Chapter 15

Stoichiometry

Student Learning Outcomes

After studying this chapter, students will be able to:

- Use the molar gas volume, 24 dm^3 at room temperature and pressure, in calculations involving gases
- Define concentration, use both g/dm^3 and $mol/\ dm^3$, and convert between them
- Calculate stoichiometric relationships between substances relationships (specifically:
- reacting masses, limiting reactants,
- volume of gasses at r.t.p.,
- volumes of solution and concentrations of solutions in $g/\ dm^3$ or $mol/\ dm^3$, including conversion between cm^3 and dm^3)
- Calculate concentration of a solution in a titration using empirical data
- Calculate empirical formula and molecular formula from appropriate data
- Calculate percentage yield, percentage composition by mass and percentage purity from appropriate data.

15.1 Molar Volume

The volume of a gas varies with changes in pressure and temperature. In order to compare the volumes occupied by different gases, we must adopt a set of standard conditions of temperature and pressure called room temperature and pressure (RTP). Accordingly, 25 $^{\circ}C$ (298.15K) and one atmosphere (760 torr) are used as standard conditions of temperature and pressure for measurement of volumes of gases at room temperature.

It has been found that one mole of all gases occupies a volume of 24 dm^3 at RTP. This is called molar volume. It is worth noting that one mole of each gas contains the same number of molecules, occupies the same volume and at the same time has different mass. In other words the masses and the sizes of the molecules of gases do not change the volume occupied by them. The concept of molar volume is useful because it enables us to calculate the number of moles or masses of the gases if we know their volumes.

15.1 Example

Calculate the number of moles of nitrogen gas present when it occupies a volume of 2.5 dm^3 at RTP.

Solution:

24 dm^3 is occupied by how many moles of nitrogen = 1 mole

1 dm^3 will be occupied by how many moles of nitrogen $= \frac{1}{24}$ mole

2.5 dm^3 will be occupied by how many moles of nitrogen $= \frac{1}{24} \times 2.5$

$= 0.10$ mole

You can use the following equation as well to solve this problem.

$$\text{No. of moles of a gas} = \frac{\text{Volume of the gas at RTP}}{\text{Molar volume of the gas}} \quad \text{..................} \quad 1$$

15.2 Example

Calculate the volume of 0.5 mol of hydrogen at RTP.

Solution:

Apply the equation (1) $\quad 0.5 = \dfrac{\text{Volume of hydrogen}}{24}$

$$\text{Volume of hydrogen} = 0.5 \times 24 = 12\,dm^3$$

15.3 Example

Calculate the volume occupied by 200 g carbon dioxide at RTP.

Solution:

$$\text{No. of moles} = \frac{\text{Volume}}{\text{Molar Volume}}$$

$$\text{Volume} = \text{No. of moles} \times \text{Molar volume}$$

$$\text{Since no. of moles} = \frac{\text{Mass}}{\text{Molar Mass}}$$

So,

$$\text{Volume} = \frac{\text{Mass}}{\text{Molar Mass}} \times \text{Molar volume}$$

$$\text{Volume} = \frac{200\text{ g}}{44\text{ g}} \times 24\text{ dm}^3 = 109\text{ dm}^3$$

15.1 Quick Check!

Calculate the amount of CO_2 in 100 dm^3 at RTP?

15.2 Concentration of a solution

Concentration of a solution is a measure of the amount of solute that has been dissolved in a given amount of solution or solvent. The amount of solute dissolved is expressed either in grams or in moles whereas the quantity of solution is expressed in terms of its volume, dm^3 or cm^3.

The unit used for the mass/volume concentration is then g/dm^3 while for the molar concentration it is mol/dm^3.

15.4 Example

Calculate concentration of a solution in g/dm^3 which contains 20 g of sodium chloride dissolved in 0.2 dm^3 of water.

Solution:

$$\text{Concentration} = \frac{\text{Mass of solute in gram}}{\text{Volume of solution in dm}^3}$$

$$= \frac{20}{0.2} = 100\text{ g/dm}^3$$

15.5 Example

500 cm^3 of hydrochloric acid contains 1.8 g of dissolved hydrogen chloride. Calculate the concentration of the acid in g/dm^3.

Solution:

Volume of hydrochloric acid in dm^3 $= \dfrac{500\ cm^3}{1000} = 0.5\ dm^3$

Concentration of hydrochloric acid $= \dfrac{1.8}{0.5} = 3.6\ g/dm^3$

15.6 Example

Calculate the mass of copper sulphate in grams dissolved in 2 dm^3 of the solution to get a solution with a concentration 15 g /dm^3.

Point to note

Molar concentration and molarity both refer to the concentration of a solute in a solution.

Solution:

Mass of solute in gram = Concentration in $g/dm^3 \times$ volume in dm^3

$= 15 \times 2 = 30\ g$

15.7 Example

A solution is prepared by dissolving 10 g of NaOH in 1.2 dm^3 of water. Calculate the concentration of the solution in mol/dm^3.

Solution:

Amount of NaOH $= 10\ g$

No. of moles of NaOH $= \dfrac{10}{40} = 0.25\ mol$

Concentration of the solution $= \dfrac{0.25}{1.2}$

$= 0.21\ mol/dm^3$

15.2 Quick Check!

What mass of $CuSO_4$ is present in 50 cm^3 of a $5 \times 10^{-2}\ mol/dm^3$ aqueous solution?

15.3 Calculate the Concentration of a Solution in Titration using Empirical Data

The concentration of a solution can be determined by titrating it with a solution of known concentration till the end point is reached. For this purpose, you need to know the balanced chemical equation for the reaction between the two solutions, the volumes of the two solutions used during the titration and the concentration of the other solution.

15.8 Example

A titration is performed between NaOH and HCl till the end point is reached. $15cm^3$ of solution of sodium hydroxide having a molar concentration $0.5\ mol/dm^3$ is used to neutralize $23\ cm^3$ of HCl. Find out the concentration of HCl in mol/dm^3.

Solution:

1. Balanced chemical equation for this reaction

$$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$

2. Convert the volume of NaOH used to dm^3

$$\text{Volume of the solution} = \frac{15}{1000} = 0.015\ dm^3$$

3. Convert the volume of HCl required to neutralize NaOH solution in dm^3

$$\text{Volume of HCl required} = \frac{23}{1000} = 0.023\ dm^3$$

4. No. of moles of NaOH $= \text{molar concentration} \times \text{Volume}$

$= 0.5\ mol/dm^3 \times 0.015\ dm^3$

$= 0.0075$

5. According to the balanced chemical equation 1 mol of NaOH has reacted with 1 mol of HCl.

6. No. of moles of HCl $= 0.0075$

7. Concertion of HCl solution $= \dfrac{\text{No. of moles}}{\text{Volume in } dm^3} = \dfrac{0.0075}{0.023}$

$= 0.326\ mol/dm^3$

Therefore, the molar concentration of the HCl solution is $0.326\ mol/dm^3$.

15.3 Quick Check!

Calculate the volume of $0.1\ mol/dm^3$ oxalic acid ($C_2O_4H_2$) solution to exactly neutralize $25\ cm^3$ of $0.1\ mol/dm^3$ KOH solution.

Interesting Information

Stoichiometry is essential for optimizing chemical reactions in many industries.

15.4 Percentage Composition by Mass, Empirical Formula and Molecular Formula

Percentage Composition by Mass

Percentage composition of a compound refers to the percentage mass of each element present in the compound. The percentage of an element in a compound is the number of grams of that element present in 100 grams of the compound. The percentage of an element can be determined by applying the following formula.

$$\text{Percentage of an element} = \frac{\text{Mass of the element in the compound}}{\text{Formula mass of the compound}} \times 100 \quad \text{.................... 2}$$

Percentage composition is a useful concept to determine the purity of a compound and its empirical and molecular formulae. Percentage composition of a compound can be determined both theoretically as well as by performing quantitative analysis. The following examples will illustrate both these methods.

15.9 Example

Find out the percentage composition of $Ca(OH)_2$

Solution:

Formula mass of $Ca(OH)_2$ = 40 + 32 + 2 = 74

Applying the formula given in equation (2)

$$\text{Percentage of an element} = \frac{\text{Mass of the element in the compound}}{\text{Formula mass of the compound}} \times 100$$

$$\text{Percentage of Ca} = \frac{40}{74} \times 100 = 54.05\ \%$$

$$\text{Percentage of Oxygen} = \frac{32}{74} \times 100 = 43.24\ \%$$

$$\text{Percentage of Hydrogen} = \frac{2}{74} \times 100 = 2.70\ \%$$

Interesting Information

The composition of all the chemical products which are in our daily use, such as shampoos, cleaners, soaps, perfumes and fertilizers are formed using stoichiometric calculations.

15.10 Example

8.657 g of a compound were decomposed into its elements. It gave 5.217 g of carbon, 0.962 g of hydrogen and 2.478 g of oxygen. Find out the percentage composition of the compound.

Solution:

Applying the following formula

$$\text{Percentage of element} = \frac{\text{Mass of element}}{\text{Mass of compound}} \times 100$$

$$\text{Percentage of carbon} = \frac{5.217\text{ g}}{8.657\text{ g}} \times 100 = 60.26\%$$

$$\text{Percentage of hydrogen} = \frac{0.962\ g}{8.657\ g} \times 100 = 11.11\%$$

$$\text{Percentage of oxygen} = \frac{2.478\ g}{8.657\ g} \times 100 = 28.62\%$$

So, 100 g of the compound under study contains 60.26 g of carbon, 11.11 g of hydrogen and 28.62 g of oxygen.

15.4 Quick Check! Calculate the percentage of sulphate ions (SO_4^{2-}) in $Al_2(SO_4)_3$.

15.5 Empirical Formula

The formula which shows the simplest whole number ratio of atoms present in a compound is called empirical formula. For example, empirical formula of hydrogen peroxide (H_2O_2) is HO, that of water (H_2O) is H_2O and for benzene (C_6H_6) is CH.

All the ionic compounds are represented by their empirical formulae. These formulae show the simplest ratio present between their ions. The formula of sodium chloride(NaCl) represents the simplest ratio between sodium and chloride ions present in it. Similarly the formula of calcium chloride ($CaCl_2$) shows the ratio present between calcium and chloride ions.

Since an empirical formula does not tell us about the actual number of atoms present in that compound rather it represents the simplest ratio of minimum ratio between atoms, it is possible that some compounds may have same empirical formula. For example, both acetylene (C_2H_2) and benzene (C_6H_6) have same empirical formula CH. In the same way acetic acid (CH_3COOH) and glucose ($C_6H_{12}O_6$) also have the same empirical formula CH_2O.

Empirical formula of a compound can be calculated following the steps shown below.

I. Determine the percentage composition of the compound.

ii. Find out the number of moles of each element. This is done by dividing the mass of each element (percentage of element) by its atomic mass.

iii. Determine the atomic ratio of each element in the compound. This ratio can be found by dividing the number of moles of each element by the smallest number of moles.

iv. If the atomic ratio obtained is simple whole number, it gives the empirical formula, otherwise multiply this ratio with suitable digit to get the whole number ratio.

v. Use this whole number ratio as a subscript for the elements to get the empirical formula.

15.11 Example

The percentage composition of a compound shows it to contain 25.26% Mg and 74.74% Chlorine. Find out its empirical formula.

Solution:

1. Change the percentage of elements to grams

$$Mg = 25.26\ \% = 25.26\ g \qquad Cl = 74.74\ \% = 74.74\ g$$

2. No. of moles of each element

$$Mg = \frac{25.26}{24} = 1.05 \qquad Cl = \frac{74.74}{35.5} = 2.10$$

3. Atomic ratio of each element

$$Mg = \frac{1.05}{1.05} = 1 \qquad Cl = \frac{2.10}{1.05} = 2$$

The empirical formula of magnesium chloride is $MgCl_2$.

15.5 Quick Check!

A 0.5 g sample of compound contains 0.418 g of antimony (Sb) and 0.082 g of oxygen. What is the empirical formula of the compound?

15.12 Example

The percentage composition of an organic compound by mass is 64.8% carbon, 15.62% hydrogen and 21.58% oxygen. Find out its empirical formula.

Solution:

1. Change the percentages into grams

$$C = 64.8\ \% = 64.8\ g \qquad H = 15.62\ \% = 15.62\ g \qquad O = 21.58\ \% = 21.58\ g$$

2. No. of moles of each element

$$C = \frac{64.8}{12} = 5.4 \qquad H = \frac{15.62}{1} = 15.62 \qquad O = \frac{21.58}{16} = 1.35$$

3. Atomic ratio of each element

$$C = \frac{5.4}{1.35} = 4 \qquad H = \frac{15.62}{1.35} = 10 \qquad O = \frac{1.35}{1.35} = 1$$

The empirical formula of the compound is $C_4H_{10}O$.

Molecular Formula

Molecular formula of an element or a compound represents the actual number of atoms present in the molecule of these substances. Water, hydrogen peroxide, ethylene and sulphur have molecular formulae H_2O, H_2O_2, C_2H_4 and S_8 respectively. Molecular formula of a compound can be calculated by multiplying a whole number with the empirical formula of that compound as shown in the following equation.

$$\text{Molecular Formula} = n\,(\text{Empirical Formula})$$

Where $$n = \frac{\text{Molecular Mass}}{\text{Empirical Formula Mass}}$$

15.13 Example

Benzene is represented by its empirical formula CH while its molecular mass is 78. Find out its molecular formula.

Solution:

Empirical formula of benzene = CH

Empirical formula mass of benzene = 12+1=13

Molecular mass of benzene = 78

$$n = \frac{78}{13}$$

Therefore the molecular formula of benzene = 6 (Empirical formula)

$= 6 \times CH = C_6H_6$

So molecular formula of benzene comes out to be C_6H_6.

Interesting Information

Stoichiometry is used to determine the optimal dosages of medications for patients based on their weight, age and other factors.

15.14 Example

Glucose has an empirical formula CH_2O and its molecular mass is 180 g mol^{1-}. What will be its molecular formula?

Solution:

Empirical formula of glucose = CH_2O

Empirical formula mass of glucose = 12 + 2 + 16 = 30

Molecular mass of glucose = 180

$$n = \frac{180}{30} = 6$$

15.6 Quick Check!

An unknown hydrocarbon contains 85.71% carbon. Its molar mass is 84 g/mol. What is its molecular formula?

Molecular formula of glucose = 6 (CH_2O) = $C_6H_{12}O_6$

So, molecular formula of glucose will be $C_6H_{12}O_6$.

Calculations Based on Chemical Equation

A complete and balanced chemical equation tells us the mole ratio or mass ratio between the reactants and the products. With the help of this ratio we can find out the masses of the products from the given masses of the reactants or the masses of the reactants from the given masses of the products.

For example, the following equation tells us that one mole (100 g) calcium carbonate reacts with two moles (73 g) of hydrochloric acid to produce one mole (111 g) of $CaCl_2$, one mole (18 g) of water and one mole (44 g) of carbon dioxide.

$$CaCO_{3\,(s)} + 2HCl_{(aq)} \longrightarrow CaCl_{2\,(aq)} + H_2O_{(l)} + CO_{2\,(g)}$$

$CaCO_3$	2HCl	$CaCl_2$	H_2O	CO_2
1 mole	2 moles	1 mole	1 mole	1 mole
100 g	73 g	111 g	18 g	44 g

15.7 Quick Check!

Why it is necessary for a chemical equation to be balanced before it can be used in calculation?

15.15 Example

25 g of lime stone ($CaCO_3$) react with an excess of hydrochloric acid according to the above equation. How much calcium chloride ($CaCl_2$) will be produced?

Solution:

According to the equation

100 g of lime stone react to produce calcium chloride = 111 g

1 g of lime stone will react to produce calcium chloride = $\frac{111}{100}$ g

25 g of lime stone will react to produce calcium chloride = $\frac{111}{100} \times 25$

= 27.75 g

15.6 Limiting Reactant

In a chemical reaction, reactants are not always mixed according to stoichiometric ratio shown in the balanced chemical equation. Most often in experimental works, one or more reactants is /are used in large excess. This is done deliberately to ensure that the other expensive reactant is used completely.

In this way the reactant which is used in excess is left behind at the end of the reaction and the other reactant is consumed completely.

This reactant which is consumed earlier is called the limiting reactant. In fact, it is the limiting reactant which controls the amount of the product formed in a chemical reaction.

How to Identify a Limiting reactant

A limiting reactant present in a chemical reaction can be identified by carrying out the following steps.

1. Calculate the number of moles of the reactants from the amounts provided.
2. Calculate the number of moles of the product which will form with the given moles of reactants.
3. Identify the reactant which gives the least amount of the product. This reactant will be the limiting reactant.

15.16 Example

Calculate the amount of carbon dioxide in grams produced when 25 g of glucose reacts with 40 g of oxygen in the following photosynthesis reaction.

$$\underset{\text{Glucose}}{C_6H_{12}O_{6\,(s)}} + 6O_{2\,(g)} \longrightarrow 6CO_{2\,(g)} + 6H_2O_{\,(l)}$$

Solution:

In order to find out the amount of CO_2 in this reaction, we will have to identify which of the reactants is a limiting reactant

1. No. of moles of the reactants

$$\text{No. of moles of glucose} = \frac{25}{180} = 0.148 \text{ mol}$$

$$\text{No. of moles of oxygen} = \frac{40}{32} = 1.25 \text{ mol}$$

2. Calculate the no. of moles of CO_2, which will be given by each of the reactants as shown in the equation:

$$\text{1 mole of glucose gives} = 6 \text{ moles of } CO_2$$

$$\text{0.138 mol of glucose will give} = \frac{6}{1} \times 0.148$$

$$= 0.848 \text{ mol}$$

15.8 Quick Check!

If you react 100 g of chlorine gas with 500 g of KI, which reactant will be the limiting reactant?

$$Cl_{2\,(g)} + 2KI_{\,(aq)} \longrightarrow KCl_{\,(aq)} + I_{2\,(g)}$$

6 moles of oxygen gives $= 6$ moles of CO_2

1.25 mol of oxygen will give $= \frac{6}{6} \times 1.25$

$= 1.25$ mol

Since CO_2 obtained from glucose is less than that of oxygen, so glucose is the limiting reactant.

Amount of CO_2 produced $= 0.878 \times 44 = 36.96$ g

15.7 Yield

The quantity of the product obtained when chemical reaction is performed is called the actual yield. This yield is always less than the theoretical yield. The theoretical yield is the amount of product which we calculate assuming that all the reactants react according to the balanced chemical equation.

A chemist is usually interested in the efficiency of the reaction which is commonly expressed by comparing the actual yield with the theoretically yield in the form of percentage yield. Percentage yield is calculated by applying the following formula.

$$\text{Percentage Yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

15.17 Example

A chemist reacts 50 g of Zn with excess of nitric acid to get 80.2 g $Zn(NO_3)_2$. Calculate the theoretical yield and the percentage yield of this reaction.

Solution:

1. Balanced chemical equation for the reaction

$$Zn_{(s)} + 2HNO_{3\,(aq)} \longrightarrow Zn(NO_3)_{2\,aq} + H_{2\,(g)}$$

2. No. of moles of Zn $= \frac{50}{65.37} = 0.76$ mol

Amount of $Zn(NO_3)_2$ $= 80.2$ g

Theoretical yield

According to the balanced chemical equation

1 mol of Zinc gives $=$ 1 mol of $Zn(NO_3)_2$

0.76 mol of Zinc gives $=$ 0.76 mol of $Zn(NO_3)_2$

Amount of $Zn(NO_3)_2$ in grams $= 0.76 \times 189.38 = 143.92$ g

Percentage yield

$$\text{Percentage Yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

$$= \frac{80.2}{143.92} \times 100 = 55.7\ \%$$

15.9 Quick Check!

Find out the theoretical yield and percent yield of oxygen generated by heating 40 g of $KClO_3$ (M = 122.5). The mass of oxygen gas produced is 14.9g

$$2KClO_{3\,(s)} \longrightarrow 2KCl_{(s)} + 3O_{2\,(g)}$$

Percentage Purity

In chemistry, purity refers to the degree to which a substance is free from impurities. Percentage purity of compound indicates the amount of pure compound present in an impure sample. Percentage purity of a sample is calculated by the following formula.

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

15.18 Example

12.0 g sample of a pharmaceutical product was found to contain 11.57 g of the active drug. Find out the percentage purity of the sample.

Solution:

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

$$= \frac{11.57}{12.0} = 96.4\ \%$$

EXERCISE

A Multiple Choice Questions

Choose and tick the correct answer from the given choices.

1. Volume occupied by 15 moles of ammonia gas at RTP:

(a) $224\ dm^3$ (b) $160\ dm^3$

(c) $360\ dm^3$ (d) $265\ dm^3$

2. How many moles of natural gas CH_4 will be present in a cylinder if its volume is 18 dm^3 at RTP?

(a) 0.75 mol (b) 0.67 mol

(c) 0.85 mol (d) 0.56 mol

3. What will the molar concentration of potassium hydroxide solution if 2.0 g of it are dissolved in 100 cm^3 of solution?

(a) 0.36 mol/dm^3 (b) 0.53 mol/dm^3

(c) 0.70 mol/dm^3 (d) 0.45 mol/dm^3

4. What will be the mass of NaCl needed to make up 15 dm^3 of its solution with a concentration of 0.7 g/dm^3?

(a) 21 g (b) 10.5 g

(c) 31.5 g (d) 5.4 g

5. Find out the concentration of H_2SO_4 in mol/dm^3 if its 10 cm^3 react with 20 cm^3 of NaOH solution having concentration 0.1 mol/dm^3.

(a) 0.01 mol/dm^3 (b) 0.001 mol/dm^3

(c) 0.1 mol/dm^3 (d) 0.15 mol/dm^3

6. If the molar mass of a compound is 178 g/mol and its empirical formula is $C_3H_5O_3$. What is its molecular formula?

(a) $C_3H_5O_3$ (b) $C_6H_{10}O_6$

(c) $C_9H_{15}O_9$ (d) $C_8H_{10}O_8$

7. When natural gas burns in air, which component is the limiting reactant?

(a) Both components

(b) Neither component is a limiting reactant

(c) Air

(d) Natural gas

8. If the actual yield of a compound is 0.198 g while the theoretical field is 0.217 g. What will be the percentage yield?

(a) 50% (b) 91.2%

(c) 80% (d) 90%

B Short Answer Questions

15.1. What is the molar volume of a gas at RTP?

15.2. How does the concept of molar volume useful for us?

15.3. How does the molar volume relate to the Avogadro's law?

15.4. Why two different compounds may show the same empirical formula?

15.5. Why is percentage yield important?

15.6. What is the concentration of a solution in mol/dm^3 if it contains 49 g of H_2SO_4 in one dm^3 of solution?

C Constructed Response Questions

15.1. How would you identify the limiting reactant in a chemical reaction?

15.2. Why is it important to find the purity of a compound which is used as a medicine?

15.3. Will there be a limiting reactant in a reversible reaction?

15.4. How do you know if a formula is empirical or not?

15.5. Write down one method to find out the molecular mass of a compound without knowing its molecular formula.

D Numerical Problems

15.1. Calculate the number of moles, volume and number of molecules in 4.0 g of CH_4 at RTP.

15.2. What mass of NaCl need to be dissolved in 150 cm^3 of a solution if the concentration of the solution is 0.4 mol/dm^3?

15.3. Find out the percentage of sodium in $NaHCO_3$.

15.4. A sulphide of iron contains 1.926 g of sulphur and 2.333 g of iron. Find out the empirical formula of this compound.

15.5. A compound contains by mass, 40.0% C, 6.71 % H and 53.3 % oxygen. A 0.320 mole of this compound weighs 28.8 g. What is the molecular formula of this compound?

15.6. The formation of ammonia gas is given by the following equation:

$$3H_{2\,(g)} + N_{2\,(g)} \longrightarrow 2NH_{3\,(g)}$$

If you start reacting 12 g of H_2 with 64 g of N_2, which will be the limiting reactant?

15.7. When 305 g of $AgNO_3$ is reacted with excess of $MgCl_2$ it produces 23.7 g of $Mg(NO_3)$. What is the percentage yield of the reaction?

$$2AgNO_{3\,(aq)} + MgCl_{2\,(aq)} \longrightarrow Mg(NO_3)_{2\,(aq)} + 2AgCl_{\,(s)}$$

15.8. A sample of metal has a total mass of 3.66 kg and contains 2.45 kg of gold. What is the percentage purity of gold in this sample?