

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 $97,44 \text{ g}\cdot\text{mol}^{-1}$.

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 $99 \text{ g}\cdot\text{mol}^{-1}$.

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:

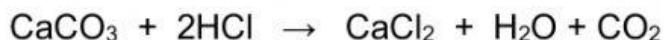
1) A batch of aspirin is prepared in the laboratory but there are concerns that it is not pure. 121,2 g of solid was prepared, but when it was analysed it

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

Solution:

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

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



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:

$$\text{Total sample mass} = 100 \text{ g}$$

$$\text{Impurity residue} = 10 \text{ g}$$

$$\text{Mass Pure CaCO}_3 = 100 - 10 = 90 \text{ g}$$

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

$$= \text{ } \times 100 = \text{ } \%$$

Exercise:

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

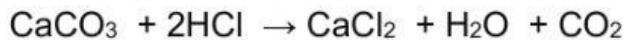


When the reaction was complete the mixture was filtered, dried and weighed and the residue had a mass of 15g. Calculate the percentage of calcium carbonate in the seashells.

$$\% \text{ CaCO}_3 = \text{_____} \times 100 = \text{_____} \%$$

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



Solution:

Step 1: Convert the volume of CO₂ produced to number of moles:

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

Step 2: Use mole ratio to calculate how many moles CaCO₃ would have been present to make this amount of CO₂ product:

From the balanced equation

1 mole CaCO₃ produces 1 mole CO₂

Therefore if we made mol CO₂, we must have started with
mol CaCO₃.

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

$$M(\text{CaCO}_3) = \text{g.mol}^{-1}$$

$$m = n \times M = \text{x} = \text{g}$$

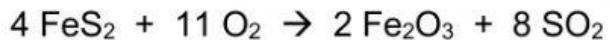
Step 4: Calculate the percent purity

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

$$\% \text{ purity} = \text{_____} \times 100 = \text{_____} \%$$

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?



Solution:

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

$$M(\text{Fe}_2\text{O}_3) = \text{g.mol}^{-1}$$

$$n = \frac{m}{M} = \text{_____} = \text{mol}$$

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:

$M(\text{FeS}_2) =$ g. mol^{-1}

$m = n \times M =$ x = g

Step 4: Calculate the percent purity

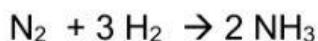
There is g of FeS_2 in the 13,9 g of impure iron pyrite.

% purity = _____ $\times 100 =$ %

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 N₂ and H₂:



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.

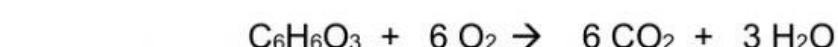
Percentage Yield: How much product is formed from a reaction (actual yield) compared to what should have formed based on the amount of reactants used (theoretical yield)

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

(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 C₆H₆O₃ produces a 39.0% yield, how many grams of H₂O would be produced ?



Given: $m = 40,8\text{g}$

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

$$= \frac{40,8}{180}$$

$$= \text{mol} \quad \begin{matrix} \div 1 \\ \times 3 \end{matrix}$$

Expected:

This is the mole ratio from the balanced equation

$$n = \text{mol}$$

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

$$= x$$

$$= g$$

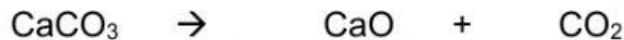
$$\% \text{ Yield} = \frac{\text{Actual mass produced}}{\text{Expected (theoretical) mass}} \times 100$$

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

$$39,0 = \frac{\text{Actual mass produced}}{x} \times 100$$

$$\text{Actual mass} = 39,0 \times \frac{x}{100} = g$$

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?



Given: $m = 20,7\text{g}$

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

$$= \underline{\quad}$$

$$= \text{mol} \quad \begin{matrix} \div 1 \\ \times 1 \end{matrix}$$

Expected:

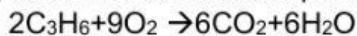
$$\begin{aligned} n(\text{CaO}) \text{ expected} &= \text{mol} \\ m(\text{expected}) &= n \times M \\ &= x \\ &= g \end{aligned}$$

$$\begin{aligned} \% \text{ Yield} &= \frac{\text{Actual mass produced}}{\text{Expected (theoretical) mass}} \times 100 \\ &= \underline{\quad} \times 100 \end{aligned}$$

$$= \% \quad \text{Red arrow from the bottom of the equation points to the equals sign.}$$

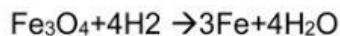
Do the remaining examples in your books and check your answers by putting them in the blocks below:

3. For the balanced equation shown below, if the reaction of 91,3 g of C₃H₆ produces a 81,3% yield, how many grams of CO₂ would be produced?



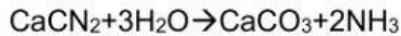
$$\text{Mass CO}_2 = \text{g}$$

4. For the balanced equation shown below, if the reaction of 0,112 g of H₂ produces 0,745 g of H₂O, what is the percent yield?



$$\% \text{ Yield} = \% \quad \text{Red arrow from the bottom of the equation points to the equals sign.}$$

5. For the balanced equation shown below, if the reaction of 77,0 g of CaCN₂ produces 27,1 g of NH₃, what is the percent yield?



$$\% \text{ Yield} = \% \quad \text{Red arrow from the bottom of the equation points to the equals sign.}$$