Empirical and Molecular formula

Watch the following video clip to help you get a better understanding of how to calculate the MOLECULAR formula from the EMPIRICAL formula, which we learnt in the last worksheet.

https://www.youtube.com/watch?v=J_MtVs0aBdU

Empirical and True formula exercise

Complete the following exercise in your book and fill in the answers below:

1. 13 g of zinc combines with 6,4 g of sulphur.

    1.1 What mass of zinc sulphide will be produced? m = g

    (Hint: remember the Law of Conservation of Mass)

    1.2 What is the percentage mass of each of the elements in zinc sulphide?

    % Zn = %

    % S = %

1.3 The molar mass of zinc sulphide is found to be 97,44 $g.mol^{-1}$. Determine the molecular formula of zinc sulphide.

Molecular formula =

2. A calcium mineral consisted of 29,4% calcium, 23,5% sulphur and 47,1% oxygen by mass. Calculate the empirical formula of the mineral.

Empirical formula =

3. A chlorinated hydrocarbon compound was analysed and found to consist of 24,24% carbon, 4,04% hydrogen and 71,72% chlorine. From another experiment the molecular mass was found to be 99 $g.mol^{-1}$. Deduce the empirical and molecular formula.

Empirical formula =

Molecular formula =

Percentage purity

Through chemical analysis we can determine the composition of minerals and therefore the purity of minerals. e.g. how much gold is present in a sample of gold ore mined from the ground.

$$\% \text{ Purity} = \frac{\text{mass of pure product}}{\text{total mass of sample}} \times 100$$

Examples:

1) A batch of aspirin is prepared in the laboratory but there are concerns that it is not pure. 121,2 g of solid was prepared, but when it was analysed it

was discovered that only 109,2 g of it was pure aspirin. Calculate the percent purity of the product.

Solution:

Percent purity $= 109{,}2 \div 121{,}2 \times 100 = 90.0\%$

2) Seashells are gathered, crushed and weighed and 100g of this is added to hydrochloric acid.

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$

When the reaction was complete the mixture was filtered, dried and weighed and the residue had a mass of 10g. Calculate the percentage of calcium carbonate in the seashells.

Hints:

- the seashells are MOSTLY $CaCO_3$
- all the $CaCO_3$ in the shells will react with the HCl
- the 10 g residue that is left after the reaction is the part of the shells that was NOT $CaCO_3$ (in other words the residue = impurities)

Solution:

Total sample mass = 100 g

Impurity residue = 10 g

Mass Pure $CaCO_3$ = 100 – 10 = 90 g

Percentage purity $= \dfrac{mass\ of\ pure\ product}{mass\ of\ the\ sample} \times 100$

$= \frac{\quad}{\quad} \times 100 = \quad \%$

Exercise:

1. Seashells are gathered, crushed and weighed and 80g of this is added to hydrochloric acid.

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$

When the reaction was complete the mixture was filtered, dried and weighed and the residue had a mass of 15g. Calculate the percentage of calcium carbonate in the seashells.

$\% \ CaCO_3 = \frac{\_\_\_\_\_\_\_\_}{} \times 100 = \quad \%$

2. Chalk is almost pure calcium carbonate. We can work out its purity by measuring how much carbon dioxide is given off. 10 g of chalk was reacted with an excess of hydrochloric acid. 2,128 liters of carbon dioxide gas was collected at standard temperature and pressure (STP).

The equation for the reaction is

$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

Solution:

Step 1: Convert the volume of CO_2 produced to number of moles:

$n = \frac{V}{V_0} = \frac{\_\_\_\_\_\_\_\_}{} = \quad$ mol (3 decimal places)

Step 2: Use mole ratio to calculate how many moles $CaCO_3$ would have been present to make this amount of CO_2 product:

From the balanced equation

1 mole $CaCO_3$ produces 1 mole CO_2

Therefore if we made mol CO_2, we must have started with mol $CaCO_3$.

Step 3: Use this number of moles to calculate the mass of $CaCO_3$ present in the chalk:

M ($CaCO_3$) = g.mol^{-1}

m = n x M = x = g

Step 4: Calculate the percent purity

There is g of calcium carbonate in the 10 g of chalk.

% purity = ________ × 100 = %

Use the same mole ratio process used in example 2 to do the following problem:

3. We have 13,9 g sample of impure iron pyrite. The sample is heated to produce iron (III) oxide and sulfur dioxide. If we obtained 8,02 g sample of iron (III) oxide, what was the percentage of iron pyrite in the original sample?

$$4\ FeS_2 + 11\ O_2 \rightarrow 2\ Fe_2O_3 + 8\ SO_2$$

Solution:

Step 1: Convert the mass of Fe_2O_3 produced to number of moles:

M(Fe_2O_3) = g.mol^{-1}

$n = \frac{m}{M}$ = ________ = mol

Step 2: Use mole ratio to calculate how many moles FeS_2 would have been present to make this amount of Fe_2O_3 product:

From the balanced equation

See mole ratio from balanced equation

moles FeS_2 produces moles Fe_2O_3

Therefore if we made mol Fe_2O_3, we must have started with mol FeS_2.

Step 3: Use this number of moles to calculate the mass of FeS_2 present in the impure iron pyrite:

M (FeS_2) = g.mol^{-1}

m = n x M = x = g

Step 4: Calculate the percent purity

There is g of FeS_2 in the 13,9 g of impure iron pyrite.

% purity = ________ × 100 = %

Percentage yield

In the laboratory or in a factory where products are being manufactured, the raw materials (reactants) are mixed together in the correct mole ratios and you can calculate how much product you expect to make.

e.g. If we wish to make ammonia by reacting N_2 and H_2:

$$N_2 + 3\,H_2 \rightarrow 2\,NH_3$$

If we react 1 mole of nitrogen gas with 3 moles of hydrogen gas, we expect to make 2 moles of ammonia. However, in reality, the experiment may not work out exactly to plan and you may only end up making 1,975 moles ammonia. We call the amount we actually made the YIELD of the reaction.

<u>Percentage Yield</u>: How much product is formed from a reaction (actual yield) compared to what should have formed based on the amount of reactants used (theoretical yield)

$$\%\ \text{Yield} = \frac{\text{actual mass of product}}{\text{theoretical mass of product}} \times 100$$

(Note that the percentage yield can only be 100% or less. If the answer is more than 100%, there has been a mistake in the calculation.)

Use mole ratio calculations to determine the expected (or theoretical) yield of product. Then calculate the % yield using expected yield and actual yield.

1. For the balanced equation shown below, if the reaction of 40,8 g $C_6H_6O_3$ produces a 39.0% yield, how many grams of H_2O would be produced ?

$$C_6H_6O_3 + 6\,O_2 \rightarrow 6\,CO_2 + 3\,H_2O$$

This is the mole ratio from the balanced equation

<u>Given</u>: $m = 40{,}8\text{g}$

$n = \dfrac{m}{M}$

$= \dfrac{40{,}8}{\quad}$

$=$ mol

÷ 1

x 3

<u>Expected</u>:

$n =$ mol

$m \text{ (expected)} = n \times M$

$=$ x

$=$ g

% Yield = $\dfrac{\text{Actual mass produced}}{\text{Expected (theoretical) mass}} \times 100$

<u>Note</u>: In this case they gave us the % Yield and we need to calculate Actual mass. Substitute the numbers in and rearrange the formula:

$39{,}0 = \dfrac{\text{Actual mass produced}}{\quad} \times 100$

Actual mass $= \dfrac{39{,}0 \times \quad}{100} =$ g

2. For the balanced equation shown below, if the reaction of 20,7 g of $CaCO_3$ produces 6,81 grams of CaO, what is the percent yield?

$$CaCO_3 \rightarrow CaO + CO_2$$

<u>Given</u>: $m = 20{,}7\text{g}$

$n = \dfrac{m}{M}$

$= \underline{\hspace{3em}}$

$=$ mol $\xrightarrow[\times 1]{\div 1}$

Expected: n (CaO) expected = mol

m (expected) = n x M

= x

= g

% Yield = $\dfrac{\text{Actual mass produced}}{\text{Expected (theoretical) mass}}$ x 100

$= \underline{\hspace{8em}}$ x 100

= %

Do the remaining examples in your books and check your answers by putting them in the blocks below:

3. For the balanced equation shown below, if the reaction of 91,3 g of C_3H_6 produces a 81,3% yield, how many grams of CO_2 would be produced?

$$2C_3H_6+9O_2 \rightarrow 6CO_2+6H_2O$$

Mass CO_2 = g

4. For the balanced equation shown below, if the reaction of 0,112 g of H_2 produces 0,745 g of H_2O, what is the percent yield?

$$Fe_3O_4+4H2 \rightarrow 3Fe+4H_2O$$

% Yield = %

5. For the balanced equation shown below, if the reaction of 77,0 g of $CaCN_2$ produces 27,1 g of NH_3, what is the percent yield?

$$CaCN_2+3H_2O \rightarrow CaCO_3+2NH_3$$

% Yield = %