Gas Stoichiometry

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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!

MASS
VOLUME
MOLES
CONC.
PRESSURE

BALANCED CHEMICAL EQUATION

$\text{mole}_1 \longrightarrow \text{mole}_2$

GIVEN
reactant or product

LOOKING FOR
reactant or product

MASS
VOLUME
MOLES
CONC.
PRESSURE

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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$$n_{NH_3} = 2.00 \text{ g } NH_3 \times \frac{1 \text{ mol } NH_3}{17.03_4 \text{ g } NH_3} = 0.117_4 \text{ mol } NH_3$$

$$n_{O_2} = 2.50 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} = 0.0780_1 \text{ mol } O_2$$

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$$n_{O_2} = 2.50\text{ g O}_2 \times \frac{1\text{ mol O}_2}{32.00\text{ g O}_2} = 0.0780_1\text{ mol O}_2$$

$$P_{total} = \frac{(n_{NH_3} + n_{O_2})RT}{V}$$

$$= \frac{(0.117_4\text{ mol} + 0.0780_1\text{ mol})\left(0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(325\text{ K})}{3.00\text{ L}}$$

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

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$$n_{O_2} = 2.50\ \mathrm{g\ O_2} \times \frac{1\ \mathrm{mol\ O_2}}{32.00\ \mathrm{g\ O_2}} = 0.0780_1\ \mathrm{mol\ O_2}$$

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Q: Which is the limiting reactant?

A: O_2 (*g*) is the limiting reactant. *I'll leave you to figure that part out.*

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Determine amount of N_2 (*g*) produced:

$$n_{N_2} = 0.0780_1 \text{ mol } O_2 \times \frac{2 \text{ mol } N_2}{3 \text{ mol } O_2} = 0.0520_8 \text{ mol } N_2$$

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Determine amount of H_2O (*g*) produced:

$$n_{H_2O} = 0.0780_1 \text{ mol } O_2 \times \frac{6 \text{ mol } H_2O}{3 \text{ mol } O_2} = 0.156_3 \text{ mol } H_2O$$

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Determine amount of NH_3 (*g*) remaining:

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Determine amount of NH_3 (*g*) remaining:

$$n_{NH_3} = \underbrace{0.117_4 \text{ mol } NH_3}_{\text{Starting Amount}} - \underbrace{\left(0.0780_1 \text{ mol } O_2 \times \frac{4 \text{ mol } NH_3}{3 \text{ mol } O_2}\right)}_{\text{Amount Reacted}} = \underbrace{0.0132_5 \text{ mol } NH_3}_{\text{Amount leftover}}$$

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Determine amount of NH_3 (*g*) remaining: $n_{NH_3} = 0.0132_5 \text{ mol } NH_3$

$$n_{total} = n_{N_2} + n_{H_2O} + n_{NH_3} = 0.221_6 \text{ mol}$$

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$$n_{total} = n_{N_2} + n_{H_2O} + n_{NH_3} = 0.221_6 \text{ mol}$$

Determine final volume:

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

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Q: Why is P = 1.00 atm now?
A: Because volume is allowed to change and pressure equilibrates with the atmosphere.

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$$n_{total} = n_{N_2} + n_{H_2O} + n_{NH_3} = 0.221_6 \text{ mol}$$

Determine final volume:

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

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

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

Consider the following balanced chemical equation:
C_3H_7SH (*l*) + $6O_2$ (*g*) → $3CO_2$ (*g*) + SO_2 (*g*) + $4H_2O$ (*g*)

What volume of oxygen is required to produce 2.33 moles of H_2O at 1.0 atm and 298 K?

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

$$2.33 \text{ mol } H_2O \times \frac{6 \text{ mol } O_2}{4 \text{ mol } H_2O} = 3.49_5 \text{ mol } O_2$$

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$C_3H_7SH\ (l) + 6O_2\ (g) \rightarrow 3CO_2\ (g) + SO_2\ (g) + 4H_2O\ (g)$

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

$$V = \frac{n_{O_2}RT}{P}$$

$$= \frac{(3.49_5\ \text{mol})\left(0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298\ \text{K})}{1.0\ \text{atm}}$$

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

Consider the following unbalanced chemical equation:

$$H_2S\ (g) + SO_2\ (g) \rightarrow S\ (s) + H_2O\ (g)$$

What pressure of H_2S gas is required to produce 55.0 g of solid sulfur?

Assume that SO_2 is in excess and that the reaction is conducted at 375 K and 29.3 L.

Consider the following unbalanced chemical equation:

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First, balance the chemical equation!

Consider the following unbalanced chemical equation:

$$2H_2S\ (g) + 1SO_2\ (g) \rightarrow 3S\ (s) + 2H_2O\ (g)$$

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Then, use mole-to-mole ratios to find the number of moles of H_2S required:

$$55.0 \text{ g S} \times \frac{1 \text{ mol S}}{32.06 \text{ g S}} \times \frac{2 \text{ mol } H_2S}{3 \text{ mol S}} = 1.14_4 \text{ mol } H_2S$$

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Now use the ideal gas law to find the pressure of H_2S:

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Assume that SO_2 is in excess and that the reaction is conducted at 375 K and 29.3 L.

First, balance the chemical equation!

Then, use mole-to-mole ratios to find the number of moles of H_2S required:

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

Now use the ideal gas law to find the pressure of H_2S:

$$P = \frac{n_{H_2S}RT}{V}$$

$$= \frac{(1.14_4\ \text{mol})\left(0.08206\ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(375\ \text{K})}{29.3\ \text{L}}$$

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