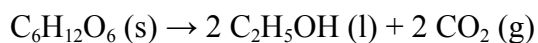


<p style="text-align: center;">16</p> <p>67. Air bags are activated when a severe impact causes a steel ball to compress a spring and electrically ignite a detonator cap. This causes sodium azide (NaN_3) to decompose explosively according to the following reaction:</p> $2\text{NaN}_3(s) \longrightarrow 2\text{Na}(s) + 3\text{N}_2(g)$ <p>What mass of $\text{NaN}_3(s)$ must be reacted to inflate an air bag to 70.0 L at STP?</p>	$\frac{70\text{L}}{22.4\frac{\text{L}}{\text{mol}}} = 3.125 \text{ mol N}_2$ $3.125 \text{ mol} \left(\frac{2}{3}\right) = 2.083 \text{ mol NaN}_3$ $2.083 \text{ mol NaN}_3 \times 65.0\frac{\text{g}}{\text{mol}} = 135 \text{ g NaN}_3$
<p style="text-align: center;">17</p> <p>10.55 The metabolic oxidation of glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, in our bodies produces CO_2, which is expelled from our lungs as a gas:</p> $\text{C}_6\text{H}_{12}\text{O}_6(aq) + 6 \text{O}_2(g) \longrightarrow 6 \text{CO}_2(g) + 6 \text{H}_2\text{O}(l)$ <p>a) Calculate the volume of dry CO_2 produced at body temperature (37°C) and 0.970 atm when 24.5 g of glucose is consumed in this reaction.</p> <p>b) Calculate the volume of oxygen at 1 atm and 298 K gas needed to completely oxidize 50.0 g of glucose.</p>	<p>a) $\frac{24.5\text{g}}{180\frac{\text{g}}{\text{mol}}} = 0.136 \text{ mol C}_6\text{H}_{12}\text{O}_6$</p> $0.136 \text{ mol C}_6\text{H}_{12}\text{O}_6 \left(\frac{6}{1}\right) = 0.817 \text{ mol CO}_2$ $V = \frac{nRT}{P} = \frac{(0.817\text{mol})(0.0821\frac{\text{Latm}}{\text{molK}})(310\text{K})}{0.970\text{atm}} = 21.4 \text{ L CO}_2$ <p>b) $\frac{50.0\text{g}}{180\frac{\text{g}}{\text{mol}}} = 0.278 \text{ mol C}_6\text{H}_{12}\text{O}_6$</p> $0.278 \text{ mol C}_6\text{H}_{12}\text{O}_6 \left(\frac{6}{1}\right) = 1.67 \text{ mol O}_2$ $V = \frac{nRT}{P} = \frac{(1.67\text{mol})(0.0821\frac{\text{Latm}}{\text{molK}})(298\text{K})}{1.00\text{atm}} = 40.8 \text{ L CO}_2$
<p style="text-align: center;">18</p> <p>Ethanol ($\text{C}_2\text{H}_5\text{OH}$) burns in air:</p> $__\text{C}_2\text{H}_5\text{OH}(l) + __\text{O}_2(g) \rightarrow __\text{CO}_2(g) + __\text{H}_2\text{O}(g)$ <p>Balance the equation and determine the volume of air in Liters at 35.0°C and 790. mm Hg required to burn 227 grams of ethanol. Assume that air is 21.0 percent O_2 by volume.</p>	$\text{C}_2\text{H}_5\text{OH}(l) + 3 \text{O}_2(g) \rightarrow 2 \text{CO}_2(g) + 3 \text{H}_2\text{O}(g)$ $\frac{227\text{g}}{46.0\frac{\text{g}}{\text{mol}}} = 4.93 \text{ mol C}_2\text{H}_5\text{OH}$ $4.93 \text{ mol C}_2\text{H}_5\text{OH} \left(\frac{3}{1}\right) = 14.8 \text{ mol O}_2$ $V = \frac{nRT}{P} = \frac{(14.8\text{mol})(62.4\frac{\text{LmmHg}}{\text{molK}})(308\text{K})}{790.\text{mmHg}} = 360. \text{ L O}_2$ <p>$360. \text{ L O}_2 = 0.21x$ where x = volume of air</p> $x = \frac{360\text{L}}{0.21} = 1714 \text{ L} \sim 1710 \text{ L air}$

19

In alcoholic fermentation, yeast converts glucose to ethanol and carbon dioxide:



If 5.97 g of glucose are reacted and 1.44 L of CO_2 gas are collected at 293 K and 0.984 atm, what is the percent yield of the reaction?

$$\frac{5.97\text{g}}{180\frac{\text{g}}{\text{mol}}} = 0.0332 \text{ mol C}_6\text{H}_{12}\text{O}_6$$

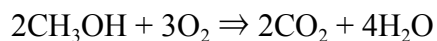
$$0.0332 \text{ mol C}_6\text{H}_{12}\text{O}_6 \left(\frac{2}{1}\right) = 0.0664 \text{ mol CO}_2$$

$$V = \frac{nRT}{P} = \frac{(0.0664\text{mol})(0.0821\frac{\text{Latm}}{\text{molK}})(293\text{K})}{0.984\text{atm}} = 1.62 \text{ L CO}_2$$

$$\frac{1.44\text{L}}{1.62\text{L}} \times 100 = 88.9\%$$

20

10.113 Consider the combustion reaction between 25.0 mL of liquid methanol (density = 0.850 g/mL) and 12.5 L of oxygen gas measured at STP. The products of the reaction are $\text{CO}_2(\text{g})$ and $\text{H}_2\text{O}(\text{g})$. Calculate the number of moles of H_2O formed if the reaction goes to completion.



$$25.0\text{mL}(0.850\frac{\text{g}}{\text{mL}}) = 21.25\text{gCH}_3\text{OH}$$

$$\frac{21.25\text{g}}{32\frac{\text{g}}{\text{mol}}} = 0.664 \text{ mol CH}_3\text{OH}$$

$$\frac{12.5 \text{ L}}{22.4\frac{\text{L}}{\text{mol}}} = 0.558 \text{ mol O}_2 \text{ LR}$$

$$0.558 \text{ mol O}_2 \left(\frac{4}{3}\right) = 0.744 \text{ mol H}_2\text{O}$$