Chemical calculations

Basic concepts:

1. Those ions that are not involved in a chemical reaction are called Spectator ions.

2. The relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction is known as Stoichiometry of the reaction.

3. The reactant that is completely used up in a reaction and determined the amount of products formed is called Limiting reactant.

4. The concentration of a solution is given by the amount of a solute dissolved in a unit volume of the solution.

5. The concentration on of a solution expressed in mol/dm^3 is known as Molar concentration.

6. The calculated amount of products that would be obtained if the reaction is completed is called Theoretical Yield.

7. The amount of pure products that is actually produced in the experiment is called Actual Yield.
8. The Percentage Composition of various elements of a compound is the mass of each element in 100 g of the compound.

ANALYSE

- Interpret and construct chemical equations (including ionic equations) with state symbols.

1.1 Interpreting Chemical Equations:

1. A chemical equation is a short way of representing a chemical reaction with the help of symbols of the reactants and products involved in the chemical reaction.It can be written in words or by using chemical formulae.

2. The substances which react together are called reactants. The new substances formed are called the products.
   An arrow ($\longrightarrow$) pointing towards the right hand side means towards the products is put between the reactants and the products. This arrow shows that the substances written on its left hand side react together and produce substances written on the right hand side of the arrow.
3. Information obtained through a chemical equation:
   The qualitative and quantitative information obtained by chemical equation as follows:
   i) The substances that react
   iii)The relative number of molecules of reactants and products
   iv)The relative weights of reactants and products
   v)Relative volumes of gaseous substances

   Example:
   The reaction between hydrochloric acid and sodium to produce sodium chloride and hydrogen can be represented as shown below.

$$\underbrace{\underset{\text{Hydrochloric acid}}{2HCl(aq)} + \underset{\text{Magnesium}}{2Na(s)}}_{\text{Reactants}} \longrightarrow \underbrace{\underset{\text{Magnesium Chloride}}{2NaCl\,(aq)} + \underset{\text{hydrogen}}{H_2(g)}}_{\text{Products}}$$

1.2 Writing Chemical Equations:

A chemical equation is balanced if the number of atoms of each element of atoms is the same on both sides of the equation.

Example:

When the compound metaldehyde($C_8H_{16}O_4$) burns in excess oxygen, carbon dioxide and water are produced. The steps in writing the balanced chemical equation for this reaction are shown below.

Step 1	Write down the chemical formulae of the reactants and products for the reaction.	$C_8H_{16}O_4 + O_2 \longrightarrow CO_2 + H_2O$ This equation is not balanced.
Step 2	Balance the number of carbon and hydrogen atoms on both sies of the equation. • 1 mol of $C_8H_{16}O_4$ contains 8 carbon atoms. 8 mol of CO_2 contain 8 carbon atoms. Hence, 8 mol of CO_2 is fomed. • 1 mol of $C_8H_{16}O_4$ contains 16 hydrogen atoms. 8 mol of H_2O contain 16 hydrogen atoms. Hence, 8 mol of H_2O is formed.	$C_8H_{16}O_4 \longrightarrow 8CO_2 + 8H_2O$
Step 3	Balance the number of oxygen atoms. • There are (16 + 8) = 24 oxygen atoms on the right side of the equation. • There are 4 oxygen atoms on the left side of the equation. • So we need to add (24 – 4) = 20 oxygen atoms or 10 O_2 molecules on the left side of the equation.	$C_8H_{16}O_4 + 10O_2 \longrightarrow 8CO_2 + 8H_2O$ This equation is now balanced
Step 4	Add the state symbols: (s) for solid, (l) for liquid, (g) for gas and (aq) for aqueous solution.	$C_8H_{16}O_4(s) + 10O_2(g) \longrightarrow 8CO_2(g) + 8H_2O(l)$

Common Errors	Actual Facts
The chemical equation for the reaction between zinc and hydrochloric acid is: $Zn + 2HCl \longrightarrow ZnCl_2 + 2H$	The reaction between zinc and hydrochloric acid produces hydrogen gas and not hydrogen ions. Thus, the correct chemical equation is: $Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$

Magnesium reacts with chlorine to produce magnesium chloride. $2Mg + Cl_2 \longrightarrow 2MgCl$	The formula for magnesium chloride is $MgCl_2$. Thus, the corret chemical equation is: $Mg(s) + Cl_2\ (g) \longrightarrow 2MgCl_2(s)$

Tip for students:
In writing chemical equation involving state symbols, the state symbol (aq) refers to a substance dissolved in water while the state symbol (l) refers to a pure liquid, e.g., liquid bromine, $Br_2(l)$, water H_2O (l) or molten sodium chloride,$NaCl(l)$.

1.3. Writing Ionic Equations:

An ionic equation is the simplified chemical equation that shows the ions taking Part in a reaction and the products formed in aqueous solution or water. It leaves out the spectator ions that do not react.

Example:
When aqueous sodium chloride is added to aqueous silver nitrate, a white precipitate of silver chloride is formed.The steps in writing the ionic equation for this reaction are shown below:

Step 1	Write the balanced chemical equation of the reaction (include state symbols).	$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$
Step 2	Identify ionic compounds that are soluble in water. These compounds become ions in water. Rewrite the chemical equation in terms of ions.	$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \longrightarrow AgCl(s) + Na^+(aq) + NO_3^-$
Step 3	Cancel out the spectator ions.	$\cancel{Na^+(aq)} + Cl^-(aq) + Ag^+(aq) \cancel{+ NO_3^-}(aq) \longrightarrow AgCl(s) \cancel{+ Na^+}(aq) \cancel{+ NO_3^-}$ In this reaction,the spectator ions are the Na^+ and NO_3^- ions.
Step 4	Write the ionic equation.	$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$

The actual reaction is between silver ions and chloride ions that form a white precipitate of silver chloride.the Spectator ions are ions that are not involved in a chemical reaction.

1.4. Calculations from Chemical Reactions:

1.the relationship between the amounts of reactants and products involved in a chemical reaction is known as the Stoichiometry of the reaction.

2. The mass of any reactant or product in a reaction can be calculated from a balanced equation.

Example:

Consider the following reaction:

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

Tip for students:

When doing chemical calculations, be careful when writing the formulae of compounds. Thus, sodium sulfate is Na_2SO_4 (not $NaSO_4$), copper(II) nitrate is $Cu(NO_3)_2$ (not $CuNO_3$) and calcium ethanoate is $(CH_3COO)_2Ca$ and not CH_3COOCa.

We can read this equation as follows:

1 mol of magnesium	Reacts with →	2 mol of hydrochloric acid	to give →	1 mol of magnesium chloride	and	1 mol of hydrogen

Hence, we can deduce that

24 g of magnesium	React with →	73 g of hydrochloric acid	to give →	95g of magnesium chloride	and	2g of hydrogen

Example1:

What is the mass of sodium oxide produced when 18.4 g of sodium is completely burnt in oxygen?

Solution:

Number of moles of sodium = $\frac{mass}{Atomic\ mass} = \frac{18.4}{23} = 0.8$ mol

The equation for the reaction is:

$$4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$$

According to the equation:

4 mol of Na produce 2 mol of Na_2O.

0.8 mol of Na produces $\frac{0.8}{4} \times 2 = 0.4$ mol of Na_2O.

Mass of sodium oxide produced = number of moles of Na_2O x M_r

= 0.4 x (23 x 2 + 16)

= 0.4 x 62 = 24.8 g

4. In a chemical calculation that involves gases, we can change "mole" of gas to "volume" of gas because the volume of gas is proportional to the number of moles of the gas, and vice versa.

Example:

Consider the following equation.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

21 mol of methane	Reacts With →	2 mol of oxygen	to give →	1 mole of carbon dioxide

This means

21 mol of methane	Reacts With →	2 mol of oxygen	to give →	1 mole of carbon dioxide

Example2:

When powdered iron is added to dilute hydrochloric acid, 720 cm^3 of hydrogen gas is produced. What is the mass of iron used?

Solution:
The equation for the reaction is:
$Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2(g)$

Number of moles of H_2 produced = $\frac{720}{24000}$ = 0.03 mol
Mass of iron used = number of moles x A_r = 0.03 x 56 = 1.68 g

1.5 Limiting reactants:

1. The reactant that is completely used up in a reaction is known as limiting reactant. It determines or limits the amount of products formed.

   Example:
   a) Consider the reaction between sodium hydrogen carbonate and dilute hydrochloric acid. The chemical equation for the equation is:
   $$NaHCO_3(s) + HCl(aq) \longrightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

   b) The equation shows that when 1 mol of $NaHCO_3$ reacts with 1 mol of HCl, 1 mol of CO_2 is produced

   c) If 1 mole of $NaHCO_3$ reacts with 2 mol of HCl, the amount of CO_2 produced is still 1 mol because $NaHCO_3$ is completely used up when it reacts with 1 mol of HCl

   In this reaction, hydrochloric acid is in excess (excess reactant) and sodium hydrogen carbonate is the limiting reactant.

2. Importance of identifying the limiting reactant :
   a) In the chemical industry, large amounts of chemical are required to manufacture a particular product.
   b) To get the maximum yield of a producer at the minimum cost, we ned to know the limiting reactant.
   c) We will generally choose the most expensive reactant to be the limiting reactant, and use excess amounts of the other reactant in a reaction. This ensures all the expensive reactant is used up.
   d) In many industrial reactions, excess reactants are recycled as far as possible in order to reduce production costs.

Example3:

The equation for the reaction between hydrogen and chlorine and chloride is show below.

$H_2(g) + Cl_2(g) \longrightarrow 2HCl\ (g)$

In a reaction, 20 cm^3 of hydrogen gas is used to react with 30 cm^3 of chlorine gas.

a) Identify the limiting reactant.
b) At the end of the reaction, what is the volume of
    a. Hydrogen?
    b. Chlorine?
    c. Hydrogen chloride?

Solution:

a) The equation shows that:
   1 mol of H_2 reacts with 1 mol of Cl_2 to give 2 mol of HCl.
   That is,
   20 cm^3 of H_2 reacts with 20cm^3 of Cl_2 to give 40 cm^3 of HCl

   In this reaction, since 30 cm^3 of Cl_2 is used, Cl_2 must be in excess.
   Thus, H_2 is the limiting reactant.

b) i) Volume of hydrogen = 0 cm^3
   ii) Volume of chlorine = original volume – volume reacted
   = 30 – 20 = 10 cm^3
   iii) Volume of hydrogen chloride = 40 cm^3

Analyse:

- Calculate the concentration of solutions (in mol/dm^3 or g/dm^3)

1.6 Concentration of Solutions:

1. The concentration of a solution is the amount dissolved in a unit volume of the solution.

2. We can express concentration in two ways:
   i) Grams of solute per litre of solution
   ii) Number of moles of solute per litre of solution

Concentration $(g/dm^3) = \frac{mass\ of\ solute\ (g)}{volume\ of\ solution\ (dm3\ )}$

Concentration $(mol/dm^3) = \frac{number\ of\ moles\ of\ solute}{volume\ of\ solution\ (dm3\ )}$

3. When the concentration of a solution is expressed in mol/dm^3,it is called molar concentration.

4. A one molar solution means the concentration of the solution is 1 mol/dm^3. A two molar solution is a solution of concentration 2 mol/dm^3.

Example4:

a) A solution of sodium hydroxide contains 3.0 g of sodium hydroxide in 75 cm_3 of solution. What is the concentration of the solution in g/dm^3.
b) A glucose solution has a concentration of 52.0 g/dm^3. What is the mass of glucose present in 250 cm^3 of the solution?

Solution:

a) 1 dm3 = $1000cm^3$

Volume of solution = 75 cm^3 = 0.075 dm^3

Concentration of sodium hydroxide solution = $\frac{mass\ of\ the\ solute\ (g)}{volume\ of\ solution\ (dm3)}$

$= \frac{3.0\ g}{0.075\ dm3}$ = 40 g/dm^3

b) 1000 cm^3 solution contains 52.0 g of glucose.

250 cm^3 contains $\frac{250}{1000}$ x 52.0 = 13 g of glucose.

Example5:

In an experiment, 3.71 g of sodium carbonate, Na_2CO_3, was dissolved in water and made up of 500 cm^3. What is the concentration of the solution in mol/dm^3?

Solution:

Relative molecular mass of Na_2CO_3 = (2 x 23) + 12 + (3 x 16) = 106

Number of mole of , $Na_2CO_3 = \frac{mass}{Mr} = \frac{3.71}{106} = 0.035$ mol

Volume of , Na_2CO_3 solution = 500 cm_3 = 0.5 dm_3

Concentration of solution $= \frac{number\ of\ moles}{volume\ of\ solution\ (dm3)} = \frac{0.035\ mol}{0.5\ dm3} = 0.07\ mol/dm^3$

5. Relationship between number of moles and concentration in g/dm^3

$$\boxed{\text{Concentration mol/dm}^3} \underset{M_r \text{ of solute}}{\overset{\text{x } M_r \text{ of solute}}{\rightleftarrows}} \boxed{\text{Concentration g/dm}^3}$$

Example6:

The concentration of a potassium hydroxide solution is 0.5 mol/dm^3

a) What is its concentration in g/dm^3
b) Calculate thevolume of the solution that contains 3.36 g of potassium hydroxide.

Solution:

a) Concentration of KOH in g/dm^3 = concentration in mol/dm^3 x Mr of KOH
= 0.5 x (39 + 16 + 1)
= 28 g/dm^3

b) 1000 cm^3 contains 28g of KOH
Volume of KOH solution that contains 3.36 g of KOH = $\frac{3.36}{28}$ x 1000
= 120 cm^3

1.7 Volumetric Analysis:

1. The concentration of an acid or alkali can be determined by titration.

2. In titration experiments, we determine the volume of a chemical solution (reagent) required to completely react with a known volume of another solution. The technique is also called volumentric analysis.

3. An indicator is used in acid-base titration.

4. The end-point is reached when the indicator changes colour. At the end-point, complete reaction between the acid and alkali occurs.

Example7:

In an experiment, 25.0 cm^3 of sodium hydroxide needed 27.5 cm^3 of 1.0 mol/dm^3 sulfuric acid for complete reaction. Calculate the concentration of sodium hydroxide in mol/dm^3.

Solution:

Number of moles of sulfuric acid used = concentration (mol/dm^3) x volume (dm^3)

$$= 1.0 \text{ x } \frac{27.5}{1000} = 0.0275 \text{ mol}$$

The equation for the reaction is:

$2NaOH(aq) + H_2SO_4\,(aq) \qquad Na_2SO_4\,(aq) + 2H_2O\,(l)$

From the equation,

Number of moles of sodium hydroxide present = 0.0275 x 2 = 0.055 mol

Volume of sodium hydroxide = 25.0 cm^3 = 0.025 dm^3

$$\text{Concentration of sodium hydroxide} = \frac{number\ of\ moles\ of\ sodium\ hydroxide}{volume\ of\ sodium\ hydroxide\ (dm3)}$$

$$= \frac{0.055\ mol}{0.025\ dm3} = 2.2 \text{ mol/dm}^3$$

Analyse:

- Calculate the percentage yield of a product in a reaction and the percentage purity of a substance

1.8 Percentage Yield of a reaction:

1. The theoretical yield is the calculated amount of products that would be obtained if the reaction is completed.

2. The actual yield is the amount of pure products that is actually produced in the experiment.

3. In many reactions, the actual yield obtained in the experiment is less than the theoretical yield.

4. There are many reasons for such differences, such as

   a. The reactant may be impure,
   b. The reaction is incomplete
   c. The reactant is volatile and is lost due to evaporation.

5. The percentage yield shows the relationship between yield and theoretical yield.

$$\text{Percentage yield} = \frac{actual\ yield}{theoretical\ yield} \times 100\%$$

Common Error	Actual Fact
When the excess reactant is an impure substance, the percentage yield of the reaction must be less than 100%	When the limiting reactant is an impure substance, the percentage yield of the reaction must be less than 100%

Example8:

When zinc nitrate is heated strongly, it decomposes to form zinc oxide.

$$2Zn(NO_3)_3(s) \longrightarrow 2ZnO(s) + 4NO_2(g) + O_2(g)$$

In the above reaction, 7.8 g of zinc oxide was obtained when 18.9 g of zinc nitrate was heated strongly. Calculate the percentage yield of this reaction.

Solution:

M_r of $Zn(NO_3)_2$ = 65 + (2 x 14) + (6 x 16) = 189

Number of moles of $Zn(NO_3)_2$ used = $\frac{mass}{Mr} = \frac{18.9}{189}$ == 0.1 mol

Based on the equation,

Ratio of number of moles of $Zn(NO_3)_2$: ZnO = 1:1

Number of moles of ZnO produced = 0.1 mol

Mass of ZnP produced = number of moles x Mr

= 0.1 x (65 + 16) = 0.1 x 81 = 8.1 g

$$\text{Percentage yield} = \frac{actual\ yield}{theoretical\ yield} \times 100\ \%$$

$$= \frac{7.8}{8.1} \times 100\% = 96.3\%$$

1.9 Percentage Purity of a Substance:

If the reactant used in the reaction is impure, we can calculate the percentage purity of the reactant by using the formula.

$$\text{Percentage purity} = \frac{mass\ of\ pure\ substance}{mass\ of\ substance\ used\ in\ teh\ reaction} \times 100\ \%$$

Example9:

The equation for the complete combustion of ethene is shown below.

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(l)$$

When 7.0g of ethene is burnt completely in air, 1.98 g of carbon dioxide was produced. What is the percentage purity of ethane?

Solution:

Number of moles of carbon produced = $\frac{mass}{Mr}$ = $\frac{1.98}{12+(2\ x\ 16)}$ = $\frac{1.98}{44}$ -= 0.045 mol

Based on the equation,
Ratio of number of moles of C2H4 : CO2 = 1:2
Number of moles of pure ethene = ½ x 0.045 = 0.0225 mol
Mass of pure ethene = number of moles x Mr
= 0.0225 x [(2 x 12) + 4 x 1)]
= 0.0225 x 28 = 0.63 g

Percentage purity = $\frac{mass\ of\ pure\ substance}{mass\ of\ substance\ used}$ x 100 %
= $\frac{0.63}{0.70}$ x 100% = 90%

STRUCTURAL AND REASONING BASED QUESTIONS & ANSWERS

1. (a) Are the following statements about balanced equations 'True' or 'False'?
   (i) The formulae of all reactants and products are shown.
   True/False
   (ii) The number of reactants and products are always equal.
   True/False
   (iii) The number of atoms in each element of the reactants is equal to the number of atoms in each element of the products.
   True/False
   (iv) There is always more than one reactant in a reaction.
   True/False
   (v) There is always more than one product in a reaction.
   True/False
   (vi) The net charge of the reactants and the net charge of the products are always equal. True/False

b) Balance the following equations by adding numbers, where necessary, in the blanks.

(i) ____I_2+ ____ N_2H_4 → ____ HI + _____ N_2
(ii) ____Li+ ____ N_2 → _____ Li_3N
(iii) ____ CS_2+ ____O_2 → ____ CO_2 + _____ SO_2
(iv) ____Mg + ____ N_2 → ____ Mg_3N_2
(v) ____CH_4 + ____ H_2O → ____ CO + _____ H_2

Answer:

1. (a) (i) True (ii) False (iii) True (iv) False (v) False (vi) True

(b) (i) $2I_2 + N_2H_4 \rightarrow 4HI + N_2$

(ii) $6Li + N_2 \rightarrow 2Li_3N$

(iii) $CS_2 + 3O_2 \rightarrow CO_2 + 2SO_2$

(iv) $3Mg + N_2 \rightarrow Mg_3N_2$

(v) $CH_4 + H_2O \rightarrow CO + 3H_2$

2.) Write balanced ionic equations, including state symbols, for each of the following reactions.

(a) Sodium hydroxide reacts with sulfuric acid to give sodium sulfate and water.
(b) Barium chloride reacts with sulfuric acid to give solid barium sulfate and hydrochloric acid.
(c) Magnesium reacts with nitric acid to give magnesium nitrate.
(d) Solid copper (II) carbonate reacts with nitric acid to give copper(II) nitrate, carbon dioxide gas and water.

Answer:

1.) (a) $H^+(aq) + OH^-(aq) \rightarrow H_2O(I)$

(b) $Ba^{+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

(c) $Mg(s) + 2H^{+}(aq) \rightarrow Mg^{2+}(aq) + H_2O(I)$

(d) $CuCO_3(S) + 2H^{+}(aq) \rightarrow Cu^{2+}(aq) + CO_2(g) + H_2O(I)$

3.) The following is an example of a balanced chemical equation with state symbols.

$Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$

(a) Identify the state symbols in the equation above, stating what each of them means.

(b) Rewrite the above chemical equation as a word equation.

(c) Write balanced chemical equations for each of the following reactions.

(i) Sodium chloride + sulfuric acid → sodium sulfate + hydrogen chloride

(ii) Sodium + water → sodium hydroxide + hydrogen gas

(iii) Calcium carbonate → calcium oxide + carbon dioxide

(iv) Potassium bromide + chlorine → bromine + potassium chloride

(v) Nitrogen + hydrogen → ammonia

(vi) Magnesium + nitric acid → magnesium nitrate + hydrogen gas

Answer:

3.)(a) The state symbols in the equation are:(s) for solids,(aq)for substances dissolved in water, (g) for gases and (l) for liquids.

(b) Sodium carbonate + hydrochloric acid → sodium chloride + carbon dioxide + water

(c) (i) $2NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2HCI$

(ii) $2Na + 2H_2O \rightarrow 2NaOH + H_2$

(iii) $CaCO_3 \rightarrow CaO + CO_2$

(iv) $2KBr + Cl_2 \rightarrow Br_2 + 2KCl$

(v) $N_2 + 3H_2 \rightarrow 2NH_3$

(vi) $Mg + 2HNO_3 \rightarrow Mg(NO_3)_2 + H_2$

4.) Find the formulae of the following:

(a) An ionic compound made up of K^+ and $Cr_2O_7^{2-}$ ions

(b) An ionic compound made up of Fe^{3+} and OH^- ions

(c) An ionic compound made up of Al^{3+} and SO_4^{2-} ions, as well as some water of crystallization (The water of crystallization accounts for 45.7% by mass of the compound.)

Answer:

(a) For an ionic compound, the charges present have to be balanced.

Since $2(+1) + 1(-2) = 0$, its formula is $K_2Cr_2O_7$.

(b) For an ionic compound, the charges present have to be balanced.

Since $1(+3) + 3((-1) = 0$, its formula is $Fe(OH)_3$.

(c) For this ionic compound, the charges have to be balanced for the ions present since H_2O is uncharged.

Since $2(+3) + (-2) = 0$, the formula of the dehydrated compound is $Al_2(SO_4)_3$.

Let the formula of the compound be $Al_2(SO_4)_3.nH_2O$.

Then % by mass of water = 45.7% = $\frac{18n}{2(27)+3[32+4(16)]+18n} \times 100\%$

$0.457(342 + 18n) = 18n \Rightarrow n = 16$

Thus, its formula is $Al_2(SO_4)_3.16H_2O$.

5.) (a.) In an experiment, 5.61 g of iron powder is burnt in excess oxygen to produce iron(III)oxide, Fe_2O_3.

(i) Construct a balanced equation, with state symbols, for the above reaction.

(ii) If 7.25 g of iron(III) oxide was obtained, calculate the percentage yield of the reaction.

(b.)When solid calcium carbonate is heated, it decomposes to form calcium oxide and carbon dioxide.

(i) Construct a balanced equation, with state symbols, for the above reaction.

ii) In an experiment, 28.12 g of calcium carbonate was heated to form 5.24 dm^3 of carbon dioxide, as measured at room temperature and pressure. Calculate the percentage yield of the reaction.

(c.)Martin performed the following experiment, and recorded all the data in his notebook. However, he accidentally spilled some ink over his notebook, covering some of his data.

Martin 14.07.2013

Zinc is reacted with copper (II) oxide, to form copper and zinc oxide, as shown in the following equation:

$Zn + CuO \rightarrow ZnO + Cu$

Mass of Zn used = 12.8q (in excess)

Mass of CuO used =

Mass of Cu formed = 1.87 g

Percentage yield of Cu = 89.7 %

using the visible information,calculate the mass of CuO that Martin had used.

Answer:

a) i) $4 Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

ii) No. of mol of Fe burnt = 5.61÷ 56 = 0.1 mol

From the equation, $\frac{\text{no.of mol of } Fe_2O_3}{\text{no.of mol of } CaCO_3} = \frac{2}{4} = 0.5$

Max. possible no. of mol of Fe_2O_3 = 0.1× 0.5 = 0.05 mol

M_r of Fe_2O_3 = 2× 56 + 3(16) = 160

Max. possible mass of Fe_2O_3 = 0.05 × 160 = 8 g

Percentage yield = (7.25 ÷ 8) × 100% = 90.6%

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

(ii) M_r of $CaCO_3$ = 40 + 12 + 3(16) = 100

No. of mol of $CaCO_3$ heated = 28.12 ÷ 100 = 0.2812 mol

From the equation, $\frac{\text{no.of mol of } CO_2}{\text{no.of mol of } CaCO_3} = \frac{1}{1} = 1$

Max. possible no. of mol of CO_2 = 0.2812 mol

Max. possible vol. of CO_2 at r.t.p. = 0.2812 × 24 = 6.749 dm^3

Percentage yield = (5.24 ÷ 6.749) × 100% = 77.6%

(c) Zn was used in excess.

Since 1.87 g of Cu corresponds to a yield of 89.7%,

Max. possible mass of Cu = (100 ÷ 89.7) × 1.87 = 2.085 g

⇒ Max. possible no. of mol of Cu = 2.085 ÷ 64 = 0.032 57 mol

From the equation, $\frac{\text{no.of mol of CuO}}{\text{no.of mol of Cu}} = \frac{1}{1} = 1$

⇒ No. of mol of CuO at the start of experiment = 0.032 57 mol

Mass of CuO at the start of experiment = 0.032 57 × (64 + 16) = 2.61g

6.) Some metals, such as iron, can form more than one type of ion in compounds. Certain chemicals are able to change one type of ion to another. For example, Fe^{2+} ions can be converted into Fe^{3+} ions by manganate(VII) (MnO_4^-) ions, according to the following equation:

$MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3} + 4H_2O$

(a) In the laboratory, potassium manganate(VII), $KMnO_4$, is most commonly used as the source of MnO_4^- ions. Which other ion is obtained when $KMnO_4$ is dissolved in water?

(b) A solution containing Fe^{2+} ions was made by dissolving 1.20 g of $FeCl_2$ in 25 cm^3 of water. Calculate the minimum mass of potassium manganate(VII) needed to completely convert all the Fe^{2+} ions into Fe^{3+} ions.

(c) (i) How can the presence of Fe^{3+} ions be tested using common reagent(s) available in the laboratory? Describe what will be observed.

(ii) Construct a balanced ionic equation, with state symbols, for this reaction.

Answer:

6.) (a) K^+

(b) Molar mass of $FeCl_2 = 56 + 2(35.5) = 127$ g/mol

No. of mol of $FeCl_2 = 1.20 \div 127 = 0.009\ 449$ mol

Since mole ratio of $FeCl_2 : Fe^{2+} = 1:1$, no. of mol of $Fe^{2+} = 0.009\ 449$ mol

From the equation, $\frac{\text{no.of mol of } Fe^{2+}}{\text{no.of mol of } MnO_4} = \frac{5}{1} = 5$

$\Rightarrow$ Min. no. of mol of MnO_4^- required $= 0.009\ 449 \div 5 = 0.001\ 890$ mol

Since mole ratio of $KMnO_4 : MnO_4^- = 1:1$,

min. no. of mol of $KMnO_4$ required $= 0.001\ 890$ mol

Molar mass of $KMnO_4 = 39 + 55 + 4(16) = 158$ g/mol

Thus, minimum mass of $KMnO_4$ required $= 0.001\ 890 \times 158 = 0.299$ g (c)(i) Dilute sodium hydroxide can be used. In the presence of Fe^{3+}, a reddish-brown precipitate of iron(III) hydroxide is formed; the precipitate is insoluble in excess sodium hydroxide.

(ii) $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3\ (s)$

7.) Chlorine and iodine are both Group VII elements. As you will learn later, chlorine is more reactive than iodine, and displaces iodine from an aqueous solution containing iodide ions. For example, when chlorine gas is bubbled into a solution of potassium iodide, iodine and potassium chloride are obtained.

(a) Write a word equation for the above reaction.
(b) Write a balanced chemical equation for the above reaction.
(c) Construct an ionic equation for the reaction.
(d) If 16.1 g of Kl reacted, what will be the mass of KCl produced?
(e) Calculate the minimum volume of chlorine gas, at r.t.p., that is required to produce 14.9 g of iodine.
(f) A solution was made by dissolving 4.15 g of potassium iodide in 50 cm^3 of deionised water. 1.6 dm^3 of chlorine gas, at r.t.p., was then bubbled into the solution.
  (i) Which of the two reactants is in excess?
  (ii) Calculate the concentration (in mol/dm^3) of the potassium chloride solution obtained at the end of the reaction.
(g) A sample of potassium iodide is contaminated by some unreactive solid. When 8.09 g of the contaminated potassium iodide was reacted with excess chlorine, 5.30 g of iodine was obtained. Calculate the percentage purity of the potassium iodide sample.

Answer:

11. (a) potassium iodide + chlorine → iodine + potassium chloride

(b) $2Kl + Cl_2 \rightarrow l_2 + 2KCl$

(c) $2l^- + Cl_2 \rightarrow l_2 + 2Cl^-$

(d) Molar mass of KI = 39 + 127 = 166 g/mol

Molar mass of KCI = 39 + 35.5 = 74.5 g/mol

No. of mol of KI reacted = 16.1 ÷ 166 = 0.096 99 mol

From the equation, $\frac{\text{no.of mol of KI}}{\text{no.of mol of KCI}} = \frac{2}{2} = 1$

⇒ No. of mol of KCI formed = 0.096 99 mol

Mass of KCI produced = 0.096 99 × 74.5 = 7.23 g (3 s.f.)

(e) Molar mass of I_2 = 2(127) = 254 g/mol

No. of mol of I_2 in 14.9 g = 14.9 ÷ 254 = 0.058 66 mol

From the equation, $\frac{\text{no.of mol of } I_2}{\text{no.of mol of } CI_2} = \frac{1}{1} = 1$

⇒ Min. no. of mol of CI_2 required = 0.058 66 mol

Min. vol. of CI_2 gas required = 0.058 66 × 24 = 1.41 dm^3 (3 s.f.)

(f) (i) No. of mol of KI in solution = 4.15 ÷ 166 = 0.025 mol

No. of mol of CI_2 bubbled into solution = 1.6 ÷ 24 = 0.066 67 mol

From the equation, $\frac{\text{no.of mol of KI}}{\text{no.of mol of } CI_2} = \frac{2}{1} = 2$

⇒ Only $\frac{0.025}{2}$ = 0.0125 mol of CI_2 is required to react with 0.025 mol of KI

Thus, chlorine gas is in excess.

(ii) From the equation, $\frac{\text{no.of mol of KI}}{\text{no.of mol of KCI}} = \frac{2}{2} = 1$

⇒ No. of mol of KCI formed = 0.025 mol

conc. of KCI solution obtained = 0.025 ÷ 0.05 = 0.50 mol/dm^3

(g) No. of mol of I_2 formed = 5.30 ÷ 254 = 0.020 87 mol

From the equation, $\frac{\text{no.of mol of } I_2}{\text{no.of mol of KI}} = \frac{1}{2}$

$\Rightarrow$ No. of mol of KI reacted $= 2 \times 0.020\ 87 = 0.041\ 73$ mol

Mass of KI reacted $= 0.041\ 73 \times 166 = 6.928$ g

Percentage purity $= (6.928 \div 8.09) \times 100\% = 85.6\%$ (3 s.f.)

8.) Ethene gas (C_2H_4) burns in excess oxygen to give carbon dioxide gas and steam.

(a) Write a balanced chemical equation with state symbols for the above reaction.

(b) In an experiment, 20 cm^3 of ethane gas was burnt in 100 cm^3 of oxygen. Calculate the volumes of carbon dioxide gas and steam that was formed. All the volume measurements were carried out at the same temperature and pressure.

Answer:

8. (a) $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$

(b) Volume of carbon dioxide gas formed $= 2 \times 20 = 40$ cm^3

Volume of steam formed $= 2 \times 20 = 40$ cm^3

9.) The following reactions involve reactants and/or products that are probably unfamiliar to you. Nonetheless, the rules for balancing equations remain the same no matter how complex the substances might be. Balance the following equations by adding numbers, where necessary, in the blanks.

(i) _____ B_2H_6 + ______ NH_3 → ____ $B_3H_6N_3$ _____ H_2

(ii) _____ Pt + ______ HNO_3 +_____ HCl → _____ H_2PtCl_6+ _____ NO_2 + _____ H_2O

(iii) _____ l_2 + ______ $HClO_4$ + _____ O_3 → _____ $l(ClO_4)_3$ + ______ H_2O

(iv)____ Ag + _____ HNO_3 → _____ $AgNO_3$ + _____ H_2O + ______ NO + _____ NO_2

(v) ______ OsO_4 + _____ CO → _____ $Os_3(CO)_{12}$ + _____ CO_2

Answer:

9.)(i) $3B_2H_6 + 6NH_3 \rightarrow 2B_3H_6N_3 + 12H_2$

(ii) $Pt + 4HNO_3 + 6HCI \rightarrow H_2PtCI_6 + 4NO_2 + 4H_2O$

(iii) $I_2 + 6HCIO_4 + O_3 \rightarrow 2I(CIO_4)_3 + 3H_2O$

(iv) $4Ag + 6HNO_3 \rightarrow 4AgNO_3 + 3H_2O + NO + NO_2$

(v) $3OsO_4 + 24CO \rightarrow Os_3(CO)_{12} + 12CO_2$

10.)Ethanol,C_2H_5OH, can be produced by the fermentation of glucose, $C_6H_{12}O_6$, by yeast in the absence of oxygen. Carbon dioxide is also produced during this reaction.

(a) Write a balanced chemical equation for the above reaction.

(b) Paul, a chemist, dissolved some glucose in 20 cm^3 of deionised water in a conical flask. If the resultant solution contained 1.50 mol/dm^3 of glucose, what was the mass of glucose added?

(c) Paul then added some yeast and sealed up the flask with a rubber bung. Next, he inserted a gas syringe through the rubber bung, as shown in the diagram below. The plunger of the gas syringe was at the zero mark at the start of the experiment. The set-up was maintained at 25 °C under atmospheric pressure. After several days, Paul noticed that the plunger of the gas syringe was at the 79.8-cm^3 mark.

(i) Calculate the concentration of glucose, in mol/dm^3, at that moment.

(ii) Calculate the mass of ethanol that had been produced.

(d) Paul carried out some other experiments to demonstrate that ethanol has some properties that are similar to those of water. Firstly, a small amount of sodium was carefully reacted with water to form sodium hydroxide and a colourless, odourless gas. When a burning splint was held at the mouth of a test tube containing this gas, a 'pop' sound was heard.

(i) Identify the gas that was produced.

(ii) Hence, write a balanced chemical equation, including state symbols, for the above reaction.

(e) Next, Paul added 124 mg of sodium to a beaker containing ethanol in excess. A product **S** was formed, and a gas having the same properties as in (d) was also given off.

(i) Based on the similar reaction of sodium with water, deduce the formula of **S**.

(ii) Write a balanced chemical equation for the above reaction.

(iii) Calculate the percentages by mass of the various elements present in **S**. Give your answers correct to one decimal place.

(iv) Given that 0.323 g of **S** was obtained, calculate the percentage yield.

Answer:

10.)(a) $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$

(b) M_r of glucose = 6(12) + 12(1) + 6(16) = 180

No. of mol of glucose in 20 cm^3 of solution = 0.020 × 1.50 = 0.030 mol

Mass of glucose added = 0.030 × 180 = 5.4 g

(c) (i) No. of mol of CO_2 produced = 79.8 ÷ 24 000 = 0.003 325 mol

= 3.325 mol

From the equation, $\frac{\text{no.of mol of glucose}}{\text{no.of mol of carbon dioxide}} = \frac{1}{2}$

⇒ No. of mol of glucose reacted = 3.325 mmol × 0.5 = 1.6625 mmol

No. of mol of glucose remaining = 30 – 1.6625 = 28.34 mmol

Conc. of glucose= (28.34 ÷ 30) × 1.50 = 1.42 mol/dm^3 (3 s.f.)

(ii) M_r of ethanol = 2(12) + 6(1) + 1(16) = 46

From the equation, $\frac{\text{no.of mol of } CO_2}{\text{no.of mol of ethanol}} = \frac{2}{2} = 1$

⇒ No. of mol of ethanol formed = 3.325 mmol

Mass of ethanol formed = 0.003 325 × 46 = 0.153 g (3 s.f.)

(d) (i) hydrogen (ii) $2H_2O(I) + 2Na(s) \rightarrow 2NAOH(aq) + H_2(g)$

(e) (i) **S** is C_2H_5ONa. (ii) $2C_2H_5OH + 2Na \rightarrow 2C_2H_5ONa + H_2$

(iii) M_r of **S** = 2(12) + 5(1) + 1(16) + 1(23) = 68

% by mass of C = (24 ÷ 68) × 100% = 35.3% (1 d.p.)

% by mass of O = (16 ÷ 68) × 100% = 23.5% (1 d.p.)

% by mass of Na = (23 ÷ 68) × 100% = 33.8% (1 d.p.)

% by mass of H = (5 ÷ 68) × 100% = 7.4% (1 d.p.)

(iv)No. of mol of Na added = 0.124 ÷ 23 = 5.391 mmol

From the equation, $\frac{\text{no.of mol of }\mathbf{S}}{\text{no.of mol of Na}} = \frac{2}{2} = 1$

⇒ Max. possible no. of mol of **S** = 5.391 mmol

Max. possible mass of **S** = 0.005 391 × 68 = 0.3666 g

Percentage yield of **S** = (0.323 ÷ 0.3666) × 100% = 88.1%

11.) Tungsten (W) is a transition metal which does not react with water and dilute acids. It is used in many high temperature application and can be extracted from tungsten (VI) oxide by reduction with hydrogen.

a. Write a balanced chemical equation for the reduction of tungsten (VI) oxide.

b. 88kg of tungsten was obtained from the reduction of 116 kg of tungsten oxide.calculate the percentage yield of tungsten.

ANSWERS:

11.)(a) $WO_3 + 3H_2 \longrightarrow W + 3H_2O$

(b) Relative formula mass of WO_3 = 184 + 3(16) = 232
No. of mol of WO_3 reacted = 116 000 ÷ 232 = 500
1 mol of WO_3 react to give 1 mol of W

Max. possible mass of W formed = 500 x 184 = 92 000 g = 92 kg

Percentage yield = $\frac{88}{92}$ x 100% = 95.7%

12.) Dolomite is a mineral composed of a double carbonate of Magnesium and Calcium.

a. Suggest whether calcium carbonate or magnesium carbonate will decompose first on heating.

b. When 2.00g of an impure sample of dolomite was completely reacted with excess dilute hydrochloric acid, 480 cm^3 of carbon dioxide gas, measured at room temperature and pressure, was given off.

i. Given that the formula of this double carbonate of magnesium and calcium can be written as $CaMg(CO_3)_2$ write a balanced chemical equation for the reaction between $CaMg(CO_3)_2$ and dilute Hydrochloric acid.

ii. Calculate the percentage purity of this sample of dolomite.

ANSWERS:

(a) Magnesium carbonate will decompose first on heating.

(b) (i) $CaMg(CO_3)_2 + 4HCl \rightarrow CaCl_2 + MgCl_2 + 2CO_2 + 2H_2O$

(ii) No. of mols of CO_2 given off = 0.48 ÷ 24 = 0.02

NO. of moles of $CaMg(CO_3)_2$ present = 0.02 ÷ 2 = 0.01

Mass of $CaMg(CO_3)_2$ in sample = 0.01 x 184 = 1.84 g

% purity = $\frac{1.84}{2.00}$ x 100% = 92%

13.) Robin carried out an experiment to determine the percentage purity of an iron wire.The procedures he followed and data he obtained are recorded in his note book as shown below.

Mass of impure iron wire = 0.145 g
Step1: The iron is placed in a conical flask. 30 cm^3 of dilute sulfuric acid is added to the conical flask.
Step 2: Iron (II) sulfate solution formed is titrated with potassium manganate(VII)($KMnO_4$) solution.

a. Calculate the number of moles of potassium manganate (VII) used in the titration.

b. Calculate the number of moles of iron (II) ions present in the conical flask.

c. Calculate the mass of iron present in the sample and hence determine the percentage purity of the iron wire.

ANSWERS:

13.)(a) No. of mol of $KMnO_4$ used = $\frac{24.50}{1000}$ x 0.0200
= 0.000 49
= 4.9 x 10^{-4}

(b)No. of mol of MnO_4^- = No. of mol of $KMnO_4$ = 4.9 x 10^{-4}

5 mol of Fe^{2+} reacts with 1 mol of MnO_4^-

No. of mol of Fe^{2+} present in conical flask = 5 x 4.9 x 10^{-4}
= 2.45 x 10^{-3}

(c) No. of mol of Fe present in wire = No. of mol of Fe^{2+} = 2.45 x 10^{-3}

Mass of Fe = $2.45 \times 10^{-3} \times 56 = 0.1472$ g

Percentage purity = $\frac{0.1372}{0.145} \times 100\% = 94.6\%$

SUMMARY AND KEY POINTS

1.) Spectator ions are ions that are not involved in a chemical reaction.

2.) Stoichiometry of the reaction is the relationship between the amounts (measured in moles) of reactants and products involved in a chemical reaction.

3.) Limiting reactant is the reactant that is completely used up in a reaction and determined (or limits) the amount of products formed.

4.) The concentration of a solution is given by the amount of a solute dissolved in a unit volume of the solution.

5.) Molar concentration is the concentrate on of a solution expressed in mol/dm^3.

6.) The theoretical yield is the calculated amount of products that would be obtained if the reaction is completed.

7.) The actual yield is the amount of pure products that is actually produced in the experiment.

8.) A chemical equation is a shorthand way of representing what occurs in a chemical reaction. It can be written in words or by using chemical formulae.

9.) The substances which react together are called reactants. The new substances formed are called the products.

10.) A chemical equation is balanced if the number of atoms of each element of atom is the same on both sides of the equation.

KEY POINTS:

a.) In writing chemical equation involving state symbols, the state symbol (aq) refers to a substance dissolved in water while the state symbol (l) refers to a pure liquid, e.g., liquid bromine, $Br_2(l)$, water H_2O (l) or molten sodium chloride , NaCl(l).

b.) When doing chemical calculations, be careful when writing the formulae of compounds. Thus, sodium sulfate is Na_2SO_4 (not $NaSO_4$), copper(II) nitrate is $Cu(NO_3)_2$ (not $CuNO_3$) and calcium ethanoate is $(CH_3COO)_2Ca$ and not CH_3COOCa.

11.) In a chemical calculation that involves gases, we can change "mole" of gas to "volume" of gas because the volume of gas is proportional to the number of moles of the gas, and vice versa.

12.) Limiting reactants:

The reactant that is completely used up in a reaction is known as limiting reactant. It determines or limits the amount of products formed.

13.) Importance of identifying the limiting reactant:

e) In the chemical industry, large amounts of chemical are required to manufacture a particular product.
f) To get the maximum yield of a producer at the minimum cost, we ned to know the limiting reactant.
g) We will generally choose the most expensive reactant to be the limiting reactant, and use excess amounts of the other reactant in a reaction. This ensures all the expensive reactant is used up.
h) In many industrial reactions, excess reactants are recycled as far as possible in order to reduce production costs.

14.) Concentration of Solutions:

A one molar solution means the concentration of the solution is 1 mol/dm^3. A two molar solution is a solution of concentration 2 mol/dm^3.

$$\text{Concentration of solution} = \frac{number\ of\ moles}{volume\ of\ solution\ (dm3)}\ \text{mol/dm}^3.$$

16.) The concentration of an acid or alkali can be determined by titration.

a.) Titration is an experimental method in which the volume of a chemical solution (reagent) that is required to completely reaction is determined with a known volume of another solution.

b.) An indicator is used in acid-base titration.

c.) The end-point is obtained when the indicator changes colour. At the end-point, reaction is completed between the acid and alkali.

17.) The theoretical yield is the calculated amount of products that would be obtained if the reaction is completed.

a.) The actual yield is the amount of pure products that is actually produced in the experiment.

b.) In many reactions, the actual yield obtained in the experiment is less than the theoretical yield.

c.) There are many reasons for such differences, such as

a. The reactant may be impure,
b. The reaction is incomplete
c. The reactant is volatile and is lost due to evaporation.

d.) The percentage yield shows the relationship between yield and theoretical yield.

$$\text{Percentage yield} = \frac{actual\ yield}{theoretical\ yield} \text{ x } 100\%$$

18.) If the reactant used in the reaction is impure, we can calculate the percentage purity of the reactant by using the formula.

$$\text{Percentage purity} = \frac{mass\ of\ pure\ substance}{mass\ of\ substance\ used\ in\ teh\ reaction} \text{ x } 100\ \%$$