

# Gas Stoichiometry

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YALE UNIVERSITY  
CHEMISTRY 161  
FALL 2019

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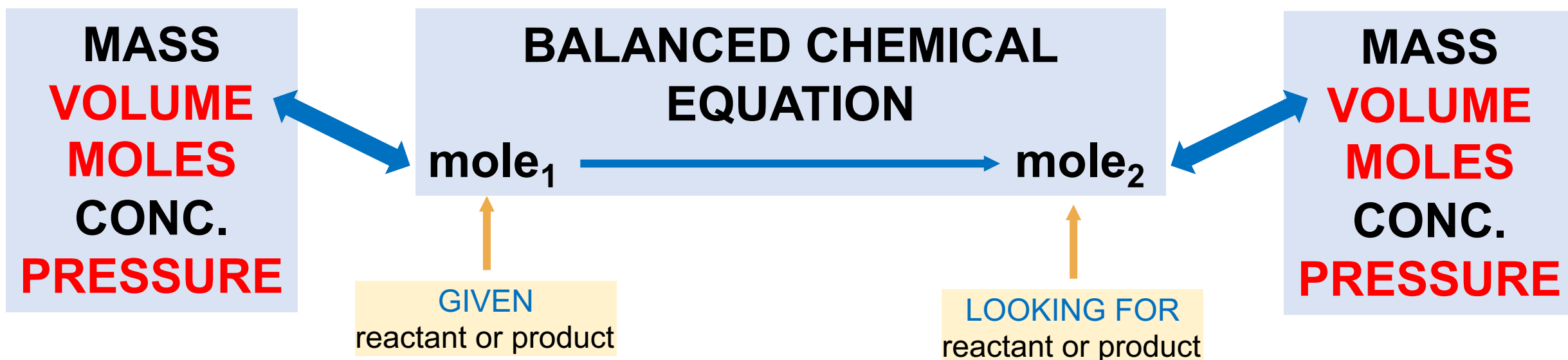
**So far we've assumed that the mixtures of gases do not react chemically with each other.**

**But we can also consider gas properties when gases do react.**

**This is just regular stoichiometry but with gases!**

# SUMMARIZING STOICHIOMETRY RELATIONSHIPS

We can add gases (volume/moles/pressure) now!



*I hope now you understand why I say to convert to moles before you do anything else. It's because a balanced chemical equation gives us **mole-to-mole ratios** that we can use to convert between one reactant/product to another reactant/product.*

## A Guided Example

A rigid (fixed volume) piston contains 2.00 g of ammonia gas and 2.50 g of oxygen gas at 325 K. The volume of the locked piston is currently 3.00 L.

Before any reaction takes place, what is the total pressure inside the piston?

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$$n_{\text{NH}_3} = 2.00 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.034 \text{ g NH}_3} = 0.117_4 \text{ mol NH}_3$$

$$n_{\text{O}_2} = 2.50 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 0.0780_1 \text{ mol O}_2$$

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$$P_{\text{total}} = \frac{(n_{\text{NH}_3} + n_{\text{O}_2})RT}{V}$$
$$= \frac{(0.117_4 \text{ mol} + 0.0780_1 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (325 \text{ K})}{3.00 \text{ L}}$$

$$P_{\text{total}} = 1.74 \text{ atm}$$

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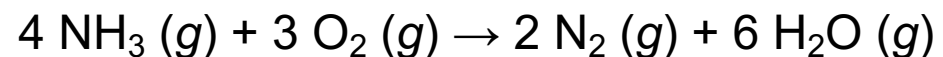


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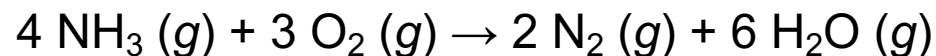


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Recall that we have the number of moles of each reactant:

$$n_{\text{NH}_3} = 2.00 \text{ g NH}_3 \times \frac{1 \text{ mol NH}_3}{17.034 \text{ g NH}_3} = 0.1174 \text{ mol NH}_3$$

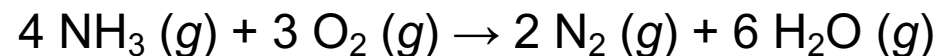
$$n_{\text{O}_2} = 2.50 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 0.07801 \text{ mol O}_2$$

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Q: Which is the limiting reactant?

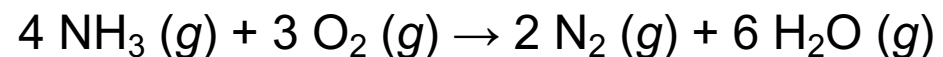
A:  $\text{O}_2 (g)$  is the limiting reactant. *I'll leave you to figure that part out.*

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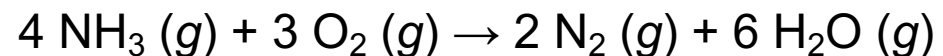
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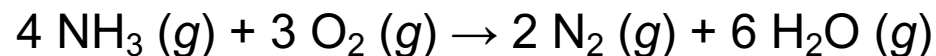
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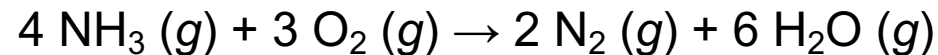
$$n_{\text{N}_2} = 0.0780_1 \text{ mol O}_2 \times \frac{2 \text{ mol N}_2}{3 \text{ mol O}_2} = 0.0520_8 \text{ mol N}_2$$

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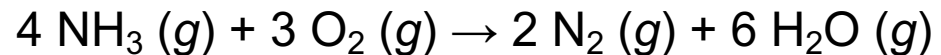
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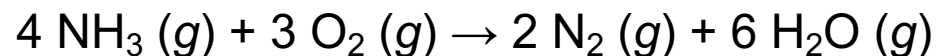


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$$n_{\text{H}_2\text{O}} = 0.0780_1 \text{ mol O}_2 \times \frac{6 \text{ mol H}_2\text{O}}{3 \text{ mol O}_2} = 0.156_3 \text{ mol H}_2\text{O}$$

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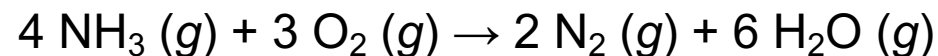
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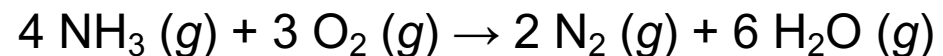
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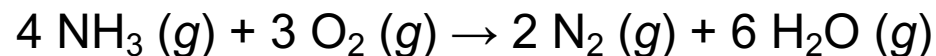
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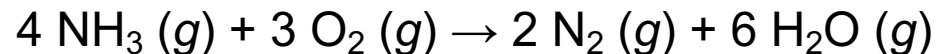
$$n_{\text{NH}_3} = \underbrace{0.117_4 \text{ mol NH}_3}_{\text{Starting Amount}} - \underbrace{\left(0.0780_1 \text{ mol O}_2 \times \frac{4 \text{ mol NH}_3}{3 \text{ mol O}_2}\right)}_{\text{Amount Reacted}} = \underbrace{0.0132_5 \text{ mol NH}_3}_{\text{Amount leftover}}$$

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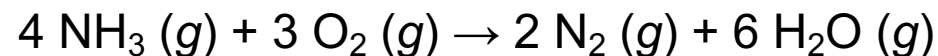
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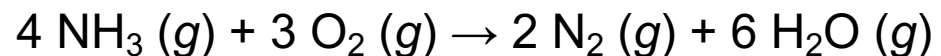
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Determine final volume:

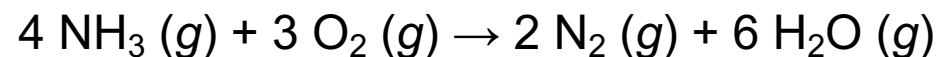
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Determine amount of NH<sub>3</sub> (g) remaining:  $n_{\text{NH}_3} = 0.0132_5 \text{ mol NH}_3$

Determine final volume:

$$V = \frac{n_{\text{total}}RT}{P}$$

$$= \frac{(0.221_6 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (325 \text{ K})}{1.00 \text{ atm}}$$

$$n_{\text{total}} = n_{\text{N}_2} + n_{\text{H}_2\text{O}} + n_{\text{NH}_3}$$

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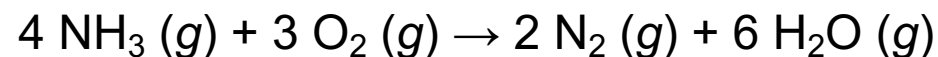


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Q: Why is  $P = 1.00 \text{ atm}$  now?

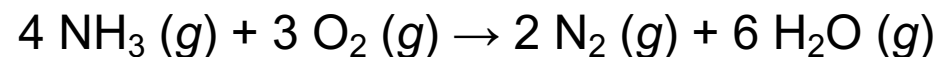
A: Because volume is allowed to change and pressure equilibrates with the atmosphere.

## A Guided Example

A rigid (fixed volume) piston contains 2.00 g of ammonia gas and 2.50 g of oxygen gas at 325 K. The volume of the locked piston is currently 3.00 L.

The piston is now unlocked so the top can move freely and the volume can change. The two reactants (ammonia and oxygen gases) react to produce nitrogen gas and water vapor. What will be the new volume after the reaction takes place at 325 K?

Start by writing a balanced chemical equation for this reaction:



Recall that we have the number of moles of each reactant: O<sub>2</sub> limiting!

Determine amount of N<sub>2</sub> (g) produced:  $n_{\text{N}_2} = 0.0520_8 \text{ mol N}_2$

Determine amount of H<sub>2</sub>O (g) produced:  $n_{\text{H}_2\text{O}} = 0.156_3 \text{ mol H}_2\text{O}$

Determine amount of NH<sub>3</sub> (g) remaining:  $n_{\text{NH}_3} = 0.0132_5 \text{ mol NH}_3$

$$\left. \begin{array}{l} n_{\text{N}_2} = 0.0520_8 \text{ mol N}_2 \\ n_{\text{H}_2\text{O}} = 0.156_3 \text{ mol H}_2\text{O} \\ n_{\text{NH}_3} = 0.0132_5 \text{ mol NH}_3 \end{array} \right\} \begin{array}{l} n_{\text{total}} = n_{\text{N}_2} + n_{\text{H}_2\text{O}} + n_{\text{NH}_3} \\ = 0.221_6 \text{ mol} \end{array}$$

Determine final volume:

$$V = \frac{n_{\text{total}}RT}{P}$$

$$= \frac{(0.221_6 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (325 \text{ K})}{1.00 \text{ atm}}$$

$$V = 5.91 \text{ L}$$



**What volume of oxygen is required to produce 2.33 moles of H<sub>2</sub>O at 1.0 atm and 298 K?**

First, use mole-to-mole ratios to find the number of moles of oxygen required:

$$2.33 \text{ mol H}_2\text{O} \times \frac{6 \text{ mol O}_2}{4 \text{ mol H}_2\text{O}} = 3.49_5 \text{ mol O}_2$$



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Now use the ideal gas law to find the volume of oxygen:

$$V = \frac{n_{\text{O}_2}RT}{P}$$
$$= \frac{(3.49_5 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})}{1.0 \text{ atm}}$$

$$V = 85 \text{ L}$$

**Consider the following unbalanced chemical equation:**



**What pressure of H<sub>2</sub>S gas is required to produce 55.0 g of solid sulfur?**

**Assume that SO<sub>2</sub> is in excess and that the reaction is conducted at 375 K and 29.3 L.**

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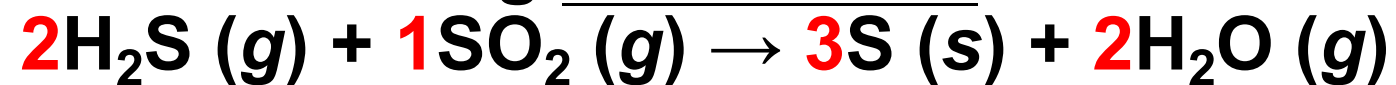


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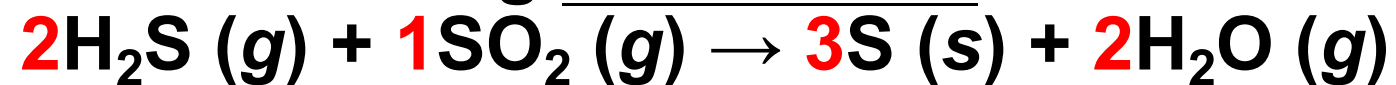


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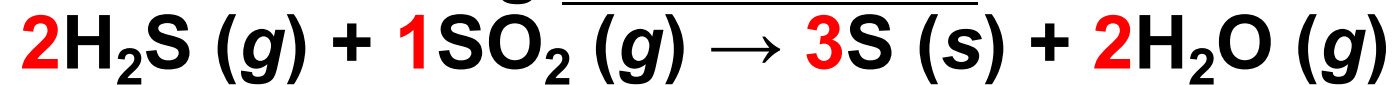
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Then, use mole-to-mole ratios to find the number of moles of  $\text{H}_2\text{S}$  required:

$$55.0 \text{ g S} \times \frac{1 \text{ mol S}}{32.06 \text{ g S}} \times \frac{2 \text{ mol H}_2\text{S}}{3 \text{ mol S}} = 1.14_4 \text{ mol H}_2\text{S}$$



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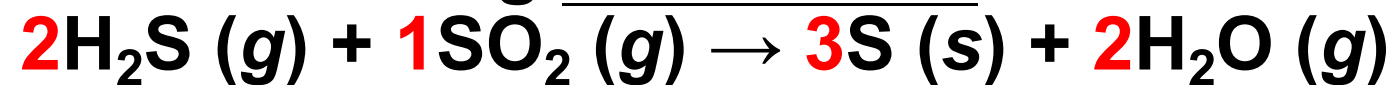
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Now use the ideal gas law to find the pressure of  $\text{H}_2\text{S}$ :

$$P = \frac{n_{\text{H}_2\text{S}}RT}{V}$$
$$= \frac{(1.14_4 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (375 \text{ K})}{29.3 \text{ L}}$$

$$P = 1.20 \text{ atm}$$