Quantitative Aspects of Chemical Change - Grade 10 (11) [CAPS][*]

Free High School Science Texts Project

Based on *Quantitative Aspects of Chemical Change - Grade 11*[†] by

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1 Quantitative Aspects of Chemical Change

An equation for a chemical reaction can provide us with a lot of useful information. It tells us what the reactants and the products are in the reaction, and it also tells us the ratio in which the reactants combine to form products. Look at the equation below:

$Fe + S \rightarrow \text{FeS}$

In this reaction, every atom of iron (Fe) will react with a single atom of sulphur (S) to form one molecule of iron sulphide (FeS). However, what the equation doesn't tell us, is the **quantities** or the **amount** of each substance that is involved. You may for example be given a small sample of iron for the reaction. How will you know how many atoms of iron are in this sample? And how many atoms of sulphur will you need for the reaction to use up all the iron you have? Is there a way of knowing what mass of iron sulphide will be produced at the end of the reaction? These are all very important questions, especially when the reaction is an industrial one, where it is important to know the quantities of reactants that are needed, and the quantity of product that will be formed. This chapter will look at how to quantify the changes that take place in chemical reactions.

2 The Mole

Sometimes it is important to know exactly how many particles (e.g. atoms or molecules) are in a sample of a substance, or what quantity of a substance is needed for a chemical reaction to take place.

You will remember from Relative atomic mass that the **relative atomic mass** of an element, describes the mass of an atom of that element relative to the mass of an atom of carbon-12. So the mass of an atom of carbon (relative atomic mass is $12\,u$) for example, is twelve times greater than the mass of an atom of

*Version 1.4: Jun 13, 2011 10:43 am -0500
†http://cnx.org/content/m35939/1.1/

hydrogen, which has a relative atomic mass of $1\,u$. How can this information be used to help us to know what mass of each element will be needed if we want to end up with the same number of *atoms* of carbon and hydrogen?

Let's say for example, that we have a sample of $12\,g$ carbon. What mass of *hydrogen* will contain the same number of atoms as $12\,g$ carbon? We know that each atom of carbon weighs twelve times more than an atom of hydrogen. Surely then, we will only need $1\,g$ of hydrogen for the number of atoms in the two samples to be the same? You will notice that the number of particles (in this case, *atoms*) in the two substances is the same when the ratio of their sample masses (12 g carbon: 1g hydrogen = 12:1) is the same as the ratio of their relative atomic masses (12 u: 1 u = 12:1).

To take this a step further, if you were to weigh out samples of a number of elements so that the mass of the sample was the same as the relative atomic mass of that element, you would find that the number of particles in each sample is $6,022 \times 10^{23}$. These results are shown in Table 1 below for a number of different elements. So, $24,31\,g$ of magnesium (relative atomic mass = $24,31\,u$) for example, has the same number of atoms as $40,08\,g$ of calcium (relative atomic mass = $40,08\,u$).

Element	**Relative atomic mass (u)**	**Sample mass (g)**	**Atoms in sample**
Hydrogen (H)	1	1	$6,022 \times 10^{23}$
Carbon (C)	12	12	$6,022 \times 10^{22}$
Magnesium (Mg)	24.31	24.31	$6,022 \times 10^{23}$
Sulphur (S)	32.07	32.07	$6,022 \times 10^{23}$
Calcium (Ca)	40.08	40.08	$6,022 \times 10^{23}$

Table 1: Table showing the relationship between the sample mass, the relative atomic mass and the number of atoms in a sample, for a number of elements.

This result is so important that scientists decided to use a special unit of measurement to define this quantity: the **mole** or 'mol'. A **mole** is defined as being an amount of a substance which contains the same number of particles as there are atoms in 12 g of carbon. In the examples that were used earlier, $24,31\,g$ magnesium is *one mole* of magnesium, while $40,08\,g$ of calcium is *one mole* of calcium. A mole of any substance always contains the same number of particles.

Definition 1: Mole

The mole (abbreviation 'n') is the SI (Standard International) unit for 'amount of substance'. It is defined as an amount of substance that contains the same number of particles (atoms, molecules or other particle units) as there are atoms in $12\,g$ carbon.

In one mole of any substance, there are $6,022 \times 10^{23}$ particles.

Definition 2: Avogadro's number

The number of particles in a mole, equal to $6,022 \times 10^{23}$. It is also sometimes referred to as the number of atoms in $12\,g$ of carbon-12.

If we were to write out Avogadro's number then it would look like: 602 200 000 000 000 000 000 000. This is a very large number. If we had this number of cold drink cans, then we could cover the surface of the earth to a depth of over $300\,km$! If you could count atoms at a rate of 10 million per second, then it would take you 2 billion years to count the atoms in one mole!

We can build up to the idea of Avogadro's number. For example, if you have 12 eggs then you have a dozen eggs. After this number we get a gross of eggs, which is 144 eggs. Finally if we wanted one mole of eggs this would be $6,022 \times 10^{23}$. That is a lot of eggs!

NOTE: The original hypothesis that was proposed by Amadeo Avogadro was that *'equal volumes of gases, at the same temperature and pressure, contain the same number of molecules'*. His ideas were not accepted by the scientific community and it was only four years after his death, that his original hypothesis was accepted and that it became known as 'Avogadro's Law'. In honour of his contribution to science, the number of particles in one mole was named *Avogadro's number*.

2.1 Moles and mass

1. Complete the following table:

Element	Relative atomic mass (u)	Sample mass (g)	Number of moles in the sample
Hydrogen	1.01	1.01	
Magnesium	24.31	24.31	
Carbon	12.01	24.02	
Chlorine	35.45	70.9	
Nitrogen		42.08	

Table 2

Click here for the solution[1]

2. How many atoms are there in...
   a. 1 mole of a substance
   b. 2 moles of calcium
   c. 5 moles of phosphorus
   d. $24,31\,g$ of magnesium
   e. $24,02\,g$ of carbon

   Click here for the solution[2]

3 Molar Mass

Definition 3: Molar mass
Molar mass (M) is the mass of 1 mole of a chemical substance. The unit for molar mass is **grams per mole** or $g \cdot \mathrm{mol}^{-1}$.

Refer to Table 1. You will remember that when the mass, in grams, of an element is equal to its relative atomic mass, the sample contains one mole of that element. This mass is called the **molar mass** of that element.

You may sometimes see the molar mass written as M_m. We will use M in this book, but you should be aware of the alternate notation.

It is worth remembering the following: On the periodic table, the relative atomic mass that is shown can be interpreted in two ways.

1. The mass of a *single, average atom* of that element relative to the mass of an atom of carbon.
2. The mass of *one mole of the element*. This second use is the molar mass of the element.

[1]http://www.fhsst.org/lgl
[2]http://www.fhsst.org/lgi

Element	Relative atomic mass (u)	Molar mass ($g{\cdot}\mathrm{mol}^{-1}$)	Mass of one mole of the element (g)
Magnesium	24,31	24,31	24,31
Lithium	6,94	6,94	6,94
Oxygen	16	16	16
Nitrogen	14,01	14,01	14,01
Iron	55,85	55,85	55,85

Table 3: The relationship between relative atomic mass, molar mass and the mass of one mole for a number of elements.

Exercise 1: Calculating the number of moles from mass *(Solution on p. 17.)*
Calculate the number of moles of iron (Fe) in a $11,7\,g$ sample.

Exercise 2: Calculating mass from moles *(Solution on p. 17.)*
You have a sample that contains 5 moles of zinc.

1. What is the mass of the zinc in the sample?
2. How many atoms of zinc are in the sample?

3.1 Moles and molar mass

1. Give the molar mass of each of the following elements:
   a. hydrogen
   b. nitrogen
   c. bromine

   Click here for the solution[3]
2. Calculate the number of moles in each of the following samples:
   a. $21,62\,g$ of boron (B)
   b. $54,94\,g$ of manganese (Mn)
   c. $100,3\,g$ of mercury (Hg)
   d. $50\,g$ of barium (Ba)
   e. $40\,g$ of lead (Pb)

   Click here for the solution[4]

4 An equation to calculate moles and mass in chemical reactions

The calculations that have been used so far, can be made much simpler by using the following equation:

$$\mathbf{n}\,(\text{number of moles}) = \frac{\mathbf{m}\,(\text{mass of substance in } g)}{\mathbf{M}\left(\text{molar mass of substance in } g \cdot \mathrm{mol}^{-1}\right)} \quad (1)$$

TIP: Remember that when you use the equation $n = \frac{m}{M}$, the mass is always in grams (g) and molar mass is in grams per mol ($g \cdot \mathrm{mol}^{-1}$).

[3]http://www.fhsst.org/lg3
[4]http://www.fhsst.org/lgO

The equation can also be used to calculate mass and molar mass, using the following equations:

$$m = n \times M \tag{2}$$

and

$$M = \frac{m}{n} \tag{3}$$

The following diagram may help to remember the relationship between these three variables. You need to imagine that the horizontal line is like a 'division' sign and that the vertical line is like a 'multiplication' sign. So, for example, if you want to calculate 'M', then the remaining two letters in the triangle are 'm' and 'n' and 'm' is above 'n' with a division sign between them. In your calculation then, 'm' will be the numerator and 'n' will be the denominator.

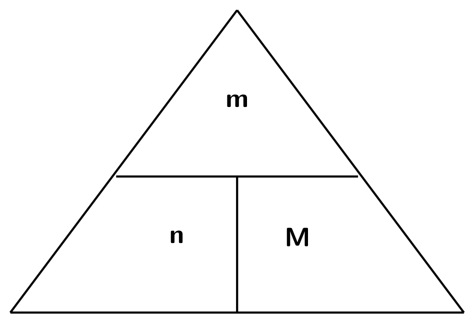

Figure 1

Exercise 3: Calculating moles from mass *(Solution on p. 17.)*
Calculate the number of moles of copper there are in a sample that weighs $127\,g$.

Exercise 4: Calculating mass from moles *(Solution on p. 17.)*
You are given a 5 mol sample of sodium. What mass of sodium is in the sample?

Exercise 5: Calculating atoms from mass *(Solution on p. 17.)*
Calculate the number of atoms there are in a sample of aluminium that weighs $80,94\,g$.

4.1 Some simple calculations

1. Calculate the number of moles in each of the following samples:
   a. $5,6\,g$ of calcium
   b. $0,02\,g$ of manganese
   c. $40\,g$ of aluminium

   Click here for the solution[5]
2. A lead sinker has a mass of $5\,g$.
   a. Calculate the number of moles of lead the sinker contains.
   b. How many lead atoms are in the sinker?

   Click here for the solution[6]

[5]http://www.fhsst.org/lgc
[6]http://www.fhsst.org/lga

3. Calculate the mass of each of the following samples:
   a. $2,5\,mol$ magnesium
   b. $12\,mol$ lithium
   c. $4,5 \times 10^{25}$ atoms of silicon

   Click here for the solution[7]

5 Molecules and compounds

So far, we have only discussed moles, mass and molar mass in relation to *elements*. But what happens if we are dealing with a molecule or some other chemical compound? Do the same concepts and rules apply? The answer is 'yes'. However, you need to remember that all your calculations will apply to the *whole molecule*. So, when you calculate the molar mass of a molecule, you will need to add the molar mass of each atom in that compound. Also, the number of moles will also apply to the whole molecule. For example, if you have one mole of nitric acid (HNO_3), it means you have $6,022 \times 10^{23}$**molecules** of nitric acid in the sample. This also means that there are $6,022 \times 10^{23}$**atoms** of hydrogen, $6,022 \times 10^{23}$**atoms** of nitrogen and $(3 \times 6,022 \times 10^{23})$ **atoms** of oxygen in the sample.

In a balanced chemical equation, the number that is written in front of the element or compound, shows the **mole ratio** in which the reactants combine to form a product. If there are no numbers in front of the element symbol, this means the number is '1'.

e.g. $N_2 + 3H_2 \rightarrow 2NH_3$

In this reaction, 1 mole of nitrogen reacts with 3 moles of hydrogen to produce 2 moles of ammonia.

Exercise 6: Calculating molar mass ***(Solution on p. 17.)***

Calculate the molar mass of H_2SO_4.

Exercise 7: Calculating moles from mass ***(Solution on p. 18.)***

Calculate the number of moles there are in $1\,kg$ of $MgCl_2$.

Exercise 8: Calculating the mass of reactants and products ***(Solution on p. 18.)***

Barium chloride and sulphuric acid react according to the following equation to produce barium sulphate and hydrochloric acid.

$BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$

If you have $2\,g$ of $BaCl_2$...

1. What quantity (in g) of H_2SO_4 will you need for the reaction so that all the barium chloride is used up?
2. What mass of HCl is produced during the reaction?

5.1 Group work : Understanding moles, molecules and Avogadro's number

Divide into groups of three and spend about 20 minutes answering the following questions together:

1. What are the units of the mole? Hint: Check the definition of the mole.
2. You have a $56\,g$ sample of iron sulphide (FeS)
   a. How many **moles** of FeS are there in the sample?
   b. How many **molecules** of FeS are there in the sample?
   c. What is the difference between a mole and a molecule?
3. The exact size of **Avogadro's number** is sometimes difficult to imagine.

[7] http://www.fhsst.org/lgx

a. Write down Avogadro's number without using scientific notation.
b. How long would it take to count to Avogadro's number? You can assume that you can count two numbers in each second.

Khan academy video on the mole - 1

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/AsqEkF7hcII&rel=0>

Figure 2

5.2 More advanced calculations

1. Calculate the molar mass of the following chemical compounds:
   a. KOH
   b. $\mathrm{FeCl_3}$
   c. $\mathrm{Mg\,(OH)_2}$

   Click here for the solution[8]
2. How many moles are present in:
   a. $10\,g$ of $Na_2\mathrm{SO_4}$
   b. $34\,g$ of $\mathrm{Ca\,(OH)_2}$
   c. $2,45 \times 10^{23}$ molecules of $\mathrm{CH_4}$?

   Click here for the solution[9]
3. For a sample of $0,2\,moles$ of potassium bromide (KBr), calculate...
   a. the number of moles of K^+ ions
   b. the number of moles of Br^- ions

   Click here for the solution[10]
4. You have a sample containing $3\,moles$ of calcium chloride.
   a. What is the chemical formula of calcium chloride?
   b. How many calcium atoms are in the sample?

   Click here for the solution[11]
5. Calculate the mass of:
   a. $3\,moles$ of $\mathrm{NH_4OH}$
   b. $4,2\,moles$ of $Ca\,(\mathrm{NO_3})_2$

   Click here for the solution[12]
6. $96,2\,g$ sulphur reacts with an unknown quantity of zinc according to the following equation: $Zn + S \rightarrow \mathrm{ZnS}$
   a. What mass of zinc will you need for the reaction, if all the sulphur is to be used up?
   b. What mass of zinc sulphide will this reaction produce?

[8] http://www.fhsst.org/lgC
[9] http://www.fhsst.org/lgr
[10] http://www.fhsst.org/lg1
[11] http://www.fhsst.org/lgY
[12] http://www.fhsst.org/lgg

Click here for the solution[13]

7. Calcium chloride reacts with carbonic acid to produce calcium carbonate and hydrochloric acid according to the following equation: $\mathrm{CaCl}_2 + H_2\mathrm{CO}_3 \rightarrow \mathrm{CaCO}_3 + 2\mathrm{HCl}$ If you want to produce $10\,g$ of calcium carbonate through this chemical reaction, what quantity (in g) of calcium chloride will you need at the start of the reaction?
Click here for the solution[14]

6 The Composition of Substances

The **empirical formula** of a chemical compound is a simple expression of the relative number of each type of atom in that compound. In contrast, the **molecular formula** of a chemical compound gives the actual number of atoms of each element found in a molecule of that compound.

Definition 4: Empirical formula
The empirical formula of a chemical compound gives the relative number of each type of atom in that compound.

Definition 5: Molecular formula
The molecular formula of a chemical compound gives the exact number of atoms of each element in one molecule of that compound.

The compound *ethanoic acid* for example, has the molecular formula $\mathrm{CH}_3\mathrm{COOH}$ or simply $C_2H_4O_2$. In one molecule of this acid, there are two carbon atoms, four hydrogen atoms and two oxygen atoms. The ratio of atoms in the compound is 2:4:2, which can be simplified to 1:2:1. Therefore, the empirical formula for this compound is CH_2O. The empirical formula contains the smallest whole number ratio of the elements that make up a compound.

Knowing either the empirical or molecular formula of a compound, can help to determine its composition in more detail. The opposite is also true. Knowing the *composition* of a substance can help you to determine its formula. There are four different types of composition problems that you might come across:

1. Problems where you will be given the formula of the substance and asked to calculate the percentage by mass of each element in the substance.
2. Problems where you will be given the percentage composition and asked to calculate the formula.
3. Problems where you will be given the products of a chemical reaction and asked to calculate the formula of one of the reactants. These are often referred to as combustion analysis problems.
4. Problems where you will be asked to find number of moles of waters of crystallisation.

Exercise 9: Calculating the percentage by mass of elements in a compound ***(Solution on p. 18.)***
Calculate the percentage that each element contributes to the overall mass of sulphuric acid ($H_2\mathrm{SO}_4$).

Exercise 10: Determining the empirical formula of a compound ***(Solution on p. 19.)***
A compound contains 52.2% carbon (C), 13.0% hydrogen (H) and 34.8% oxygen (O). Determine its empirical formula.

Exercise 11: Determining the formula of a compound ***(Solution on p. 19.)***
$207\,g$ of lead combines with oxygen to form $239\,g$ of a lead oxide. Use this information to work out the formula of the lead oxide (Relative atomic masses: $Pb = 207\,u$ and $O = 16\,u$).

Exercise 12: Empirical and molecular formula ***(Solution on p. 20.)***
Vinegar, which is used in our homes, is a dilute form of acetic acid. A sample of acetic acid has the following percentage composition: 39,9% carbon, 6,7% hyrogen and 53,4% oxygen.

[13]http://www.fhsst.org/lg4
[14]http://www.fhsst.org/lg2

1. Determine the empirical formula of acetic acid.
2. Determine the molecular formula of acetic acid if the molar mass of acetic acid is $60\,g \cdot \mathrm{mol}^{-1}$.

Exercise 13: Waters of crystallisation ***(Solution on p. 20.)***

Aluminium trichloride ($AlCl_3$) is an ionic substance that forms crystals in the solid phase. Water molecules may be trapped inside the crystal lattice. We represent this as: $AlCl_3 \cdot nH_2O$. A learner heated some aluminium trichloride crystals until all the water had evaporated and found that the mass after heating was $2,8\,g$. The mass before heating was $5\,g$. What is the number of moles of water molecules in the aluminium trichloride?

Khan academy video on molecular and empirical formulae - 1

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/gfBcM3uvWfs&rel=0>

Figure 3

Khan academy video on mass composition - 1

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/xatVrAh2U0E&rel=0>

Figure 4

6.1 Moles and empirical formulae

1. Calcium chloride is produced as the product of a chemical reaction.
   a. What is the formula of calcium chloride?
   b. What percentage does each of the elements contribute to the mass of a molecule of calcium chloride?
   c. If the sample contains $5\,g$ of calcium chloride, what is the mass of calcium in the sample?
   d. How many moles of calcium chloride are in the sample?

   Click here for the solution[15]
2. $13\,g$ of zinc combines with $6,4\,g$ of sulphur. What is the empirical formula of zinc sulphide?
   a. What mass of zinc sulphide will be produced?
   b. What percentage does each of the elements in zinc sulphide contribute to its mass?
   c. Determine the formula of zinc sulphide.

   Click here for the solution[16]
3. A calcium mineral consisted of 29,4% calcium, 23,5% sulphur and 47,1% oxygen by mass. Calculate the empirical formula of the mineral.
   Click here for the solution[17]

[15]http://www.fhsst.org/lgT
[16]http://www.fhsst.org/lgb
[17]http://www.fhsst.org/lgj

4. 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 $99 g \cdot \mathrm{mol}^{-1}$. Deduce the empirical and molecular formula.
Click here for the solution[18]

7 Molar Volumes of Gases

It is possible to calculate the volume of one mole of gas at STP using what we know about gases.

1. *Write down the ideal gas equation*pV = nRT, therefore $V = \frac{nRT}{p}$
2. *Record the values that you know, making sure that they are in SI units* You know that the gas is under STP conditions. These are as follows: $p = 101,3\,\mathrm{kPa} = 101300\,Pan = 1\,\mathrm{mol}R = 8,31\,J \cdot K^{-1} \cdot \mathrm{mol}^{-1}T = 273\,K$
3. *Substitute these values into the original equation.*

$$V = \frac{nRT}{p} \qquad (4)$$

$$V = \frac{1\mathrm{mol} \times 8,31 J \cdot K^{-1} \cdot \mathrm{mol}^{-1} \times 273K}{101300 Pa} \qquad (5)$$

4. *Calculate the volume of* 1 *mole of gas under these conditions* The volume of 1 *mole* of gas at STP is $22,4 \times 10^{-3}\,m^3 = 22,4\mathrm{dm}^3$.

TIP: The standard units used for this equation are P in Pa, V in m^3 and T in K. Remember also that $1\,000\,\mathrm{cm}^3 = 1\mathrm{dm}^3$ and $1\,000\,\mathrm{dm}^3 = 1\,m^3$.

Exercise 14: Ideal Gas *(Solution on p. 20.)*
A sample of gas occupies a volume of $20\,\mathrm{dm}^3$, has a temperature of $200\,K$ and has a pressure of $105\,Pa$. Calculate the number of moles of gas that are present in the sample.

7.1 Using the combined gas law

1. An enclosed gas(i.e. one in a sealed container) has a volume of $300\,\mathrm{cm}^3$ and a temperature of $300\,K$. The pressure of the gas is $50\,kPa$. Calculate the number of moles of gas that are present in the container.
Click here for the solution[19]
2. What pressure will $3\,mol$ of gaseous nitrogen exert if it is pumped into a container that has a volume of $25\,\mathrm{dm}^3$ at a temperature of $29\,^0C$?
Click here for the solution[20]
3. The volume of air inside a tyre is 19 litres and the temperature is $290\,K$. You check the pressure of your tyres and find that the pressure is $190\,kPa$. How many moles of air are present in the tyre?
Click here for the solution[21]
4. Compressed carbon dioxide is contained within a gas cylinder at a pressure of 700 kPa. The temperature of the gas in the cylinder is $310\,K$ and the number of moles of gas is $13\,mols$ of carbon dioxide. What is the volume of the gas inside the cylinder?
Click here for the solution[22]

[18]http://www.fhsst.org/lgD
[19]http://www.fhsst.org/lgW
[20]http://www.fhsst.org/lgZ
[21]http://www.fhsst.org/lgB
[22]http://www.fhsst.org/lgK

8 Molar concentrations of liquids

A typical solution is made by dissolving some solid substance in a liquid. The amount of substance that is dissolved in a given volume of liquid is known as the **concentration** of the liquid. Mathematically, concentration (C) is defined as moles of solute (n) per unit volume (V) of solution.

$$C = \frac{n}{V} \tag{6}$$

For this equation, the units for volume are dm^3. Therefore, the unit of concentration is $\mathrm{mol} \cdot \mathrm{dm}^{-3}$. When concentration is expressed in $\mathrm{mol} \cdot \mathrm{dm}^{-3}$ it is known as the **molarity** (M) of the solution. Molarity is the most common expression for concentration.

TIP: Do not confuse molarity (M) with molar mass (M). Look carefully at the question in which the M appears to determine whether it is concentration or molar mass.

Definition 6: Concentration
Concentration is a measure of the amount of solute that is dissolved in a given volume of liquid. It is measured in $\mathrm{mol} \cdot \mathrm{dm}^{-3}$. Another term that is used for concentration is **molarity (M)**

Exercise 15: Concentration Calculations 1 ***(Solution on p. 21.)***
If $3,5\,g$ of sodium hydroxide (NaOH) is dissolved in $2,5\,\mathrm{dm}^3$ of water, what is the concentration of the solution in $\mathrm{mol} \cdot \mathrm{dm}^{-3}$?

Exercise 16: Concentration Calculations 2 ***(Solution on p. 21.)***
You have a $1\,\mathrm{dm}^3$ container in which to prepare a solution of potassium permanganate ($KMnO_4$). What mass of $KMnO_4$ is needed to make a solution with a concentration of $0,2\,M$?

Exercise 17: Concentration Calculations 3 ***(Solution on p. 21.)***
How much sodium chloride (in g) will one need to prepare $500\,\mathrm{cm}^3$ of solution with a concentration of $0,01\,M$?

8.1 Molarity and the concentration of solutions

1. $5,95\,g$ of potassium bromide was dissolved in $400\,\mathrm{cm}^3$ of water. Calculate its molarity.
   Click here for the solution[23]
2. $100\,g$ of sodium chloride (NaCl) is dissolved in $450\,\mathrm{cm}^3$ of water.
   a. How many moles of NaCl are present in solution?
   b. What is the volume of water (in dm^3)?
   c. Calculate the concentration of the solution.
   d. What mass of sodium chloride would need to be added for the concentration to become $5,7\,\mathrm{mol} \cdot \mathrm{dm}^{-3}$?

   Click here for the solution[24]
3. What is the molarity of the solution formed by dissolving $80\,g$ of sodium hydroxide ($NaOH$) in $500\,\mathrm{cm}^3$ of water?
   Click here for the solution[25]
4. What mass (g) of hydrogen chloride (HCl) is needed to make up $1000\,\mathrm{cm}^3$ of a solution of concentration $1\,\mathrm{mol} \cdot \mathrm{dm}^{-3}$?
   Click here for the solution[26]

[23]http://www.fhsst.org/lgk
[24]http://www.fhsst.org/lg0
[25]http://www.fhsst.org/lg8
[26]http://www.fhsst.org/lg9

5. How many moles of H_2SO_4 are there in $250\,cm^3$ of a $0,8\,M$ sulphuric acid solution? What mass of acid is in this solution?
   Click here for the solution[27]

9 Stoichiometric calculations

Stoichiometry is the calculation of the quantities of reactants and products in chemical reactions. It is also the numerical relationship between reactants and products. In representing chemical change showed how to write balanced chemical equations. By knowing the ratios of substances in a reaction, it is possible to use stoichiometry to calculate the amount of either reactants or products that are involved in the reaction. The examples shown below will make this concept clearer.

Exercise 18: Stoichiometric calculation 1 ***(Solution on p. 21.)***

What volume of oxygen at S.T.P. is needed for the complete combustion of $2\,dm^3$ of propane (C_3H_8)? (Hint: CO_2 and H_2O are the products in this reaction (and in all combustion reactions))

Exercise 19: Stoichiometric calculation 2 ***(Solution on p. 21.)***

What mass of iron (II) sulphide is formed when $5,6\,g$ of iron is completely reacted with sulphur?

When we are given a known mass of a reactant and are asked to work out how much product is formed, we are working out the theoretical yield of the reaction. In the laboratory chemists never get this amount of product. In each step of a reaction a small amount of product and reactants is 'lost' either because a reactant did not completely react or some of the product was left behind in the original container. Think about this. When you make your lunch or supper, you might be a bit hungry, so you eat some of the food that you are preparing. So instead of getting the full amount of food out (theoretical yield) that you started preparing, you lose some along the way.

Exercise 20: Industrial reaction to produce fertiliser ***(Solution on p. 22.)***

Sulphuric acid (H_2SO_4) reacts with ammonia (NH_3) to produce the fertiliser ammonium sulphate ($(NH_4)_2SO_4$) according to the following equation:

$H_2SO_4\,(aq) + 2NH_3\,(g) \rightarrow (NH_4)_2SO_4\,(aq)$

What is the maximum mass of ammonium sulphate that can be obtained from $2,0\,kg$ of sulphuric acid?

Khan academy video on stoichiometry - 1

This media object is a Flash object. Please view or download it at
<http://www.youtube.com/v/jFv6k2OV7IU&rel=0>

Figure 5

9.1 Stoichiometry

1. Diborane, B_2H_6, was once considered for use as a rocket fuel. The combustion reaction for diborane is: $B_2H_6\,(g) + 3O_2\,(g) \rightarrow 2HBO_2\,(g) + 2H_2O\,(l)$ If we react $2,37\,grams$ of diborane, how many grams of water would we expect to produce?
   Click here for the solution[28]

[27]http://www.fhsst.org/lgX
[28]http://www.fhsst.org/lgI

2. Sodium azide is a commonly used compound in airbags. When triggered, it has the following reaction: $2\mathrm{NaN}_3\,(s) \rightarrow 2Na\,(s) + 3N_2\,(g)$ If $23,4\,grams$ of sodium azide is used, how many moles of nitrogen gas would we expect to produce?
   Click here for the solution[29]
3. Photosynthesis is a chemical reaction that is vital to the existence of life on Earth. During photosynthesis, plants and bacteria convert carbon dioxide gas, liquid water, and light into glucose ($C_6H_{12}O_6$) and oxygen gas.
   a. Write down the equation for the photosynthesis reaction.
   b. Balance the equation.
   c. If $3\,moles$ of carbon dioxide are used up in the photosynthesis reaction, what mass of glucose will be produced?

   Click here for the solution[30]

This media object is a Flash object. Please view or download it at
<http://static.slidesharecdn.com/swf/ssplayer2.swf?doc=moles-100512054508-phpapp02&stripped_title=moles-4065434&userName=kwarne>

Figure 6

10 Summary

- It is important to be able to quantify the changes that take place during a chemical reaction.
- The **mole (n)** is a SI unit that is used to describe an amount of substance that contains the same number of particles as there are atoms in $12\,g$ of carbon.
- The number of particles in a mole is called the **Avogadro constant** and its value is $6,022 \times 10^{23}$. These particles could be atoms, molecules or other particle units, depending on the substance.
- The **molar mass (M)** is the mass of one mole of a substance and is measured in grams per mole or $g \cdot \mathrm{mol}^{-1}$. The numerical value of an element's molar mass is the same as its relative atomic mass. For a compound, the molar mass has the same numerical value as the molecular mass of that compound.
- The relationship between moles (n), mass in grams (m) and molar mass (M) is defined by the following equation:

$$n = \frac{m}{M} \qquad (7)$$

- In a balanced chemical equation, the number in front of the chemical symbols describes the **mole ratio** of the reactants and products.
- The **empirical formula** of a compound is an expression of the relative number of each type of atom in the compound.
- The **molecular formula** of a compound describes the actual number of atoms of each element in a molecule of the compound.
- The formula of a substance can be used to calculate the **percentage by mass** that each element contributes to the compound.
- The **percentage composition** of a substance can be used to deduce its chemical formula.
- One mole of gas occupies a volume of $22,4\,\mathrm{dm}^3$.

[29]http://www.fhsst.org/lg5
[30]http://www.fhsst.org/lgN

- The **concentration** of a solution can be calculated using the following equation,

$$C = \frac{n}{V} \tag{8}$$

where C is the concentration (in $\text{mol} \cdot \text{dm}^{-3}$), n is the number of moles of solute dissolved in the solution and V is the volume of the solution (in dm^{-3}).
- **Molarity** is a measure of the concentration of a solution, and its units are $\text{mol} \cdot \text{dm}^{-3}$.
- **Stoichiometry** is the calculation of the quantities of reactants and products in chemical reactions. It is also the numerical relationship between reactants and products.
- The theoretical yield of a reaction is the maximum amount of product that we expect to get out of a reaction

11 End of chapter exercises

1. Write only the word/term for each of the following descriptions:
   a. the mass of one mole of a substance
   b. the number of particles in one mole of a substance

   Click here for the solution[31]
2. Multiple choice: Choose the one correct answer from those given.
   a. $5\,g$ of magnesium chloride is formed as the product of a chemical reaction. Select the **true** statement from the answers below:
      a. 0.08 moles of magnesium chloride are formed in the reaction
      b. the number of atoms of Cl in the product is $0,6022 \times 10^{23}$
      c. the number of atoms of Mg is 0,05
      d. the atomic ratio of Mg atoms to Cl atoms in the product is 1:1

      Click here for the solution[32]
   b. 2 moles of oxygen gas react with hydrogen. What is the mass of oxygen in the reactants?
      a. 32 g
      b. 0,125 g
      c. 64 g
      d. 0,063 g

      Click here for the solution[33]
   c. In the compound potassium sulphate (K_2SO_4), oxygen makes up x% of the mass of the compound. x = ...
      a. 36.8
      b. 9,2
      c. 4
      d. 18,3

      Click here for the solution[34]
   d. The molarity of a $150\,\text{cm}^3$ solution, containing 5 g of $NaCl$ is...
      a. $0,09\,M$
      b. $5,7\, \times 10^{-4}\,M$
      c. $0,57\,M$
      d. $0,03\,M$

[31] http://www.fhsst.org/lgR
[32] http://www.fhsst.org/lgn
[33] http://www.fhsst.org/lgQ
[34] http://www.fhsst.org/lgU

Click here for the solution[35]

3. Calculate the number of moles in:
   a. 5 g of methane (CH_4)
   b. 3,4 g of hydrochloric acid
   c. 6,2 g of potassium permanganate ($KMnO_4$)
   d. 4 g of neon
   e. 9,6 kg of titanium tetrachloride ($TiCl_4$)

   Click here for the solution[36]
4. Calculate the mass of:
   a. 0,2 mols of potassium hydroxide (KOH)
   b. 0,47 mols of nitrogen dioxide
   c. 5,2 mols of helium
   d. 0,05 mols of copper (II) chloride ($CuCl_2$)
   e. $31,31 \times 10^{23}$ molecules of carbon monoxide (CO)

   Click here for the solution[37]
5. Calculate the percentage that each element contributes to the overall mass of:
   a. Chloro-benzene (C_6H_5Cl)
   b. Lithium hydroxide (LiOH)

   Click here for the solution[38]
6. CFC's (chlorofluorocarbons) are one of the gases that contribute to the depletion of the ozone layer. A chemist analysed a CFC and found that it contained 58,64% chlorine, 31,43% fluorine and 9,93% carbon. What is the empirical formula?
   Click here for the solution[39]
7. 14 g of nitrogen combines with oxygen to form 46 g of a nitrogen oxide. Use this information to work out the formula of the oxide.
   Click here for the solution[40]
8. Iodine can exist as one of three oxides (I_2O_4; I_2O_5; I_4O_9). A chemist has produced one of these oxides and wishes to know which one they have. If he started with 508 g of iodine and formed 652 g of the oxide, which form has he produced?
   Click here for the solution[41]
9. A fluorinated hydrocarbon (a hydrocarbon is a chemical compound containing hydrogen and carbon.) was analysed and found to contain 8,57% H, 51,05% C and 40,38% F.
   a. What is its empirical formula?
   b. What is the molecular formula if the molar mass is $94,1 g \cdot \text{mol}^{-1}$?

   Click here for the solution[42]
10. Copper sulphate crystals often include water. A chemist is trying to determine the number of moles of water in the copper sulphate crystals. She weighs out 3 g of copper sulphate and heats this. After heating, she finds that the mass is 1,9 g. What is the number of moles of water in the crystals? (Copper sulphate is represented by $CuSO_4 \cdot xH_2O$).Click here for the solution[43]
11. $300\,\text{cm}^3$ of a $0,1\,\text{mol} \cdot \text{dm}^{-3}$ solution of sulphuric acid is added to $200\,\text{cm}^3$ of a $0,5\,\text{mol} \cdot \text{dm}^{-3}$ solution of sodium hydroxide.

[35] http://www.fhsst.org/lgy
[36] http://www.fhsst.org/lT3/
[37] http://www.fhsst.org/lTO
[38] http://www.fhsst.org/lTc
[39] http://www.fhsst.org/lTx
[40] http://www.fhsst.org/lTa
[41] http://www.fhsst.org/lTC
[42] http://www.fhsst.org/lT1
[43] http://www.fhsst.org/lTr

a. Write down a balanced equation for the reaction which takes place when these two solutions are mixed.
b. Calculate the number of moles of sulphuric acid which were added to the sodium hydroxide solution.
c. Is the number of moles of sulphuric acid enough to fully neutralise the sodium hydroxide solution? Support your answer by showing all relevant calculations. (IEB Paper 2 2004)

Click here for the solution[44]

12. A learner is asked to make $200\,\mathrm{cm}^3$ of sodium hydroxide ($NaOH$) solution of concentration $0,5\,\mathrm{mol}\cdot\mathrm{dm}^{-3}$.

a. Determine the mass of sodium hydroxide pellets he needs to use to do this.
b. Using an accurate balance the learner accurately measures the correct mass of the $NaOH$ pellets. To the pellets he now adds exactly $200\,\mathrm{cm}^3$ of pure water. Will his solution have the correct concentration? Explain your answer.
c. The learner then takes $300\,\mathrm{cm}^3$ of a $0,1\,\mathrm{mol}\cdot\mathrm{dm}^{-3}$ solution of sulphuric acid ($H_2\mathrm{SO}_4$) and adds it to $200\,\mathrm{cm}^3$ of a $0,5\,\mathrm{mol}\cdot\mathrm{dm}^{-3}$ solution of $NaOH$ at 25^0C.
d. Write down a balanced equation for the reaction which takes place when these two solutions are mixed.
e. Calculate the number of moles of $H_2\mathrm{SO}_4$ which were added to the NaOH solution.
f. Is the number of moles of $H_2\mathrm{SO}_4$ calculated in the previous question enough to fully neutralise the $NaOH$ solution? Support your answer by showing all the relevant calculations. (IEB Paper 2, 2004)

Click here for the solution[45]

[44]http://www.fhsst.org/lgP
[45]http://www.fhsst.org/lgm

Solutions to Exercises in this Module

Solution to Exercise (p. 4)

Step 1. If we look at the periodic table, we see that the molar mass of iron is $55,85\,g \cdot \text{mol}^{-1}$. This means that 1 mole of iron will have a mass of $55,85\,g$.

Step 2. If 1 mole of iron has a mass of $55,85\,g$, then: the number of moles of iron in $111,7\,g$ must be:

$$\frac{111,7g}{55,85g \cdot \text{mol}^{-1}} = 2\,\text{mol} \tag{9}$$

There are 2 moles of iron in the sample.

Solution to Exercise (p. 4)

Step 1. Molar mass of zinc is $65,38\,g \cdot \text{mol}^{-1}$, meaning that 1 mole of zinc has a mass of $65,38\,g$.

Step 2. If 1 mole of zinc has a mass of $65,38\,g$, then 5 moles of zinc has a mass of: $65,38\,g \times 5\,\text{mol} = 326,9\,g$ (answer to a)

Step 3.

$$5 \times 6,022 \times 10^{23} = 30,115 \times 10^{23} \tag{10}$$

(answer to b)

Solution to Exercise (p. 5)

Step 1.

$$n = \frac{m}{M} \tag{11}$$

Step 2.

$$n = \frac{127}{63,55} = 2 \tag{12}$$

There are 2 moles of copper in the sample.

Solution to Exercise (p. 5)

Step 1.

$$m = n \times M \tag{13}$$

Step 2. $\text{M}_{Na} = 22,99\,g \cdot \text{mol}^{-1}$
Therefore,

$$m = 5 \times 22,99 = 114,95\,g \tag{14}$$

The sample of sodium has a mass of $114,95\,g$.

Solution to Exercise (p. 5)

Step 1.

$$n = \frac{m}{M} = \frac{80,94}{26,98} = 3\,\text{moles} \tag{15}$$

Step 2. Number of atoms in 3 mol aluminium $= 3 \times 6,022 \times 10^{23}$
There are $18,069 \times 10^{23}$ aluminium atoms in a sample of $80,94\,g$.

Solution to Exercise (p. 6)

Step 1. Hydrogen $= 1,008\,g \cdot \text{mol}^{-1}$; Sulphur $= = 32,07\,g \cdot \text{mol}^{-1}$; Oxygen $= 16\,g \cdot \text{mol}^{-1}$

Step 2.

$$M_{(H_2SO_4)} = (2 \times 1,008) + (32,07) + (4 \times 16) = 98,09\,g \cdot \text{mol}^{-1} \quad (16)$$

Solution to Exercise (p. 6)

Step 1.

$$n = \frac{m}{M} \quad (17)$$

Step 2. a. Convert mass into grams

$$m = 1\text{kg} \times 1000 = 1000\,g \quad (18)$$

b. Calculate the molar mass of MgCl_2.

$$M_{(\text{MgCl}_2)} = 24,31 + (2 \times 35,45) = 95,21\,g \cdot \text{mol}^{-1} \quad (19)$$

Step 3.

$$n = \frac{1000}{95,21} = 10,5\,\text{mol} \quad (20)$$

There are $10,5$ moles of magnesium chloride in a $1\,kg$ sample.

Solution to Exercise (p. 6)

Step 1.

$$n = \frac{m}{M} = \frac{2}{208,24} = 0,0096\,\text{mol} \quad (21)$$

Step 2. According to the balanced equation, 1 mole of BaCl_2 will react with 1 mole of $H_2\text{SO}_4$. Therefore, if $0,0096$ mol of BaCl_2 react, then there must be the same number of moles of $H_2\text{SO}_4$ that react because their mole ratio is 1:1.

Step 3.

$$m = n \times M = 0,0096 \times 98,086 = 0,94\,g \quad (22)$$

(answer to 1)

Step 4. According to the balanced equation, 2 moles of HCl are produced for every 1 mole of the two reactants. Therefore the number of moles of HCl produced is $(2 \times 0,0096)$, which equals $0,0096$ moles.

Step 5.

$$m = n \times M = 0,0192 \times 35,73 = 0,69\,g \quad (23)$$

(answer to 2)

Solution to Exercise (p. 8)

Step 1. Hydrogen $= 1,008 \times 2 = 2,016\,u$
Sulphur $= 32,07\,u$
Oxygen $= 4 \times 16 = 64\,u$

Step 2. Use the calculations in the previous step to calculate the molecular mass of sulphuric acid.

$$\text{Mass} = 2,016 + 32,07 + 64 = 98,09\,u \quad (24)$$

Step 3. Use the equation:
Percentage by mass $= \frac{\text{atomic mass}}{\text{molecular mass of H}_2\text{SO}_4} \times 100\%$
Hydrogen

$$\frac{2,016}{98,09} \times 100\% = 2,06\% \quad (25)$$

Sulphur

$$\frac{32,07}{98,09} \times 100\% = 32,69\% \quad (26)$$

Oxygen

$$\frac{64}{98,09} \times 100\% = 65,25\% \quad (27)$$

(You should check at the end that these percentages add up to 100%!)
In other words, in one molecule of sulphuric acid, hydrogen makes up 2,06% of the mass of the compound, sulphur makes up 32,69% and oxygen makes up 65,25%.

Solution to Exercise **(p. 8)**

Step 1. Carbon $= 52,2\,g$, hydrogen $= 13,0\,g$ and oxygen $= 34,8\,g$
Step 2.

$$n = \frac{m}{M} \quad (28)$$

Therefore,

$$n\,(\text{Carbon}) = \frac{52,2}{12,01} = 4,35\text{mol} \quad (29)$$

$$n\,(\text{Hydrogen}) = \frac{13,0}{1,008} = 12,90\text{mol} \quad (30)$$

$$n\,(\text{Oxygen}) = \frac{34,8}{16} = 2,18\text{mol} \quad (31)$$

Step 3. In this case, the smallest number of moles is 2.18. Therefore...
Carbon

$$\frac{4,35}{2,18} = 2 \quad (32)$$

Hydrogen

$$\frac{12,90}{2,18} = 6 \quad (33)$$

Oxygen

$$\frac{2,18}{2,18} = 1 \quad (34)$$

Therefore the empirical formula of this substance is: C_2H_6O. Do you recognise this compound?

Solution to Exercise **(p. 8)**

Step 1.

$$239 - 207 = 32\,g \quad (35)$$

Step 2.

$$n = \frac{m}{M} \quad (36)$$

Lead

$$\frac{207}{207} = 1\,\text{mol} \quad (37)$$

Oxygen

$$\frac{32}{16} = 2\,\text{mol} \tag{38}$$

Step 3. The mole ratio of $Pb : O$ in the product is 1:2, which means that for every atom of lead, there will be two atoms of oxygen. The formula of the compound is PbO_2.

Solution to Exercise (p. 8)

Step 1. In 100 g of acetic acid, there is $39,9\,g\,C$, $6,7\,g\,H$ and $53,4\,g\,O$
Step 2. $n = \frac{m}{M}$

$$\begin{array}{rcl} n_C & = & \frac{39,9}{12} = 3,33\,\text{mol} \\ n_H & = & \frac{6,7}{1} = 6,7\,mol \\ n_O & = & \frac{53,4}{16} = 3,34\,mol \end{array} \tag{39}$$

Step 3. Empirical formula is CH_2O
Step 4. The molar mass of acetic acid using the empirical formula is $30\,g \cdot \text{mol}^{-1}$. Therefore the actual number of moles of each element must be double what it is in the empirical formula.
The molecular formula is therefore $C_2H_4O_2$ or CH_3COOH

Solution to Exercise (p. 9)

Step 1. We first need to find n, the number of water molecules that are present in the crystal. To do this we first note that the mass of water lost is $5 - 2,8 = 2,2$.
Step 2. The next step is to work out the mass ratio of aluminium trichloride to water and the mole ratio. The mass ratio is:

$$2,8 : 2,2 \tag{40}$$

To work out the mole ratio we divide the mass ratio by the molecular mass of each species:

$$\frac{2,8}{133} : \frac{2,2}{18} = 0,021 : 0,12 \tag{41}$$

Next we do the following:

$$0,021\frac{1}{0,021} = 1 \tag{42}$$

and

$$\frac{0,12}{0,021} = 6 \tag{43}$$

So the mole ratio of aluminium trichloride to water is:

$$1 : 6 \tag{44}$$

Step 3. And now we know that there are 6 moles of water molecules in the crystal.

Solution to Exercise (p. 10)

Step 1. The only value that is not in SI units is volume. $V = 0,02\,m^3$.
Step 2. We know that pV = nRT
Therefore,

$$n = \frac{pV}{RT} \tag{45}$$

Step 3.

$$n = \frac{105 \times 0,02}{8,31 \times 280} = \frac{2,1}{2326,8} = 0,0009\,\text{moles} \quad (46)$$

Solution to Exercise (p. 11)

Step 1.

$$n = \frac{m}{M} = \frac{3,5}{40} = 0,0875\,\text{mol} \quad (47)$$

Step 2.

$$C = \frac{n}{V} = \frac{0,0875}{2,5} = 0,035 \quad (48)$$

The concentration of the solution is $0,035\,\text{mol} \cdot \text{dm}^{-3}$ or $0,035\,M$

Solution to Exercise (p. 11)

Step 1.

$$C = \frac{n}{V} \quad (49)$$

therefore

$$n = C \times V = 0,2 \times 1 = 0,2\,\text{mol} \quad (50)$$

Step 2.

$$m = n \times M = 0,2 \times 158,04 = 31,61\,g \quad (51)$$

The mass of KMnO_4 that is needed is $31,61\,g$.

Solution to Exercise (p. 11)

Step 1.

$$V = \frac{500}{1000} = 0,5\,\text{dm}^3 \quad (52)$$

Step 2.

$$n = C \times V = 0,01 \times 0,5 = 0,005\,\text{mol} \quad (53)$$

Step 3.

$$m = n \times M = 0,005 \times 58,45 = 0,29\,g \quad (54)$$

The mass of sodium chloride needed is $0,29\,g$

Solution to Exercise (p. 12)

Step 1. $C_3H_8\,(g) + 5O_2\,(g) \rightarrow 3CO_2\,(g) + 4H_2O\,(g)$
Step 2. From the balanced equation, the ratio of oxygen to propane in the reactants is 5:1.
Step 3. 1 volume of propane needs 5 volumes of oxygen, therefore $2\,\text{dm}^3$ of propane will need $10\,\text{dm}^3$ of oxygen for the reaction to proceed to completion.

Solution to Exercise (p. 12)

Step 1. $Fe\,(s) + S\,(s) \rightarrow \text{FeS}\,(s)$
Step 2.

$$n = \frac{m}{M} = \frac{5,6}{55,85} = 0,1\text{mol} \quad (55)$$

Step 3. From the equation 1 *mole* of Fe gives 1 *mole* of FeS. Therefore, $0,1$ *moles* of iron in the reactants will give $0,1$ *moles* of iron sulphide in the product.

Step 4.

$$m = n \times M = 0,1 \times 87,911 = 8,79\,g \quad (56)$$

The mass of iron (II) sulphide that is produced during this reaction is $8,79\,g$.

Solution to Exercise (p. 12)

Step 1.

$$n\left(H_2\mathrm{SO}_4\right) = \frac{m}{M} = \frac{2000\,g}{98,078\,g \cdot \mathrm{mols}^{-1}} = 20,39\,\mathrm{mols} \quad (57)$$

Step 2. From the balanced equation, the mole ratio of $H_2\mathrm{SO}_4$ in the reactants to $(\mathrm{NH}_4)_2\mathrm{SO}_4$ in the product is 1:1. Therefore, $20,39\,mols$ of $H_2\mathrm{SO}_4$ of $(\mathrm{NH}_4)_2\mathrm{SO}_4$.

The maximum mass of ammonium sulphate that can be produced is calculated as follows:

$$m = n \times M = 20,41\mathrm{mol} \times 132\,g \cdot \mathrm{mol}^{-1} = 2694\,g \quad (58)$$

The maximum amount of ammonium sulphate that can be produced is $2,694\,kg$.