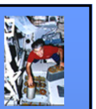
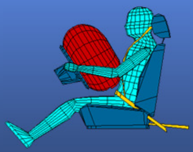
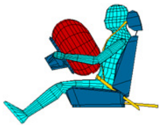


Stoichiometry in the Real World



Air Bag Design



- Exact quantity of nitrogen gas must be produced in an instant.
- Use a *catalyst* to speed up the reaction

$$2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$$

$$6\ Na(s) + Fe_2O_3(s) \rightarrow 3\ Na_2O(s) + 2\ Fe\ (s)$$

$$2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$$

$$6\ Na(s) + Fe_2O_3(s) \rightarrow 3\ Na_2O(s) + 2\ Fe(s)$$

Airbag Design

Assume that 65.1 L of N_2 gas are needed to inflate an air bag to the proper size. How many grams of NaN_3 must be included in the gas generant to generate this amount of N_2?
(Hint: The density of N_2 gas at this temperature is about 0.916 g/L).

$$\begin{array}{r} 65.1\ \text{L}\ N_2 \\ \underline{\times\ 0.916\ \text{g/L}\ N_2} \\ 59.6\ \text{g}\ N_2 \end{array}$$

$$\text{x g } NaN_3 = 59.6\ \cancel{\text{g } N_2}\left(\frac{1\ \cancel{\text{mol } N_2}}{28\ \cancel{\text{g } N_2}}\right)\left(\frac{2\ \cancel{\text{mol } NaN_3}}{3\ \cancel{\text{mol } N_2}}\right)\left(\frac{65\ \text{g } NaN_3}{1\ \cancel{\text{mol } NaN_3}}\right)$$

X = 92.2 g NaN_3

How much Fe_2O_3 must be added to the gas generant for this amount of NaN_3?

$$\text{x g } Fe_2O_3 = 92.2\ \text{g } \cancel{NaN_3}\left(\frac{1\ \cancel{\text{mol } NaN_3}}{65\ \cancel{\text{g } NaN_3}}\right)\left(\frac{2\ \cancel{\text{mol Na}}}{2\ \cancel{\text{mol } NaN_3}}\right)\left(\frac{1\ \cancel{\text{mol } Fe_2O_3}}{6\ \cancel{\text{mol Na}}}\right)\left(\frac{159.6\ \text{g } Fe_2O_3}{1\ \cancel{\text{mol } Fe_2O_3}}\right)$$

X = 37.7 g Fe_2O_3

Water from a Camel

Camels store the fat tristearin ($C_{57}H_{110}O_6$) in the hump. As well as being a source of energy, the fat is a source of water, because when it is used the reaction

$$2\ C_{57}H_{110}O_6(s) + 163\ O_2(g) \rightarrow 114\ CO_2(g) + 110\ H_2O(l)$$

takes place. What mass of water can be made from 1.0 kg of fat?

$$\text{x g } H_2O = 1\ \cancel{\text{kg 'fat''}}\left(\frac{1000\ \cancel{\text{g ``fat''}}}{1\ \cancel{\text{kg ``fat''}}}\right)\left(\frac{1\ \cancel{\text{mol ``fat''}}}{890\ \cancel{\text{g ``fat''}}}\right)\left(\frac{110\ \cancel{\text{mol } H_2O}}{2\ \cancel{\text{mol ``fat''}}}\right)\left(\frac{18\ \text{g } H_2O}{1\ \cancel{\text{mol } H_2O}}\right)$$

X = 1112 g H_2O

or 1.112 liters water

Rocket Fuel

The compound diborane (B_2H_6) was at one time considered for use as a rocket fuel. How many grams of liquid oxygen would a rocket have to carry to burn 10 kg of diborane completely?
(The products are B_2O_3 and H_2O).

Chemical equation $B_2H_6 + O_2 \longrightarrow B_2O_3 + H_2O$

Balanced chemical equation $\mathbf{B_2H_6 + 3\ O_2 \longrightarrow B_2O_3 + 3\ H_2O}$

10 kg *x g*

$$\text{x g } O_2 = 10\ \cancel{\text{kg } B_2H_6}\left(\frac{1000\ \cancel{\text{g } B_2H_6}}{1\ \cancel{\text{kg } B_2H_6}}\right)\left(\frac{1\ \cancel{\text{mol } B_2H_6}}{28\ \cancel{\text{g } B_2H_6}}\right)\left(\frac{3\ \cancel{\text{mol } O_2}}{1\ \cancel{\text{mol } B_2H_6}}\right)\left(\frac{32\ \text{g } O_2}{1\ \cancel{\text{mol } O_2}}\right)$$

X = 34,286 g O_2

Water in Space

Click Here

In the space shuttle, the CO_2 that the crew exhales is removed from the air by a reaction within canisters of lithium hydroxide. On average, each astronaut exhales about 20.0 mol of CO_2 daily. What volume of water will be produced when this amount of CO_2 reacts with an excess of LiOH?
(Hint: The density of water is about 1.00 g/mL.)

$$\mathbf{CO_2(g) + 2\ LiOH(s) \rightarrow Li_2CO_3(aq) + H_2O(l)}$$

20.0 mol *excess* x g

Water is NOT at STP!

$$\text{x mL } H_2O = 20.0\ \cancel{\text{mol } CO_2}\left(\frac{1\ \cancel{\text{mol } H_2O}}{1\ \cancel{\text{mol } CO_2}}\right)\left(\frac{\text{22.4 L / 18 g}\ \cancel{H_2O}}{1\ \cancel{\text{mol } H_2O}}\right)\left(\frac{1\ \text{mL } H_2O}{1\ \cancel{\text{g } H_2O}}\right)$$

X = 360 mL H_2O

Lithium Hydroxide Scrubber

Modified by Apollo 13 Mission

Astronaut John L. Swigert holds the jury-rigged lithium hydroxide scrubber used to remove excess carbon dioxide from the damaged Apollo 13 spacecraft.

Water in Space

In the space shuttle, the CO_2 that the crew exhales is removed from the air by a reaction within canisters of lithium hydroxide. On average, each astronaut exhales about 20.0 mol of CO_2 daily. What volume of water will be produced when this amount of CO_2 reacts with an excess of LiOH? (Hint: The density of water is about 1.00 g/mL.)

$$\mathbf{CO_2(g) + 2\,LiOH(s) \rightarrow Li_2CO_3(aq) + H_2O(l)}$$

20.0 mol — *excess* — x g

Water is NOT at STP!

$$x\text{ mL }H_2O = 20.0\ \cancel{\text{mol }CO_2}\left(\frac{1\ \cancel{\text{mol }H_2O}}{1\ \cancel{\text{mol }CO_2}}\right)\left(\frac{\text{22.4 L }H_2O\ \ \text{18 g }H_2O}{1\ \cancel{\text{mol }H_2O}}\right)\left(\frac{1\text{ mL }H_2O}{1\ \cancel{\text{g }H_2O}}\right)$$

$$X = 360\text{ mL }H_2O$$

Real Life Problem Solving

Determine the amount of LiOH required for a seven-day mission in space for three astronauts and one 'happy' chimpanzee. Assume each passenger expels 20 mol of CO_2 per day.

Note: The lithium hydroxide scrubbers are only 85% efficient.

(4 passengers) x (10 days) x (20 mol/day) = 800 mol CO_2

Plan for a delay

$$\mathbf{CO_2(g) + 2\,LiOH(s) \rightarrow Li_2CO_3}$$

800 mol — X g

$$\mathbf{CO_2(g) + 2\,LiOH(s) \rightarrow Li_2CO_3(aq) + H_2O(l)}$$

$$\begin{tikzcd}
800\text{ mol} \arrow[d] & 38{,}240\text{ g} \\
800\text{ mol} \arrow[r, dash, "1:2"] & 1600\text{ mol} \arrow[u, "\times 23.9\text{ g/mol}"']
\end{tikzcd}$$

Needed (actual yield)

$$X\text{ g LiOH} = 800\ \cancel{\text{mol }CO_2}\left(\frac{2\ \cancel{\text{mol LiOH}}}{1\ \cancel{\text{mol }CO_2}}\right)\left(\frac{23.9\text{ g LiOH}}{1\ \cancel{\text{mol LiOH}}}\right) = 38{,}240\text{ g LiOH}$$

Note: The lithium hydroxide scrubbers are only 85% efficient.

$$\%\text{ Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \qquad 0.85 = \frac{38{,}240\text{ g LiOH}}{x\text{ g LiOH}}$$

Amount of LiOH to be taken into space → x = 44,988 g LiOH

Careers in Chemistry: Farming

Farming is big business in the United States with profits for the lucky and possible bankruptcy for the less fortunate. *Farmers should not be ignorant of chemistry*. For instance, to be profitable, a farmer must know when to plant, harvest, and sell his/her crops to maximize profit. In order to get the greatest yield farmers often add *fertilizers* to the soil to replenish vital nutrients removed by the previous season's crop.

Corn is one product that removes a tremendous amount of phosphorous from the soil. For this reason, farmers will rotate crops and/or add fertilizer to the ground before planting crops for the following year. On average, an acre of corn will remove 6 kilograms of phosphorous from the ground.

Assume you inherit a farm and must now have to purchase fertilizer for the farm. The farm is 340 acres and had corn planted the previous year. You must add fertilizer to the soil before you plant this years' crop. You go to the local fertilizer store and find **SuperPhosphate**™ brand fertilizer. You read the fertilizer bag and can recognize from your high school chemistry class a molecular formula $Ca_3P_2H_{14}S_2O_{21}$ (*you don't understand anything else written on the bag because it is imported fertilizer from Japan*). You must decide how much fertilizer to buy for application to your corn fields. If each bag costs $54.73; how many bags of fertilizer must you purchase and how much will it cost you to add the necessary fertilizer to your fields?

Given: 1 bag of fertilizer weighs 10,000 g [454 g = 1 pound]

Careers in Chemistry: Farming

How much fertilizer will you need?

Conversion Factor: 1 acre corn = 6 kg phosphorous

$$x\text{ g P} = 340\ \cancel{\text{acres}}\left(\frac{6\ \cancel{\text{kg P}}}{1\ \cancel{\text{acre}}}\right)\left(\frac{1000\text{ g P}}{1\ \cancel{\text{kg P}}}\right) = 2.04 \times 10^6\text{ g P}$$

If a bag of fertilizer has the formula $Ca_3P_2H_{14}S_2O_{21}$,
The molar mass of it is 596 g/mol.

3 Ca @ 40g/mol = 120 g
2 P@ 31 g/mol = 62 g
14 H@ 1 g/mol = 14 g
2 S@ 32 g/mol = 64 g
21 O @ 16 g/mol = 335 g
$Ca_3P_2H_{14}S_2O_{21}$ = 596 g

$$\%\text{ P} = \frac{\text{part } 62\text{ g}}{\text{whole } 596\text{ g}} \times 100\ \%$$

10.4 % Phosphorous

In a bag of fertilizer you have 10.4 % (by mass) phosphorous.
A bag of fertilizer weighs 10,000 g (about 22 pounds).

10.4 % of 10,000 g = 1040 g phosphorous / bag of fertilizer

$$\frac{2.04 \times 10^6\text{ g P}}{1040\text{ g/bag}} = 1962\text{ bags of fertilizer}$$

Total Cost (1962 bags of fertilizer)($54.73 / bag) = $107,380

Careers in Chemistry: Dentistry

We learned that fluoride is an essential element to be taken to reduce teeth cavities. Too much fluoride can produce yellow spots on the teeth and too little will have no effect. After years of study it was determined that a quantity of 1 part per million (ppm) fluoride in the water supply is enough to significantly reduce cavities and not stain teeth yellow.

Measure the mass of the mineral *fluorite* (chemically, CaF_2). Use this sample to determine how much water must be added to yield a 1 ppm fluoride solution. Sounds difficult? Lets apply what we've learned this unit to solve this problem.

1 part per million = 1 atom of fluorine per 999,999 water molecules

What information do we know:

1 mol CaF_2 = 78.08 g CaF_2 = 6.02 x 10^{23} molecules of CaF_2
1 molecules of CaF_2 = 2 atoms of F
1 mol H_2O = 18 g H_2O Density of water is 1 g/mL
1000 mL = 1 L and 3.78 L = 1 gallon
mass of sample of CaF_2 = 92.135 g

Careers in Chemistry: Dentistry


Calcium Fluoride

$$\text{x atoms F} = 92.135 \text{ g} \cancel{CaF_2}\left(\frac{1\ \cancel{\text{mol } CaF_2}}{78\ \cancel{\text{g } CaF_2}}\right)\left(\frac{6.02 \text{ x } 10^{23}\ \cancel{\text{molecules } CaF_2}}{1\ \cancel{\text{mol } CaF_2}}\right)\left(\frac{2 \text{ atoms F}}{1\ \cancel{\text{molecules } CaF_2}}\right) = 1.42 \text{ x } 10^{24} \text{ atoms F}$$

$$\text{x gallons } H_2O = 1.42 \text{ x } 10^{24}\ \cancel{\text{F atoms}}\left(\frac{999{,}999\ \cancel{H_2O \text{ molecules}}}{1\ \cancel{\text{F atom}}}\right)\left(\frac{1\ \cancel{\text{mol } H_2O}}{6.02 \text{ x } 10^{23}\ \cancel{H_2O \text{ molecules}}}\right)\left(\frac{18\ \cancel{\text{g } H_2O}}{1\ \cancel{\text{mol } H_2O}}\right)\left(\frac{1\ \cancel{\text{mL } H_2O}}{1\ \cancel{\text{g } H_2O}}\right)\left(\frac{1\ \cancel{\text{L } H_2O}}{1000\ \cancel{\text{mL } H_2O}}\right)\left(\frac{1 \text{ gallon } H_2O}{3.78\ \cancel{\text{L } H_2O}}\right) =$$

Need 11,238 gallons of water needed to dissolve 91.235 g CaF_2 to yield a 1 ppm F^{1-} solution.

Energy with Stoichiometry

Given: 1 mol O_2 yields 350 kJ

methane + $\underline{\text{oxygen}}$ ⟶ carbon dioxide + water + energy

Excess Limiting

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O + 700 \text{ kJ}$$

100 g 100 g ? kJ

/16 g/mol /32 g/mol

$$\frac{6.25 \text{ mol } CH_4}{1} \qquad \frac{3.125 \text{ mol } O_2}{2}$$

6.25 1.56

smaller number
is limiting reactant

$$\text{x kJ} = 3.125\ \cancel{\text{mol } O_2}\left(\frac{700 \text{ kJ}}{2\ \cancel{\text{mol } O_2}}\right) = 1094 \text{ kJ}$$