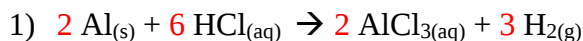
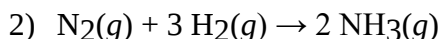


Stoichiometry Practice – Complete Answers



a. $11.5 \text{ g} \times \frac{1 \text{ mol Al}}{26.982 \text{ g}} \times \frac{2 \text{ mol AlCl}_3}{2 \text{ mol Al}} \times \frac{133.341 \text{ g}}{1 \text{ mol}} = 56.8 \text{ g}$

b. $\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$
 $= (55.0 / 56.83) \times 100$
 $= 96.8\%$



5.0 g 5.0 g

Amount of product given each reactant:

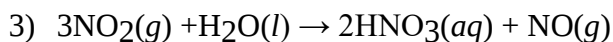
$5.0 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.014 \text{ g N}_2} \times \frac{2 \text{ mol NH}_3}{1 \text{ mol N}_2} \times \frac{17.031 \text{ g NH}_3}{1 \text{ mol NH}_3} = 6.079 \text{ g NH}_3$

$5.0 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} \times \frac{17.031 \text{ g NH}_3}{1 \text{ mol NH}_3} = 28.160 \text{ g NH}_3$

- a. Hydrogen is the excess reactant. TRUE – more NH_3 is made given the amount of hydrogen.
- b. Nitrogen is the limiting reactant. TRUE – less NH_3 is made given the amount of nitrogen.
- c. 2.8 grams of hydrogen are left over. FALSE –
 $5.0 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.014 \text{ g N}_2} \times \frac{3 \text{ mol H}_2}{1 \text{ mol N}_2} \times \frac{2.016 \text{ g H}_2}{1 \text{ mol H}_2} = 1.079 \text{ g H}_2 \text{ needed.}$

Given $5.0 \text{ g} - 1.079 = 3.9 \text{ g}$ left over.

- d. The theoretical yield of ammonia is 6.1 g. TRUE – see calculations above.

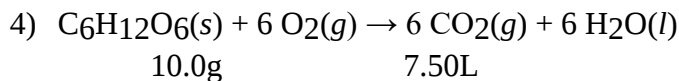


45.0 g 60%

Theoretical yield:

$45.0 \text{ g NO}_2 \times \frac{1 \text{ mol NO}_2}{46.007 \text{ g}} \times \frac{2 \text{ mol HNO}_3}{3 \text{ mol NO}_2} \times \frac{63.012 \text{ g}}{1 \text{ mol HNO}_3} = 41.089 \text{ g}$

$60\% = \frac{\text{X}}{41.089 \text{ g}} \times 100$
 $= 24.7 \text{ g}$



Theoretical

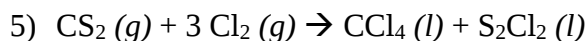
$$10.0 \text{ g} \times \frac{1 \text{ mol glucose}}{180.156 \text{ g}} \times \frac{6 \text{ mol CO}_2}{1 \text{ mol glucose}} \times \frac{44.009 \text{ g}}{1 \text{ mol}} = 14.657 \text{ g CO}_2$$

Actual

$$7.50 \text{ L} \times \frac{1.26 \text{ g CO}_2}{1 \text{ L}} = 9.576 \text{ g}$$

$$\% \text{ yield} = (9.576 \text{ g} / 14.657 \text{ g}) \times 100$$

$$65.3\%$$



$$6.00 \text{ g CS}_2 \times \frac{1 \text{ mol CS}_2}{76.143 \text{ g}} \times \frac{3 \text{ mol Cl}_2}{1 \text{ mol CS}_2} = 0.23639 \text{ mol Cl}_2 \text{ reacts with } 6.0 \text{ g CS}_2$$

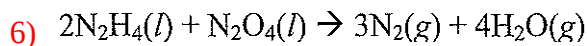
$$= 0.0788 \text{ mol}$$

$$10.0 \text{ g Cl}_2 \times \frac{1 \text{ mol Cl}_2}{70.906 \text{ g}} = 0.1410 \text{ mol Cl}_2 \text{ available.}$$

Cl₂(g) is the limiting reactant because more moles of Cl₂ (0.236 mol) are needed to react with the available moles of CS₂ than what is available (0.141 mol).

$$0.1410 \text{ mol Cl}_2 \times \frac{1 \text{ mol CS}_2}{3 \text{ mole Cl}_2} \times \frac{76.143 \text{ g CS}_2}{1 \text{ mol CS}_2} = 3.578 \text{ g reacted CS}_2$$

$$6.0 \text{ g} - 3.578 \text{ g} = 2.42 \text{ g}$$



Determine how many moles of one of the products (ex. water) are produced using the initial amounts of each reactant.

$$\text{a. } 1.40 \text{ kg N}_2\text{H}_4 \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol N}_2\text{H}_4}{32.046 \text{ g N}_2\text{H}_4} \times \frac{4 \text{ mol H}_2\text{O}}{2 \text{ mol N}_2\text{H}_4} = 87.374 \text{ mol H}_2\text{O}$$

b. Using the # of moles of dinitrogen tetraoxide, N₂O₄:

$$\text{i. } 2.80 \text{ kg N}_2\text{O}_4 \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol N}_2\text{O}_4}{92.01 \text{ g N}_2\text{O}_4} \times \frac{4 \text{ mol H}_2\text{O}}{1 \text{ mol N}_2\text{O}_4} = 121.7 \text{ mol H}_2\text{O}$$

c. Since hydrazine (N₂H₄) produces fewer moles of water, it is the limiting reagent.