculated.

(e) Pressure will be larger, therefore number of moles will be larger

molar mass = \( \frac{\text{mass}}{\text{moles}} \), therefore calculated molar mass will be smaller  \( 1 \) pt

**OR,** \( M = \frac{mRT}{PV} \) (or \( \frac{DRT}{P} \)), and the denominator, \( PV \), will be too large.  \( 2 \) pts

Therefore, the value of the molar mass (\( \frac{mRT}{PV} \) or \( \frac{DRT}{P} \)) will be too small.

**OR,** The pressure is larger, or the number of moles is larger, or since \( P_{\text{total}} = (P_{\text{unknown}} - P_{\text{water}}) \)  \( 1 \) pt

we know that \( P_{\text{total}} > P_{\text{unknown}} \).

**Note:** If \( n = \frac{m}{M} \) is missing in part (a) but present in part (e), 1 point is earned for part (a).

8 points:
Which of the following gases contains the fewest number of molecules?

(A) $6.0 \text{ g of } \text{H}_2$
(B) $14 \text{ g of } \text{N}_2$
(C) $17 \text{ g of } \text{NH}_3$
(D) $34 \text{ g of } \text{H}_2\text{S}$

Which of the following gases is NOT diatomic?

(A) HCl
(B) Hydrogen
(C) Carbon monoxide
(D) Neon

Questions 53-55: Select from the gases below at 25.0°C and 1 atm.

(A) CO$_2$ (molar mass 44)
(B) NH$_3$ (molar mass 17)
(C) H$_2$O (molar mass 18)
(D) C$_4$H$_{10}$ (molar mass 58)

Which molecule would move the fastest in the gas state at 25.0°C and 1 atm?

Which gas would dissolve in distilled water to form a basic solution?

Which gas is an odorless gas frequently used in fire extinguishers?
What is the ratio of the average speed of gaseous Ne atoms to gaseous SO₃ molecules, if equal volumes of both gases are measured at the same temperature and pressure?

(A) 8 to 1  
(B) 4 to 1  
(C) 2 to 1  
(D) 1 to 1

The temperature of 36.0 mL of CH₄ gas is raised from 27.0°C to 327°C at constant pressure. What is the final volume of the CH₄ gas after heating?

(A) 18.0 mL  
(B) 36.0 mL  
(C) 72.0 mL  
(D) 327 mL

A real gas closely approaches the behavior of an ideal gas under conditions of

(A) low pressure and high temperature  
(B) low pressure and low temperature  
(C) high pressure and low temperature  
(D) high pressure and high temperature

A mixture of 0.40 mole of He gas and 0.50 mole of N₂ gas exerts a total pressure of 0.90 atm at 25°C. What is the partial pressure of the He gas?

(A) 0.36 atm  
(B) 0.40 atm  
(C) 0.45 atm  
(D) 0.81 atm

What is the final gas pressure exerted by a sample of gas if the absolute (Kelvin) temperature of a gas is doubled and the volume is tripled?

(A) 1/6 the original pressure  
(B) 2/3 the original pressure  
(C) 3/2 the original pressure  
(D) 6 times the original pressure
An inverse relationship is to Boyle's Law as a direct relationship is to

(A) Ideal Gas Law
(B) Graham's Law of Gas Diffusion
(C) Hess's Law
(D) Charles' Law

When water is heated to boiling in an open soda can and the can is then inverted into a bucket containing cold water,

(A) the can expands due to rapid expansion of the water vapor inside
(B) the can implodes due to the rapid change in room air pressure
(C) the can implodes due to rapid condensation of the inside water vapor
(D) the can implodes due to the rapid change in room air pressure and rapid condensation of the inside water vapor

When a sample of an ideal gas is heated from 30°C to 60°C, the average kinetic energy of the gas changes. Which factor describes this change?

(A) 1/2
(B) 333/303
(C) 333/273
(D) 2/1

The dissolving of CO₂ gas in room temperature distilled water will typically

(A) increase if the temperature of the water is increased
(B) decrease if the temperature of the water is decreased
(C) result in a change of pH from that of the pure water
(D) result in the formation of a precipitate as the CO₂ dissolves

A rigid metal contains nitrogen gas. Which of the following statements applies to the N₂ gas in the tank if 10% of the gas is allowed to escape? (Assume the temperature remains constant.)

(A) The volume of the gas decreases.
(B) The pressure of the gas decreases.
(C) The average kinetic energy of the N₂ gas decreases.
(D) The mass of the gas remains the same.
Questions 66-67: Hydrogen gas is produced and then collected over water at 24°C in a long gas-collecting tube, which is inverted in a battery jar. The water levels in the battery jar and the tube are made equal, and the volume of the gas is 37.20 mL. The gas, mixed with evaporated water vapor, exerts a total pressure of 751.0 mm.

66. If the water vapor pressure at this temperature is 23.1 mm, what is the pressure of the "dry" hydrogen at 24°C?

(A) 774.1 mm Hg  
(B) 751.0 mm Hg  
(C) 727.9 mm Hg  
(D) 23.1 mm Hg

67. The purpose of leveling the gas measuring tube in the above experiment is to

(A) determine the volume of the gas at room temperature  
(B) ensure that the gas pressure inside the tube is equal to room pressure  
(C) prevent further water molecules from evaporating in the tube  
(D) prevent the gas molecules from escaping from the tube

Questions 68-69: Use the balanced equation \(2 \text{C}_2\text{H}_6(\text{g}) + 7 \text{O}_2(\text{g}) \rightarrow 6 \text{H}_2\text{O}(\text{g}) + 4 \text{CO}_2(\text{g})\) to answer the following questions.

68. What minimum volume of oxygen will be required to completely react with 6.0 liters of \(\text{C}_2\text{H}_6\)? (Assume room pressure and temperature remain constant.)

(A) 21.0 L  
(B) 14.0 L  
(C) 6.0 L  
(D) 3.5 L

69. How many molecules of \(\text{H}_2\text{O}(\text{g})\) will result if \(1 \times 10^{23}\) molecules of \(\text{C}_2\text{H}_6\) completely react with excess oxygen?

(A) \(1 \times 10^{23}\) molecules  
(B) \(3 \times 10^{23}\) molecules  
(C) \(4 \times 10^{23}\) molecules  
(D) \(6 \times 10^{23}\) molecules
20. When a balloon filled with helium gas is released, it rises and floats away. Which statement below is the best explanation for this observed behavior?

(A) The helium density inside the balloon is less than that of the surrounding air.
(B) The temperature of the surrounding air is less than the temperature of the helium.
(C) Air pressure is greater than the pressure exerted by the helium.
(D) The rate of diffusion of cooler air is less than that of warmer air.

Questions 21-22: The density of a mystery gas is 2.0 g/Liter at 0°C and 1 atm (STP) conditions.

21. What is the molar mass of this gas?

(A) 2.0 × 22.4
(B) 22.4/2.0
(C) 2.0/22.4
(D) 2.0 × 24.5

22. What is the most likely formula of this mystery gas?

(A) Ne
(B) CH₄
(C) CO
(D) C₃H₈

23. 2NH₃(g) → 3H₂(g) + N₂(g). Consider the decomposition of NH₃ gas in the equation above at constant room temperature and pressure. Compared to the initial pressure, the final pressure exerted by the products will be

(A) 2 times greater
(B) 3 times greater
(C) the same as the initial pressure
(D) one-half the initial pressure

24. Which statement is typically true about gaseous molecules?

(A) The heavier the molecules, the faster they travel.
(B) The heavier the molecules, the greater their average kinetic energy.
(C) The hotter the molecules, the faster they travel.
(D) The larger the volume of the container, the greater the pressure exerted by the gaseous molecules.
A sample of N₂ gas is heated from 200 K to 400 K in a rigid container. Which of the following does NOT double?

(A) the average \( \sqrt{\text{speed}} \) of the N₂ molecules
(B) the pressure of the N₂ gas
(C) the density of the N₂ gas
(D) the average kinetic energy of the N₂ gas
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Answer EITHER Question 2 below OR Question 3 printed on page 8. Only one of these two questions will be graded. If you start both questions, be sure to cross out the question you do not want graded. The Section II score weighting for the question you choose is 20 percent.

2. A rigid 8.20 L flask contains a mixture of 2.50 moles of H₂, 0.500 mole of O₂, and sufficient Ar so that the partial pressure of Ar in the flask is 2.00 atm. The temperature is 127°C.

(a) Calculate the total pressure in the flask.

(b) Calculate the mole fraction of H₂ in the flask.

(c) Calculate the density (in g L⁻¹) of the mixture in the flask.

The mixture in the flask is ignited by a spark, and the reaction represented below occurs until one of the reactants is entirely consumed.

\[ 2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g}) \]

(d) Give the mole fraction of all species present in the flask at the end of the reaction.
10 points

2. A rigid 8.20 L flask contains a mixture of 2.50 moles of H₂, 0.500 mole of O₂, and sufficient Ar so that the partial pressure of Ar in the flask is 2.00 atm. The temperature is 127°C.

(a) Calculate the total pressure in the flask.

\[ P_{H_2} = \left( \frac{n_{H_2} RT}{V} \right) = \left( \frac{(2.50 \text{ mol})(0.0821 \text{ L·atm/mol·K})(400 \text{ K})}{8.20 \text{ L}} \right) = 10.0 \text{ atm} \]

\[ P_{O_2} = \left( \frac{n_{O_2} RT}{V} \right) = \left( \frac{(0.500 \text{ mol})(0.0821 \text{ L·atm/mol·K})(400 \text{ K})}{8.20 \text{ L}} \right) = 2.00 \text{ atm} \]

\[ P_{Ar} = 2.0 \text{ atm} \]

\[ P_T = P_{H_2} + P_{O_2} + P_{Ar} = 10.0 \text{ atm} + 2.0 \text{ atm} + 2.0 \text{ atm} = 14.0 \text{ atm} \]

1 point earned for the partial pressure of H₂

1 point earned for the partial pressure of O₂

1 point earned for the total pressure

(b) Calculate the mole fraction of H₂ in the flask.

\[ \text{Mol fraction}_{H_2} = \left( \frac{\text{mol}_{H_2}}{\text{mol}_{H_2} + \text{mol}_{O_2} + \text{mol}_{Ar}} \right) \]

\[ \text{mol}_{H_2} = 2.50 \text{ mol} \]

\[ \text{mol}_{O_2} = 0.500 \text{ mol} \]

\[ \text{mol}_{Ar} = \left( \frac{PV}{RT} \right) = \left( \frac{(2.00 \text{ atm})(8.20 \text{ L})}{(0.0821 \text{ L·atm/mol·K})(400 \text{ K})} \right) = 0.500 \text{ mol Ar} \]

\[ \text{mol}_{H_2} + \text{mol}_{O_2} + \text{mol}_{Ar} = 2.50 \text{ mol} + 0.500 \text{ mol} + 0.500 \text{ mol} = 3.50 \text{ mol total} \]

\[ \text{Mol fraction}_{H_2} = \left( \frac{\text{mol}_{H_2}}{\text{mol}_{H_2} + \text{mol}_{O_2} + \text{mol}_{Ar}} \right) = \left( \frac{2.50 \text{ mol}}{3.50 \text{ mol}} \right) = 0.714 \]

1 point earned for mol Ar

1 point earned for mol fraction of H₂
(c) Calculate the density (in g L\(^{-1}\)) of the mixture in the flask

\[
\begin{align*}
2.50 \text{ mol H}_2 \left( \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} \right) &= 5.04 \text{ g H}_2 \\
0.500 \text{ mol O}_2 \left( \frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2} \right) &= 16.0 \text{ g O}_2 \\
0.500 \text{ mol Ar} \left( \frac{40.0 \text{ g Ar}}{1 \text{ mol Ar}} \right) &= 20.0 \text{ g Ar}
\end{align*}
\]

\[
\text{total mass} = 5.04 \text{ g} + 16.0 \text{ g} + 20.0 \text{ g} = 41.0 \text{ g}
\]

\[
\text{density} = \left( \frac{\text{total mass}}{\text{volume}} \right) = \left( \frac{41.0 \text{ g}}{8.20 \text{ L}} \right) = 5.00 \text{ g L}^{-1}
\]

1 point earned for mass of all species

1 point earned for density

The mixture in the flask is ignited by a spark, and the reaction represented below occurs until one of the reactants is entirely consumed.

\[
2 \text{ H}_2(\text{g}) + \text{ O}_2(\text{g}) \rightarrow 2 \text{ H}_2\text{O}(\text{g})
\]

(d) Give the mole fraction of all species present in the flask at the end of the reaction.

\[
\begin{array}{c|c|c|c}
 & \text{2 H}_2(\text{g}) & + & \text{O}_2(\text{g}) & \rightarrow & 2 \text{ H}_2\text{O}(\text{g}) \\
\hline
\text{I} & 2.50 & 0 & 0.500 & 0 & \text{1 point earned for 1.00 mol H}_2\text{O} \\
\text{C} & -1.00 & -0.500 & 2 & (+0.500) & \text{1 point earned for total moles} \\
\text{E} & 1.50 & 0 & 1.00 & \text{1 point earned for any two mol fractions, excluding O}_2
\end{array}
\]

\[
\text{total moles after reaction} = \text{mol H}_2 + \text{mol H}_2\text{O} + \text{mol Ar} = 1.50
\]

\[
\text{mol} + 1.00 \text{ mol} + 0.500 \text{ mol} = 3.00 \text{ mol total}
\]

\[
\text{mol fraction H}_2 = \left( \frac{1.50 \text{ mol H}_2}{3.00 \text{ mol}} \right) = 0.500
\]

\[
\text{mol fraction O}_2 = \left( \frac{0 \text{ mol O}_2}{3.00 \text{ mol}} \right) = 0 \text{ (not necessary)}
\]

\[
\text{mol fraction Ar} = \left( \frac{0.500 \text{ mol Ar}}{3.00 \text{ mol}} \right) = 0.167
\]

\[
\text{mol fraction H}_2\text{O} = \left( \frac{1.00 \text{ mol H}_2\text{O}}{3.00 \text{ mol}} \right) = 0.333
\]
8. Answer the following questions about carbon monoxide, CO(g), and carbon dioxide, CO₂(g). Assume that both gases exhibit ideal behavior.

(a) Draw the complete Lewis structure (electron-dot diagram) for the CO molecule and for the CO₂ molecule.

(b) Identify the shape of the CO₂ molecule.

(c) One of the two gases dissolves readily in water to form a solution with a pH below 7. Identify the gas and account for this observation by writing a chemical equation.

(d) A 1.0 mole sample of CO(g) is heated at constant pressure. On the graph below, sketch the expected plot of volume versus temperature as the gas is heated.

![Graph showing V (liters) vs. T (kelvins)]

(e) Samples of CO(g) and CO₂(g) are placed in 1 L containers at the conditions indicated in the diagram below.

![Diagram showing CO gas at 2 atm, 25°C and CO₂ gas at 1 atm, 25°C)]

(i) Indicate whether the average kinetic energy of the CO₂(g) molecules is greater than, equal to, or less than the average kinetic energy of the CO(g) molecules. Justify your answer.

(ii) Indicate whether the root-mean-square speed of the CO₂(g) molecules is greater than, equal to, or less than the root-mean-square speed of the CO(g) molecules. Justify your answer.

(iii) Indicate whether the number of CO₂(g) molecules is greater than, equal to, or less than the number of CO(g) molecules. Justify your answer.
Question 8

Answer the following questions about carbon monoxide, CO(g), and carbon dioxide, CO₂(g). Assume that both gases exhibit ideal behavior.

(a) Draw the complete Lewis structure (electron-dot diagram) for the CO molecule and for the CO₂ molecule.

\[ \cdot\text{O} = \cdot\text{C} = \cdot\text{O} \quad \text{1 point for each correct, complete Lewis structure} \]

(b) Identify the shape of the CO₂ molecule.

CO₂ has a linear molecular geometry \hspace{1cm} 1 point for correct molecular geometry

(c) One of the two gases dissolves readily in water to form a solution with a pH below 7. Identify the gas and account for this observation by writing a chemical equation.

The gas that produces a pH less than 7 when added to water is CO₂. The reaction that accounts for this is

\[ \text{CO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{HCO}_3^-(aq) + \text{H}^+(aq) \quad \text{1 point for identifying CO}_2 \text{ in a correct chemical equation} \]

\[ \text{OR} \quad \text{CO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{CO}_3(aq) \]

(d) A 1.0 mole sample of CO(g) is heated at constant pressure. On the graph below, sketch the expected plot of volume versus temperature as the gas is heated.

The graph should have a straight line with a positive slope. \hspace{1cm} 1 point for drawing a correct line
(e) Samples of CO(g) and CO₂(g) are placed in 1 L containers at the conditions indicated in the diagram below.

```
CO gas
2 atm
25°C

CO₂ gas
1 atm
25°C
```

(i) Indicate whether the average kinetic energy of the CO₂(g) molecules is greater than, equal to, or less than the average kinetic energy of the CO(g) molecules. Justify your answer.

| The average kinetic energy is the same for both samples because the temperature is the same for both samples. Average kinetic energy is proportional to temperature. | 1 point for correct answer and explanation |

(ii) Indicate whether the root-mean-square speed of the CO₂(g) molecules is greater than, equal to, or less than the root-mean-square speed of the CO(g) molecules. Justify your answer.

| The root-mean-square speed for CO₂ is lower than the root-mean-square speed for CO. The molar mass of CO₂ is higher than the molar mass of CO. The root-mean-square speed is inversely proportional to the square root of the molar mass of the gas. | 1 point for correct answer and explanation |

(iii) Indicate whether the number of CO₂(g) molecules is greater than, equal to, or less than the number of CO(g) molecules. Justify your answer.

| There are fewer CO₂ molecules than CO molecules. The CO₂ molecules exert half the pressure of the CO molecules (at the same T and V), so there must be half as many molecules present. | 1 point for correct answer and explanation |
2004 AP® CHEMISTRY FREE-RESPONSE QUESTIONS (Form B)

Answer EITHER Question 2 below OR Question 3 printed on page 8. Only one of these two questions will be graded. If you start both questions, be sure to cross out the question you do not want graded. The Section II score weighting for the question you choose is 20 percent.

2. Answer the following questions related to hydrocarbons.

(a) Determine the empirical formula of a hydrocarbon that contains 85.7 percent carbon by mass.

(b) The density of the hydrocarbon in part (a) is 2.0 g L\(^{-1}\) at 50°C and 0.948 atm.
   (i) Calculate the molar mass of the hydrocarbon.
   (ii) Determine the molecular formula of the hydrocarbon.

(c) Two flasks are connected by a stopcock as shown below. The 5.0 L flask contains CH\(_4\) at a pressure of 3.0 atm, and the 1.0 L flask contains C\(_2\)H\(_6\) at a pressure of 0.55 atm. Calculate the total pressure of the system after the stopcock is opened. Assume that the temperature remains constant.

(c) Two flasks are connected by a stopcock as shown below. The 5.0 L flask contains CH\(_4\) at a pressure of 3.0 atm, and the 1.0 L flask contains C\(_2\)H\(_6\) at a pressure of 0.55 atm. Calculate the total pressure of the system after the stopcock is opened. Assume that the temperature remains constant.

(d) Octane, C\(_8\)H\(_{18}\)(l), has a density of 0.703 g mL\(^{-1}\) at 20°C. A 255 mL sample of C\(_8\)H\(_{18}\)(l) measured at 20°C reacts completely with excess oxygen as represented by the equation below.

\[
2 \text{ C}_8\text{H}_{18}(l) + 25 \text{ O}_2(g) \rightarrow 16 \text{ CO}_2(g) + 18 \text{ H}_2\text{O}(g)
\]

Calculate the total number of moles of gaseous products formed.
Question 2

2. Answer the following questions related to hydrocarbons.

(a) Determine the empirical formula of a hydrocarbon that contains 85.7 percent carbon by mass.

\[
\begin{align*}
n_C &= \frac{85.7 \text{ g C}}{12.01 \text{ g C}} \times 1 \text{ mol C} = 7.14 \text{ mol C} \\
n_H &= \frac{14.3 \text{ g H}}{1.008 \text{ g H}} \times 1 \text{ mol H} = 14.2 \text{ mol H} \\
\frac{7.14 \text{ mol C}}{7.14} : \frac{14.2 \text{ mol H}}{7.14} \\
1 \text{ mol C} : 1.99 \text{ mol H} \\
The \text{ empirical formula is CH}_2
\end{align*}
\]

1 point for moles of C and moles of H

1 point for ratio of moles of C to moles of H

1 point for correct formula

(b) The density of the hydrocarbon in part (a) is 2.0 g L\(^{-1}\) at 50°C and 0.948 atm.

(i) Calculate the molar mass of the hydrocarbon.

\[
\begin{align*}
PV &= nRT \\
molar mass &= \frac{\text{mass}}{\text{molar mass}} \times \frac{RT}{P} = \text{density} \times \frac{RT}{P} \\
molar mass &= \frac{2.0 \text{ g L}^{-1} \times \frac{0.0821 \text{ L atm}}{\text{mol K}} \times 323 \text{ K}}{0.948 \text{ atm}} = 56 \text{ g mol}^{-1}
\end{align*}
\]

1 point for correct substitution

1 point for the answer

(ii) Determine the molecular formula of the hydrocarbon.

\[
\begin{align*}
\text{empirical mass} \times n &= \text{molar mass} \\
\text{empirical mass for CH}_2 &= 14 \text{ g mol}^{-1} \\
14 \text{ g mol}^{-1} \times n &= 56 \text{ g mol}^{-1} \\
n &= 4 \\
The \text{ molecular formula is C}_4\text{H}_8.
\end{align*}
\]

1 point for correct formula
(c) Two flasks are connected by a stopcock as shown below. The 5.0 L flask contains CH₄ at a pressure of 3.0 atm, and the 1.0 L flask contains C₂H₆ at a pressure of 0.55 atm. Calculate the total pressure of the system after the stopcock is opened. Assume that the temperature remains constant.

\[
P_f \text{ of CH}_4 = \frac{P_V}{V_f} = \frac{(3.0 \text{ atm})(5.0 \text{ L})}{6.0 \text{ L}} = 2.5 \text{ atm CH}_4
\]

1 point for final pressure of CH₄ or C₂H₆

\[
P_f \text{ of C}_2\text{H}_6 = \frac{P_V}{V_f} = \frac{(0.55 \text{ atm})(1.0 \text{ L})}{6.0 \text{ L}} = 0.092 \text{ atm C}_2\text{H}_6
\]

1 point for the total pressure

\[
P_T = P_f \text{CH}_4 + P_f \text{C}_2\text{H}_6 = 2.5 \text{ atm} + 0.092 \text{ atm} = 2.6 \text{ atm}
\]

(d) Octane, C₈H₁₈(l), has a density of 0.703 g mL⁻¹ at 20°C. A 255 mL sample of C₈H₁₈(l) measured at 20°C reacts completely with excess oxygen as represented by the equation below.

\[
2 \text{ C}_8\text{H}_{18}(l) + 25 \text{ O}_2(g) \rightarrow 16 \text{ CO}_2(g) + 18 \text{ H}_2\text{O}(g)
\]

Calculate the total number of moles of gaseous products formed.

\[
n_{\text{products}} = 255 \text{ mL C}_8\text{H}_{18} \times \frac{0.703 \text{ g C}_8\text{H}_{18}}{1 \text{ mL C}_8\text{H}_{18}} \times \frac{1 \text{ mol C}_8\text{H}_{18}}{114 \text{ g C}_8\text{H}_{18}} \times \frac{34 \text{ mol products}}{2 \text{ mol C}_8\text{H}_{18}} = 26.7 \text{ mol products}
\]

1 point for substitution of any of these conversion factors

1 point for the correct answer
6. Consider two containers of volume 1.0 L at 298 K, as shown above. One container holds 0.10 mol \( \text{N}_2(g) \) and the other holds 0.10 mol \( \text{H}_2(g) \). The average kinetic energy of the \( \text{N}_2(g) \) molecules is \( 6.2 \times 10^{-21} \) J. Assume that the \( \text{N}_2(g) \) and the \( \text{H}_2(g) \) exhibit ideal behavior.

(a) Is the pressure in the container holding the \( \text{H}_2(g) \) less than, greater than, or equal to the pressure in the container holding the \( \text{N}_2(g) \) ? Justify your answer.

(b) What is the average kinetic energy of the \( \text{H}_2(g) \) molecules?

(c) The molecules of which gas, \( \text{N}_2 \) or \( \text{H}_2 \), have the greater average speed? Justify your answer.

(d) What change could be made that would decrease the average kinetic energy of the \( \text{N}_2(g) \) molecules in the container?

(e) If the volume of the container holding the \( \text{H}_2(g) \) was decreased to 0.50 L at 298 K, what would be the change in each of the following variables? In each case, justify your answer.

   (i) The pressure within the container

   (ii) The average speed of the \( \text{H}_2(g) \) molecules
Consider two containers of volume 1.0 L at 298 K, as shown above. One container holds 0.10 mol N\textsubscript{2}(g) and the other holds 0.10 mol H\textsubscript{2}(g). The average kinetic energy of the N\textsubscript{2}(g) molecules is 6.2 \times 10^{-21} \text{ J}. Assume that the N\textsubscript{2}(g) and the H\textsubscript{2}(g) exhibit ideal behavior.

(a) Is the pressure in the container holding the H\textsubscript{2}(g) less than, greater than, or equal to the pressure in the container holding the N\textsubscript{2}(g)? Justify your answer.

The pressure in the container holding the H\textsubscript{2}(g) is equal to the pressure in the container holding the N\textsubscript{2}(g) because there is an equal number of moles of both gases at the same temperature and volume (\( P = nK \), where the constant \( K = \frac{RT}{V} \)).

One point is earned for the correct choice.
One point is earned for the correct explanation.

(b) What is the average kinetic energy of the H\textsubscript{2}(g) molecules?

The average kinetic energy of the H\textsubscript{2}(g) molecules is 6.2 \times 10^{-21} \text{ J} because both gases are at the same temperature.

One point is earned for the correct energy.

(c) The molecules of which gas, N\textsubscript{2} or H\textsubscript{2}, have the greater average speed? Justify your answer.

H\textsubscript{2}(g) molecules will have the greater average speed. Both gases have the same average kinetic energy, but H\textsubscript{2}(g) has the smaller molar mass. Therefore, the H\textsubscript{2}(g) molecules will have a greater average speed because, at a given temperature, the average (root-mean-square) speed of gas molecules is inversely proportional to the square root of the molar mass of the gas:

\[
\nu_{rms} = \left(\sqrt{\frac{3RT}{M}}\right)^{\frac{1}{2}}
\]

One point is earned for the correct answer with an explanation.
(d) What change could be made that would decrease the average kinetic energy of the molecules in the container?

The average kinetic energy of a gas particle depends on the temperature of the gas sample. To decrease the average kinetic energy of the gas particles in a gas sample, the temperature of the $\text{N}_2(g)$ would have to be lowered.

One point is earned for the correct answer with an explanation.

(e) If the volume of the container holding the $\text{H}_2(g)$ was decreased to 0.50 L at 298 K, what would be the change in each of the following variables? In each case, justify your answer.

(i) The pressure within the container

The pressure would be doubled. $PV$ is a constant when the temperature and number of moles of gas are held constant. Therefore, if the volume is halved the pressure is doubled.

$$P_1V_1 = P_2V_2$$

If $V_2 = \frac{1}{2}V_1$, then $P_1V_1 = P_2 \left(\frac{1}{2}V_1\right) \Rightarrow P_1 = P_2 \left(\frac{1}{2}\right) \Rightarrow P_2 = 2P_1$

One point is earned for the correct answer.

One point is earned for the correct explanation.

(ii) The average speed of the $\text{H}_2(g)$ molecules

The average speed is unchanged when the volume of the gas sample is halved. Average speed depends on changes in temperature, not changes in volume.

One point is earned for the correct answer with an explanation.
(d) A buffer solution is prepared by dissolving some solid NaOCl in a solution of HOCl at 298 K. The pH of the buffer solution is determined to be 6.48.

(i) Calculate the value of [H$_3$O$^+$] in the buffer solution.

(ii) Indicate which of HOCl(aq) or OCl$^-$(aq) is present at the higher concentration in the buffer solution.
Support your answer with a calculation.

2. A student was assigned the task of determining the molar mass of an unknown gas. The student measured the mass of a sealed 843 mL rigid flask that contained dry air. The student then flushed the flask with the unknown gas, resealed it, and measured the mass again. Both the air and the unknown gas were at 23.0°C and 750. torr.
The data for the experiment are shown in the table below.

<table>
<thead>
<tr>
<th>Volume of sealed flask</th>
<th>843 mL</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of sealed flask and dry air</td>
<td>157.70 g</td>
</tr>
<tr>
<td>Mass of sealed flask and unknown gas</td>
<td>158.08 g</td>
</tr>
</tbody>
</table>

(a) Calculate the mass, in grams, of the dry air that was in the sealed flask. (The density of dry air is 1.18 g L$^{-1}$ at 23.0°C and 750. torr.)

(b) Calculate the mass, in grams, of the sealed flask itself (i.e., if it had no air in it).

(c) Calculate the mass, in grams, of the unknown gas that was added to the sealed flask.

(d) Using the information above, calculate the value of the molar mass of the unknown gas.

After the experiment was completed, the instructor informed the student that the unknown gas was carbon dioxide (44.0 g mol$^{-1}$).

(e) Calculate the percent error in the value of the molar mass calculated in part (d).

(f) For each of the following two possible occurrences, indicate whether it by itself could have been responsible for the error in the student’s experimental result. You need not include any calculations with your answer. For each of the possible occurrences, justify your answer.

**Occurrence 1:** The flask was incompletely flushed with CO$_2$(g), resulting in some dry air remaining in the flask.

**Occurrence 2:** The temperature of the air was 23.0°C, but the temperature of the CO$_2$(g) was lower than the reported 23.0°C.

(g) Describe the steps of a laboratory method that the student could use to verify that the volume of the rigid flask is 843 mL at 23.0°C. You need not include any calculations with your answer.
A student was assigned the task of determining the molar mass of an unknown gas. The student measured the mass of a sealed 843 mL rigid flask that contained dry air. The student then flushed the flask with the unknown gas, resealed it, and measured the mass again. Both the air and the unknown gas were at 23.0°C and 750. torr. The data for the experiment are shown in the table below.

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</tbody>
</table>

(a) Calculate the mass, in grams, of the dry air that was in the sealed flask. (The density of dry air is 1.18 g L⁻¹ at 23.0°C and 750. torr.)

\[
m = D \times V = (1.18 \text{ g L}^{-1})(0.843 \text{ L}) = 0.995 \text{ g}
\]

One point is earned for the correct setup and calculation of mass.

(b) Calculate the mass, in grams, of the sealed flask itself (i.e., if it had no air in it).

\[
157.70 \text{ g} - 0.995 \text{ g} = 156.71 \text{ g}
\]

One point is earned for subtracting the answer in part (a) from 157.70 g.

(c) Calculate the mass, in grams, of the unknown gas that was added to the sealed flask.

\[
158.08 \text{ g} - 156.71 \text{ g} = 1.37 \text{ g}
\]

One point is earned for subtracting the answer in part (b) from 158.08 g.

(d) Using the information above, calculate the value of the molar mass of the unknown gas.

\[
n = \frac{PV}{RT} = \frac{\left(\frac{750}{760} \text{ atm}\right)(0.843 \text{ L})}{(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1})(296 \text{ K})} = 0.0342 \text{ mol}
\]

molar mass = \[
\frac{1.37 \text{ g}}{0.0342 \text{ mol}} = 40.1 \text{ g mol}^{-1}
\]

OR

molar mass = \[
\frac{DRT}{P} = \frac{\left(\frac{1.37 \text{ g}}{0.843 \text{ L}}\right)(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1})(296 \text{ K})}{\left(\frac{750}{760} \text{ atm}\right)}
\]

= 40.0 \text{ g mol}^{-1}

One point is earned for the conversion of pressure (if necessary) and temperature and the use of the appropriate R.

One point is earned for the correct setup and calculation of moles of gas.

One point is earned for the correct setup and calculation of molar mass.

OR

If calculation is done in a single step, 1 point is earned for the correct \( P \) and \( T \), 1 point is earned for the correct density, and 1 point is earned for the correct answer.
After the experiment was completed, the instructor informed the student that the unknown gas was carbon dioxide (44.0 g mol⁻¹).

(e) Calculate the percent error in the value of the molar mass calculated in part (d).

\[
\text{percent error} = \left| \frac{44.0 \text{ g mol}^{-1} - 40.1 \text{ g mol}^{-1}}{44.0 \text{ g mol}^{-1}} \right| \times 100 = 8.9\%
\]

One point is earned for the correct setup and answer.

(f) For each of the following two possible occurrences, indicate whether it by itself could have been responsible for the error in the student’s experimental result. You need not include any calculations with your answer. For each of the possible occurrences, justify your answer.

**Occurrence 1**: The flask was incompletely flushed with CO₂(g), resulting in some dry air remaining in the flask.

This occurrence could have been responsible.

The dry air left in the flask is less dense (or has a lower molar mass) than CO₂ gas at the given T and P. This would result in a lower mass of gas in the flask and a lower result for the molar mass of the unknown gas.

One point is earned for the correct reasoning and conclusion.

**Occurrence 2**: The temperature of the air was 23.0°C, but the temperature of the CO₂(g) was lower than the reported 23.0°C.

This occurrence could not have been responsible.

The density of CO₂ is greater at the lower temperature. A larger mass of CO₂ would be in the flask than if the CO₂ had been at 23.0°C, resulting in a higher calculated molar mass for the unknown gas.

One point is earned for the correct reasoning and conclusion.

(g) Describe the steps of a laboratory method that the student could use to verify that the volume of the rigid flask is 843 mL at 23.0°C. You need not include any calculations with your answer.

Valid methods include the following:

1. Find the mass of the empty flask. Fill the flask with a liquid of known density (e.g., water at 23°C), and measure the mass of the liquid-filled flask. Subtract to find the mass of the liquid. Using the known density and mass, calculate the volume.

   One point is earned for a valid method.

2. Measure 843 mL of a liquid (e.g., water) in a 1,000 mL graduated cylinder and transfer the liquid quantitatively into the flask to see if the water fills the flask completely.
3. The mass of an aqueous solution of H₂O₂ is 6.951 g. The H₂O₂ in the solution decomposes completely according to the reaction represented above. The O₂(g) produced is collected in an inverted graduated tube over water at 23.4°C and has a volume of 182.4 mL when the water levels inside and outside of the tube are the same. The atmospheric pressure in the lab is 762.6 torr, and the equilibrium vapor pressure of water at 23.4°C is 21.6 torr.

(a) Calculate the partial pressure, in torr, of O₂(g) in the gas-collection tube.

(b) Calculate the number of moles of O₂(g) produced in the reaction.

(c) Calculate the mass, in grams, of H₂O₂ that decomposed.

(d) Calculate the percent of H₂O₂, by mass, in the original 6.951 g aqueous sample.

(e) Write the oxidation number of the oxygen atoms in H₂O₂ and the oxidation number of the oxygen atoms in O₂ in the appropriate cells in the table below.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Oxidation Number of Oxygen Atoms</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₂O₂</td>
<td></td>
</tr>
<tr>
<td>O₂</td>
<td></td>
</tr>
</tbody>
</table>

(f) Write the balanced oxidation half-reaction for the reaction.

STOP

If you finish before time is called, you may check your work on this part only. Do not turn to the other part of the test until you are told to do so.
2 H₂O₂(aq) → 2 H₂O(l) + O₂(g)

The mass of an aqueous solution of H₂O₂ is 6.951 g. The H₂O₂ in the solution decomposes completely according to the reaction represented above. The O₂(g) produced is collected in an inverted graduated tube over water at 23.4°C and has a volume of 182.4 mL when the water levels inside and outside of the tube are the same. The atmospheric pressure in the lab is 762.6 torr, and the equilibrium vapor pressure of water at 23.4°C is 21.6 torr.

(a) Calculate the partial pressure, in torr, of O₂(g) in the gas-collection tube.

\[ P_{\text{atm}} = P_{O_2} + P_{H_2O} \Rightarrow P_{O_2} = P_{\text{atm}} - P_{H_2O} \]
\[ P_{O_2} = 762.6 \text{ torr} - 21.6 \text{ torr} = 741.0 \text{ torr} \]

(b) Calculate the number of moles of O₂(g) produced in the reaction.

\[ PV = nRT \Rightarrow n = \frac{PV}{RT} \]
\[ P = 741.0 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.9750 \text{ atm} \]
\[ T = 273.15 + 23.4^\circ \text{C} = 296.6 \text{ K} \]
\[ V = 182.4 \text{ mL} \times \frac{1 \text{ L}}{1,000 \text{ mL}} = 0.1824 \text{ L} \]
\[ n_{O_2} = \frac{PV}{RT} = \frac{(0.9750 \text{ atm})(0.1824 \text{ L})}{(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1})(296.6 \text{ K})} = 7.304 \times 10^{-3} \text{ mol} \]

(c) Calculate the mass, in grams, of H₂O₂ that decomposed.

\[ (7.304 \times 10^{-3} \text{ mol O}_2) \times \frac{2 \text{ mol H}_2\text{O}_2}{1 \text{ mol O}_2} \times \frac{34.0 \text{ g H}_2\text{O}_2}{1 \text{ mol H}_2\text{O}_2} = 0.497 \text{ g H}_2\text{O}_2 \]

(d) Calculate the percent of H₂O₂, by mass, in the original 6.951 g aqueous sample.

\[ \frac{0.497 \text{ g H}_2\text{O}_2}{6.951 \text{ g sample}} \times 100 = 7.15\% \]
(e) Write the oxidation number of the oxygen atoms in H$_2$O$_2$ and the oxidation number of the oxygen atoms in O$_2$ in the appropriate cells in the table below.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Oxidation Number of Oxygen Atoms</th>
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<tbody>
<tr>
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<td></td>
</tr>
<tr>
<td>O$_2$</td>
<td></td>
</tr>
</tbody>
</table>

In H$_2$O$_2$, the oxidation number of O is -1. In O$_2$, the oxidation number of O is 0.

Two points are earned for the correct oxidation numbers (1 point each).

(f) Write the balanced oxidation half-reaction for the reaction.

\[ \text{H}_2\text{O}_2(aq) \rightarrow \text{O}_2(g) + 2 \text{H}^+(aq) + 2 e^- \]

One point is earned for the correct reactant and products. One point is earned for correct balancing.
2. A sample of a pure, gaseous hydrocarbon is introduced into a previously evacuated rigid 1.00 L vessel. The pressure of the gas is 0.200 atm at a temperature of 127°C.

(a) Calculate the number of moles of the hydrocarbon in the vessel.

(b) O₂(g) is introduced into the same vessel containing the hydrocarbon. After the addition of the O₂(g), the total pressure of the gas mixture in the vessel is 1.40 atm at 127°C. Calculate the partial pressure of O₂(g) in the vessel.

The mixture of the hydrocarbon and oxygen is sparked so that a complete combustion reaction occurs, producing CO₂(g) and H₂O(g). The partial pressures of these gases at 127°C are 0.600 atm for CO₂(g) and 0.800 atm for H₂O(g). There is O₂(g) remaining in the container after the reaction is complete.

(c) Use the partial pressures of CO₂(g) and H₂O(g) to calculate the partial pressure of the O₂(g) consumed in the combustion.

(d) On the basis of your answers above, write the balanced chemical equation for the combustion reaction and determine the formula of the hydrocarbon.

(e) Calculate the mass of the hydrocarbon that was combusted.

(f) As the vessel cools to room temperature, droplets of liquid water form on the inside walls of the container. Predict whether the pH of the water in the vessel is less than 7, equal to 7, or greater than 7. Explain your prediction.
A sample of a pure, gaseous hydrocarbon is introduced into a previously evacuated rigid 1.00 L vessel. The pressure of the gas is 0.200 atm at a temperature of 127°C.

(a) Calculate the number of moles of the hydrocarbon in the vessel.

\[
\begin{align*}
 n &= \frac{PV}{RT} \\
 &= \frac{(0.200 \text{ atm})(1.00 \text{ L})}{(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1})(400. \text{ K})} \\
&= 6.09 \times 10^{-3} \text{ mol}
\end{align*}
\]

1 point is earned for the setup.

1 point is earned for the numerical answer.

(b) O\textsubscript{2}(g) is introduced into the same vessel containing the hydrocarbon. After the addition of the O\textsubscript{2}(g), the total pressure of the gas mixture in the vessel is 1.40 atm at 127°C. Calculate the partial pressure of O\textsubscript{2}(g) in the vessel.

\[P_{O_2} = 1.40 \text{ atm} - 0.200 \text{ atm} = 1.20 \text{ atm}\]

1 point is earned for the correct pressure.

The mixture of the hydrocarbon and oxygen is sparked so that a complete combustion reaction occurs, producing CO\textsubscript{2}(g) and H\textsubscript{2}O(g). The partial pressures of these gases at 127°C are 0.600 atm for CO\textsubscript{2}(g) and 0.800 atm for H\textsubscript{2}O(g). There is O\textsubscript{2}(g) remaining in the container after the reaction is complete.

(c) Use the partial pressures of CO\textsubscript{2}(g) and H\textsubscript{2}O(g) to calculate the partial pressure of the O\textsubscript{2}(g) consumed in the combustion.

\[
\begin{align*}
\text{before rxn:} & \quad 0.200 \text{ atm} \quad 1.20 \text{ atm} \quad - \quad - \\
\text{after rxn:} & \quad 0 \text{ atm} \quad ? \text{ atm} \quad 0.600 \text{ atm} \quad 0.800 \text{ atm} \\
0.600 \text{ atm CO}_2 \left( \frac{1 \text{ atm O}_2}{1 \text{ atm CO}_2} \right) &= 0.600 \text{ atm O}_2 \\
0.800 \text{ atm H}_2\text{O} \left( \frac{1 \text{ atm O}_2}{2 \text{ atm H}_2\text{O}} \right) &= 0.400 \text{ atm O}_2 & \text{Total O}_2 \text{ consumed} = 1.000 \text{ atm}
\end{align*}
\]

\[\text{OR, based on } PV = nRT \text{ and mole calculations:}
\]

\[
\begin{align*}
 n_{\text{H}_2\text{O}} &= \frac{PV}{RT} = \frac{(0.800 \text{ atm})(1.00 \text{ L})}{(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1})(400. \text{ K})} = 0.0244 \text{ mol H}_2\text{O} \times \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2\text{O}} \\
&= 0.0122 \text{ mol O}_2 \\
n_{\text{CO}_2} &= \frac{PV}{RT} = \frac{(0.600 \text{ atm})(1.00 \text{ L})}{(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1})(400. \text{ K})} = 0.0183 \text{ mol CO}_2 \times \frac{1 \text{ mol O}_2}{1 \text{ mol CO}_2} \\
&= 0.0183 \text{ mol O}_2
\end{align*}
\]

Total moles O\textsubscript{2} = 0.0305; \[P = \frac{nRT}{V} = \frac{(0.0305 \text{ mol})(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1})(400. \text{ K})}{1.00 \text{ L}}\]

\[P = 1.00 \text{ atm O}_2\]
(d) On the basis of your answers above, write the balanced chemical equation for the combustion reaction and determine the formula of the hydrocarbon.

The partial pressures occur in the same proportions as the number of moles.

\[ P_{\text{hydrocarbon}} : P_{O_2} : P_{CO_2} : P_{H_2O} \]

0.200 atm : 1.00 atm : 0.600 atm : 0.800 atm

= 1 : 5 : 3 : 4

\[ C_3H_8 + 5 O_2 \rightarrow 3 CO_2 + 4 H_2O \]

OR

\[ n_{H_2O} = \frac{PV}{RT} = \frac{(0.800 \text{ atm})(1.00 \text{ L})}{(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1})(400. \text{ K})} = 0.0244 \text{ mol H}_2O \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2O} = 0.0487 \text{ mol H} \]

\[ n_{CO_2} = \frac{PV}{RT} = \frac{(0.600 \text{ atm})(1.00 \text{ L})}{(0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1})(400. \text{ K})} = 0.0183 \text{ mol CO}_2 \times \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.0183 \text{ mol C} \]

\[ \frac{0.0487 \text{ mol H}}{0.0183 \text{ mol C}} = \frac{2.66 \text{ mol H}}{1 \text{ mol C}} \times \frac{3}{3} = \frac{8 \text{ mol H}}{3 \text{ mol C}} \Rightarrow C_3H_8 \]

\[ C_3H_8 + 5 O_2 \rightarrow 3 CO_2 + 4 H_2O \]

1 point is earned for the formula of the hydrocarbon.

1 point is earned for a balanced equation with the correct proportions among reactants and products.

(e) Calculate the mass of the hydrocarbon that was combusted.

\[ \text{mass} = (\text{number of moles})(\text{molar mass}) \]

\[ = (6.09 \times 10^{-3} \text{ mol})(44.1 \text{ g/mol}) = 0.269 \text{ g} \]

1 point is earned for using the number of moles combusted from part (a).

1 point is earned for the calculated mass.

(f) As the vessel cools to room temperature, droplets of liquid water form on the inside walls of the container. Predict whether the pH of the water in the vessel is less than 7, equal to 7, or greater than 7. Explain your prediction.

The pH will be less than 7 because CO\textsubscript{2} is soluble in water, with which it reacts to form H\textsuperscript{+} ions.

1 point is earned for the correct choice and explanation.