Chemistry 6A F2007

Dr. J.A. Mack

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Stoichiometric balancing coefficients

They are the numbers in front of the chemical formulas.

They give the ***ratio*** of reactants and products.

ratio $\longrightarrow$ ***Conversion factors***

The balancing coefficients allow us to convert between numbers of reactants and products.

Stoichiometry: The branch of chemistry that deals with the mole proportions of chemical reactions.

Stoichiometric ratio: The ratio of any two species (reactants or products) in a balanced chemical reaction.

$$2A + 3B \longrightarrow A_2B_3$$

2 A's combine with 3B's

CONVERSION FACTORS!!!

$$\frac{2A}{3B} \qquad \frac{3B}{2A}$$

Since the individual ratios must scale to moles, we can write:

$$\frac{2A}{3B} \longrightarrow \frac{2\text{ mols A}}{3\text{ mols B}}$$

The molar ratios allow us to relate the amounts of reactants and products in a chemical equation.

i.e.

$$3\ \cancel{\text{mols}}\ \text{B} \times \frac{2\text{ mols A}}{3\ \cancel{\text{mols}}\ \text{B}} = 2\text{ mols A}$$

Quantitative calculations: Mass and moles

Consider the following reaction:

$$HCl\ (aq) + Ba(OH)_2\ (aq) \rightarrow H_2O\ (l) + $$

Balacing:

$$\mathbf{2}\ HCl\ (aq) + Ba(OH)_2\ (aq) \rightarrow \mathbf{2}H_2O\ (l) + B$$

This means that there is more than enough $Ba(OH)_2$ to react all of the HCl.

How many moles of HCl are consumed if 1.50 g of $BaCl_2$ are produced assuming that $Ba(OH)_2$ is in excess?

Solution:

convert g $BaCl_2$ *to* → mol $BaCl_2$ *then convert* → *to* mol HCl

using the Molar mass of $BaCl_2$ *using the Molar ratio for equation*

$$\mathbf{2}HCl\ (aq) + Ba(OH)_2\ (aq) \rightarrow \mathbf{2}H_2O\ (l) + BaCl_2(aq)$$

the calculation follows the sequence…

g $BaCl_2$ → mol $BaCl_2$ → mol HCl

$$1.50\ g\ BaCl_2 \times \frac{1\ mol\ BaCl_2}{208.24\ g\ BaCl_2} \times \frac{2\ mol\ HCl}{1\ mol\ BaCl_2} = 0.0144\ mol\ HCl$$

Molar mass $BaCl_2$ Molar ratio for equation

Calcium carbonate decomposes to form carbon dioxide and calcium oxide

If 3.55 mol of calcium carbonate decomposes, how many mols of carbon dioxide will form?

Step 1: Put your calculator down!

Step 2: Write the balanced chemical equation.

$$\boxed{1\ CaCO_3(s)} \longrightarrow \boxed{1\ CO_2(g)} + CaO(s)$$

Step 3: use the molar ratios to convert the moles

$$3.55mol\ CaCO_3 \times \frac{1mol\ CO_2}{1mol\ CaCO_3} = 3.55mol\ CO_2$$

We can also determine the number of grams of each product that will be produced:

mols reactant → mols product → mass product

stoichiometric molar ratios *molar mass*

$$3.55\cancel{mol\ CaCO_3} \times \frac{1\cancel{mol\ CO_2}}{1\cancel{mol\ CaCO_3}} \times \frac{44.01g\ CO_2}{1\cancel{mol\ CO_2}} =$$

$$= 156g\ CO_2$$

If 3.50g of calcium oxide remain after the decomposition of calcium carbonate, how many moles of calcium carbonate must have been initially present?

Step 1: Write the balanced chemical equation

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Step 2: Outline your conversion

$$\text{g CaO} \longrightarrow \text{mols CaO} \longrightarrow \text{mols } CaCO_3$$

$$3.50\cancel{\text{g CaO}} \times \frac{1\cancel{\text{mol CaO}}}{56.08\cancel{\text{g CaO}}} \times \frac{1\text{mol } CaCO_3}{1\cancel{\text{mol CaO}}} = 0.0624\text{mol } CaCO_3$$

(a) How many grams of nitrogen will be needed to produce 0.384 mols of ammonia?

$$N_2(g) + 3\ H_2(g) \longrightarrow 2\ NH_3(g)$$

$$0.384 \text{ mols } NH_3 \times \frac{1\text{mol } N_2}{2\text{mol } NH_3} \times \frac{28.02\text{g } N_2}{1\text{mol } N_2} = 5.38\text{g } N_2$$

(b) How many moles of hydrogen are needed to combine with 5.84g of nitrogen?

$$5.38\text{g } N_2 \times \frac{1\text{mol } N_2}{28.02\text{ g}N_2} \times \frac{3\text{mol } H_2}{1\text{mol } N_2} = 0.576 \text{ mols } H_2$$

How many grams of nitrogen are needed to completely react with 0.525 g of hydrogen in the formation of ammonia?

To answer this question, one must go through moles.

$$N_2(g) + 3\ H_2(g) \longrightarrow 2\ NH_3(g)$$

The equation relates moles, not mass.

$$\text{g } H_2 \longrightarrow \text{mols } H_2 \longrightarrow \text{mols } N_2 \longrightarrow \text{g } N_2$$

How many grams of nitrogen are needed to completely react with 0.525 g of hydrogen in the formation of ammonia?

$$N_2(g) + 3\ H_2(g) \longrightarrow 2\ NH_3(g)$$

$$0.525\text{g}H_2 \times \frac{1 \text{ mol } H_2}{2.02\text{g } H_2} \times \frac{1\text{mol } N_2}{3\text{mol } H_2} \times \frac{28.02\text{g } N_2}{1\text{mol } N_2} = 2.43\text{g } N_2$$

How many grams of ammonia form?

$$0.525\text{g}H_2 \times \frac{1 \text{ mol } H_2}{2.02\text{g } H_2} \times \frac{2 \text{ mol } NH_3}{3\text{mol } H_2} \times \frac{17.04\text{g } NH_3}{1\text{mol } NH_3} = 2.95\text{g } NH_3$$

Stoichiometry

Excess reagent (reactant): The reactant that is not completely consumed during a chemical reaction.

Limiting reagent (reactant): The reactant that is completely consumed in a chemical reaction, while one or more other reactants is not consumed.

When propane (C_3H_8) burns the products are carbon dioxide and water. How many grams of water will result form the burning of 5.05g propane in excess oxygen.?

Where do you start?

Burning = combustion

combustion = reacting with oxygen

Write and balance the equation!

$$C_3H_8(g) + 5\ O_2(g) \longrightarrow 3\ CO_2(g) + 4\ H_2O(l)$$

When propane (C_3H_8) burns the products are carbon dioxide and water. How many grams of water will result form the burning of 5.05g propane in excess oxygen.?

$$C_3H_8(g) + 5\ O_2(g) \longrightarrow 3CO_2(g) + 4\ H_2O(l)$$

Excess oxygen means more than enough to react all of the propane.

The amount of water produced is determined by the amount of propane that reacts.

$$\text{g } C_3H_8 \longrightarrow \text{mols } C_3H_8 \longrightarrow \text{mols } H_2O \longrightarrow \text{g } H_2O$$

$$5.05\text{g}C_3H_8 \times \frac{1\text{mol } C_3H_8}{44.11\text{g } C_3H_8} \times \frac{4\text{mol } H_2O}{1\text{mol } C_3H_8} \times \frac{18.02\text{g } H_2O}{1\text{mol } H_2O} = 8.25\text{g } H_2O$$

What happens when we we don't have an excess of oxygen? Will the same amount of water be formed?

The answer to that question is: Depends…

It depends upon the relative amounts of each reactant.
If one reactant is used up before the other is completely consumed,
Then it ***limits*** the reaction: it is the ***limiting reactant!***

When propane (C_3H_8) burns the products are carbon dioxide and water. How many grams of water will result form the burning of 5.05g propane with 10.1g of oxygen.?

Which one limits, the propane or the oxygen!

When propane (C_3H_8) burns the products are carbon dioxide and water. How many grams of water will result form the burning of 5.05g propane with 10.1g of oxygen.?

The limiting reactant will be that reactant which yields the smallest amount of an in common product.

$$C_3H_8(g) + 5\ O_2(g) \longrightarrow 3CO_2(g) + 4\ H_2O(l)$$

Convert: g C_3H_8 $\longrightarrow$ mols or grams H_2O or CO_2

then convert: g O_2 $\longrightarrow$ mols or grams H_2O or CO_2

Which produces the least is the limiting reactant.

When propane (C_3H_8) burns the products are carbon dioxide and water. How many grams of water will result form the burning of 5.05g propane with 10.1g of oxygen.?

$$C_3H_8(g) + 5\ O_2(g) \longrightarrow 3CO_2(g) + 4\ H_2O(l)$$

$$5.05gC_3H_8 \times \frac{1mol\ C_3H_8}{44.11g\ C_3H_8} \times \frac{4mol\ H_2O}{1mol\ C_3H_8} \times \frac{18.02g\ H_2O}{1mol\ H_2O} = 8.25g\ H_2O$$

$$10.1g\ O_2 \times \frac{1mol\ O_2}{32.00g\ O_2} \times \frac{4mol\ H_2O}{5mol\ O_2} \times \frac{18.02g\ H_2O}{1mol\ H_2O} = 4.55\ gH_2O$$

Even though we start with more O_2 by mass, it is used up first by moles!

Oxygen is the limiting reactant.

Consider the reaction of aluminum and oxygen:

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

Which is the limiting reactant if we start with 50.0 g Al and 50.0 g O_2?

$$50.0g\ Al \times \frac{1\ mol\ Al}{26.98g\ Al} \times \frac{3mol\ O_2}{4\ mol\ Al} \times \frac{32.00\ g\ O_2}{1mol\ O_2} = 44.5g\ O_2\ \textbf{needed}$$

Since we start with 50.0 g of O_2 and we need 44.5g of O_2, one can conclude that: ***Al limits***

i.e. there is more than enough O_2 to react with all of the Al.

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

Which is the limiting reactant if we start with 50.0 g Al and 50.0 g O_2?

How much O_2 remains?
How much aluminum oxide will form?

Solution: *since Al limits, the amount of O_2 left over and the amount of product formed are determined by the moles of Al.*

g Al $\rightarrow$ mol Al $\rightarrow$ mol O_2 $\rightarrow$ g O_2 reacted

g O_2 initial – g O_2 reacted = g O_2 left

g Al $\rightarrow$ mol Al $\rightarrow$ mol Al_2O_3 $\rightarrow$ g Al_2O_3 produced

$$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$

$$50.0 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol } O_2}{4 \text{ mol Al}} \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} \quad \boxed{= 44.5 \text{ g } O_2 \text{ reacted}}$$

$$50.0 \text{ g } O_2 - 44.5 \text{ g } O_2 \quad \boxed{= 5.5 \text{ g } O_2 \text{ left over}}$$

$$50.0 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{2 \text{ mol } Al_2O_3}{4 \text{ mol Al}} \times \frac{101.96 \text{ g } Al_2O_3}{1 \text{ mol } Al_2O_3} \boxed{= 94.5 \text{ g } Al_2O_3 \text{ forms}}$$

note that 94.5g of Al_2O_3 + 5.5 g O_2 = 100.0 g

= 50.0 g of Al + 50.0g O_2 = 100.0g