MOLES & STOICHIOMETRY WS 2

1 When an excess of chlorine was reacted with 0.72 g of titanium, 2.85 g of a chloride **A** was formed.

(i) Calculate the amount, in moles, of titanium used.

$$n\,Ti = \frac{0.72}{47.9} = 15.03 \times 10^{-3}\,mol.$$

(ii) Calculate the amount, in moles, of chlorine atoms that reacted.

Mass of Cl_2 = 2.85 − 0.72 = 2.13g n Cl Atoms = $n\,Cl_2 \times 2$

$$n \text{ of } Cl_2 = \frac{2.13}{71} = 30 \times 10^{-3}\,mol. \qquad = 60 \times 10^{-3}\,mol.$$

iii) Hence, determine the empirical formula of **A**.

Ti	:	Cl
15	:	60
1	:	4

$TiCl_4$

iv) Construct a balanced equation for the reaction between titanium and chlorine.

$$Ti + 2Cl_2 \longrightarrow TiCl_4.$$

[4]

Question say Ti reacts with Cl_2 to form $TiCl_x$. We dont know the equation for this but, its pretty simple to know 1 Ti is going to give you 1 $TiCl_x$, just by looking at Ti.

$$Ti + \frac{x}{2}Cl_2 \longrightarrow TiCl_x$$

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2 Ammonium sulfate, $(NH_4)_2SO_4$, is widely used as a fertiliser.

In order to determine its percentage purity, a sample of ammonium sulfate fertiliser was analysed by reacting a known amount with an excess of NaOH(aq) and then titrating the unreacted NaOH with dilute HC*l*.

(a) Ammonium sulfate reacts with NaOH in a 1 : 2 ratio.
Complete and balance the equation for this reaction.

$(NH_4)_2SO_4$ + 2NaOH → ..2...NH_3 + ..Na_2SO_4... + 2H_2O.........

[2]

(b) A 5.00 g sample of a fertiliser containing $(NH_4)_2SO_4$ was warmed with 50.0 cm^3 (an excess) of 2.00 mol dm^{-3} NaOH.

132.1

40.1

When all of the ammonia had been driven off, the solution was cooled.

The remaining NaOH was then titrated with 1.00 mol dm^{-3} HC*l* and 31.2 cm^3 were required for neutralisation.

(i) Write a balanced equation for the reaction between NaOH and HC*l*.

$NaOH + HCl \longrightarrow NaCl + H_2O.$

(ii) Calculate the amount, in moles, of HC*l* in 31.2 cm^3 of 1.00 mol dm^{-3} HC*l*.

$$\eta\, HCl = \frac{31.2 \times 1}{1000} = 0.0312 \text{ mol}$$

(iii) Calculate the amount, in moles, of NaOH in 50.0 cm^3 of 2.00 mol dm^{-3} NaOH.

$$\eta\, NaOH = 50 \times 2 / 1000 = 0.05 \times 2 = 0.1 \text{ mol.}$$

(iv) Use your answers to **(i)**, **(ii)** and **(iii)** to calculate the amount, in moles, of NaOH used up in the reaction with $(NH_4)_2SO_4$.

$$\eta \text{ excess } NaOH = \eta\, HCl$$
$$= 0.0312 \text{ mol.}$$

$$\eta\, NaOH \text{ used} = \eta\, NaOH_{(initially)} + \eta\, NaOH_{(excess)}$$

$$= 0.1 - 0.0312$$

$$= 0.0688 \text{ mol.}$$

(v) Use your answer to **(iv)** and the equation in **(a)** to calculate the amount, in moles, of $(NH_4)_2SO_4$ that reacted with NaOH.

$$n(NH_4)_2SO_4 = \frac{n\ NaOH}{2} = 0.0344\ mol.$$

(vi) Use your answer to **(v)** to calculate the mass of $(NH_4)_2SO_4$ that reacted with NaOH.

$$Mass = 132.1 \times 0.0344 = 4.54\ g$$

(vii) Hence, calculate the percentage purity of the ammonium sulfate fertiliser.

$$Purity = \frac{4.54}{5} \times 100 = 90.9\ \%$$

[7]

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3 Washing soda is hydrated sodium carbonate, $Na_2CO_3.xH_2O$.

A student wished to determine the value of x by carrying out a titration, with the following results.

> 5.13 g of washing soda crystals were dissolved in water and the solution was made up to 250 cm^3 in a standard volumetric flask.
>
> 25.0 cm^3 of this solution reacted exactly with 35.8 cm^3 of 0.100 mol dm^{-3} hydrochloric acid and carbon dioxide was produced.

(a) **(i)** Write a balanced equation for the reaction between Na_2CO_3 and HCl.

$Na_2CO_3 + 2HCl \longrightarrow 2NaCl + H_2O + CO_2$

(ii) Calculate the amount, in moles, of HCl in the 35.8 cm^3 of solution used in the titration.

$35.8 \times \frac{0.1}{1000} = 3.58 \times 10^{-3}$ mol.

(iii) Use your answers to **(i)** and **(ii)** to calculate the amount, in moles, of Na_2CO_3 in the 25.0 cm^3 of solution used in the titration.

$n\ Na_2CO_3$ in $25cm^3 = \frac{n\ HCl}{2} = \frac{3.58 \times 10^{-3}}{2} = 1.79 \times 10^{-3}$ mol.

(iv) Use your answer to **(iii)** to calculate the amount, in moles, of Na_2CO_3 in the 250 cm^3 of solution in the standard volumetric flask.

$n\ Na_2CO_3 = 1.79 \times 10^{-3} \times \frac{250}{25} = 1.79 \times 10^{-2}$ mol.

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(v) Hence calculate the mass of Na_2CO_3 present in 5.13 g of washing soda crystals.

$Mass = 1.79 \times 10^{-2} \times (46 + 12 + 48)$

$= 1.897 g$

$= 1.90 g$

[6]

(b) Use your calculations in **(a)** to determine the value of x in $Na_2CO_3.xH_2O$.

Mass of $xH_2O = 5 - 1.897$

$= 3.102 g$

η of $H_2O = \frac{3.102}{18} = 0.172$ mol

$x = \frac{\eta\ H_2O}{\eta\ NaOH} = \frac{0.172}{0.0179} \cong 10.$

[2]

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4 Compounds of phosphorus have many uses in everyday life, e.g. fertilisers, matches and in water softeners.

Don't worry about this

(a) State the full electronic configuration of phosphorus.

$1s^2, 2s^2, 2p^6, 3s^2, 3p^3$ [1]

(b) Phosphoric acid, H_3PO_4, is used in the manufacture of phosphate fertilisers.

Deduce the oxidation number of phosphorus in H_3PO_4.

Cause of $3H^+ = -3$, Cause of $4O^{-2} = +8$, $8-3 = 5$. [1]

(c) The salt sodium phosphate, Na_3PO_4, is a water-softening agent.

(i) Write the equation for the complete neutralisation of phosphoric acid with aqueous sodium hydroxide.

$3NaOH + H_3PO_4 \longrightarrow Na_3PO_4 + 3H_2O.$

Sodium phosphate was prepared from 50.0 cm^3 of 0.500 mol dm^{-3} H_3PO_4 and an excess of aqueous sodium hydroxide.

(ii) How many moles of H_3PO_4 were used?

$n\ Na_3PO_4 = \frac{50 \times 0.5}{1000} = 0.025$ mol.

(iii) Use your equation in **(c)(i)** to calculate how many moles of sodium hydroxide are required.

$n\ NaOH = 3(n \text{ of } Na_3PO_4) = 0.075$ mol.

[3]

(d) Phosphorus sulphide, P_4S_3, is used in small amounts in the tip of a match. On striking a match, this compound burns.

(i) Construct an equation for this reaction.

$P_4S_3 + 8O_2 \longrightarrow 2P_2O_5 + 3SO_2$

(ii) Both oxides formed in **(i)** dissolve in water to give acidic solutions. Construct an equation for the reaction of each oxide with water.

$P_4O_{10} + 6H_2O \longrightarrow 4H_3PO_4$

$SO_2 + H_2O \longrightarrow H_2SO_3$ [4]

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5 Methanoic acid, HCO_2H, was formerly known as formic acid because it is present in the sting of ants and the Latin name for ant is *formica*. It was first isolated in 1671 by John Ray who collected a large number of dead ants and extracted the acid from them by distillation.

In this question, you should give all numerical answers to two significant figures.

At room temperature, pure methanoic acid is a liquid which is completely soluble in water.

When we are stung by a 'typical' ant a solution of methanoic acid, **A**, is injected into our skin.

Solution **A** contains 50% by volume of pure methanoic acid.

A 'typical' ant contains 7.5×10^{-6} dm^3 of solution **A**.

(a) **(i)** Calculate the volume, in cm^3, of solution **A** in one ant.

$7.5 \times 10^{-6} \times 1000 = 7.5 \times 10^{-3}$ cm^3

volume = 7.5×10^{-3} cm^3

(ii) Use your answer to **(i)** to calculate the volume, in cm^3, of pure methanoic acid in one ant.

$V, HCO_2H = 7.5 \times 10^{-3} \times \frac{50}{100} = 3.75 \times 10^{-3}$

volume = 3.8×10^{-3} cm^3

(iii) Use your answer to **(ii)** to calculate how many ants would have to be distilled to produce 1 dm^3 of pure methanoic acid.

$\# \text{ of Ants} = \frac{1000}{3.8 \times 10^{-3}} = 263157.8$

number = 2.6×10^5 Ants

[3]

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When we are stung by an ant, the amount of solution **A** injected is 80% of the total amount of solution **A** present in one ant.

The density of pure methanoic acid is $1.2\,g\,cm^{-3}$.

(b) **(i)** Calculate the volume, in cm^3, of **pure** methanoic acid injected in one ant sting.

$3.8 \times 10^{-3} \times \frac{80}{100} = 3.04 \times 10^{-3}$

volume = 3×10^{-3} cm^3

(ii) Use your answer to **(i)** to calculate the mass of methanoic acid present in one ant sting.

$Mass = (1.2\,g\,cm^{-3})(3 \times 10^{-3}\,cm^3)$

$= 3.6 \times 10^{-3}\,g$

mass = 3.6×10^{-3} g

[3]

Bees also sting us by using methanoic acid. One simple treatment for ant or bee stings is to use sodium hydrogencarbonate, $NaHCO_3$.

(c) **(i)** Construct a balanced equation for the reaction between methanoic acid and sodium hydrogencarbonate.

$HCO_2H + NaHCO_3 \longrightarrow HCO_2Na + H_2O + CO_2$

(ii) In a typical bee sting, the mass of methanoic acid injected is $5.4 \times 10^{-3}\,g$. Calculate the mass of $NaHCO_3$ needed to neutralise one bee sting.

$HCO_2H : NaHCO_3$

$46 : 84$

$5.4\,mg : x$

$x = \frac{84}{46} \times 5.4\,m$

mass = 9.9×10^{-3} g

[3]

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6 Zinc is an essential trace element which is necessary for the healthy growth of animals and plants. Zinc deficiency in humans can be easily treated by using zinc salts as dietary supplements.

(a) One salt which is used as a dietary supplement is a hydrated zinc sulfate, $ZnSO_4.xH_2O$, which is a colourless crystalline solid.

Crystals of zinc sulfate may be prepared in a school or college laboratory by reacting dilute sulfuric acid with a suitable compound of zinc.

Give the formulae of **two** simple compounds of zinc that could **each** react with dilute sulfuric acid to produce zinc sulfate.

$ZnCO_3$ and ZnO [2]

(b) A simple experiment to determine the value of *x* in the formula $ZnSO_4.xH_2O$ is to heat it carefully to drive off the water.

$$ZnSO_4.xH_2O(s) \rightarrow ZnSO_4(s) + xH_2O(g)$$

A student placed a sample of the hydrated zinc sulfate in a weighed boiling tube and reweighed it. He then heated the tube for a short time, cooled it and reweighed it when cool. This process was repeated four times. The final results are shown below.

mass of empty tube / g	mass of tube + hydrated salt / g	mass of tube + salt after fourth heating / g
74.25	77.97	76.34

(i) Why was the boiling tube heated, cooled and reweighed four times?

To ensure the water has evaporated and the sample is at a constant mass.

(ii) Calculate the amount, **in moles**, of the anhydrous salt produced.

$ZnSO_4 = 76.34 - 74.25 = 2.09g$

$Mr\ ZnSO_4 = 161.5$

$n = \frac{2.09}{161.5} = 1.29 \times 10^{-2}$

(iii) Calculate the amount, **in moles**, of water driven off by heating.

Mass of $H_2O = 77.97 - 76.34 = 1.63g$

$n\ H_2O = \frac{1.63}{18} = 0.0905 = 9.1 \times 10^{-2}$ mol.

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(iv) Use your results to **(ii)** and **(iii)** to calculate the value of x in $ZnSO_4.xH_2O$.

$x = \frac{9.1 \times 10^{-2}}{1.29 \times 10^{-2}} = 7.054$

$x = 7.$

[7]

(c) For many people, an intake of approximately 15 mg per day of zinc will be sufficient to prevent deficiencies.

Zinc ethanoate crystals, $(CH_3CO_2)_2Zn.2H_2O$, may be used in this way.

(i) What mass of pure crystalline zinc ethanoate (M_r = 219.4) will need to be taken to obtain a dose of 15 mg of zinc?

$n\,Zn = \frac{0.015}{65.4}$

$= 2.29 \times 10^{-4}$ mol

Mass $= 2.29 \times 10^{-4} \times 219.4$

$= 0.05026g = 50mg$

(ii) If this dose is taken in solution as 5 cm^3 of aqueous zinc ethanoate, what would be the concentration of the solution used?
Give your answer in mol dm^{-3}.

C of crystal $= \frac{2.29 \times 10^{-4}}{0.05} = 0.0458$

$= 4.58 \times 10^{-2}$ moldm^{-3}

[4]

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7 A sample of a fertiliser was known to contain ammonium sulfate, $(NH_4)_2SO_4$, and sand only.

A 2.96 g sample of the solid fertiliser was heated with 40.0 cm^3 of NaOH(aq), an excess, and all of the ammonia produced was boiled away.

After cooling, the remaining NaOH(aq) was exactly neutralised by 29.5 cm^3 of 2.00 mol dm^{-3} HC*l*.

In a separate experiment, 40.0 cm^3 of the original NaOH(aq) was exactly neutralised by 39.2 cm^3 of the 2.00 mol dm^{-3} HC*l*.

(a) **(i)** Write balanced equations for the following reactions.

NaOH with HC*l*

$NaOH + HCl \longrightarrow NaCl + H_2O$

$(NH_4)_2SO_4$ with NaOH

$(NH_4)_2SO_4 + 2NaOH \longrightarrow 2NH_3 + Na_2SO_4 + 2H_2O$

(ii) Calculate the amount, in moles, of NaOH present in the 40.0 cm^3 of the original NaOH(aq) that was neutralised by 39.2 cm^3 of 2.00 mol dm^{-3} HC*l*.

$n\,NaOH = n\,HCl = \frac{39.2 \times 2}{1000} = 0.0784$ mol.

(iii) Calculate the amount, in moles, of NaOH present in the 40.0 cm^3 of NaOH(aq) that remained after boiling the $(NH_4)_2SO_4$.

$n\,NaOH = 29.5 \times 2 \times \frac{1}{1000} = 0.059$ mol.

(iv) Use your answers to **(ii)** and **(iii)** to calculate the amount, in moles, of NaOH that reacted with the $(NH_4)_2SO_4$.

$0.0784 - 0.059 = 0.0194$ mol

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(v) Use your answers to **(i)** and **(iv)** to calculate the amount, in moles, of $(NH_4)_2SO_4$ that reacted with the NaOH.

$\frac{0.0194}{2} = 9.7\times10^{-3}$ mol.

(vi) Hence calculate the mass of $(NH_4)_2SO_4$ that reacted.

Mass $= 9.7\times10^{-3} \times [(14+4)2 + 32.1 + 64] = 1.2814$

1.28g.

(vii) Use your answer to **(vi)** to calculate the percentage, by mass, of $(NH_4)_2SO_4$ present in the fertiliser.
Write your answer to a suitable number of significant figures.

$\% = \frac{1.2814}{2.96} \times 100 = 43.3\%$

[9]

(b) The uncontrolled use of nitrogenous fertilisers can cause environmental damage to lakes and streams. This is known as *eutrophication*.

What are the processes that occur when excessive amounts of nitrogenous fertilisers get into lakes and streams?

Fertilizers in the river causes excessive growth of aquatic plants and algae, when they die O_2 is used up leading the aquatic life to die. [2]

(c) Large quantities of ammonia are manufactured by the Haber process.
Not all of this ammonia is used to make fertilisers.
State **one** large-scale use for ammonia, **other than** in the production of nitrogenous fertilisers.

Explosives, Nylon, cleaning agent [1]

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8 Chile saltpetre is a mineral found in Chile and Peru, and which mainly consists of sodium nitrate, $NaNO_3$. The mineral is purified to concentrate the $NaNO_3$ which is used as a fertiliser and in some fireworks.

In order to find the purity of a sample of sodium nitrate, the compound is heated in NaOH(aq) with Devarda's alloy which contains aluminium. This reduces the sodium nitrate to ammonia which is boiled off and then dissolved in acid.

$$3NaNO_3(aq) + 8Al(s) + 5NaOH(aq) + 18H_2O(l) \rightarrow 3NH_3(g) + 8NaAl(OH)_4(aq)$$

The ammonia gas produced is dissolved in an excess of H_2SO_4 of known concentration.

$$2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$$

The amount of unreacted H_2SO_4 is then determined by back-titration with NaOH of known concentration.

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

(a) A 1.64 g sample of impure $NaNO_3$ was reacted with an excess of Devarda's alloy.
The NH_3 produced was dissolved in 25.0 cm^3 of 1.00 mol dm^{-3} H_2SO_4.
When all of the NH_3 had dissolved, the resulting solution was titrated with NaOH(aq).
For neutralisation, 16.2 cm^3 of 2.00 mol dm^{-3} NaOH were required.

(i) Calculate the amount, in moles, of H_2SO_4 present in the 25.0 cm^3 of 1.00 mol dm^{-3} H_2SO_4.

$n\ H_2SO_4 = \frac{25 \times 1}{1000} = 0.025$ mol.

(ii) Calculate the amount, in moles, of NaOH present in 16.2 cm^3 of 2.00 mol dm^{-3} NaOH.

$n\ NaOH = \frac{16.2 \times 2}{1000} = 0.0324$ mol.

(iii) Use your answer to **(ii)** to calculate the amount, in moles, of H_2SO_4 that reacted with 16.2 cm^3 of 2.00 mol dm^{-3} NaOH.

$\frac{0.0324}{2} = 0.0162$ mol.

(iv) Use your answers to **(i)** and **(iii)** to calculate the amount, in moles, of H_2SO_4 that reacted with the NH_3.

$0.025 - 0.0162 = 0.0088$ mol

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(v) Use your answer to **(iv)** to calculate the amount, in moles, of NH_3 that reacted with the H_2SO_4.

$2 \times 0.0088 = 0.0176$ mol.

(vi) Use your answer to **(v)** to calculate the amount, in moles, of $NaNO_3$ that reacted with the Devarda's alloy.

n NaOH that reacted = n NH_3 produced = 0.0176 mol.

(vii) Hence calculate the mass of $NaNO_3$ that reacted.

Mass = $0.0176 \times 85 = 1.496$g

(viii) Use your answer to **(vii)** to calculate the percentage by mass of $NaNO_3$ present in the impure sample.
Write your answer to a suitable number of significant figures.

$\% \ NaNO_3 = \dfrac{1.496 \times 100}{1.64} = 91.219 = 91.2$

[9]

(b) The above reaction is an example of a redox reaction.
What are the oxidation numbers of nitrogen in $NaNO_3$ and in NH_3?

$NaNO_3$ ....5.......... NH_3 ....-3.......... [1]

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9 A sample of a hydrated double salt, $Cu(NH_4)_x(SO_4)_2.6H_2O$, was boiled with an excess of sodium hydroxide. Ammonia was given off.

The ammonia produced was absorbed in 40.0 cm³ of 0.400 mol dm⁻³ hydrochloric acid. The resulting solution required 25 cm³ of 0.12 mol dm⁻³ sodium hydroxide to neutralise the excess acid.

(a) Write the ionic equation for the reaction between ammonium ions and hydroxide ions.

$NH_4^+ + OH^- \longrightarrow NH_3 + H_2O.$ [1]

(b) **(i)** Calculate the amount, in moles, of hydrochloric acid in 40.0 cm³ of 0.400 mol dm⁻³ solution.

$40 \times 0.4 / 1000 = 0.016$ mol

[1]

(ii) Calculate the amount, in moles, of sodium hydroxide needed to neutralise the excess acid. This will be equal to the amount of hydrochloric acid left in excess.

$\frac{25 \times 0.12}{1000} = 3 \times 10^{-3}$ mol

[1]

(iii) Calculate the amount, in moles, of hydrochloric acid that reacted with ammonia.

excess = 0.003 mol

reacted = 0.016 − 0.003 = 0.013 mol

[1]

(iv) Calculate the amount, in moles, of ammonium ions in the sample of the double salt.

$NH_4^+ : H^+$

1 : 1 so $n(NH_4^+) = 0.013$ mol

[1]

(v) The sample contained 0.413 g of copper. Use this information and your answer to **(iv)** to calculate the value of *x* in $Cu(NH_4)_x(SO_4)_2.6H_2O$.

$n\,Cu = \frac{0.413}{63.5} = 6.5 \times 10^{-3}$ mol

Cu : NH₄

0.0065 : 0.013

1 : 2

$x = 2.$

[2]

(vi) Calculate the M_r of $Cu(NH_4)_x(SO_4)_2.6H_2O$.

$M_r = 399.7$

[1]

[Total: 8]

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115.8

10 Spathose is an iron ore that contains iron(II) carbonate, $FeCO_3$. The percentage of iron(II) carbonate in spathose can be determined by titration with acidified potassium dichromate(VI) solution using a suitable indicator.

The ionic equation is shown below.

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6Fe^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(l)$$

(a) A 5.00 g sample of spathose was reacted with excess concentrated hydrochloric acid and then filtered.
The filtrate was made up to 250 cm³ in a volumetric flask with distilled water.

A 25.0 cm³ sample of the standard solution required 27.30 cm³ of 0.0200 mol dm⁻³ dichromate(VI) solution for complete reaction.

(i) Calculate the amount, in moles, of dichromate(VI) ions used in the titration.

η Chromate $= \frac{27.3 \times 0.02}{1000} = 5.46 \times 10^{-4}$ mol.

amount = 5.46×10^{-4} mol [1]

(ii) Use your answer to **(i)** to calculate the amount, in moles, of Fe^{2+} present in the 25.0 cm³ sample.

$Cr_2O_7^{2-} : Fe^{2+}$

$1 : 6$ $\quad \eta Fe^{2+} = 3.28 \times 10^{-3}$ mol

in 25 cm³.

amount = 3.28×10^{-3} mol [1]

(iii) Use your answer to **(ii)** to calculate the amount, in moles, of Fe^{2+} present in the 250 cm³ volumetric flask.

$\eta(Fe^{2+}) = 3.28 \times 10^{-3} \times \frac{250}{25}$

in 250 cm³

amount = 3.28×10^{-2} mol [1]

(iv) Use your answer to **(iii)** to calculate the mass of iron(II) carbonate present in the sample of spathose.

$\eta(Fe^{2+}) = \eta(FeCO_3) = 3.28 \times 10^{-2}$ mol

Mass $= \frac{3.28}{100} \times 115.8 = 3.798g \approx 3.80g$

mass = 3.80 g [2]

(v) Calculate the percentage of iron(II) carbonate in the sample of spathose.

$\frac{3.80}{5.00} \times 100$

percentage of iron(II) carbonate = 76 % [1]

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