



**AP<sup>®</sup> Chemistry**  
**2002 Sample Student Responses**  
**Form B**

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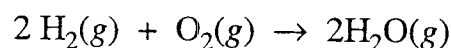
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Answer EITHER Question 2 below OR Question 3 printed on page 12. Only one of these two questions will be graded. If you start both questions, be sure to cross out the question you do not want graded. The Section II score weighting for the question you choose is 20 percent.

2. A rigid 8.20 L flask contains a mixture of 2.50 moles of  $\text{H}_2$ , 0.500 mole of  $\text{O}_2$ , 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  $\text{H}_2$  in the flask.  
 (c) Calculate the density (in  $\text{g L}^{-1}$ ) 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.



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

$$\begin{aligned} \text{a) } PV &= nRT \\ P &= \frac{nRT}{V} = \frac{(2.5 + 0.500) \times 0.08206 \times (127 + 273)}{8.2} = 12.01 \text{ atm } (P_{\text{H}_2} + P_{\text{O}_2}) \\ 12.01 \text{ atm} + 2.00 \text{ atm} &= 14.01 \text{ atm} \end{aligned}$$

$$\text{total pressure} = \boxed{14.0 \text{ atm}}$$

$$\text{b) moles Ar} = \frac{PV}{RT} = \frac{2 \times 8.2}{0.08206 \times 400} = 0.4996 \text{ mol}$$

$$\text{mole fraction H}_2 = \frac{\text{mol H}_2}{\text{total moles}} = \frac{2.50}{2.50 + 0.500 + 0.4996}$$

$$\text{mole fraction H}_2 = \boxed{.714}$$

c) 8.20 L.

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

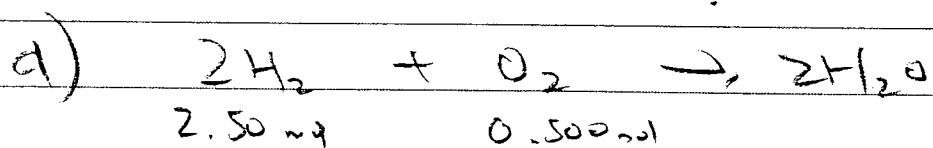
$$0.500 \text{ mol O}_2 \times 32.00 \frac{\text{g}}{\text{mol}} = 16.00 \text{ g}$$

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

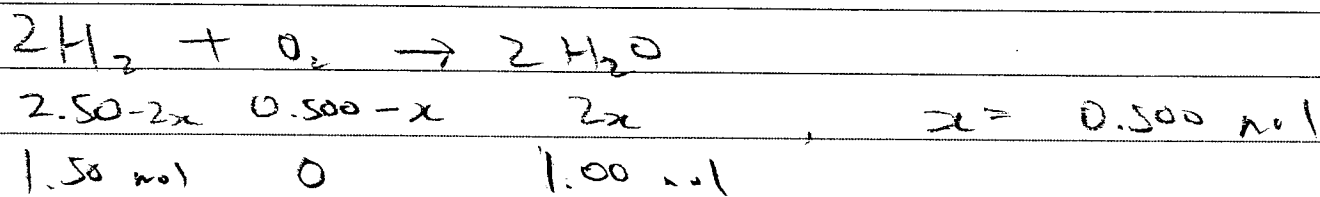
$$\text{total grams} = 41.0 \text{ g}$$

$$\text{Density} = \frac{\text{g}}{\text{L}} = \frac{41.0}{8.20} = 5.00 \frac{\text{g}}{\text{L}}$$

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$2 \times \text{mol O}_2 < \text{mol H}_2$ , therefore  $\text{O}_2$  is limiting reactant



$$\text{mol H}_2\text{O} = 1.00 \text{ mol}$$

$$\text{mol H}_2 = 1.50 \text{ mol}$$

$$\text{mol Ar} = .4496 \text{ mol}$$

$$\text{mol fraction H}_2\text{O} = \frac{1.00}{1 + 1.5 + .4496} = \boxed{.333}$$

$$\text{mol fraction H}_2 = \frac{1.5}{1 + 1.5 + .4496} = \boxed{.500}$$

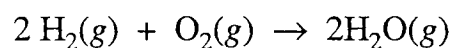
$$\text{mol fraction Ar} = \frac{.4496}{1 + 1.5 + .4496} = \boxed{.167}$$

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2. A rigid 8.20 L flask contains a mixture of 2.50 moles of  $\text{H}_2$ , 0.500 mole of  $\text{O}_2$ , and sufficient Ar so that the partial pressure of Ar in the flask is 2.00 atm. The temperature is  $127^\circ\text{C}$ .

- (a) Calculate the total pressure in the flask.
- (b) Calculate the mole fraction of  $\text{H}_2$  in the flask.
- (c) Calculate the density (in  $\text{g L}^{-1}$ ) 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.



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

$$V = 8.20 \text{ L}$$

$$T = 127^\circ\text{C} \text{ or } 400.15 \text{ K}$$

$$\text{H}_2 = 2.50 \text{ moles}$$

$$\text{O}_2 = 0.500 \text{ moles}$$

$$P_{\text{Ar}} = 2.00 \text{ atm}$$

$$a) PV = nRT$$

$$P = \frac{nRT}{V}$$

$$n = n_{\text{H}_2} + n_{\text{O}_2}$$

$$P_{\text{H}_2 + \text{O}_2} = \frac{(2.50 \text{ moles H}_2 + 0.500 \text{ moles O}_2) \left(0.0821 \frac{\text{L atm}}{\text{mol K}}\right) (400.15 \text{ K})}{8.20 \text{ L}}$$

$$= 12.019 \text{ L}$$

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

$$P_{\text{total}} = 12.019 \text{ atm} + 2 \text{ atm}$$

$$= \boxed{14.0 \text{ atm}}$$

$$b) P_{\text{H}_2} = P_{\text{total}} \times X_{\text{H}_2}$$

$$\frac{P_{\text{H}_2} RT}{V P_{\text{total}}} = X_{\text{H}_2}$$

$$\frac{2.50 \text{ moles H}_2 \left(0.0821 \frac{\text{L atm}}{\text{mol K}}\right) (400.15 \text{ K})}{8.20 \text{ L} (14.019 \text{ atm})} = X_{\text{H}_2}$$

$$\boxed{0.714} = X_{\text{H}_2}$$

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c)  $D = \frac{m}{V}$

$PV = nRT$   
 $n = \frac{PV}{RT}$   
 $n = \frac{m}{M}$

$D = \frac{PVM_{total}}{RT}$   
 $m = \frac{PVM}{RT}$   
 $M_{total} = M_{H_2} + M_{O_2} + M_{Ar}$   
 $M_{H_2} = 2.016$   
 $M_{O_2} = 32.00$   
 $M_{Ar} = 39.948$

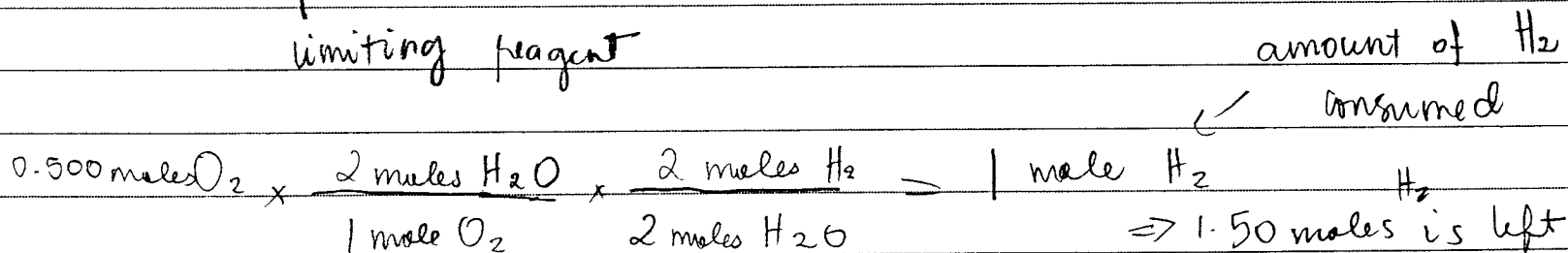
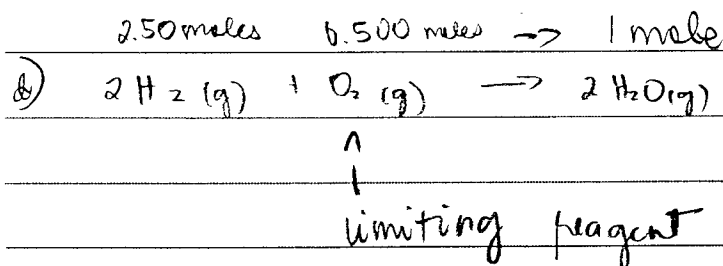
$= \frac{(14.019 \text{ atm}) (2.016 + 32.00 + 39.948 \text{ g/mol})}{(0.0821 \frac{\text{L atm}}{\text{mol K}}) (400.15 \text{ K})}$

$D = 258, 81$

$(\text{atm}) (\text{L}) (\frac{\text{g}}{\text{mol}}) (\frac{\text{mol K}}{\text{L atm}}) (\frac{1}{\text{K}})$

$= \frac{(14.019 \text{ atm}) (2.016 + 32.00 + 39.948 \text{ g/mol})}{(0.0821 \frac{\text{L atm}}{\text{mol K}}) (400.15 \text{ K})}$

$D = 31.6 \frac{\text{g}}{\text{L}}$



since 1:1 ratio of  $O_2$  and  $H_2O$ , only 1 mole of  $H_2O$  is produced

When  $O_2$  is consumed

$X_{O_2} = 0$

$X_{H_2} = \frac{1.50 \text{ moles } H_2}{2.50 \text{ moles total}} = 0.60 H_2$

$X_{H_2O} = \frac{1 \text{ mole } H_2O}{2.50 \text{ moles total}} = 0.40 H_2O$

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