

Worksheet 1.6

Supplemental Instruction
Iowa State University

Leader: Lea

Course: Chem 177

Instructor: Dr. Schewe-Miller

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- 1.) A piece of aluminum foil 1.00 cm square and 0.550 mm thick is allowed to react with bromine to form aluminum bromide.

a.) How many moles of aluminum were used? (density Al = 2.699 g/cm³)

$$.550 \text{ mm} \times \frac{1 \text{ cm}}{10 \text{ mm}} = .055 \text{ cm}$$

$$1 \text{ cm} \times .055 \text{ cm} = .055 \text{ cm}^2 \times \frac{2.699 \text{ g}}{\text{cm}^3} = .148445 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = \boxed{5.50 \times 10^{-3} \text{ mol Al}}$$

b.) How many grams of aluminum bromide form, assuming the aluminum reacts completely? What type of reaction is this?



$$5.50 \times 10^{-3} \text{ mol Al} \times \frac{2 \text{ mol AlBr}_3}{2 \text{ mol Al}} \times \frac{266.69 \text{ g AlBr}_3}{1 \text{ mol AlBr}_3} = \boxed{1.47 \text{ g}}$$

composition reaction

2.) The chemical formula for benzene is C₆H₆.

a. Determine the percent composition of H and C for benzene.

$$6(12.01) + 6(1.01) = 78.12$$

$$\text{C: } \frac{72.06}{78.12} \times 100 = 92\% \quad \text{H: } 100 - 92\% = 8\%$$

$$\boxed{\begin{array}{l} \text{C: } 92\% \\ \text{H: } 8\% \end{array}}$$

b. Write a balanced combustion reaction for benzene.



c. Assuming you have 364g of benzene, how many moles of O₂ will be required to completely combust this amount of benzene?

$$364 \text{ g} \times \frac{1 \text{ mol C}_6\text{H}_6}{78.12 \text{ g C}_6\text{H}_6} \times \frac{15 \text{ mol O}_2}{2 \text{ mol C}_6\text{H}_6} = \boxed{35.0 \text{ mol}}$$

d. How many grams of CO₂ will this produce? How many grams of H₂O?

$$\text{CO}_2: 364 \text{ g C}_6\text{H}_6 \times \frac{1 \text{ mol C}_6\text{H}_6}{78.12 \text{ g C}_6\text{H}_6} \times \frac{12 \text{ mol CO}_2}{2 \text{ mol C}_6\text{H}_6} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = \boxed{1.23 \times 10^3 \text{ g CO}_2}$$

$$\text{H}_2\text{O}: 364 \text{ g C}_6\text{H}_6 \times \frac{1 \text{ mol C}_6\text{H}_6}{78.12 \text{ g C}_6\text{H}_6} \times \frac{6 \text{ mol H}_2\text{O}}{2 \text{ mol C}_6\text{H}_6} \times \frac{18 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = \boxed{2.52 \times 10^3 \text{ g H}_2\text{O}}$$

e. You perform the combustion, and 956g of CO₂ are produced. What is the percent yield?

$$\frac{\text{Actual}}{\text{Theoretical}} \times 100 = \frac{956 \text{ g CO}_2}{1230 \text{ g CO}_2} \times 100 = \boxed{77.7\%}$$

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Julia Burdge



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