

I. Methanol, CH₃OH, is used as a fuel. Recall that combustion reactions produce carbon dioxide and water.

1. Write a balanced chemical equation for the combustion reaction.



2. How many moles of carbon dioxide are produced in each of the following cases? Include calculations quantifying the number of moles left over for any excess reactants.

a. 2 mols of CH₃OH react with 3 mols of O₂.

$$2 \text{ mol CH}_3\text{OH} \times \frac{2 \text{ mol CO}_2}{2 \text{ mol CH}_3\text{OH}} = 2 \text{ mol CO}_2 \text{ max}$$

These reagents are in stoichiometric equivalence, so both reagents react completely.
0 moles of excess reagent

$$3 \text{ mol O}_2 \times \frac{2 \text{ mol CO}_2}{3 \text{ mol O}_2} = 2 \text{ mol CO}_2 \text{ max}$$

b. 2 mols of CH₃OH react with 2 mols of O₂

$$2 \text{ mol CH}_3\text{OH} \times \frac{2 \text{ mol CO}_2}{2 \text{ mol CH}_3\text{OH}} = 2 \text{ mol CO}_2 \text{ max}$$

CH₃OH excess reagent

$$2 \text{ mol O}_2 \times \frac{2 \text{ mol CO}_2}{3 \text{ mol O}_2} = 1.33 \text{ mol CO}_2 \text{ max}$$

O₂ limiting reagent

$$2 \text{ mol O}_2 \times \frac{2 \text{ mol CH}_3\text{OH}}{3 \text{ mol O}_2} = 1.33 \text{ mol CH}_3\text{OH used}$$

2 - 1.33 = **0.67 moles excess methanol**

c. 3 mols of CH₃OH react with 3 mols of O₂

Perform hypothetical maximum calculations as above. Another way of determining the limiting reagent is to say..." 3 moles of methanol would require 4.5 moles of oxygen. We only have 3 moles of O₂, so oxygen must be limiting."

$$3 \text{ mol CH}_3\text{OH} \times \frac{3 \text{ mol O}_2}{2 \text{ mol CH}_3\text{OH}} = 4.5 \text{ mol O}_2 \text{ required}$$

... O₂ is limiting

All of the oxygen will react...

$$3 \text{ mol O}_2 \times \frac{2 \text{ mol CO}_2}{3 \text{ mol O}_2} = 2 \text{ mol CO}_2 \text{ produced}$$

$$3 \text{ mol O}_2 \times \frac{2 \text{ mol CH}_3\text{OH}}{3 \text{ mol O}_2} = 2 \text{ mol CH}_3\text{OH used};$$

1 mole of excess methanol

d. 88 g of CH₃OH react with 88 g of O₂

These problems require you to convert grams using the molar mass before performing steps above. Note: methanol and oxygen have different molecular formulas but happen to have similar molar masses.

$$88 \text{ g CH}_3\text{OH} \times \frac{1 \text{ mol CH}_3\text{OH}}{32 \text{ g CH}_3\text{OH}} \times \frac{2 \text{ mol CO}_2}{2 \text{ mol CH}_3\text{OH}} = 2.75 \text{ mol CO}_2 \text{ max}$$

CH₃OH excess reagent

$$88 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32 \text{ g O}_2} \times \frac{2 \text{ mol CO}_2}{3 \text{ mol O}_2} = 1.83 \text{ mol CO}_2 \text{ max}$$

O₂ limiting reagent

$$1.83 \text{ mol } \text{CO}_2 \times \frac{2 \text{ mol } \text{CH}_3\text{OH}}{2 \text{ mol } \text{CO}_2} = 1.83 \text{ mol } \text{CH}_3\text{OH} \text{ used}$$

$$88 \text{ g } \text{CH}_3\text{OH} \times \frac{1 \text{ mol } \text{CH}_3\text{OH}}{32 \text{ g } \text{CH}_3\text{OH}} = 2.75 \text{ mol } \text{CH}_3\text{OH} \text{ available} - 1.83 \text{ mol used} = \mathbf{0.92 \text{ mol excess } \text{CH}_3\text{OH}}$$

e. 15 g of CH₃OH react with 12 g of O₂

$$15 \text{ g } \text{CH}_3\text{OH} \times \frac{1 \text{ mol}}{32 \text{ g}} = 0.47 \text{ mol } \text{CH}_3\text{OH} \text{ available} \dots \times (3 \text{ mol } \text{O}_2 / 2 \text{ mol } \text{CH}_3\text{OH}) = 0.71 \text{ mol } \text{O}_2 \text{ required}$$

$$12 \text{ g } \text{O}_2 \times \frac{1 \text{ mol}}{32 \text{ g}} = \text{only } 0.375 \text{ mol } \text{O}_2 \text{ available} \dots \text{ So } \text{O}_2 \text{ is limiting.}$$

All of the oxygen will react...

$$0.375 \text{ mol } \text{O}_2 \times \frac{2 \text{ mol } \text{CO}_2}{3 \text{ mol } \text{O}_2} = \mathbf{0.25 \text{ mol } \text{CO}_2 \text{ produced}}$$

$$0.375 \text{ mol } \text{O}_2 \times \frac{2 \text{ mol } \text{CH}_3\text{OH}}{3 \text{ mol } \text{O}_2} = 0.25 \text{ mol } \text{CH}_3\text{OH} \text{ used; } (0.47 - 0.25) = \mathbf{0.22 \text{ mole excess methanol}}$$

f. 25 g of CH₃OH react with 35 g of O₂

$$25 \text{ g } \text{CH}_3\text{OH} \times \frac{1 \text{ mol}}{32 \text{ g}} = 0.78 \text{ mol } \text{CH}_3\text{OH} \text{ available} \dots \times (3 \text{ mol } \text{O}_2 / 2 \text{ mol } \text{CH}_3\text{OH}) = 1.17 \text{ mol } \text{O}_2 \text{ required}$$

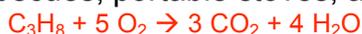
$$35 \text{ g } \text{O}_2 \times \frac{1 \text{ mol}}{32 \text{ g}} = \text{only } 1.09 \text{ mol } \text{O}_2 \text{ available} \dots \text{ So } \text{O}_2 \text{ is limiting.}$$

All of the oxygen will react...

$$1.09 \text{ mol } \text{O}_2 \times \frac{2 \text{ mol } \text{CO}_2}{3 \text{ mol } \text{O}_2} = \mathbf{0.73 \text{ mol } \text{CO}_2 \text{ produced}}$$

$$1.09 \text{ mol } \text{O}_2 \times \frac{2 \text{ mol } \text{CH}_3\text{OH}}{3 \text{ mol } \text{O}_2} = 0.73 \text{ mol } \text{CH}_3\text{OH} \text{ used; } (0.78 - 0.73) = \mathbf{0.05 \text{ mole excess methanol}}$$

II. Propane is by-product of natural gas processing and petroleum refining. It is commonly used as a fuel for engines, oxy-gas torches, barbecues, portable stoves, and residential central heating.



Note: C₃H₈ and CO₂ have different molecular formulas but happen to have similar molar masses.

a. What mass of CO₂ is produced when 6.5 g of propane is reacted with 14.2 g of O₂?

$$6.5 \text{ g } \text{C}_3\text{H}_8 \times \frac{1 \text{ mol}}{44 \text{ g}} = 0.15 \text{ mol } \text{C}_3\text{H}_8 \text{ available} \dots \times (5 \text{ mol } \text{O}_2 / 1 \text{ mol } \text{C}_3\text{H}_8) = 0.74 \text{ mol } \text{O}_2 \text{ required}$$

$$14.2 \text{ g } \text{O}_2 \times \frac{1 \text{ mol}}{32 \text{ g}} = \text{only } 0.44 \text{ mol } \text{O}_2 \text{ available} \dots \text{ So } \text{O}_2 \text{ is limiting.}$$

All of the oxygen will react...

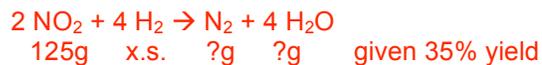
$$0.44 \text{ mol } \text{O}_2 \times \frac{3 \text{ mol } \text{CO}_2}{5 \text{ mol } \text{O}_2} \times \frac{44 \text{ g } \text{CO}_2}{1 \text{ mol } \text{CO}_2} = \mathbf{11.7 \text{ g } \text{CO}_2 \text{ produced (theoretical yield)}}$$

b. The actual yield of the reaction described above is 8.0 g of carbon dioxide. What is the percent yield?

Percent yield is how much you got out of how much you could *theoretically* get...

$$\frac{8.0 \text{ g } \text{CO}_2}{11.7 \text{ g } \text{CO}_2} \times 100\% = \mathbf{68\% \text{ yield}}$$

III. Nitrogen dioxide reacts with hydrogen to produce nitrogen and water. When 125g of nitrogen dioxide are allowed to react with excess hydrogen, the percent yield is 35%. How many grams of each product are actually formed during this process?



Convert limiting NO₂ to moles...

$$125\text{g NO}_2 \times \frac{1 \text{ mol}}{46\text{g}} = 2.72 \text{ mol NO}_2 \text{ reacts}$$

Calculate theoretical yield of nitrogen, in grams, and take 35% of that value.

$$2.72 \text{ mol NO}_2 \times \frac{1 \text{ mol N}_2}{2 \text{ mol NO}_2} \times \frac{28 \text{ g N}_2}{1 \text{ mol N}_2} \times 0.35 = \mathbf{13.3 \text{ g N}_2 \text{ produced}}$$

Same for water

$$2.72 \text{ mol NO}_2 \times \frac{4 \text{ mol H}_2\text{O}}{2 \text{ mol NO}_2} \times \frac{18 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \times 0.35 = \mathbf{34.3 \text{ g H}_2\text{O produced}}$$