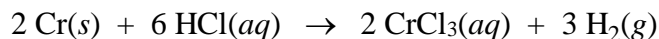
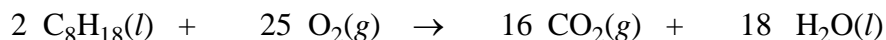


Chapters 8 Study Questions

1. Chromium reacts with hydrochloric acid in a single replacement reaction. The balanced equation is:



- How many moles of HCl are needed to produce 1.60 moles of CrCl₃?
 - How many grams of Cr are needed to produce 3.20 g H₂?
 - In an experiment, 10.2 grams of CrCl₃ are produced starting from 8.30 grams of HCl. What was the theoretical yield and the percent yield in this experiment?
 - When 6.0 moles of Cr are combined with 12.0 moles of HCl, which reactant is limiting? How many moles of excess reactant are left over?
 - How many grams of CrCl₃ are produced starting from 13.0 g of Cr and 43.8 g of HCl?
2. Octane undergoes complete combustion to form carbon dioxide and water.



- How many moles of oxygen are required to burn 1.00 mole of octane?
- What mass of CO₂ is produced when 4.77 grams of oxygen gas are used up?
- How many grams of CO₂ are produced from 11.4 g C₈H₁₈ and 32.0 g O₂?
- In an experiment, 2.28 g C₈H₁₈ produced 2.43 g of H₂O. What is the theoretical yield (the amount of H₂O expected from 2.28 g C₈H₁₈)? What is the percent yield?

Summary of Chapter 8: Quantities in Chemical Reactions

Calculations from a balanced chemical equation:

 mole relationships between reactants and products

 mass relationships between reactants and products

limiting reactant

theoretical yield

actual yield

calculating percent yield

enthalpy of reactions

Answers to Chapters 8 Study Questions

1. a) $1.60 \text{ mol CrCl}_3 \times \frac{6 \text{ mol HCl}}{2 \text{ mol CrCl}_3} = 4.80 \text{ moles HCl}$

b) $3.20 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{2 \text{ mol Cr}}{3 \text{ mol H}_2} \times \frac{52.0 \text{ g Cr}}{1 \text{ mol Cr}} = 55.0 \text{ g Cr}$

c) $8.30 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \times \frac{2 \text{ mol CrCl}_3}{6 \text{ mol HCl}} \times \frac{158.35 \text{ g CrCl}_3}{1 \text{ mol CrCl}_3} = 12.0 \text{ g CrCl}_3$

theoretical yield = 12.0 g CrCl₃

$$\% \text{ Yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{10.2 \text{ g}}{12.0 \text{ g}} \times 100\% = 85.0\%$$

d) $6.0 \text{ moles Cr} \times \frac{2 \text{ mol CrCl}_3}{2 \text{ mol Cr}} = 6.0 \text{ moles CrCl}_3$

$$12.0 \text{ moles HCl} \times \frac{2 \text{ mol CrCl}_3}{6 \text{ mol HCl}} = 4.0 \text{ moles CrCl}_3; \text{ therefore, HCl is limiting}$$

$$4.0 \text{ moles CrCl}_3 \times \frac{2 \text{ mol Cr}}{2 \text{ mol CrCl}_3} = 4.0 \text{ moles Cr used up.}$$

6.0 - 4.0 = 2.0 moles Cr left over.

e) $13.0 \text{ g Cr} \times \frac{1 \text{ mol Cr}}{52.0 \text{ g Cr}} \times \frac{2 \text{ mol CrCl}_3}{2 \text{ mol Cr}} \times \frac{158 \text{ g CrCl}_3}{1 \text{ mol CrCl}_3} = 39.5 \text{ g CrCl}_3$

$$43.8 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.5 \text{ g HCl}} \times \frac{2 \text{ mol CrCl}_3}{6 \text{ mol HCl}} \times \frac{158 \text{ g CrCl}_3}{1 \text{ mol CrCl}_3} = 63.2 \text{ g CrCl}_3$$

since 39.5 g < 63.2 g, 39.5 g CrCl₃ is produced.

2. a) $1.00 \text{ mol C}_8\text{H}_{18} \times \frac{25 \text{ mole O}_2}{2 \text{ mole C}_8\text{H}_{18}} = 12.5 \text{ mol O}_2$

b) $4.77 \text{ g O}_2 \times \frac{1 \text{ mole O}_2}{32.0 \text{ g O}_2} \times \frac{16 \text{ mole CO}_2}{25 \text{ mole O}_2} \times \frac{44.0 \text{ g CO}_2}{1 \text{ mole CO}_2} = 4.20 \text{ g CO}_2$

c) $11.4 \text{ g C}_8\text{H}_{18} \times \frac{1 \text{ mole C}_8\text{H}_{18}}{114 \text{ g C}_8\text{H}_{18}} \times \frac{16 \text{ mole CO}_2}{2 \text{ mole C}_8\text{H}_{18}} \times \frac{44.0 \text{ g CO}_2}{1 \text{ mole CO}_2} = 35.2 \text{ g CO}_2$

$$32.0 \text{ g O}_2 \times \frac{1 \text{ mole O}_2}{32.0 \text{ g O}_2} \times \frac{16 \text{ mole CO}_2}{25 \text{ mole O}_2} \times \frac{44.0 \text{ g CO}_2}{1 \text{ mole CO}_2} = 28.2 \text{ g CO}_2$$

Since 28.2 g is less than 35.2 g, 28.2 g CO₂ are produced.

$$d) 2.28 \text{ g } C_8H_{18} \times \frac{1 \text{ mole } C_8H_{18}}{114 \text{ g } C_8H_{18}} \times \frac{18 \text{ mole } H_2O}{2 \text{ mole } C_8H_{18}} \times \frac{18.0 \text{ g } H_2O}{1 \text{ mole } H_2O} = 3.24 \text{ g } H_2O$$

The theoretical yield is 3.24 g H_2O

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{2.43 \text{ g}}{3.24 \text{ g}} \times 100\% = 75.0\%$$