AP® Chemistry
2002 Sample Student Responses
Form B

The materials included in these files are intended for use by AP teachers for course and exam preparation in the classroom; permission for any other use must be sought from the Advanced Placement Program®. Teachers may reproduce them, in whole or in part, in limited quantities, for face-to-face teaching purposes but may not mass distribute the materials, electronically or otherwise. These materials and any copies made of them may not be resold, and the copyright notices must be retained as they appear here. This permission does not apply to any third-party copyrights contained herein.
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.

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

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

\[ \text{a) } \frac{PV}{nRT} = \frac{P}{n} = \frac{(2.5 + 0.50) \times 0.08206 \times (127 + 273)}{2.01 \text{ atm} + 2.00 \text{ atm}} = 12.0 \text{ atm} \left( P_{\text{H}_2} + P_{\text{O}_2} \right) \]

\[ \text{total pressure} = 14.0 \text{ atm} \]

\[ \text{b) moles Ar} = \frac{PV}{RT} = \frac{2 \times 8.2}{0.08206 \times 273} = 0.4996 \text{ mol} \]

\[ \text{mole fraction H}_2 = \frac{\text{mol H}_2}{\text{total moles}} = \frac{2.50}{2.50 + 0.50 + 0.4996} \]

\[ \text{mol fraction H}_2 = 0.714 \]

\[ \text{c) } 8.20 \text{ L.} \]

\[ 2.50 \text{ mol H}_2 \times 2.016 \frac{\text{g}}{\text{mol}} = 5.04 \text{ g} \]

\[ 0.500 \text{ mol O}_2 \times 32.00 \frac{\text{g}}{\text{mol}} = 16.00 \text{ g} \]

\[ 0.4996 \text{ mol Ar} \times 39.95 \frac{\text{g}}{\text{mol}} = 19.96 \text{ g} \]

\[ \text{total grams} = 41.0 \text{ g} \]

\[ \text{Density} = \frac{\text{g}}{\text{L}} = \frac{41.0}{8.2} = 5.00 \frac{\text{g}}{\text{L}} \]
a) \[ 2H_2 + O_2 \rightarrow 2H_2O \]

\[ 2.50 \text{ mol} \quad 0.500 \text{ mol} \]

2 x mol O₂ < mol H₂, therefore O₂ is the limiting reactant.

\[ 2H_2 + O_2 \rightarrow 2H_2O \]

\[ 2.50 - 2x \quad 0.500 - x \quad 2x \]

\[ x = 0.500 \text{ mol} \]

\[ 1.50 \text{ mol} \quad 0 \quad 1.00 \text{ mol} \]

mol \[ H_2O = 1.00 \text{ mol} \]

mol \[ H_2 = 1.50 \text{ mol} \]

mol \[ Ar = 0.4976 \text{ mol} \]

\[ \text{mol fraction } H_2O = \frac{1.00}{1.50 + 0.4976} = 0.333 \]

\[ \text{mol fraction } H_2 = \frac{1.50}{1.50 + 0.4976} = 0.500 \]

\[ \text{mol fraction } Ar = \frac{0.4976}{1.50 + 0.4976} = 0.167 \]
Answer EITHER Question 2 below OR Question 3 printed on page 12. 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.

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

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

\[ V = 8.20 \text{ L} \quad T = 127^\circ \text{C} \quad \text{of} \quad 400.15 \text{ K} \]

\[ \text{H}_2 = 2.50 \text{ moles} \]

\[ \text{O}_2 = 0.500 \text{ moles} \]

\[ P_{\text{Ar}} = 2.00 \text{ atm} \]

\[ \text{F)} \quad \frac{PV}{n} = RT \]

\[ P = \frac{nRT}{V} \]

\[ n = n_{\text{H}_2} + n_{\text{O}_2} \]

\[ P_{\text{H}_2}O_2 = \left( 2.50 \text{ moles} \times \frac{0.0821 \text{ L atm}}{\text{mol K}} \right) \left( 400.15 \text{ K} \right) \]

\[ 8.20 \text{ L} \]

\[ = 12.019 \text{ L} \]

\[ P_{\text{total}} = P_{\text{H}_2} + P_{\text{O}_2} + P_{\text{Ar}} \]

\[ P_{\text{total}} = 12.019 \text{ atm} + 2 \text{ atm} \]

\[ = 14.0 \text{ atm} \]

\[ \text{b)} \quad \frac{n_{\text{H}_2}RT}{\sqrt{P_{\text{total}}}} = X_{\text{H}_2} \]

\[ \frac{2.50 \text{ moles} \times 0.0821 \text{ L atm/mol K}}{14.019 \text{ atm}} = X_{\text{H}_2} \]

\[ = 0.714 \]

GO ON TO THE NEXT PAGE.
c) \( D = \frac{m \cdot \sqrt{\frac{PV}{n \cdot RT}}} \)

\[ n = \frac{PV}{RT} \]

\[ \gamma = \frac{m}{M} \]

\[ D = \frac{PV \cdot M_{total}}{RT} \]

\[ m = \frac{PV}{RT} \]

\[ M_{total} = M_{H_2} + M_{O_2} + M_{H_2O} \]

\[ M_{H_2} = 2.016 \]

\[ M_{O_2} = 32.00 \]

\[ = \left( \frac{14.019 \text{ atm}}{0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}} \right) \left( \frac{2.016 + 32.00 + 39.948 \text{g/mol}}{400.15 \text{K}} \right) \]

\[ (a_1V_1)(a_2V_2) \left( \frac{9}{\text{mol}} \right) \left( \frac{\text{mol}}{2.016 \text{ atm} \cdot \text{K}} \right) \left( \frac{1}{\text{L}} \right) \]

\[ D = 258.81 \]

\[ = \left( \frac{14.019 \text{ atm}}{0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}} \right) \left( \frac{2.016 + 32.00 + 39.948 \text{g/mol}}{400.15 \text{K}} \right) \]

\[ D = 31.6 \text{ g/L} \]

\[ 2.50 \text{ mole } 1.50 \text{ mole } \rightarrow 1 \text{ mole } \]

8) \( 2 \text{H}_2(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g) \)

\[ \frac{\text{ limiting reagent}}{\text{ amount of } \text{H}_2 \text{ consumed}} \]

\[ 0.300 \text{ mole } \text{O}_2 \times \frac{2 \text{ mole } \text{H}_2\text{O}}{2 \text{ mole } \text{H}_2} \times \frac{1 \text{ mole } \text{H}_2}{1 \text{ mole } \text{O}_2} \]

\[ 1 \text{ mole } \text{O}_2 \times 2 \text{ mole } \text{H}_2\text{O} \rightarrow 1.50 \text{ mole is left} \]

Since 1:1 ratio of \( \text{H}_2 \) and \( \text{H}_2\text{O} \), only 1 mole of \( \text{H}_2\text{O} \) is produced when \( \text{O}_2 \) is consumed.

\[ X_{\text{O}_2} = 0 \]

\[ \frac{X_{\text{H}_2}}{2.50 \text{ mole total}} = 0.60 \text{ mole } \text{H}_2 \]

\[ X_{\text{H}_2\text{O}} = \frac{1 \text{ mole } \text{H}_2}{2.50 \text{ mole total}} = 0.40 \text{ mole } \text{H}_2\text{O} \]

---

Copyright © 2002 by College Entrance Examination Board. All rights reserved. Available at apcentral.collegeboard.com.