

YEAR 12 CHEMISTRY

3+4 CALCULATIONS WORKBOOK 2016



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EXPLANATORY NOTES

In completing these problems you will need to use some different values other than those on your WACE Data Sheet

Volume of 1 mole of an ideal gas at S.T.P. = 22.41 L
where S.T.P. = 0.0 °C and 101.3 kPa

Universal Gas constant (R) = 8.315 J K⁻¹ mol⁻¹

Density = mass of sample ÷ volume of sample

101.3 kPa = 1 atm (atmosphere) = 760 mmHg

STRATEGY FOR WORKING STOICHIOMETRIC PROBLEMS

Quantities in Chemistry: grams (g)
litres of gas (L)
the concentration (molL⁻¹) of a known volume (L) of solution

Strategy:

Step 1: Convert quantity of known to moles of known

Step 2: Convert moles of known to moles of unknown (using the stoichiometric ratio)

$n(\text{unknown}) = \text{unknown coefficient} \div \text{known coefficient}$

or

through the equation from moles of known to moles of unknown

Note: This step requires a balanced chemical equation

Step 3: Convert the moles of unknown to quantity of unknown

Step 1: from the WACE Data Sheet

Grams → moles

$$n = m/M$$

where n = number of moles

m = mass in grams

M = Molar mass (formula mass in grams)

Litres of gas → moles

$$n = V/22.41 \quad \text{at STP where } n = \text{number of moles}$$

V = volume of gas in litres

22.41 L = Molar volume of gas @ STP

$$PV = nRT \quad \text{for gases @ non-standard conditions}$$

$$n = PV/RT \quad \text{where } n = \text{number of moles}$$

P = Pressure in kPa (1 atm = 101.3 kPa)

V = Volume of gas in litres ($L = \text{mL} \div 1000$)

R = 8.315 (Universal Gas Constant)

T = Temperature in Kelvin (Kelvin = $^{\circ}\text{C} + 273.1$)

Concentration (molL^{-1}) of a known volume (L) of solution → moles

$$n = cV$$

where n = number of moles

c = concentration in molL^{-1}

V = volume of solution in litres ($L = \text{mL} \div 1000$)

Whenever quantities of two or more reactants are given in a stoichiometric problem, you must identify the limiting reagent.

LIMITING REAGENTS

There is a systematic method for finding the limiting reagent (reactant) and determining the maximum amounts of products formed.

- Take each reactant in turn, assume that it is used up completely in the reaction, and calculate the number of moles (or the mass) of *one* of the products that will be formed.
- Whichever reactant gives the *smallest* number of moles (or mass) of this product is the limiting reagent. Once it has reacted fully, no further product can be formed.
- The limiting reagent determines (limits) the amount of products formed and the amount of reagent in excess which reacts with it.
- The limiting reagent determines the amount of the reagent in excess.

$$(n_{\text{inxs}} = n_{\text{initially}} - n_{\text{reacting}})$$

CALCULATIONS

Set 1

1. A chloride of titanium dissolves in water to give titanium ions and chloride ions. 3.40 g of the titanium chloride was dissolved in water and all the chloride ions were precipitated as silver chloride by addition of excess silver nitrate solution.

After filtration and drying, 9.47 g of silver chloride was obtained.
What is the empirical formula of this chloride of titanium?

2. To reduce water pollution, you are told to neutralize 200 litres of waste 2.0 mol L^{-1} sulfuric acid solution prior to filtration and subsequent disposal in a river.
Calculate the mass of $\text{CaCO}_{3(s)}$ required to neutralise this sample of waste acid.

4. Oxygen was prepared by the decomposition of hydrogen peroxide using manganese(IV) oxide as a catalyst and the gas was collected over water.

The gas collected occupied 5.00 L at 24°C when the barometric pressure was 99.7 kPa. What would be the volume of dry oxygen under standard conditions of temperature and pressure?

(Vapour pressure of water at 24°C is 2.98 kPa)

5. The percentage by mass of chromium in a compound is determined by converting a sample of known mass into sodium dichromate, and titrating (reacting) an acidified solution of the sodium dichromate with a standard solution of iron(II) sulfate.

Using this method, 1.70 g of a chromium containing mineral is converted into an acidified solution of sodium dichromate, which required 50.0 mL of 0.400 mol L^{-1} iron(II) sulfate to each an end point.

Calculate the percentage by mass of chromium in the mineral.

ANSWERS Calculations Set 1

- | | | | |
|----|-----------------|----|---|
| 1. | TiCl_3 | 2. | $4.00 \times 10^4 \text{ g}$ or 40.0 kg |
| 4. | 4.37 L | 5. | 20.4 % |

CALCULATIONS

Set 2

1. A white crystalline compound containing only carbon, hydrogen and oxygen was analysed as follows. A 2.76 g sample of the compound when burned in excess oxygen produced 1.93 g carbon dioxide and 1.18 g of water.
 - A. Calculate the empirical formula of the compound.
 - B. From further analysis it was established that the relative molecular mass was 126.0. What is the compound's molecular formula?
 - C. Given that the compound was a dihydrate, rewrite the formula to indicate this.

2. A 2.05 g sample of haematite (mainly Fe_2O_3) was dissolved in sulfuric acid and diluted to 250 mLs. A 25.0 mL sample (aliquot) of the diluted solution was treated with just enough zinc to reduce the iron(III) ions to iron(II) ions. The 25.0 mL sample was then titrated with $0.0212 \text{ mol L}^{-1}$ potassium permanganate solution. A volume of 23.4 mL KMnO_4 solution was needed to reach the end point.
 - A. Using the table of standard reduction potentials on the data sheet, write the half equation for the reduction of MnO_4^- to Mn^{2+} .
 - B. Using the table of standard reduction potentials on the data sheet, write the half equation for the oxidation of Fe^{2+} back to Fe^{3+} .
 - C. Write the balanced chemical equation for the reaction.
 - D. Calculate the number of moles of permanganate.
 - E. Calculate the number of moles of Fe^{2+} in the 25.0 mL sample.
 - F. Calculate the number of moles of iron in 250 mL of solution.
 - G. Using your answer in E., calculate the mass of Fe_2O_3 which was dissolved.
 - H. Calculate the % purity of the haematite sample.

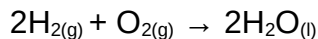
3. A pure sample of magnesium was treated with excess hydrochloric acid to produce hydrogen gas which was collected by the displacement of water. After adjusting the pressure to 101.3 kPa, the volume of gas collected over water at 20.0°C was 1.16 litres.
 - A. Write a balanced net ionic equation for the reaction of magnesium with hydrochloric acid.
 - B. Given that the water has a partial pressure of 2.33 kPa at 20.0°C , calculate the partial pressure of the hydrogen gas collected.
 - C. Calculate the volume this hydrogen gas would occupy if it was dry and under STP.
 - D. Calculate the number of moles of hydrogen gas produced.
 - E. Calculate the number of moles of magnesium consumed in the reaction.
 - F. Calculate the mass of magnesium consumed in the reaction.

5. A 6 L flask containing hydrogen gas at 200.0 kPa is connected via a stopcock to a 2 L flask containing oxygen gas at 400.0 kPa. The temperature of the system being 0.0 °C.

The stopcock was opened and the gases allowed to mix. Assuming the temperature of the system remained constant, calculate

- A. the part(ial) pressures of hydrogen and oxygen in the system and the total pressure.
B. The number of moles of hydrogen and oxygen.

The mixture was sparked and the following and the following reaction occurred:



- C. Calculate the number of moles of water produced.
D. Calculate the mass of water produced.

ANSWERS Calculations Set 2

- | | | | | | | | |
|----|----|---|--------------------------------------|---|------------|--|--------|
| 1. | A. | CH_3O_3 | B. | $(\text{CH}_3\text{O}_3)_2$ | C. | $\text{C}_2\text{H}_6\text{O}_6 \cdot 2\text{H}_2\text{O}$ | |
| 2. | C. | $\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ = \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O}$ | D. | $n(\text{MnO}_4^-) = 4.961 \times 10^{-4}$ | | | |
| | E. | 2.480×10^{-3} | F. | 2.480×10^{-2} | G. | 1.240×10^{-2} | |
| 3. | A. | $\text{Mg} + 2\text{H}^+ = \text{Mg}^{2+} + \text{H}_2$ | B. | 98.97 kPa | C. | 1.056 L | |
| | D. | 0.04714 mol | E. | 0.04714 mol | F. | 1.15 g | |
| 5. | A. | $P(\text{H}_2) = 150 \text{ kPa}$ | $P(\text{O}_2) = 100 \text{ kPa}$ | $P_T = P_{\text{H}_2} + P_{\text{O}_2} = 150 + 100 = 250 \text{ kPa}$ | | | |
| | B. | $n(\text{H}_2) = 0.5284 \text{ mol}$ | $n(\text{O}_2) = 0.3522 \text{ mol}$ | C. | 0.5284 mol | D. | 9.52 g |

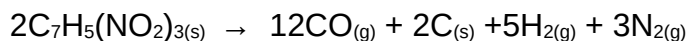
CALCULATIONS

Set 3

1. In a laboratory experiment, 2.12 g sodium carbonate, Na_2CO_3 , was dissolved in 10 mL of 1.00 mol L^{-1} copper(II) nitrate, $\text{Cu}(\text{NO}_3)_2$. The resulting precipitate of insoluble copper(II) carbonate, CuCO_3 , was washed and redissolved in a slight excess of dilute sulfuric acid. Copper(II) sulfate was recovered by evaporation as the pentahydrate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$.
 - A. Calculate the number of moles of Na_2CO_3 and $\text{Cu}(\text{NO}_3)_2$.
 - B. What mass of CuCO_3 was precipitated?
 - C. Assuming complete recovery of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$, what was the mass of this product?
3. A 0.5064 g sample of a liquid known to contain only carbon, hydrogen, sulfur and oxygen was thoroughly burned in excess oxygen. The resultant carbon dioxide and water were separated and their mass determined. The sulfur dioxide was catalytically converted into sulfur trioxide which was dissolved in excess aqueous barium chloride precipitating the sulfur as barium sulfate.
The mass of carbon dioxide was found to be 0.5702 g, that of the water was 0.3501, and that of the precipitated barium sulfate was 1.5120 g.

Calculate the empirical formula of the compound analysed.

4. The explosive TNT (trinitrotoluene), when detonated, decomposes according to the equation.



A 25.00 g charge of TNT is detonated in an evacuated 5.00 L container.

- A. Calculate the mass of carbon deposited.
- B. Calculate the final pressure of the system if the temperature reaches 230°C .
- C. Calculate the partial pressure of nitrogen under these conditions.

5. A hydrogen peroxide solution (solution A) was analysed by titration with standard permanganate solution. 25 mLs of solution A was diluted to 250.0 mL with water. 25.0 mL of this diluted peroxide solution (acidified with sulfuric acid) required 23.65 mL of $0.0220 \text{ mol L}^{-1}$ permanganate solution for reaction.
- Using the table of standard reduction potentials on the data sheet, write the half equation for the reduction of MnO_4^- to Mn^{2+} .
 - Using the table of standard reduction potentials on the data sheet, write the half equation for the oxidation of H_2O_2 to O_2 .
 - Write the balanced chemical equation for the reaction.
 - Calculate the number of moles of permanganate in 23.65 mL of $0.0220 \text{ mol L}^{-1}$ permanganate solution.
 - Calculate the number of moles of hydrogen peroxide in 25.0 mL of the diluted peroxide solution.
 - Calculate the concentration (molarity) of hydrogen peroxide in the original solution (solution A).
 - Calculate to the nearest litre the volume of oxygen, at STP, obtained if the hydrogen peroxide in 1 litre of solution A were decomposed into water and oxygen.

ANSWERS Calculations Set 3

- $0.0200 \text{ mol Na}_2\text{CO}_3$
 - 1.24 g
- $0.0100 \text{ mol Cu(NO}_3)_2$
 - 2.50 g
- Excess Na_2CO_3 ; $\text{Cu(NO}_3)_2$ limiting
- $\text{C}_2\text{H}_6\text{SO}$
- 1.32 g
 - 921 kPa
 - 138 kPa
- $2\text{MnO}_4^- + 6\text{H}^+ + 5\text{H}_2\text{O}_2 = 2\text{Mn}^{2+} + \text{O}_2 + 8\text{H}_2\text{O}$
 - 5.203×10^{-4}
 - 1.301×10^{-3}
 - $0.05204 \text{ mol L}^{-1}$ in diluted solution; $0.5204 \text{ mol L}^{-1}$ in original solution
 - 6 L

CALCULATIONS

Set 4

1. Oxygen was generated by heating 2.030 g of potassium chlorate, KClO_3 using a suitable catalyst. After allowing for the system to come to the room temperature of 22.5°C it was found that 485 mL of oxygen was collected over water at a pressure of 100.5 kPa. The vapour pressure of the water (i.e the part(ial) pressure of the water vapour) at 22.5°C is 2.7 kPa. The flask and contents were weighed before and after heating and a weight loss of 0.621 g was recorded.
 - A. Write an equation for the thermal decomposition of potassium chlorate into potassium chloride and oxygen.
 - B. Based on the results of this experiment and the molar mass of oxygen, calculate the molar volume of oxygen corrected to STP.
 - C. Not all the potassium chlorate was decomposed in the above experiment. Calculate the volume of gas which would have been obtained (after cooling and pressure adjustment as before) if **ALL** the potassium chlorate had been decomposed to give potassium chloride and oxygen.
2. The chromium content in a sample can be determined by oxidizing the chromium to sodium dichromate with an excess of persulfate solution, boiling to destroy excess oxidant and then titrating the acidified solution with standard iron(II) sulfate solution. Using this method a 8.405 g sample of a chromium-containing mineral was converted into an acidified solution of sodium dichromate and diluted to 250.0 mL. A 25.00 mL sample of this solution was then titrated against 0.415 mol L^{-1} iron(II) sulfate and required 24.8 mL to reach the end point.
 - A. Using the table of standard reduction potentials on the data sheet, write the half equation for the reduction of $\text{Cr}_2\text{O}_7^{2-}$ to Cr^{3+} .
 - B. Using the table of standard reduction potentials on the data sheet, write the half equation for the oxidation of Fe^{2+} to Fe^{3+} .
 - C. Write the balanced chemical equation for the reaction.
 - D. Calculate the number of moles of Fe^{2+} used in the titration.
 - E. Calculate the number of moles of $\text{Cr}_2\text{O}_7^{2-}$ in the 25.00 mL sample.
 - F. Calculate the mass of chromium in 250.0 mL of the dichromate solution.
 - G. Calculate the percentage by mass of chromium in the mineral.
3. A compound of carbon, hydrogen, nitrogen and oxygen was analysed as follows.

A 1.342 g sample was burned in excess oxygen to produce 0.288 g of water and 2.109 g of carbon dioxide.

A second 1.061 g sample was treated to convert all the nitrogen into ammonia which was then analysed by titration and found to be $1.263 \times 10^{-2}\text{ mol}$.

 - A. Calculate the empirical formula of the compound.
 - B. Given that the compound is aromatic, what is the simplest molecular formula possible for it?

5. A sample of commercial vinegar contains 4.13% ethanoic acid (acetic acid) by mass. A 15.121 g sample of vinegar was diluted to 100.0 mL and 25.00 mL of 0.242 mol L⁻¹ sodium hydroxide solution was then added.
- A. Calculate the number of moles of each of the two reagents before mixing.
B. Calculate the concentration of the ethanoate (acetate) ion in the final solution.

ANSWERS Calculations Set 4

- | | | | | | | |
|----|----|---|--|--|----|-----------------------------|
| 1. | A. | $2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2$ | B. | 22.3 L | C. | 0.621 L |
| 2. | C. | $\text{Cr}_2\text{O}_7^{2-} + 6\text{Fe}^{2+} + 14\text{H}^+ = 2\text{Cr}^{3+} + 6\text{Fe}^{3+} + 7\text{H}_2\text{O}$ | | | | |
| | D. | 0.01029 mol | E. | 0.001715 mol | F. | 1.784 g |
| 3. | A. | $\text{C}_3\text{H}_2\text{NO}_2$ | B. | $\text{C}_6\text{H}_4\text{N}_2\text{O}_4$ | G. | 21.2% |
| 5. | A. | $n(\text{CH}_3\text{COOH}) = 0.01041 \text{ mol};$ | $n(\text{OH}^-) = 0.00605 \text{ mol}$ | | B. | $0.0484 \text{ mol L}^{-1}$ |

CALCULATIONS

Set 5

1. A solid alkanolic (carboxylic) acid has two acidic hydrogens (i.e. it is a diprotic acid) and is hydrated in its crystalline form.
 - A. On heating a 0.808 g sample of the acid at 110 °C to constant mass, the mass of anhydrous alkanolic acid remaining is 0.576 g. What is the percentage by mass of water of crystallization in the hydrated alkanolic acid.
 - B. A 2.050 g sample of the hydrated alkanolic acid was dissolved in water and made up to 250.0 mL in a volumetric flask. Titration of both acidic hydrogen atoms in a 20.00 mL aliquot (portion) of this solution against 0.110 mol L⁻¹ sodium hydroxide solution, using phenolphthalein as indicator, required a volume of 23.7 mL.
 - (i) Calculate the number of moles of sodium hydroxide used in the titration.
 - (ii) Calculate the number of moles of the alkanolic acid in the 20.00 mL aliquot.
 - (iii) Calculate the number of moles of the alkanolic acid in the 2.050 g sample.
 - (iv) Calculate the molecular mass of the hydrated alkanolic acid.
 - C. Analysis of the anhydrous alkanolic acid gave an empirical formula of CHO₂ and a molecular mass of 90. What is the molecular formula of the anhydrous alkanolic acid?
 - D. Determine the molecular formula of the hydrated alkanolic acid.
 - E. From the molecular masses of the hydrated and anhydrous alkanolic acids, calculate the theoretical percentage by mass of water of crystallization in the hydrated alkanolic acid.
2. Petrol, although a mixture of hydrocarbons, may be considered to be octane.
 - A. Write a balanced chemical equation for the complete combustion of octane.

Petrol has a density of 702.5 g L⁻¹. For each litre of petrol consumed, calculate

 - B. the mass of oxygen required
 - C. the volume of carbon dioxide emitted, measured at 100.7 kPa and 22.0 °C.
3. An organic compound used as a herbicide contains only carbon, hydrogen, oxygen and chlorine. When a 0.9112 g sample of the compound was completely burned in excess oxygen it produced 1.451 g of carbon dioxide and 0.223 g of water. The chlorine was recovered by treatment of the combustion mixture with silver nitrate solution and 1.182 g of silver chloride was isolated.

Calculate the empirical formula of the herbicide.

5. In a laboratory experiment 14.00 g of barium chloride dihydrate, $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$ was dissolved in water and made up to 100.0 mL in a volumetric flask. A 25.00 mL sample of this solution was added to 50.00 mL of 0.245 mol L^{-1} sulfuric acid solution and the precipitate of barium sulfate, BaSO_4 was collected by filtration.
- A. Calculate the mass of BaSO_4 collected.
- B. Calculate the concentration of chloride ion in the filtrate.

ANSWERS Calculations Set 5

- | | | | | | | | |
|----|-------|--|------|------------------------------|------------------------------------|----------------------------------|------------------------------------|
| 1. | A. | 28.7% | B. | (i) | $2.607 \times 10^{-3} \text{ mol}$ | (ii) | $1.304 \times 10^{-3} \text{ mol}$ |
| | (iii) | $1.630 \times 10^{-3} \text{ mol}$ | (iv) | $M_r = 126$ | C. | $\text{C}_2\text{H}_2\text{O}_4$ | |
| | D. | $\text{C}_2\text{H}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$ | E. | 28.6% | | | |
| 2. | A. | $2\text{C}_8\text{H}_{18} + 25\text{O}_2 \rightarrow 16\text{CO}_2 + 18\text{H}_2\text{O}$ | B. | $2.46 \times 10^3 \text{ g}$ | C. | $1.20 \times 10^3 \text{ L}$ | |
| 3. | | $\text{C}_8\text{H}_6\text{Cl}_2\text{O}_3$ | | | | | |
| 5. | A. | H_2SO_4 is limiting $m(\text{BaSO}_4) = 2.86 \text{ g}$ | B. | 0.382 mol L^{-1} | | | |

CALCULATIONS

Set 6

1. Calcium hydroxide, Ca(OH)_2 , is a strong base which dissociates completely in solution. 500.0 mL of $0.00500 \text{ mol L}^{-1}$ Ca(OH)_2 solution was prepared by dissolving Ca(OH)_2 in water and making the solution up to 500.0 mL.
 - A. Calculate the mass of Ca(OH)_2 required to make up 500.0 mL of $0.00500 \text{ mol L}^{-1}$ solution.
 - B. In the $0.00500 \text{ mol L}^{-1}$ Ca(OH)_2 solution calculate
 - (i) the concentration of the OH^- . (i.e. $[\text{OH}^-]$ where $[]$ = concentration)
 - (ii) the concentration of the H^+ . (i.e. $[\text{H}^+]$)
 - (iii) the pH

Use the relationships:

$$[\text{H}^+] \times [\text{OH}^-] = 10^{-14}$$

$$\text{pH} = -\log_{10} [\text{H}^+]$$

$$\text{or } \text{pH} + \text{pOH} = 14 \text{ (where } \text{pOH} = -\log_{10} [\text{OH}^-] \text{)}$$

2. A chemist sets out to determine the identity of a food preservative. The preservative is an organic compound which contains only carbon, hydrogen and oxygen. Complete combustion of 1.161 g of the compound produced 2.932 g of carbon dioxide and 0.514 g of water.
 - A. Calculate the empirical formula of the preservative.
 - B. The relative molecular mass of the preservative is found to be 122 and further analysis indicates that the substance is an aromatic alkanoic (carboxylic) acid. Write the molecular formula and draw a structural formula for the preservative.

3. The concentration of a sample of battery acid (sulfuric acid) was determined in the following way.

Step 1: The concentration of a sodium hydroxide solution was determined by titration against 0.105 mol L^{-1} hydrochloric acid, using phenolphthalein as the indicator. 20.00 mL of hydrochloric acid required 22.35 mL of sodium hydroxide to reach the end point.

Step 2: 5.00 mL of battery acid was diluted to 500.0 mL in a volumetric flask.

Step 3: The diluted battery acid was titrated with the sodium hydroxide solution which was standardized (whose concentration was determined) in Step 1. 19.15 mL of the standard sodium hydroxide solution was required to react completely with 20.00 mL of the diluted battery acid using phenolphthalein as the indicator.

- A. Using the information in Step 1, calculate the concentration of the sodium hydroxide solution.
 - B. Using the information in Step 3, calculate the concentration of the diluted battery acid.
 - C. Using the information in Step 2, calculate the concentration of the battery acid.
4. The equations below represent the reactions that occur in the manufacture of nitric acid from ammonia



If $5.00 \times 10^3 \text{ kg}$ of ammonia (NH_3) is consumed in this process, calculate

- A. the mass of HNO_3 produced.
- B. the final volume of NO produced at STP.

ANSWERS Calculations Set 6

1. A. 0.185 g B. (i) $0.0100 \text{ mol L}^{-1}$ (ii) $1.00 \times 10^{-12} \text{ mol L}^{-1}$ (iii) 12.0
2. A. $\text{C}_7\text{H}_6\text{O}_2$ B. $\text{C}_7\text{H}_6\text{O}_2$
3. A. $9.40 \times 10^{-2} \text{ mol L}^{-1}$ B. $4.50 \times 10^{-2} \text{ mol L}^{-1}$ C. 4.50 mol L^{-1}
4. A. $1.24 \times 10^7 \text{ g}$ or $1.24 \times 10^4 \text{ kg}$ B. $2.20 \times 10^6 \text{ L}$ Note: if the $\text{NO}_{(\text{g})}$ in the third step is recycled then zero moles of NO is produced. This is an equally valid answer to part B.

CALCULATIONS

Set 7

2. A sample of steel from a cannon was analysed to determine the percentage of iron present. A 1.120 g sample was carefully removed from the underside of the cannon and dissolved in excess 6.00 mol L^{-1} sulfuric acid and the resulting solution made up to 50.00 mL. A 20.00 mL aliquot (portion) of this solution was titrated against $0.1030 \text{ mol L}^{-1}$ solution of potassium permanganate. 15.49 mL of the permanganate solution was needed to reach the end-point.
- Write a balanced chemical equation for the reaction of permanganate ion with iron(II) ion.
 - Calculate the number of moles of permanganate ion in 15.49 mL of $0.1030 \text{ mol L}^{-1}$ potassium permanganate.
 - Calculate the number of moles of $\text{Fe}^{2+}_{(\text{aq})}$ in the 20.00 mL aliquot.
 - Calculate the number of moles of Fe in the original sample.
 - Calculate the percentage of iron in the original sample.
3. Three solutions, A, B and C have the following composition:
- Solution A: 50.0 mL of 0.500 mol L^{-1} sulfuric acid was diluted with distilled water and the resulting solution made up to 250.0 mL.
- Solution B: 560.0 mL of hydrogen chloride gas at S.T.P. was dissolved in distilled water and the resulting solution made up to 500 mL.
- Solution C: 100.0 mL contains 3.20 g of sodium hydroxide.
- Calculate the concentration of each solution.
 - 20.0 mL of solution A and 50.0 mL of solution B are mixed. Calculate the concentration of hydrogen ions.
 - What volume of solution C must be added to the mixture from part B in order to adjust the pH of the solution to 7.00 (i.e to neutralise the H^+ ions in part B)?
4. 12.00 L of a gas mixture known to contain butane and nitrogen was mixed with an excess of oxygen and ignited. The products were cooled to 25.00°C and dried. The volume of the products was 77.00 L. The products were then passed through a solution of potassium hydroxide to absorb the carbon dioxide and the remaining gas mixture dried. The volume of this final mixture was found to be 37.00 L. All gas measurements were carried out at 25.00°C and 101.3 kPa.
- Write the equation for the combustion of butane.
 - Determine the volume of butane in the original sample.
 - Determine the volume of oxygen mixed with the original butane/nitrogen mixture (measured at 25.00°C and 101.3 kPa)
 - Calculate the percentage by mass of the butane in the original sample.

5. Substance X, containing the elements C, H and O only, may be converted to the monoprotic acid Y by the action of an excess of dilute aqueous acidified potassium permanganate.

18.70 mL of a solution of 1.0165 gL^{-1} of Y was required to neutralize 20.00 mL of 0.0108 molL^{-1} potassium hydroxide.

Substance Z may be prepared by heating X with an excess of ethanoic acid in the presence of a small amount of sulfuric acid, followed by purification of the product.

When 0.2870 g of Z was completely burned in an excess of oxygen the products were 0.6532 g of carbon dioxide and 0.2672 g of water.

- A. Calculate the relative molecular mass of Y.
- B. Calculate the empirical formula of Z.
- C. Given that the carbon chain of X is branched, draw structural formulae for X, Y and Z which are consistent with the data.

ANSWERS Calculations Set 7

- | | | | | | | |
|----|----|---|----|--------------------------------|----|--|
| 2. | A. | $\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ = \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O}$ | B. | 0.001595 | C. | 0.007977 |
| | D. | 0.01994 | E. | 99.4% | | |
| 3. | A. | $0.1 \text{ molL}^{-1} \text{ H}_2\text{SO}_4$; $0.050 \text{ molL}^{-1} \text{ HCl}$; $0.800 \text{ molL}^{-1} \text{ NaOH}$ | B. | 0.0929 molL^{-1} | | |
| | C. | 8.13 mL | | | | |
| 4. | A. | $\text{C}_4\text{H}_{10} + 13\text{O}_2 = 8\text{CO}_2 + 10\text{H}_2\text{O}$ | B. | 10 L | C. | 100 L |
| 5. | A. | 88.2 | B. | $\text{C}_3\text{H}_6\text{O}$ | C. | X: 2-methylpropanol
Y: 2-methylpropanoic acid
Z: 2-methylpropylethanoate |

CALCULATIONS

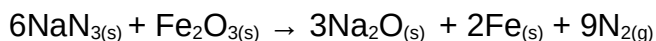
Set 8

1. A solution containing 5.00 g of aluminium chloride is mixed with a solution containing 4.00 g of sodium sulfide.
 - A. Write a net ionic equation for the reaction.
 - B. Identify the limiting reagent.
 - C. What mass of precipitate is produced?
 - D. Calculate the mass of excess reactant remaining after the reaction.

2. Although the hydrolysis of esters involves a reversible reaction, adequate precautions will result in a reaction which is sufficiently quantitative to permit the determination of relative molecular mass.

In a certain experiment, 0.306 g of an ester is reacted, under reflux, with 20.00 mL of 1.00 mol L⁻¹ sodium hydroxide solution. After cooling phenolphthalein is added and the excess sodium hydroxide is back titrated with 0.513 mol L⁻¹ hydrochloric acid, 32.2 mL being required.

- A. Assuming the compound to be saturated and to contain only one ester group, calculate the relative molecular mass of the ester.
 - B. Determine the molecular formula.
 - C. Draw one possible structure and assign it a systematic (IUPAC) name.
3. The automobile air bag is a remarkable device. It is deceptively simple in concept, but the design requirements of the system are actually very demanding. Most cars are equipped with air bags that inflate on collision to protect the occupants from injury. Many such air bags are inflated with nitrogen, N₂, using the rapid reaction of sodium azide, NaN₃, and iron(III) oxide, Fe₂O₃, which is initiated by a spark. The overall reaction is



Sodium azide is a stable solid and a small pellet can be easily stored in air bag compartments. Sodium azide, itself, is pretty nasty and so it is important that it is completely consumed when an air bag is deployed.

- A. Draw a Lewis (electron dot) diagram for the azide ion N₃⁻.
 - B. Why is a spark required to initiate the reaction?
 - C. Give one reason as to why you think nitrogen gas is used to inflate the air bag.
 - D. What type (class) of reaction is this?
 - E. How many grams of sodium azide would be required to provide 75.0 L of nitrogen gas at 25 °C and 748 mmHg.

4. The iron content of steel can be determined by dissolving the steel in acid and titrating with acidified potassium dichromate ($\text{K}_2\text{Cr}_2\text{O}_7$).

1.00 g steel was dissolved in acid to give a solution of Fe^{2+} and the solution made up to 250 mL with water. Duplicate 25.0 mL samples were titrated with standard potassium dichromate solution ($0.0100 \text{ mol L}^{-1}$). The average titration volume was 28.4 mL.

- A. Identify the oxidant in the titration reaction.
 B. Write the balanced oxidation half-equation.
 C. Write the balanced reduction half-equation.
 D. Write the overall redox equation for this reaction.
 E. Calculate the number of moles of potassium dichromate in the average titration.
 F. Calculate the number of moles of Fe^{2+} present in the 250 mL sample.
 G. Calculate the percentage (g/100 g) of iron in the steel.

5. 'Free sulfur dioxide' may be present in wines in a number of different forms, including the sulfite ion. Its concentration can be determined using the following reaction in aqueous solution:



- A. What is the oxidation number of sulfur in sulfite ion and in sulfate ion?
 B. From the equation, identify a reactant that is behaving as an oxidant.
 C. A titration was performed to determine the amount of 'free sulfur dioxide' (as sulfite ion) in a 50.0 mL sample of white wine. It was found that 12.5 mL of a $0.00206 \text{ mol L}^{-1}$ iodine solution was required to obtain the end-point. What concentration of 'free sulfur dioxide' was present in the sample?
 D. This type of wine is considered to be 'a good wine' if the value for 'free sulfur dioxide' is greater than 30 ppm SO_2 (ppm = mg L^{-1}). Does this sample of wine rate as 'a good wine'? Show your working.

ANSWERS Calculations Set 8

1. A. $2\text{Al}^{3+}_{(\text{aq})} + 3\text{S}^{2-}_{(\text{aq})} \rightarrow \text{Al}_2\text{S}_3(\text{s})$ B. Na_2S is limiting C. 2.56 g
 D. 0.444g
 2. A. 87.9 B. $\text{C}_4\text{H}_8\text{O}_2$ C. methylpropanoate or ethylethanoate or propylmethanoate
 3. B. To overcome the Activation Energy (E_a) C. non-toxic, odourless & cool enough to avoid burning the passengers D. Red/Ox E. 131 g
 4. A. $\text{Cr}_2\text{O}_7^{2-}/\text{H}^+$ B. $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$ C. $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- = 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$
 D. $\text{Cr}_2\text{O}_7^{2-} + 6\text{Fe}^{2+} + 14\text{H}^+ = 2\text{Cr}^{3+} + 6\text{Fe}^{3+} + 7\text{H}_2\text{O}$ E. 0.000284 mol
 F. 0.01704 mol G. 95.2%
 5. A. sulfite +4; sulfate +6 B. I_2 C. $0.000515 \text{ mol L}^{-1}$
 D. Yes; because ppm (SO_2) 33.0 ppm > 30 ppm

CALCULATIONS

Set 9

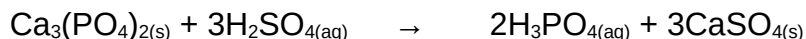
2. A double sulfate of potassium and chromium(III) has the general formula:



36.50 g of the pure hydrated compound is treated with excess sodium carbonate solution and it is found that 10.38 g of highly insoluble chromium(III) carbonate is precipitated. A further 4.700 g sample of the compound is heated strongly to drive off all the water of crystallization. A constant mass of 2.665 g of anhydrous powder remains after several heatings.

Calculate the empirical formula of the compound.

3. The production of the fertilizer di-ammonium hydrogen phosphate, $(\text{NH}_4)_2\text{HPO}_4$, from rock phosphate (impure calcium phosphate) can be represented by the following reaction sequence:



- A. A particular sample of rock phosphate weighing 60.0 kg (6.00×10^4 g) yielded 46.12 kg of diammonium hydrogen phosphate. Calculate the percentage calcium phosphate in the rock phosphate deposit. Assume 100% efficiency in the process.
- B. What volume of ammonia gas measured at S.T.P. would be needed to bring about the above conversion, given that 10.0% by volume of the ammonia gas escapes and is not used in the process?

4. A solution X is a mixture of sulfuric acid and ethandioic (oxalic) acid. It was analysed as follows:
- (i) 25.0 mL of X required 29.2 mL of a 0.602 mol L^{-1} sodium hydroxide for complete reaction. In this titration, all the available $\text{H}^+_{(\text{aq})}$ from both diprotic acids was neutralized.
 - (ii) A further 25.0 mL of X required 12.10 mL of a 0.207 mol L^{-1} acidified potassium permanganate solution for complete reaction. In this case, the permanganate solution was titrated into X until the first permanent pink colour remained in the solution. (You may assume that in this second titration, there is no chemical reaction between the sulfuric acid and the permanganate solution).

From the above data, calculate the masses of H_2SO_4 and HOOCCOOH present in 1.00 L of solution X.

Using the table of standard reduction potentials on the data sheet, write the half equation for the reduction of MnO_4^- to Mn^{2+} .
 Using the table of standard reduction potentials on the data sheet, write the half equation for the oxidation of HOOCCOOH to CO_2 .
 Write the balanced chemical equation for the reaction.

5. A chemistry student investigating the formation of complex ions decided to prepare some copper(II) ammonia complex ions. She added 150 mL of a 0.502 mol L^{-1} copper(II) sulfate solution to 500 mL of 3.50 mol L^{-1} ammonia solution.
- A. Write the equation for the formation of the complex ion.
 - B. Determine which of the two reagents was in excess, and by how many moles.
 - C. Calculate the final concentrations of the sulfate ions and the complex ions in the final solution.

ANSWERS Calculations Set 9

- 2. $\text{KCr}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$
- 3. A. 90.3% B. $1.74 \times 10^4 \text{ L}$
- 4. $n(\text{H}^+) = 0.01758 \text{ mol}$
 $2\text{MnO}_4^- + 5\text{H}_2\text{C}_2\text{O}_4 + 6\text{H}^+ = 2\text{Mn}^{2+} + 10\text{CO}_2 + 8\text{H}_2\text{O}$ 9.91 g H_2SO_4 ; 22.6 g $\text{H}_2\text{C}_2\text{O}_4$
- 5. A. $\text{Cu}^{2+} + 2\text{OH}^- \rightarrow \text{Cu}(\text{OH})_2 + 4\text{NH}_3 \rightarrow [\text{Cu}(\text{NH}_3)_4]^{2+} + 2\text{OH}^-$
 B. NH_3 inXS, Cu^{2+} is limiting; Ammonia inXS by 1.45 moles
 C. Concentration of both sulfate and complex ions is $1.16 \times 10^{-1} \text{ mol L}^{-1}$

CALCULATIONS

Set 10

1. The active ingredient of a drain cleaning powder is sodium hydroxide. A sample of this powder weighing 11.40 g dissolved in water and made up to 250.0 mL in a volumetric flask. A 20.00 mL sample of the resulting solution required 18.75 mL of 0.320 mol L^{-1} sulfuric acid for complete neutralization. You may assume that none of the other ingredients of the drain cleaning powder react with the sulfuric acid.
 - A. How many moles of sulfuric acid were used?
 - B. How many moles of sodium hydroxide were in the 20.00 mL sample?
 - C. What is the molar concentration of the sodium hydroxide solution?
 - D. How many grams of sodium hydroxide are contained in 1.00 kg of the drain cleaning powder?

3. 3.20 g of cast iron was dissolved in dilute sulfuric acid and the resulting solution made up to 100.0 mL with distilled water. 25.00 mL portions (aliquots) of this solution were then acidified with dilute sulfuric acid and titrated against 0.101 mol L^{-1} potassium permanganate solution. The volume of potassium permanganate solution to reach the end point was 26.45 mL.

Calculate the percentage by mass of iron in the cast iron. You may assume that none of the impurities in the cast iron react during the analysis.

4. The owner of a fresh water swimming pool of volume 31500 L wishes to chlorinate it for the first time using calcium hypochlorite, $\text{Ca}(\text{ClO})_2$.
 - A. What mass of calcium hypochlorite should be added such that the final concentration of the hypochlorite ion is 10.0 parts per million (ppm)?
(You may assume that $1 \text{ ppm} = 1 \text{ mg L}^{-1}$)
 - B. The pool owner actually more calcium hypochlorite than was necessary.

After some days **ALL** of the hypochlorite ion had been reduced to chloride according to the equation;



A sample of the pool water was analysed for chloride ion by adding excess silver nitrate and weighing the silver chloride which was precipitated. A 200.0 mL sample of pool water produced 0.0362 g of dry silver chloride. What mass of calcium hypochlorite did the pool owner actually put into the pool?

5. When 1.000 g of an organic compound containing only carbon, hydrogen and nitrogen was completely burned in air, the carbon dioxide produced had a volume of 1.250 L when measured at 101.3 kPa and 27 °C.

All of the nitrogen from a further 1.000 g sample was converted into ammonia, and this was dissolved in water and titrated with a standard 0.500 mol L⁻¹ hydrochloric acid solution. The volume of hydrochloric acid necessary was 33.90 mL.

- A. What mass of nitrogen is in 1.000 g of the organic compound?
B. Determine the empirical formula of the compound.
C. Given that the empirical formula and the molecular formula are identical, draw on possible structure for the compound.

ANSWERS CALCULATIONS Set 10

- | | | | | | | | | |
|----|--|---------------------------------|----|--------------------------------|----|--|----|------------------------------|
| 1. | A. | 6.00×10^{-3} | B. | 1.20×10^{-2} | C. | $6.00 \times 10^{-1} \text{ mol L}^{-1}$ | D. | $5.26 \times 10^2 \text{ g}$ |
| 3. | $\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ = \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O};$ | | | | | 93.2% | | |
| 4. | A. | 438 g | B. | $2.84 \times 10^3 \text{ g}$ | | | | |
| 5. | A. | $2.37 \times 10^{-1} \text{ g}$ | B. | $\text{C}_3\text{H}_9\text{N}$ | C. | Propanamine or isomers of propanamine | | |

CALCULATIONS

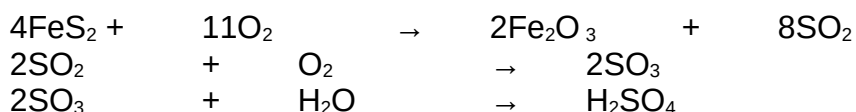
Set 11

- 2.440 g of a mixture of anhydrous sodium chloride and anhydrous sodium hydrogencarbonate required 18.5 mL of 0.705 mol L⁻¹ sulfuric acid for complete neutralization. Calculate the percentage by mass of sodium chloride in the mixture.
- "Volume strength" of a hydrogen peroxide solution is defined as the number of volumes of oxygen, measured at S.T.P., obtainable from the decomposition of one volume of hydrogen peroxide solution into water and oxygen.

A hydrogen peroxide (bleach) solution was analysed as follows:

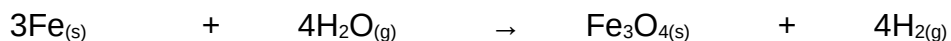
Bleach solution (25.0 mL) was diluted to 1.000 L with water. Samples of 20.0 mL of the diluted solution were titrated with 0.0204 mol L⁻¹ potassium permanganate, requiring 17.1 mL to reach the end point.

- Write the half-equation for the oxidation of hydrogen peroxide to oxygen and combine it with the half-equation for reduction of MnO₄⁻ to Mn²⁺ to establish the balanced equation for the titration reaction.
 - Calculate the concentration, in mol L⁻¹, of the diluted solution of hydrogen peroxide.
 - Calculate the concentration of hydrogen peroxide, in mol L⁻¹, in the bleach solution.
 - Write a balanced equation for the decomposition of hydrogen peroxide to water and oxygen, and hence calculate the volume strength of the bleach solution.
- The preparation of sulfuric acid from iron pyrites (FeS₂) can be represented by the following series of reactions.



What mass of iron pyrites would be needed to produce 125 kg of 90.0% sulfuric acid if the contact process converts 95.0% of the sulfur dioxide and the other reactions achieve 100% conversion?

5. Hydrogen can be prepared by passing superheated steam over iron:



A closed vessel of capacity 675 L contained a mixture of steam and hydrogen resulting from an incomplete reaction in the above process. The pressure was 42.3 atmospheres and the temperature was 253 °C. When the temperature was lowered to 10 °C, the water condensed and the pressure fell to 19.4 atmospheres. The vapour pressure of water at 10 °C is negligible and hydrogen can be taken to be insoluble in water.

Calculate the mass of water in the vessel.

ANSWERS CALCULATIONS Set 11

1. 10.2%
2. A. $2\text{MnO}_4^- + 6\text{H}^+ + 5\text{H}_2\text{O}_2 = 2\text{Mn}^{2+} + \text{O}_2 + 8\text{H}_2\text{O}$ B. 0.0436 molL^{-1} C. 1.74 molL^{-1}
D. $2\text{H}_2\text{O}_2 = \text{O}_2 + 2\text{H}_2\text{O}$; 19.5
3. $7.24 \times 10^4 \text{ g}$ or 72.4 kg
5. $1.76 \times 10^3 \text{ g}$ or 1.76 kg

CALCULATIONS

Set 12

1. A sample of industrial waste water is analysed and found to contain magnesium sulfate as the only significant impurity.

A 5.00 L sample of this waste water is evaporated to about 200 mL and sufficient barium nitrate solution is added to precipitate all the sulfate ion. The precipitate is washed with distilled water, filtered and dried. The mass of the precipitate is found to be 1.70 g.

- A. Determine the concentration in mgL^{-1} of magnesium sulfate in the waste water.
 - B. What mass of washing soda, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, would need to be added to 100.0 L of the waste water to precipitate all the magnesium ion?
3. A pure substance 'A' is a colourless pungent liquid which is found by qualitative analysis to contain carbon, hydrogen and chlorine. With no simple test for oxygen available, it can be assumed oxygen might be present.

A sample of 'A' is burnt in a current of dry oxygen and the resulting gases passed over silver foil to remove chlorine-containing compounds. When 1.036 g of 'A' is so burnt, 0.705 g of carbon dioxide and 0.145 g of water are produced.

In a second experiment, a different sample of 'A' is heated under pressure with excess concentrated nitric acid and silver nitrate. When 0.945 g of 'A' is so treated, 2.108 g of silver chloride is produced.

Calculate the empirical formula of 'A'

4. A pure substance 'A' is a liquid boiling at 194°C . 1.650 g of 'A' is vaporized at 200°C and 101.3 kPa, and the gas produced occupies 250 mL.

Calculate the relative molecular mass of 'A'.

5. Experiments have established that a pure substance 'A' is a soluble monoprotic acid.

1.0308 g of 'A' is dissolved in water in a 250.0 mL volumetric flask and the volume made up to the mark. 20.00 mL portions (aliquots) of this solution are titrated with $0.1031 \text{ mol L}^{-1}$ sodium hydroxide *in the burette* using phenolphthalein as indicator, and the following titration figures are recorded.

Final reading (mL)	10	17	23.32	29.5	35.70
Initial reading (mL)	0.07	9.96	17.10	23.32	29.50

- Calculate the appropriate value for the volume of sodium hydroxide solution required for titration.
- From the titration results determine the number of moles of 'A' in the original 1.0308 g sample of 'A' and hence calculate the relative molecular mass of 'A'.
- Suggest one way in which the experiment could be modified to involve less experimental error, so that a more accurate value for the molecular mass might be obtained.
- The pure substance 'A' referred to in this question is the same substance that was used in Questions 3 and 4. An experienced chemist recognizes that the molecular mass calculated in Question 5 is the correct value, while the Question 4 molecular mass is wrong. Determine the structural formula for 'A' and comment briefly on the evidence (conflicting or otherwise) provided by the three experiments in Questions 3, 4 and 5.

ANSWERS CALCULATIONS Set 12

- 175 mg L^{-1}
 - 41.7 g
- CHC/O
- 256
- 6.20 mL
 - 0.00799 mol; 129
 - Modify the experiment so that a titre closer to 20 mL might be obtained; e.g. use a more dilute solution of NaOH
 - Mass of CHC/O=64.5 hence $129/64.5 = 2$; therefore the molecular formula is $(\text{CHC/O})_2$
 Since 'A' is an acid it probably contains alkanoic acid $-\text{COOH}$ group. Since the gas volume experiment gives double the correct molecular mass, the molecules must exist as dimers where they are held together by hydrogen bonds between the two $-\text{COOH}$ groups.

CALCULATIONS

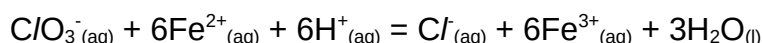
Set 13

1. An unknown organic compound X contains only the elements carbon, hydrogen and oxygen.

When 1.360 g of X is completely burned in an excess of oxygen the products are found to be 1.113 g of water and 1.384 L of carbon dioxide measured at a temperature of 273 K and 101.3 kPa.

When a further 2.500 g sample of X is vapourised the volume of the vapour produced is found to be 868.9 mL when measured at a temperature of 373 K and a pressure of 101.3 kPa.

- A. Determine the empirical formula of X.
B. Determine the molecular formula of X.
C. Given that X will react with sodium carbonate to produce carbon dioxide, draw and name two possible structures for X.
2. Iron(II) solutions can be used to determine the concentration of chlorate ion in a particular sample. The balanced equation is as follows:



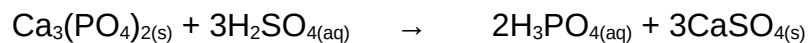
0.721 g of the impure potassium chlorate was dissolved in the minimum amount of water and then made up to a volume of 250 mL in a volumetric flask using distilled water. 20.0 mL of this solution was titrated with 0.106 mol L⁻¹ iron(II) solution.

The burette readings were as follows

Titration Number	1	2	3	4	5
Initial volume mL	0.00	23.60	3.10	25.60	1.00
Final volume mL	29.60	46.00	25.60	48.05	23.30

- A. What volume of Fe²⁺ solution is required to react fully with the 20.0 mL of the chlorate solution?
B. Calculate the number of moles of Fe²⁺ used.
C. Calculate the number of moles of potassium chlorate in a 20.0 mL sample.
D. Calculate the percentage purity of the potassium chlorate.

3. The production of the fertilizer di-ammonium hydrogen phosphate, $(\text{NH}_4)_2\text{HPO}_4$, from rock phosphate (impure calcium phosphate) can be represented by the following reaction sequence:



- A. What mass of rock phosphate would be required to produce 1.00 tonne (1000 kg) of di-ammonium hydrogen phosphate if the rock phosphate contains 70.0% by mass calcium phosphate $\text{Ca}_3(\text{PO}_4)_2$? You may assume that the impurities in the rock phosphate do not react with the sulfuric acid or take part in the reaction.
- B. What volume of ammonia gas measured at S.T.P. would be needed to bring about the above conversion?
4. Solution A is a mixture of sulfuric and hydrochloric acids of unknown concentration.

20.0 mL of A required 10.5 mL of a 0.715 mol L^{-1} sodium hydroxide for complete neutralisation.

An excess of barium chloride solution was added to a separate 20.0 mL sample of A, and this resulted in a precipitate of 0.500 g of barium sulfate (BaSO_4).

Calculate the concentration of hydrochloric acid in solution A in grams per litre.

ANSWERS CALCULATIONS Set 13

- | | | | | | | |
|----|----|--------------------------------|----|----------------------------------|----|--|
| 1. | A. | $\text{C}_2\text{H}_4\text{O}$ | B. | $\text{C}_4\text{H}_8\text{O}_2$ | C. | butanoic acid & 2-methylpropanoic acid |
| 2. | A. | 22.4125 mL | B. | 0.002375 mol | C. | 0.0003958 mol |
| 3. | | A. 1.68 tonne | B. | 339 kL or m^3 | D. | 84.1% |
| 4. | | 5.87 g L^{-1} | | | | |

CALCULATIONS

Set 14

1. Silicon carbide, SiC, a very hard material used as an abrasive on sandSet and in other applications, is prepared by the reaction of pure sand, SiO₂, with carbon at high temperature. Carbon monoxide, CO, is the other product of this reaction.

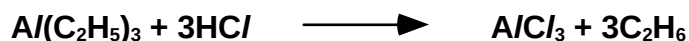
- A. Write the balanced equation for the reaction.
B. Calculate how much SiO₂ is required to produce 1.00 Kg of SiC.

C.

What type of chemical bonding would you expect to find in SiC?

- D. Use your answer in C. above to explain why SiC is a very hard material.

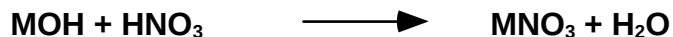
2. A 25 mL volume of a solution of triethylaluminium Al(C₂H₅)₃, a substance used in the production of the plastic polyethene, was allowed to react with 25.0 mL of a 0.103 mol L⁻¹ HCl solution.



The Al(C₂H₅)₃ was the limiting reagent, so some HCl remained unreacted. The unreacted HCl was titrated with 0.142 mol L⁻¹ solution of NaOH; 16.75 mL being required.

What was the concentration of the Al(C₂H₅)₃ in the solution?

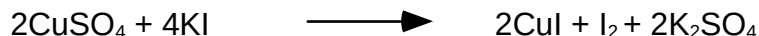
3. A common oven cleaner contains a very caustic metal hydroxide, which we can represent as MOH. Titration of 0.7134 g of MOH with 1.000 M HNO₃ requires 17.85 mL to reach end point.



- A. What do you understand by the term end point.
B. What is the element M?
C. What is the formula for the hydroxide of M?
D. In the main, the dirt in ovens consists of fats/oils and grease. Explain how an oven cleaner, whose active constituent is MOH, actually works. (An equation may help you in explaining your answer)

4. Bronzes used in bearings are often alloys (solid solutions) of copper and aluminium.

A. A 1.953 g sample of one such bronze was analysed for its copper content by the sequence of reactions given below. The aluminium also dissolves, but we need not consider it because aluminium sulphate, $\text{Al}_2(\text{SO}_4)_3$, does not react with KI.



If 35.06 mL of 0.837 M $\text{Na}_2\text{S}_2\text{O}_3$ is required to react with the I_2 formed, what is the percentage by mass of Cu in the sample?

B. Why does the aluminium sulphate not react with the potassium iodide?

5. When 1.5173 g of an organic iron compound containing Fe, C, H and O was burned in oxygen, 2.838 g of CO_2 and 0.8122 g of H_2O were produced.

In a separate experiment to determine the mass % of iron, 0.3355 g of the compound yielded 0.0758 g of Fe_2O_3 .

What is the empirical formula of the compound?

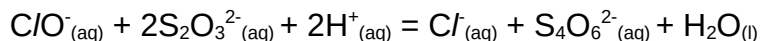
ANSWERS CALCULATIONS Set 14

1. A. $\text{SiO}_2 + 3\text{C} \rightarrow \text{SiC} + 2\text{COB}$. 1500 g or 1.50 kg C. network covalent
D. The atoms in SiC are covalently bound in a lattice. These strong covalent bonds are hard to break and hence the atoms are not easily displaced (dislodged).
2. $2.62 \times 10^{-3} \text{ M}$ or $0.00262 \text{ mol L}^{-1}$
3. A. Point at which the indicator changes colour for a small amount of added titrant
B. $A_r = 22.99$ hence element M = Na C. NaOH D. Fats/oils are triglycerides with ester linkages. NaOH hydrolyses this linkage and forms 1,2,3-propantriol and soap. The fats/oils are insoluble in water whereas the 1,2,3-propantriol and soap are.
4. A. 95.5% B. Aluminium ions and iodide ions do not react together (either in an electron transfer reaction (Red/Ox) or to form a precipitate). All potassium salts are soluble in aqueous solution.
5. $\text{C}_{15}\text{H}_{21}\text{O}_6\text{Fe}$

CALCULATIONS

Set 15

1. A sample of impure calcium carbonate weighing 0.2450 g is dissolved in 100 mL of standardized 0.0905 molL⁻¹ hydrochloric acid solution. After the reaction is complete, it is found that the excess acid required 17.0 mL of 0.250 molL⁻¹ sodium hydroxide for complete neutralization using phenolphthalein indicator. Find the percentage calcium carbonate in the sample.
2. A textile company uses a solution of sodium hypochlorite to bleach sheets. An industrial chemist employed by the company is asked to investigate the change in the concentration of the bleaching solution after the sheets have been soaked in this solution for a time of 2 hours. The original solution is known to contain 30.0 g of sodium hypochlorite per litre of solution. [The bleaching process slowly reduces hypochlorite ions to chloride ions while at the same time changing double bonds in the sheet fibres into single bonds – an oxidation process]. The chemist removes a sample of the bleaching solution after two hours of bleaching and tests it with sodium thiosulfate solution to determine the final concentration of hypochlorite ions. The redox reaction between hypochlorite ion and thiosulfate ion in an acidified solution may be represented as follows:



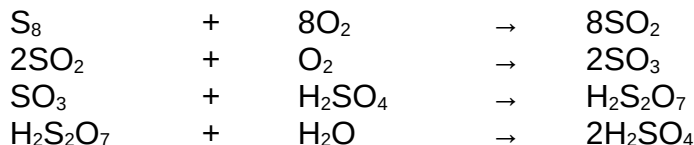
Using a starch-iodine indicator to moderate the equivalence point of the reaction, it is found that an average of 47.5 mL of 0.250 molL⁻¹ sodium thiosulfate solution is needed to react with 20.0 mL samples of the final bleaching solution.

- A. Calculate the concentration of sodium hypochlorite in molL⁻¹ in the original solution.
 - B. Calculate, from the given redox reaction, the concentration of hypochlorite ions in the final bleaching solution.
4. An organic compound containing only carbon, hydrogen and nitrogen was analysed as follows. A sample of the compound of mass 2.435 g was burnt in excess oxygen and produced 4.753 g of carbon dioxide and 3.405 g of water.

When vapourised in the absence of air, a further 2.435 g sample of the compound was found to occupy 1.321 L at 25 °C and 101.3 kPa.

- A. Calculate the empirical formula of the compound.
- B. Calculate the molecular formula of the compound.
- C. Given that the compound is an amine, draw a structural formula for the compound and give it a systematic or IUPAC name.

5. In the manufacture of sulfuric acid, liquid sulfur is burned in air to produce sulfur dioxide. The sulfur dioxide is then catalytically converted into sulfur trioxide and then to fuming sulfuric acid, $\text{H}_2\text{S}_2\text{O}_7$, by adding the sulfur trioxide to pure sulfuric acid. Finally, sulfuric acid is produced by adding a stoichiometric amount of water to the fuming sulfuric acid.



- A. Given that the liquid sulfur is 98.0% by mass pure, and that the impurity does not react, find the mass of liquid sulfur needed to produce 1.00 tonne (10^6 g) of sulfuric acid.
- B. Calculate the volume of oxygen measured at 350°C and 98.0 kPa needed to produce the above amount of sulfuric acid.
- C. Find the mass of water required in the production of 1.00 tonne of sulfuric acid.

ANSWERS CALCULATIONS Set 15

- | | | | | |
|----|-------|--------------------------------|----|--------------------------------|
| 1. | 98.0% | | | |
| 2. | A. | 0.403 mol L^{-1} | B. | 0.297 mol L^{-1} |
| 4. | A. | $\text{C}_2\text{H}_7\text{N}$ | B. | $\text{C}_2\text{H}_7\text{N}$ |
| | | | C. | Ethanamine |
| 5. | A. | $3.34 \times 10^5 \text{ g}$ | B. | $8.08 \times 10^5 \text{ L}$ |
| | | | C. | $1.84 \times 10^2 \text{ kg}$ |

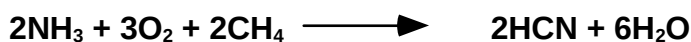
CALCULATIONS

Set 16

1. A 0.608 g sample of fertiliser containing nitrogen as ammonium sulphate, $(\text{NH}_4)_2\text{SO}_4$, was analysed for nitrogen by heating with sodium hydroxide, the ammonium sulphate being converted to sodium sulphate, ammonia gas and water. The ammonia was collected in 46.3 mL of 0.213 M HCl with which it reacted. The resultant excess hydrochloric acid was titrated with 0.128 M NaOH, 44.3 mL being required for neutralisation.

What is the percentage nitrogen in the fertiliser?

2. Hydrogen cyanide, HCN, is prepared from ammonia, air and natural gas (CH_4) by the following process.



Hydrogen cyanide is used to prepare sodium cyanide, which is used to obtain gold from gold-bearing rock. If a reaction vessel contains 11.5 g NH_3 , 10.0 g O_2 and 10.5 g CH_4 , what is the maximum volume of hydrogen cyanide gas that could be made at 300 °C and 99.5 kPa, assuming the reaction goes to completion?

3. A typical formulation for window glass is 75% SiO_2 , 15% Na_2O and 10% CaO by mass.

What masses of sand, sodium carbonate and calcium carbonate must be combined to produce 1.00 kg glass after the carbon dioxide is driven off by thermal decomposition of the carbonates.

4. 25 mLs of a solution containing chromium(III) ions was treated with H_2O_2 in alkaline solution to oxidise the Cr^{3+} to CrO_4^{2-} . The excess peroxide was removed by boiling and the solution acidified with sulphuric acid. 25 mLs of 0.100 mol L^{-1} FeSO_4 solution was added, and the residual(excess) Fe^{2+} was back titrated with 0.01680 mol L^{-1} $\text{K}_2\text{Cr}_2\text{O}_7$, giving a 10.06 mL back titre.

- A. Write balanced net ionic half equations for the reactions involved, including the initial oxidation of chromium (four in total).
- B. Calculate the chromium content of the original solution in mol L^{-1}

5. In most of its ionic compounds, cobalt is either Co(II) or Co(III). One such compound, containing chloride ion and water of hydration (water of crystallisation), was analysed, and the following results were obtained.

A 0.256 g sample of the compound was dissolved in water, and excess silver nitrate was added. The silver chloride was filtered, dried and when weighed had a mass of 0.308 g.

A second sample of 0.416 g of the compound was dissolved in water and an excess of sodium hydroxide was added. The hydroxide salt was filtered and heated in a flame, forming cobalt(III) oxide. The mass of cobalt(III) oxide formed was 0.145 g.

- A. Assuming the compound contains one cobalt atom per formula unit, what is the formula?
- B. Write balanced net ionic equations for
- the dissolution of the cobalt compound in water
 - the reaction of silver nitrate solution with the solution of the cobalt compound.

ANSWERS CALCULATIONS Set 16

- 9.66%
- O₂ is the limiting reagent 9.98 L
- m(SiO₂) = 750 g m(Na₂CO₃) = 257 g m(CaCO₃) = 178 g
- $$\begin{array}{lcl} \text{Cr}^{3+} + 4\text{H}_2\text{O} & \rightarrow & \text{CrO}_4^{2-} + 8\text{H}^+ + 3\text{e}^- \\ \text{H}_2\text{O}_2 + 2\text{e}^- & \rightarrow & 2\text{OH}^- \\ \text{Fe}^{2+} & \rightarrow & \text{Fe}^{3+} + \text{e}^- \\ \text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- & \rightarrow & 2\text{Cr}^{3+} + 7\text{H}_2\text{O} \end{array}$$
 - 0.0198 molL⁻¹
- CoCl₂·6H₂O
 - $\text{CoCl}_{2,6}\text{H}_2\text{O}_{(\text{s})} \rightarrow \text{Co}^{2+}_{(\text{aq})} + 2\text{Cl}^{-}_{(\text{aq})} + 6\text{H}_2\text{O}_{(\text{l})}$
 - $\text{Ag}^{+}_{(\text{aq})} + \text{Cl}^{-}_{(\text{aq})} \rightarrow \text{AgCl}_{(\text{s})}$

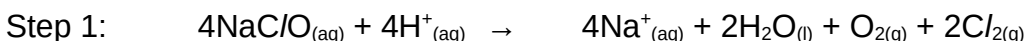
CALCULATIONS

Set 17

- 3.53 g of an organic compound was completely burnt and analysis of the two products showed that 5.16 g of carbon dioxide and 2.118 g of water had been produced.

Another 7.32 g of the compound was vapourised in a 2.00 L container and produced a pressure of 200.0 kPa at a temperature of 120.0 °C.

- Determine the empirical formula of the compound.
 - Determine the molecular formula of the compound.
 - Draw 2 possible structural formulae for the compound and assign each an IUPAC or systematic name.
- The quantity of sodium hypochlorite in house bleach can be determined through a series of three steps, the final step being a titration between iodine and sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$) solutions.



In one such analysis, a 20.0 mL sample of household bleach was placed in a 250.0 mL volumetric flask and then enough distilled water was added to correctly fill the flask. 20.0 mL of this diluted bleach solution was then subjected to the 3 steps outlined above. In step 3, 35.0 mL of $0.0612 \text{ mol L}^{-1}$ sodium thiosulfate was required to reach the end point of the titration.

- Determine the number of moles of sodium thiosulfate consumed in the titration.
 - Determine the number of moles of sodium hypochlorite in 20.0 mL of the diluted bleach solution.
 - Calculate the concentration of the original bleach solution.
- A 225 L aquarium was set up to have a pH of 7.00. 31.0 mL of 12.0 mol L^{-1} hydrochloric acid was accidentally added to the aquarium. In an attempt to return the pH to 7.00, 1.00 L of a 0.250 mol L^{-1} calcium hydroxide solution was added to the aquarium.
- Calculate the pH of the aquarium after the HCl was added.
 - Calculate the final pH of the aquarium.

5. The following procedure was used to accurately determine the volume strength of 30 volume hydrogen peroxide.

- Step 1: 25.00 mL of the peroxide was pipetted into a 1.000 L volumetric flask which was subsequently filled with distilled water.
- Step 2: 20.00 mL of this diluted peroxide solution was then added to 20.00 mL of sulfuric acid in a conical flask.
- Step 3: 35.6 mL of a $0.0195 \text{ mol L}^{-1}$ standardised solution of potassium permanganate was required to reach the end point.
- A. Use half equations to write the full net ionic equation for the oxidation of hydrogen peroxide.
- B. Determine the number of moles of KMnO_4 used.
- C. Determine the number of moles of H_2O_2 in the 20.0 mL of the diluted solution.
- D. Determine the concentration (in mol L^{-1}) of the original H_2O_2 solution.
- E. Determine the volume strength of the original H_2O_2 solution.

ANSWERS CALCULATIONS Set 17

- | | | | | | | |
|----|----|---|----|----------------------------------|------------|--------------------------------------|
| 1. | A. | CH_2O | B. | $\text{C}_2\text{H}_4\text{O}_2$ | C. | Ethanoic acid or Methylmethanoate |
| 2. | A. | 2.142×10^{-3} | B. | 1.34 mol L^{-1} | | |
| 3. | A. | 2.78 | B. | 10.8 | | |
| 5. | A. | $2\text{MnO}_4^- + 5\text{H}_2\text{O}_2 + 6\text{H}^+ = 2\text{Mn}^{2+} + 5\text{O}_2 + 8\text{H}_2\text{O}$ | | | | |
| | B. | 6.942×10^{-4} | C. | 1.736×10^{-3} | in 20.0 mL | D. 3.47 mol L^{-1} E. 38.9 |

CALCULATIONS

Set 18

1. The hardness of a bore water sample was determined by analysis for calcium. An excess of sodium oxalate [sodium ethandioate) [$\text{Na}_2\text{C}_2\text{O}_4$ or $\text{Na}_2(\text{O}_2\text{CCO}_2)$], was added to 1.000 L of the bore water, and the pH adjusted with ammonia to ensure complete precipitation of the calcium ions as calcium oxalate.

The resulting slurry was filtered, the precipitate washed, and the oxalate ion was brought into solution as oxalic acid (ethandioic acid) (HOOCOOH), using dilute sulfuric acid.

This solution was titrated with potassium permanganate solution, and required 10.20 mL of $0.1000 \text{ mol L}^{-1}$ permanganate before the first faint permanent pink colour indicated the end point.

- A. Use half-equations from your table of standard reduction potentials to write a balanced equation for the reaction permanganate ion with oxalic acid in acid solution, and state how many moles of oxalic acid are oxidized by 1 mole of permanganate.
- B. Calculate the number of moles of oxalic acid oxidized in the titration with 10.20 mL of $0.1000 \text{ mol L}^{-1}$ permanganate.
- C. Hardness in water is expressed as a number of “degrees of hardness” where a degree of hardness is defined as the equivalent of 1 mg of CaCO_3 in solution in 1 L of the hard water.

Use the answer to B. to state the number of moles of Ca^{2+} in 1.000 L of the hard water. Hence calculate the number of mg of CaCO_3 corresponding to this amount of Ca^{2+} .

2. A. A pure compound, which is a colourless liquid at room temperature, boils at 138°C , and analysis shows it to contain 90.49% carbon and 9.48% hydrogen by mass. Calculate the empirical formula of the compound.
- B. Explain why the empirical formula calculated in A. cannot be its molecular formula?
- C. A known mass of the liquid is vapourised, and from the volume of gas produced the molecular mass of the compound is calculate to be about 105. What is the molecular formula of the compound?
- D. The odour of the liquid, and the smoky flame when the sample is burned, indicates that compound is aromatic. Draw one of four possible structural formulae for the compound.

3. A sample of brass (consisting only of copper and zinc) is dissolved in concentrated nitric acid and the volume made up to 500.0 mL with distilled water.

200.0 mL of this solution is treated with an excess of sodium hydroxide solution to neutralize the excess nitric acid and to precipitate all the copper(II) ion and to convert all the zinc ion into $[\text{Zn}(\text{OH})_4]^{2-}$. The mixture is filtered and the residue washed with distilled water. The residue is then mixed with 100 mL of distilled water and boiled until all the residue is converted to copper(II) oxide. The oxide is collected by filtration, washed with distilled water and dried. The mass of copper(II) oxide is found to be 3.005 g.

A further 200.0 mL of the original solution is treated with an excess of sodium carbonate solution to neutralize the excess nitric acid and to precipitate all the copper(II) ion and zinc ion. The resulting mixture is filtered and the precipitate washed, dried and heated strongly to convert it into a mixture of copper(II) oxide and zinc oxide. The mass of this solid is 4.996 g.

- A. Calculate the mass of copper in the sample.
- B. Calculate the percentage by mass of zinc in the sample.

4. A bottle of drain cleaner, the active ingredient of which is sodium hydroxide, is accidentally tipped into a farmer's storage tank. The farmer recognises that the hydroxide ion is likely to corrode the tank and looks desperately around for something to react with it. The only available useful substance seems to be a drum of bluestone (copper sulfate) solution which is waiting to be turned into Bordeaux Mixture. The farmer hastily tips the contents into the tank.

If either corrosive hydroxide ion or poisonous copper ion is left over in the mixture, this will need to be removed before the tank is pumped out.

In summary

Tipped into the 10000 L of water in the tank is

Drain cleaner: 4.00 kg of 20.0% (by mass) NaOH solution, and

Bluestone solution: 45.0 kg of 5.00% (by mass) $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ solution.

- A. Calculate which of OH^- or Cu^{2+} is left over, and how much there is of it in excess in the tank.
- B. Suggest a substance that could be added to the water to remove the excess component so that the tank can be safely drained. Without doing any further calculations, state how to determine the necessary amount of material to add.

ANSWERS

CALCULATIONS Set 18

1. A. $2\text{MnO}_4^- + 5\text{H}_2\text{C}_2\text{O}_4 + 6\text{H}^+ = 2\text{Mn}^{2+} + 10\text{CO}_2 + 8\text{H}_2\text{O}$
B. 0.00255 C. 255 mg
2. A. C_4H_5 B. It is impossible to tie up all the bonds. C. C_8H_{10}
D. Ethyl benzene or 1.2- dimethyl benzene or 1.3- dimethyl benzene or 1.4- dimethyl benzene
3. A. 6.002 g B. 40.0%
4. A. OH^- left over = 1.98 mol
B. Add CH_3COOH until pH reaches 7 or
Add HCl until the pH reaches just 7 (or until blue colour appears in solution) and then add back base or
Add the calculated quantity on HCl or
Add CuSO_4 solution with stirring, watching until no more precipitate forms or
Add the calculated quantity of CuSO_4 in solution to precipitate the 1.98 mol of hydroxide.

CALCULATIONS

Set 19

2. The yellow mineral mendipite is a pure compound which contains only lead, oxygen and chlorine.
- A. 1.205 g of mendipite is heated strongly in a current of hydrogen gas until constant mass is reached. The grey blob which remains consists of 1.034 g of lead metal. Calculate the percentage of lead in mendipite.
- B. 1.803 g of mendipite is boiled with concentrated nitric acid for 1 hour; the solution is then cooled and diluted, and dilute silver nitrate solution added until no further precipitate forms. The mixture is left to stand in the dark for 1 hour, and when it is filtered, and the precipitate washed and dried, 0.713 g of silver chloride is obtained. Calculate the percentage of chlorine in mendipite.
- C. Calculate the empirical formula of mendipite.
- D. The crystal lattice of mendipite consists of monoatomic ions. Write the formula of these ions.
3. A sample of Perth tap water contains 12.13 mg L^{-1} of magnesium and 5.01 mg L^{-1} of calcium.

In order to soften 200.0 L of tap water, a scientist adds 50.00 g of sodium carbonate decahydrate ($\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$) to precipitate the magnesium ion and the calcium ion.

Calculate whether sufficient sodium carbonate has been added, and determine the mass of additional sodium carbonate decahydrate that needs to be added or the mass that is in excess.

4. Permanganate ion may be used in a titration to analyse hydrogen peroxide solutions.

A. From your E^0 Table obtain

- The half equation for the reduction of permanganate ion (in acid conditions) to give manganese(II) ion
- The half equation for the oxidation of hydrogen peroxide (in acid conditions) to give oxygen.

Use these half equations to construct the redox equation for the reaction that occurs when potassium permanganate solution is added to a solution containing hydrogen peroxide and sulfuric acid.

B. Hairdressers use hydrogen peroxide to bleach hair. An analyst uses a pipette to transfer 20.00 mL of commercial hairdressers' hydrogen peroxide to a 250.0 mL volumetric flask, and she makes the volume up to the mark with distilled water. She places 20.00 mL of this solution in a titration flask, adds 5 mLs of 2 mol L^{-1} sulfuric acid, and titrates the mixture with $0.02056 \text{ mol L}^{-1}$ potassium permanganate. She records the following titration figures.

Titration Number	1	2	3	4	5
Initial volume mL	0.11	1.55	0.41	0.81	0.66
Final volume mL	29.5	29.53	29.35	28.74	28.62

Calculate the approximate value for the volume of potassium permanganate solution required for titration.

C. Calculate the concentration, in mol L^{-1} , of the original commercial hairdressers' hydrogen peroxide.

5. A consignment of pure chalcocite, Cu_2S , has over a period of time become partially oxidized and hydrated, giving $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ and mixed oxides of copper. This mixture of Cu_2S , $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$, and oxides of copper was treated as follows.
- A. A representative 1025 g sample was washed with water to remove the copper sulfate, and then dried and reweighed. The mass of the washed and dried sample (consisting of Cu_2S and copper oxides) was 919 g. Calculate the percentage by mass of copper sulfate pentahydrate in the consignment.
- B. A 0.536 g sample of the dried washed material (Cu_2S and copper oxides) was roasted in an excess of oxygen and the gases treated so that all the sulfur in the sample was converted into sulfuric acid.
- The sulfuric acid solution from the 0.536 g sample was titrated with sodium hydroxide solution, and required 27.53 mL of $0.2000 \text{ mol L}^{-1}$ NaOH solution for an end point with phenolphthalein.
- (i) Determine the number of moles of sulfur in the 0.536 g sample.
- (ii) Calculate the percentage by mass of Cu_2S in the dried sample.
- (iii) Calculate the percentage by mass of Cu_2S in the original consignment.
- C. Calculate the percentage by mass of mixed copper oxides in the consignment.

ANSWERS CALCULATIONS Set 19

2. A. 85.81% B. 9.779% C. $\text{Pb}_3\text{Cl}_2\text{O}_2$ D. Pb^{2+} , Cl^- , O^{2-}
3. $m(\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}) = 35.71$
Extra $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} = 14.3 \text{ g}$
4. A. $2\text{MnO}_4^- + 5\text{H}_2\text{O}_2 + 6\text{H}^+ = 2\text{Mn}^{2+} + 5\text{CO}_2 + 8\text{H}_2\text{O}$
B. 27.96 mL C. 0.898 mol L^{-1}
5. A. 10.34%
B. (i) 0.002753 mol (ii) 81.75% (iii) 73.30%
C. 16.4%

CALCULATIONS

Set 20

1. A pure substance 'A' is a white solid, melting at 153°C , which is found by qualitative analysis to contain carbon and hydrogen. With no simple test for oxygen available, it can be assumed oxygen might be present.

When 0.8062 g of 'A' is burnt in a current of dry oxygen, 1.110 g of carbon dioxide and 0.303 g of water are produced.

Calculate the empirical formula of 'A'

2. An analyst has a pure substance 'A' which is a weak acid and she decides to determine its molecular mass by titration.

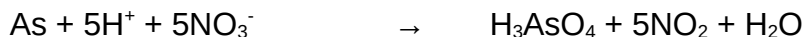
She dissolves 1.3867 g of 'A' in water and makes the solution up to 250.0 mL in a volumetric flask. She then titrates 20.00 mL aliquots (portions) of the solution of 'A' with $0.09826\text{ mol L}^{-1}$ sodium hydroxide *from the burette* using phenolphthalein as indicator, and records the following titration figures.

Titration Number	1	2	3	4	5
Initial volume mL	0.11	1.32	0.67	0.08	1.20
Final volume mL	20	19.5	18.32	17.77	18.87

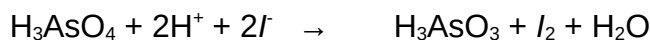
- A. Calculate the appropriate value for the volume of sodium hydroxide solution required for titration.
- B. Assume that 'A' is a monoprotic acid and calculate the number of moles of 'A' in the 1.3867 g of 'A'.
- C. Use this value to calculate the molecular mass of 'A'.
- D. Assume that 'A' is a diprotic acid and calculate its molecular mass.
- E. Assume that 'A' is a triprotic acid and calculate its molecular mass.
- F. The pure substance 'A' referred to in this question is the same substance that was used in Question 1. Suggest a likely molecular formula for 'A'.
- G. Suggest a structural formula for 'A'.

3. The amount of arsenic in a specimen may be determined through the following series of reactions:

A sample is boiled with excess concentrated nitric acid, and all the arsenic (whether it is as the element or in a compound) is converted into arsenic acid. Arsenic metal, for example reacts as follows:



The arsenic acid is then reacted with concentrated hydroiodic acid:

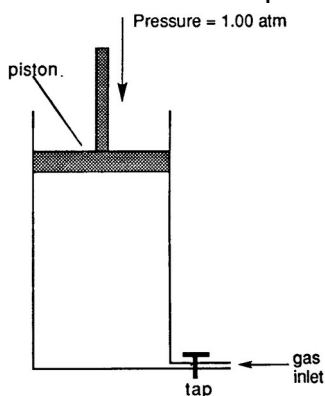


The iodine so produced is titrated with sodium thiosulfate solution using starch as indicator:



A 1.0346 g sample of an arsenic-containing substance is treated as described above and ultimately when the sample is titrated, it requires 44.25 mL of 0.1000 molL⁻¹ sodium thiosulfate solution.

- A. Use the above equations to work out how many moles of thiosulfate ion will be needed to react with the iodine produced from 1 mole of arsenic?
 - B. Use this value to calculate the percentage by mass of arsenic metal in the sample.
4. 2.00 L of propane (C₃H₈), 30.0 L of oxygen and 15.0 L of nitrogen, all at 1.00 atm and 25 °C, were introduced into the cylinder illustrated. A spark was used to ignite the gas mixture.



- A. Write the equation for the reaction that has occurred.

The water in the mixture is frozen out, and removed from the cylinder.

- B. The original conditions of temperature and pressure are then restored. Calculate the volume and composition (by volume) of the gas mixture.
- C. Calculate the volume of water produced, at 1.00 atm and 25 °C.
(1.00 mL of water has a mass of 1.00 g)

5. Experiments show that a grayish blue salt of the metal M has the composition $MCI_3 \cdot xH_2O$.

1.5674 g of the salt was dissolved in water, and silver nitrate solution added until no further precipitation occurred. The washed and dried precipitate of silver chloride was found to weigh 2.527 g.

A second (and much larger) sample of the salt was dissolved in water and the solution electrolysed with inert electrodes. When a current of 0.750 ampere was passed through the solution for 6 hours 18 minutes and 30 seconds, the mass of the cathode increased by 3.047 g.

- A. Calculate the atomic mass of M.
- B. Calculate the value of x.
- C. What element could M be?

ANSWERS CALCULATIONS Set 20

1. $C_6H_8O_7$
2. A. 17.67 mL B. 0.02170 mol C. 63.9 D. 127.8
E. 191.7 F. 192.1
Any structural formula with molecular formula $C_6H_8O_7$ and 3 $-COOH$ groups. (Citric acid has 3 $-COOH$ groups and 1 $-OH$ group)
3. A. 1 mol As produces 1 mol H_3AsO_4 produces 1 mol $H_3AsO_3 + I_2$ which reacts with 2 mol of $S_2O_3^{2-}$
B. 16.0%
4. A. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
B. 41.0 L C. 5.88 mL
5. A. 22710 s 51.79 g B. 6
C. Cr [It has atomic mass of 52.00, well within experimental error of 51.8 and has an oxidation state +3]
[The original grayish-blue salt then is $CrCl_3 \cdot 6H_2O$ or $Cr(H_2O)_6Cl_3$]

CALCULATIONS

Set 21

1. A pure substance 'A' is a colourless liquid which boils at 57 °C and has a strong odour. 'A' burns readily in air leaving no ash, and qualitative analysis shows that nitrogen, sulfur and the halogens are absent. It may thus be concluded that 'A' contains carbon and hydrogen, and possibly oxygen.
 - A. In an experiment, 0.6047 g of 'A' is burnt in a current of dry air, and 1.078 g of carbon dioxide and 0.441 g of water are produced. Calculate the empirical formula of 'A'.
 - B. At 100 °C and 1.00 atm pressure, 0.1376 g of 'A' evaporates to occupy a volume of 57 mL. Calculate the molecular mass of 'A'.
 - C. What is the molecular formula of 'A'?
 - D. A nuclear magnetic resonance (NMR) spectrum of 'A' indicates that all the H atoms in 'A' are in methyl groups. Draw the structural formula for 'A'.
2. Cerium(IV) ion is an important industrial oxidizing agent, but no cerium(IV) compound is a primary standard. Cerium(IV) solutions can be standardized as follows:

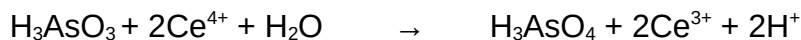
The primary standard arsenic(III) oxide is dissolved in sodium hydroxide solution which converts it to arsenite ion:



Dilute sulfuric acid is then added to give arsenious acid:



The arsenious acid is then titrated with cerium(IV) sulfate solution, using osmium tetroxide as catalyst and ferroin as indicator:



A 0.2476 g sample of arsenic(III) oxide is treated as described above, and ultimately when the sample is titrated, it requires 46.34 mL of unknown cerium(IV) sulfate solution.

- A. use the above equations to work out how many moles of cerium(IV) ion will react with the arsenious acid produced from 1 mole of arsenic(III) oxide.
- B. Use this value to calculate the concentration of the $\text{Ce}(\text{SO}_4)_2$ solution.

3. A brand of toilet cleaner claims on the label to contain

Sodium hydrogensulfate NaHSO_4 , 410 g/kg.

A chemist wishes to verify this claim by carrying out a volumetric analysis, using a standard solution of $0.04343 \text{ mol L}^{-1}$ sodium hydroxide.

- A. She decides to make a solution of toilet cleaner in a 500.0 mL volumetric flask that will give a titer of 20.00 mL when titrated against 20.00 mL of the sodium hydroxide solution. Assume that the claimed contents of the cleaner are correct, and calculate the mass of the cleaner she should take to make the 500.0 mL.
- B. By incredible good fortune she weighs out exactly the mass of cleaner she has calculated in part A. Using phenolphthalein as indicator, she then finds that a titre of 18.00 mL of the cleaner solution is needed for reaction with 20.00 mL of the sodium hydroxide solution. What is the actual concentration of the cleaner in g/kg of sodium hydrogensulfate.
5. Copper sulfate is dissolved in water, and treated with excess concentrated ammonia solution until the initial precipitate of copper hydroxide dissolves to yield a deep blue solution. When ethanol is added to the solution deep blue crystals precipitate. The isolated crystals smell of ammonia, and are presumably a complex salt with formula $\text{Cu}(\text{NH}_3)_x\text{SO}_4 \cdot y\text{H}_2\text{O}$.
- A. 1.0087 g of the complex salt is heated at 300°C to constant mass, and the white anhydrous copper(II) sulfate which remains weighs 0.652 g. Calculate the molecular mass of the complex salt.
- B. 1.4009 g of the complex salt is heated at 300°C and the ammonia so produced passed into water and made up to 250.0 mL in a volumetric flask. 20.00 mL portions of the ammonia solution are titrated with $0.09880 \text{ mol L}^{-1}$ hydrochloric acid using methyl orange as indicator. The average titre is 18.33 mL. Calculate the number of moles of ammonia in the 1.4009 g sample of the complex.
- C. Use your molecular mass from Part A and your answer to Part B to calculate the value of x in the formula $\text{Cu}(\text{NH}_3)_x\text{SO}_4 \cdot y\text{H}_2\text{O}$.
- D. Calculate the value of y, and write the full formula for the complex salt.

ANSWERS CALCULATIONS Set 21

- | | | | | | | | | |
|----|----|--|----|-------------|----|----------------------------------|----|--|
| 1. | A. | $\text{C}_3\text{H}_6\text{O}_2$ | B. | 73 g | C. | $\text{C}_3\text{H}_6\text{O}_2$ | D. | methylethanoate |
| 2. | A. | 1 mol As_2O_3 produces 2 mol AsO_3^{3-} which produces 2 mol H_3AsO_3 which reacts with 4 mol of Ce^{4+} | | | | | | |
| | B. | $[\text{Ce}^{4+}] = 0.1080 \text{ mol L}^{-1} = [\text{Ce}(\text{SO}_4)_2]$ | | | | | | |
| 3. | A. | 6.359 g | B. | 456 g/kg | | | | |
| 5. | A. | 246.9 g | B. | 0.02264 mol | C. | 4 | D. | $\text{Cu}(\text{NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}$ |

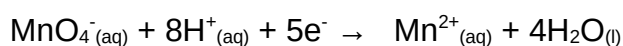
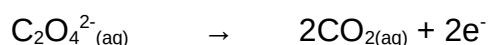
CALCULATIONS

Set 22

1. A pure substance 'A' is a white solid which sublimes at 76°C and melts under pressure at 158°C . 'A' burns readily in air leaving no ash, and qualitative analysis shows that nitrogen, sulfur, and the halogens are absent. It may thus be concluded that 'A' contains the elements carbon and hydrogen, and possibly oxygen.
 - A. In an experiment, 0.5991 g of 'A' is burnt in a current of dry air, and 1.334 g of carbon dioxide and 0.233 g of water are produced. Calculate the empirical formula of 'A'.
 - B. At 100°C and 1.00 atm pressure, 0.3301 of 'A' evaporates to occupy a volume of 72 mL. Calculate the molecular mass of 'A'.
 - C. What is the molecular formula of 'A'?
 - D. 'A' burns with a smoky flame, and thus it is probably aromatic. 'A' is only slightly soluble in water, but dissolves readily in sodium hydroxide solution. On the basis of this series of experiments there are three possible structural formulae for 'A'; write one of them.
3. Magnesium metal reacts with dilute hydrochloric acid to form hydrogen gas.
 - A. Write the equation for this reaction.
 - B. 0.1046 g of magnesium is added to 50.00 mL of 0.1000 mol L^{-1} hydrochloric acid. Which reactant is the limiting reactant?
 - C. State or calculate each of the following when the reaction has reached completion:
 - (i) the mass of magnesium remaining
 - (ii) the number of moles of hydrogen ion remaining
 - (iii) the volume of hydrogen gas produced, dried and measured at 25°C and 1.000 atm
 - (iv) the concentration of magnesium ion in the 50.0 mL of solution.

4. A technician is to determine the level of calcium in the bloodstream of an elderly hospital patient who is suffering from osteoporosis (a disease where insufficient calcium causes weakened bones). She adopts the following procedure.
- She takes a 15.00 mL sample of blood from the patient, and dilutes it to 100.00 mL.
 - She then takes a 10.00 mL sample of the diluted solution, and the calcium is precipitated as calcium ethandioate (calcium oxalate, CaC_2O_4).
 - The precipitate is washed and dissolved in excess dilute sulfuric acid, and then titrated at 60 °C with 0.0005200 mol L⁻¹ potassium permanganate. After several repetitions of the experiment, the average volume of potassium permanganate solution required to reach the end point is 9.88 mL.

A. The half equations for the titration reactions are as follows:

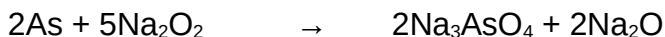


Write the equation for the titration reaction.

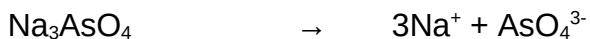
- B. Calculate the number of moles of calcium ethandioate (calcium oxalate) in the 10.00 mL.
- C. Calculate the mass of calcium present in the original 15.00 mL sample of the patient's blood.

5. You may analyse a sample for arsenic in the following way.

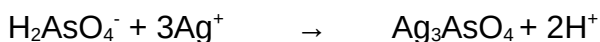
Heat the sample strongly with excess sodium peroxide so the mixture melts; the arsenic is converted to sodium arsenate:



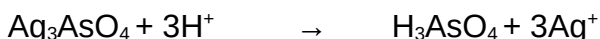
Allow the product to cool, and dissolve it in water; the sodium arsenate dissolves:



Then add dilute nitric acid, boil the solution, and adjust the pH to 8 with sodium hydroxide solution and acetic acid (ethanoic acid) solution; the arsenate ion is converted to dihydrogenarsenate ion:



Filter the silver arsenate, wash it, and dissolve it in dilute nitric acid to produce silver ion:



Titrate the silver ion with potassium thiocyanate solution, using iron(III) ion as indicator:



A 0.0603 g sample of commercial arsenic metal is treated as described, and ultimately when the processed sample is titrated, it requires 23.38 mL of 0.10034 mol L⁻¹ potassium thiocyanate solution.

- A. Use the above equations to work out how many moles of thiocyanate ion will react with the silver ion from 1 mole of arsenic.
- B. Use this value to calculate the percentage by mass of arsenic in the sample.

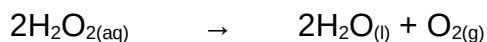
ANSWERS CALCULATIONS Set 22

1. A. C₇H₆O₃ B. 140 g C. C₇H₆O₃
 D. 1-hydroxybenzoic acid or 2-hydroxybenzoic acid or 3-hydroxybenzoic acid
3. A. Mg + 2H⁺ → H₂ + Mg²⁺ B. H⁺ is limiting reagent
4. C. (i) 0.04382 g (ii) No H⁺ remaining (iii) 0.0612 L (iv) 0.0500 mol L⁻¹
4. A. C₂O₄²⁻ (aq) + 2MnO₄⁻ (aq) + 16H⁺ (aq) → 10CO₂(aq) + 2Mn²⁺ (aq) + 8H₂O(l)
 B. 1.284 x 10⁻⁵ mol C. 5.15 x 10⁻³ g
5. A. As → Na₃AsO₄ → AsO₄³⁻ → H₂AsO₄⁻ → Ag₃AsO₄ → Ag⁺
 1 mol 1mol 1mol 1mol 1mol 3mol
 Ag⁺ reacts with SCN⁻ in a ratio of 3mol : 3mol or 1mol : 1 mol
 B. 97.2%

CALCULATIONS

Set 23

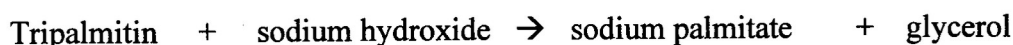
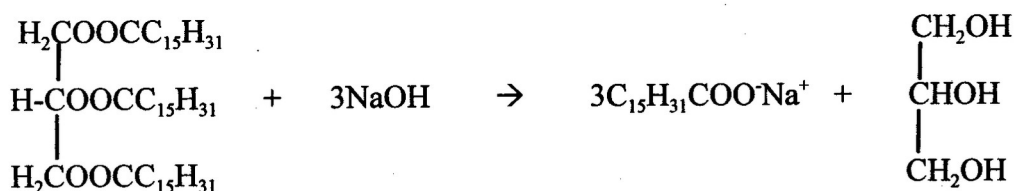
1. The crystalline mineral carnallite has the general formula $(\text{MgCl}_2)_x(\text{KCl})_y \cdot z\text{H}_2\text{O}$. When a sample of pure carnallite of mass 5.830 g is heated, all the water of crystallization is driven off and the remaining anhydrous powder weighs 3.561 g. All the magnesium in a further 10.270 g sample of the mineral is converted into “insoluble” magnesium hydroxide which required 35.1 mL of a 2.105 mol L^{-1} hydrochloric acid to dissolve it completely.
 - A. Determine the empirical formula of carnallite from the above information, i.e. find the values of x, y and z.
 - B. What total number of moles of ions are present in 1.00 L of solution which contains 5.830 g of dissolved carnallite.
3. An analytic chemist is required to check a bottle of commercial hydrogen peroxide solution which is labeled “ 10.0 mol L^{-1} ” to determine if the concentration has been reduced due to the decomposition reaction:



She takes 20.0 mL of the commercial hydrogen peroxide solution and dilutes it to 250 mL in a volumetric flask. 20.0 mL aliquots of the diluted solution are then added to separate conical flasks, acidified with a little concentrated sulfuric acid and titrated against a standardized potassium permanganate solution of concentration $0.2360 \text{ mol L}^{-1}$. The titrations yielded an average titre volume of 23.05 mL of the potassium permanganate solution required for complete reaction.

- A. determine the concentration of the hydrogen peroxide solution which the chemist took out of the bottle.
- B. Show by clearly indicating the relevant changes in oxidation numbers that the spontaneous decomposition of hydrogen peroxide solution is actually a self oxidation-reduction (i.e. disproportionation).

- Q4. After a laboratory acid-base titration, a student inadvertently poured an unwanted 50.00 mL sample of a standardized 1.00 mol L⁻¹ hydrochloric acid solution into a bottle containing exactly 400 mL of a standardized 0.500 mol L⁻¹ sulfuric acid solution.
- Determine the pH of the resulting solution after it is thoroughly mixed to ensure uniform concentration. Assume full ionization of both acids has occurred.
 - What volume of carbon dioxide gas at S.T.P. would this 450 mL of acid mixture produce if it was all added to 120.0 g of pure solid calcium carbonate.
- Q5. A group of Year 12 chemistry students decided to prepare some sodium palmitate soap (C₁₅H₃₁COO⁻Na⁺) in the laboratory by the reaction between pure tripalmitin oil and sodium hydroxide.



The students weighed out 42.0 g of the tripalmitin oil.

- Calculate
 - The mass of sodium hydroxide needed to react with the tripalmitin oil.
 - The number of moles of glycerol which are produced as a by-product of this reaction.
- The soap which was prepared was then dissolved in excess “hard” water [assumed to contain only Ca²⁺_(aq) ions and HCO₃⁻_(aq) ions]. This process produced a grey insoluble deposit of calcium palmitate.

What mass of calcium palmitate would be formed by the dissolution of 10.0 g of sodium palmitate soap in excess “hard” water?

ANSWERS CALCULATIONS Set 23

- | | | | | |
|----|----|---|-----------------|---|
| 1. | A. | MgCl ₂ ·KCl·6H ₂ O (i.e. x=1, y=1, z=6) | B. | 0.105 mol |
| 3. | A. | 8.50 mol L ⁻¹ | B. | 2H ₂ O ₂ (O = -1) → 2H ₂ O (O = -2) + O ₂ (O = 0) |
| 4. | A. | 0 as 10 ⁰ = 1 | B. | 5.04 L |
| 5. | A. | (i) 6.24 g | (ii) 0.0520 mol | B. 9.89 g |

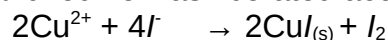
CALCULATIONS

Set 24

1. Copper(II) sulfate may be converted into copper(II) chloride by treatment of a copper(II) sulfate solution with the calculated quantity of barium chloride solution, removing the precipitate and evaporating the filtrate. A 2.136 g sample of copper(II) sulfate pentahydrate was converted into copper(II) chloride by this procedure.
 - A. Calculate the mass of barium chloride dehydrate that must be dissolved in water and added quantitatively to the copper(II) sulfate solution in order to bring about this conversion. Show your working.
 - B. Calculate the mass of the precipitate. Show your working.
 - C. The solution is then evaporated until the crystals just begin to form. The mixture is then allowed to cool and the blue-green crystals filtered off, washed with a little water and allowed to dry. In the experiment 0.785 g of pure copper(II) chloride dehydrate was obtained. Calculate the percentage yield. Show your working.
 - C. If the solution were evaporated to dryness you might get a 100% yield of copper(II) chloride dehydrate. Suggest one reason why this is not done in practice.
 - D. Another method of bringing about this conversion is to add the copper(II) sulfate solution to sodium carbonate solution. The precipitate is filtered, washed with water, and dissolved in dilute hydrochloric acid. When the resulting solution is filtered and evaporated, copper(II) chloride dehydrate is obtained. Is it necessary to weigh the sodium carbonate accurately to prepare this solution? Explain your answer.

- Q3. A copper sulfide ore was analysed by iodometry in the following way. The ore was finely crushed and a representative sample of 6.056 g was brought into solution with a mixture of concentrated hydrochloric acid and concentrated nitric acid. The excess hydrochloric acid and nitric acids were removed by heating with a little concentrated sulfuric acid. The remaining solution was quantitatively transferred into a 250.0 mL volumetric flask and the solution made up to the mark with water.

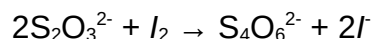
25.00 mL portions of the above solution were each treated as follows. Sodium acetate (ethanoate) was added to bring the pH to 4. Then an excess of potassium iodide solution was added, and iodine was liberated according to the following equation.



Each portion was titrated with $0.05103 \text{ mol L}^{-1}$ sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$) using starch indicator (which was added towards the end of the titration). The following figures were obtained.

Final reading (mL)	17	16.5	15.68	15.12	16.28
Initial reading (mL)	0.22	1.10	0.66	0.06	1.24

The equation for the reaction of iodine with sodium thiosulfate is as follows:



Starch is blue in the presence of I_2 . The end point is taken to be the loss of all blue colour.

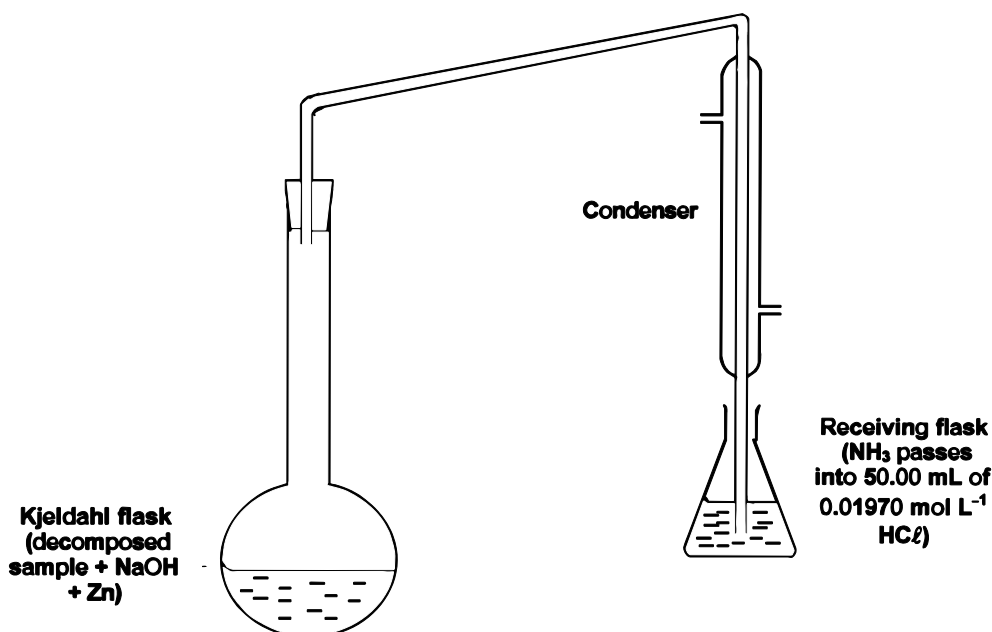
- A. (i) How many moles of iodine are liberated from one mole of copper(II) ion?
(ii) How many moles of thiosulfate ion are required for one mole of iodine?
(iii) How many moles of thiosulfate ion are required for one mole of copper(II) ion?
- B. What titration volume would you use in the calculation? Show your working.
- C. Calculate the percentage by mass of copper in the ore. Show your working.
- Q4. The Kjeldahl method is used to analyse for nitrogen in an organic substance. The substance is treated with concentrated H_2SO_4 , using anhydrous CuSO_4 as a catalyst; all nitrogen is converted into NH_4^+ ion.

The mixture is then treated with excess OH^- to convert the NH_4^+ ion into NH_3 . the NH_3 is boiled off and absorbed in an excess of dilute HCl .

In Kjeldahl determination 1.2540 g of dried pet food was heated for an hour with concentrated H_2SO_4 and anhydrous CuSO_4 (together with K_2SO_4 to raise the boiling point of the reaction mixture).

On cooling the reaction mixture, Zn pieces and an excess of concentrated NaOH solution were added and the flask quickly attached to a distillation apparatus as shown below. The mixture was gently boiled to drive the NH_3 into the receiving flask. (The Zn dissolves to give

$\text{An}(\text{OH})_4^{2-}$ and H_2 gas; the H_2 gas helps sweep out all the NH_3 .) The NH_3 was distilled into 50.00 mL of $0.01970 \text{ mol L}^{-1} \text{ HCl}$.



- Calculate the original number of moles of H^+ in the solution in the receiving flask before any NH_3 was absorbed. Show your working.
- After the NH_3 had been absorbed by the HCl solution, the excess HCl was titrated with $0.1000 \text{ mol L}^{-1} \text{ NaOH}$ (in the burette). Methyl orange was used as indicator; 5.62 mL of NaOH solution was needed for the colour change. Calculate the number of moles of H^+ in this solution after the absorption of the NH_3 . Show your working.
- Calculate the number of moles of NH_3 absorbed by the HCl solution, and hence the percentage by mass of nitrogen in the 1.2540 g of dried pet food. Show your working.
- During the decomposition of the pet food (with H_2SO_4 , CuSO_4 and K_2SO_4) there is no stopper on the flask. Why is nitrogen not lost?

Q5. The mineral azurite is azure blue in colour and has the formula $\text{Cu}_x(\text{CO}_3)_y(\text{OH})_z$.

When 2.0881 g of finely crushed pure azurite is heated at 600°C . 0.531 g of carbon dioxide and 0.109 g of water are produced; the black residue of copper(II) oxide has a mass of 1.444 g.

- Use these data to calculate the empirical formula of azurite. Show your working.
- Assuming azurite to be an ionic crystal, what species are in the crystal lattice? Show how these species are consistent with your formula.

ANSWERS

CALCULATIONS Set 24

1. A. 2.089 g (4sig figs) B. 1.996 g (4 sig figs)
 C. 1.458 g (theoretical yield) 53.8% (actual yield)
 D. Some water of crystallization may be lost or the product may be impure.
 E. No, it is not necessary to weigh the sodium carbonate accurately as long as the sodium carbonate is in excess, so that all the Cu^{2+} is precipitated as CuCO_3 .
3. A. (i) 0.5 mol iodine (ii) 2.0 mol thiosulfate (iii) 1.0 mol thiosulfate
 B. 15.04 mL; ignore readings 1 & 2 as they are discrepant (not consistent with the others)
 C. 8.053%
4. A. 9.85×10^{-4} mol B. 5.62×10^{-4} mol C. 0.473%
 D. Nitrogen is not lost because the acid keeps the N in as NH_4^+ ion or the nitrogen is combined in a compound.
5. A. $\text{Cu}_3(\text{CO}_3)_2(\text{OH})_2$
 B. $3 \times \text{Cu}^{2+} = +6$; $2 \times \text{CO}_3^{2-} = -4$ and $2 \times \text{OH}^- = -2$ giving a total -6 maintaining charge neutrality.

ANSWERS

CALCULATIONS Set 1

1. TiCl_3
2. $4.00 \times 10^4 \text{ g}$ or 40.0 kg
3. 697 mLs
4. 4.37 L
5. 20.4%

CALCULATIONS Set 2

1. A. CH_3O_3 B. $(\text{CH}_3\text{O}_3)_2$ C. $\text{C}_2\text{H}_6\text{O}_6 \cdot 2\text{H}_2\text{O}$
2. C. $\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ = \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O}$ D. 4.961×10^{-4}
E. 2.480×10^{-3} F. 1.240×10^{-3} G. 1.979 g H 96.5%
3. A. $\text{Mg} + 2\text{H}^+ = \text{Mg}^{2+} + \text{H}_2$ B. 98.97 kPa C. 1.056 L
D. 0.04714 mol E. 0.04714 mol F. 1.15 g
4. A. $Q = 5.00 \times 24 \times 60 \times 60 = 4.320 \times 10^5 \text{ C}$, $n(e^-) = 4.477 \text{ mol}$
B. 125 g C. 100 L
5. A. $P(\text{H}_2) = 150 \text{ kPa}$, $P(\text{O}_2) = 250 \text{ kPa}$
B. $n(\text{H}_2) = 0.5288 \text{ mol}$, $n(\text{O}_2) = 0.3526 \text{ mol}$ C. 0.5288 mol D. 9.53 g

CALCULATIONS Set 3

1. A. $0.0200 \text{ mol Na}_2\text{CO}_3$ $0.0100 \text{ mol Cu}(\text{NO}_3)_2$ Excess Na_2CO_3 ; $\text{Cu}(\text{NO}_3)_2$ limiting
B. 1.24 g C. 2.50 g
2. A. 1.889 tonnes for 100% recovery; 2.099 tonnes for 90% recovery;
 7.00 tonnes for 30% recovery B. $2.14 \times 10^5 \text{ s}$
3. $\text{C}_2\text{H}_6\text{SO}$
4. A. 1.32 g B. 921 kPa C. 138 kPa
5. C. $2\text{MnO}_4^- + 6\text{H}^+ + 5\text{H}_2\text{O}_2 = 2\text{Mn}^{2+} + \text{O}_2 + 8\text{H}_2\text{O}$
D. 5.203×10^{-4} E. 1.301×10^{-3}
F. $0.05204 \text{ mol L}^{-1}$ in diluted solution; $0.5204 \text{ mol L}^{-1}$ in original solution
G. 6 L

CALCULATIONS Set 4

1. A. $2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2$ B. 22.3 L C. 0.621 L
2. C. $\text{Cr}_2\text{O}_7^{2-} + 6\text{Fe}^{2+} + 14\text{H}^+ = 2\text{Cr}^{3+} + 6\text{Fe}^{3+} + 7\text{H}_2\text{O}$
D. 0.01029 mol E. 0.001715 mol F. 1.784 g G 21.2%
3. A. $\text{C}_3\text{H}_2\text{NO}_2$ B. $\text{C}_6\text{H}_4\text{N}_2\text{O}_4$
4. A. 65.4 a.m.u. B. 0.329 A
5. A. $n(\text{CH}_3\text{COOH}) = 0.01041 \text{ mol}$; $n(\text{OH}^-) = 0.00605 \text{ mol}$ B. $0.0484 \text{ mol L}^{-1}$

CALCULATIONS Set 5

1. A. 28.7% B. (i) $2.607 \times 10^{-3} \text{ mol}$ (ii) $1.304 \times 10^{-3} \text{ mol}$
(iii) $1.630 \times 10^{-3} \text{ mol}$ (iv) $M_r = 126$ C. $\text{C}_2\text{H}_2\text{O}_4$
D. $\text{C}_2\text{H}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$ E. 28.6%
2. A. $2\text{C}_8\text{H}_{18} + 25\text{O}_2 \rightarrow 16\text{CO}_2 + 18\text{H}_2\text{O}$ B. $2.46 \times 10^3 \text{ g}$ C. $1.20 \times 10^3 \text{ L}$
3. $\text{C}_8\text{H}_6\text{Cl}_2\text{O}_3$
4. A. 1.48 g B. 43.8 min
5. A. H_2SO_4 is limiting $m(\text{BaSO}_4) = 2.86 \text{ g}$ B. 0.382 mol L^{-1}

CALCULATIONS Set 6

1. A. 0.185 g B. (i) 0.0100 mol L⁻¹ (ii) 1.00 x 10⁻¹² mol L⁻¹ (iii) 12.0
2. A. C₇H₆O₂ B. C₇H₆O₂
3. A. 9.40 x 10⁻² mol L⁻¹ B. 4.50 x 10⁻² mol L⁻¹ C. 4.50 mol L⁻¹
4. A. 1.24 x 10⁷ g or 1.24 x 10⁴ kg B. 2.20 x 10⁶ L Note: if the NO_(g) in the third step is recycled then zero moles of NO is produced. This is an equally valid answer to part B.
5. A. 4.32 x 10⁶ C B. 501 L C. 4.48 mol L⁻¹

CALCULATIONS Set 7

1. A. 4.56 g B. 78.4 mL
2. A. MnO₄⁻ + 5Fe²⁺ + 8H⁺ = Mn²⁺ + 5Fe³⁺ + 4H₂O B. 0.001595 C. 0.007977
D. 0.01994 E. 99.4%
3. A. 0.2 mol L⁻¹ H₂SO₄; 0.05 mol L⁻¹ HCl; 0.8 mol L⁻¹ NaOH B. 0.15 mol L⁻¹ C. 13.1 mL
4. A. C₄H₁₀ + 13O₂ = 8CO₂ + 10H₂O B. 10 L C. 100 L D. 85.9%
5. A. 88.2
B. C₃H₆O C. X: 2-methylpropanol Y: 2-methylpropanoic acid Z: 2-methylpropylethanoate

CALCULATIONS Set 8

1. A. 2Al³⁺_(aq) + 3S²⁻_(aq) → Al₂S_{3(s)} B. Na₂S is limiting C. 2.56 g
D. 0.444g
2. A. 87.9 B. C₄H₈O₂ C. ethylethanoate or propylethanoate
3. A. Cr₂O₇²⁻/H⁺ B. Fe²⁺ → Fe³⁺ + e⁻ C. Cr₂O₇²⁻ + 14H⁺ + 6e⁻ = 2Cr³⁺ + 7H₂O
D. Cr₂O₇²⁻ + 6Fe²⁺ + 14H⁺ = 2Cr³⁺ + 6Fe³⁺ + 7H₂O E. 0.000284 mol
F. 0.01704 mol G. 95.2%
4. A. sulfite +4; sulfate +6 B. I₂ C. 0.000515 mol L⁻¹
D. Yes; because ppm (SO₂) 33.0 ppm > 30 ppm
5. B. To overcome the Activation Energy (E_a) C. non-toxic, odourless & cool enough to avoid burning the passengers D. Red/Ox E. 131 g

ANSWERS

CALCULATIONS Set 9

1. A. $6.76 \times 10^4 \text{ s}$ B. $1.37 \times 10^3 \text{ L}$
2. $\text{KCr}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$
3. A. 90.3% B. $1.74 \times 10^4 \text{ L}$
4. $n(\text{H}^+) = 0.01758 \text{ mol}$
 $2\text{MnO}_4^- + 5\text{H}_2\text{C}_2\text{O}_4 + 6\text{H}^+ = 2\text{Mn}^{2+} + 10\text{CO}_2 + 8\text{H}_2\text{O}$ 9.91 g H_2SO_4 ; 22.6 g $\text{H}_2\text{C}_2\text{O}_4$
5. A. $\text{Cu}^{2+} + 2\text{OH}^- \rightarrow \text{Cu}(\text{OH})_2 + 4\text{NH}_3 \rightarrow [\text{Cu}(\text{NH}_3)_4]^{2+} + 2\text{OH}^-$
 B. NH_3 inXS, Cu^{2+} is limiting; Ammonia inXS by 1.45 moles
 C. Concentration of both sulfate and complex ions is $1.16 \times 10^{-1} \text{ mol L}^{-1}$

CALCULATIONS Set 10

1. A. 6.00×10^{-3} B. 1.20×10^{-2} C. $6.00 \times 10^{-1} \text{ mol L}^{-1}$ D. $5.26 \times 10^2 \text{ g}$
2. A. 71.5 s B. 668 L
3. $\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ = \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O}$; 93.2%
4. A. 656 g B. $2.84 \times 10^3 \text{ g}$
5. A. $2.37 \times 10^{-1} \text{ g}$ B. $\text{C}_3\text{H}_9\text{N}$ C. Propanamine or isomers of propanamine

CALCULATIONS Set 11

1. 10.2%
2. A. $2\text{MnO}_4^- + 6\text{H}^+ + 5\text{H}_2\text{O}_2 = 2\text{Mn}^{2+} + \text{O}_2 + 8\text{H}_2\text{O}$ B. $0.0436 \text{ mol L}^{-1}$ C. 1.74 mol L^{-1}
 D. $2\text{H}_2\text{O}_2 = \text{O}_2 + 2\text{H}_2\text{O}$; 19.5
3. $7.24 \times 10^4 \text{ g}$ or 72.4 kg
4. A. $0.0760 \text{ mol L}^{-1}$ B. $n(e^-) = 0.001897 + 2y \times 0.025 = 1.90 \times 10^{-3} + 5.00 \times 10^{-2}y$
 C. 4.92×10^{-2} D. 4.13%
5. $1.76 \times 10^3 \text{ g}$ or 1.76 kg

CALCULATIONS Set 12

1. A. 75 mg L^{-1} B. 41.7 g
2. A. 1.04 g B. (i) $0.0950 \text{ mol L}^{-1}$ (ii) still 0.200 mol L^{-1} C. Although Pb^{2+} has been lost, NO_3^- has not been involved in the electrolysis and so its concentration is unchanged. [The solution remains neutral because H^+ is produced by $2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4e^-$ the final solution is 0.010 M with respect to H^+]
3. CHC/O
4. 256
5. A. 6.20 mL B. 0.00799 mol ; 129 C. Modify the experiment so that a titre closer to 20 mL might be obtained; e.g. use a more dilute solution of NaOH
 D. Mass of $\text{CHC/O} = 64.5$ hence $129/64.5 = 2$; therefore the molecular formula is $(\text{CHC/O})_2$
 Since 'A' is an acid it probably contains alkanoic acid $-\text{COOH}$ group. Since the gas volume experiment gives double the correct molecular mass, the molecules must exist as dimers where they are held together by hydrogen bonds between the two $-\text{COOH}$ groups.

CALCULATIONS Set 13

1. A. $\text{C}_2\text{H}_4\text{O}$ B. $\text{C}_4\text{H}_8\text{O}_2$ C. butanoic acid & 2-methylpropanoic acid
2. A. 22.4125 mL B. 0.002375 mol C. 0.0003958 mol D. 84.1%
3. A. 1.68 tonne B. 339 kL or m^3
4. 5.87 g L^{-1}
5. 63.53

CALCULATIONS Set 14

- A. $\text{SiO}_2 + 3\text{C} \rightarrow \text{SiC} + 2\text{COB}$. 1500 g or 1.50 kg C. network covalent
 D. The atoms in SiC are covalently bound in a lattice. These strong covalent bonds are hard to break and hence the atoms are not easily displaced (dislodged).
- $2.62 \times 10^{-3} \text{ M}$ or $0.00262 \text{ mol L}^{-1}$
- A. Point at which the indicator changes colour for a small amount of added titrant
 B. $A_r = 22.99$ hence element M = Na C. NaOH D. Fats/oils are triglycerides with ester linkages. NaOH hydrolyses this linkage and forms 1,2,3-propantriol and soap. The fats/oils are insoluble in water whereas the 1,2,3-propantriol and soap are.
- A. 95.5% B. Aluminium ions and iodide ions do not react together (either in an electron transfer reaction (Red/Ox) or to form a precipitate). All potassium salts are soluble in aqueous solution.
- $\text{C}_{15}\text{H}_{21}\text{O}_6\text{Fe}$

CALCULATIONS Set 15

- 98.0%
- A. 0.403 mol L^{-1} B. 0.297 mol L^{-1}
- A. 271 g B. 218 mL
- A. $\text{C}_2\text{H}_7\text{N}$ B. $\text{C}_2\text{H}_7\text{N}$ C. Ethanamine
- A. $3.34 \times 10^5 \text{ g}$ B. $8.08 \times 10^5 \text{ L}$ C. $1.84 \times 10^2 \text{ kg}$

CALCULATIONS Set 16

- 9.66%
- O_2 is the limiting reagent 9.98 L
- $m(\text{SiO}_2) = 750 \text{ g}$ $m(\text{Na}_2\text{CO}_3) = 257 \text{ g}$ $m(\text{CaCO}_3) = 178 \text{ g}$
- A. $\text{Cr}^{3+} + 4\text{H}_2\text{O} \rightarrow \text{CrO}_4^{2-} + 8\text{H}^+ + 3\text{e}^-$
 $\text{H}_2\text{O}_2 + 2\text{e}^- \rightarrow 2\text{OH}^-$
 $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$
 $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$
- B. $0.0198 \text{ mol L}^{-1}$
- A. $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$
- (i) $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}_{(\text{s})} \rightarrow \text{Co}^{2+}_{(\text{aq})} + 2\text{Cl}^{-}_{(\text{aq})} + 6\text{H}_2\text{O}_{(\text{l})}$
 (ii) $\text{Ag}^{+}_{(\text{aq})} + \text{Cl}^{-}_{(\text{aq})} \rightarrow \text{AgCl}_{(\text{s})}$

ANSWERS

CALCULATIONS Set 17

1. A. CH_2O B. $\text{C}_2\text{H}_4\text{O}_2$ C. Ethanoic acid or Methylmethanoate
2. A. 2.142×10^{-3} B. 1.34 mol L^{-1}
3. A. 2.78 B. 10.8
4. A. $1.30 \times 10^3 \text{ g}$ B. 107 kPa C. 0.130%
5. A. $2\text{MnO}_4^- + 5\text{H}_2\text{O}_2 + 6\text{H}^+ = 2\text{Mn}^{2+} + 5\text{O}_2 + 8\text{H}_2\text{O}$
B. 6.942×10^{-4} C. 1.736×10^{-3} in 20.0 mL D. 3.47 mol L^{-1} E. 38.9

CALCULATIONS Set 18

1. A. $2\text{MnO}_4^- + 5\text{H}_2\text{C}_2\text{O}_4 + 6\text{H}^+ = 2\text{Mn}^{2+} + 10\text{CO}_2 + 8\text{H}_2\text{O}$
B. 0.00255 C. 255 mg
2. A. C_4H_5 B. It is impossible to tie up all the bonds. C. C_8H_{10}
D. Ethyl benzene or 1.2- dimethyl benzene or 1.3- dimethyl benzene or 1.4- dimethyl benzene
3. A. 6.002 g B. 40.0%
4. A. OH^- left over = 1.98 mol
B. Add CH_3COOH until pH reaches 7 or
Add HCl until the pH reaches just 7 (or until blue colour appears in solution) and then add back base or
Add the calculated quantity on HCl or
Add CuSO_4 solution with stirring, watching until no more precipitate forms or
Add the calculated quantity of CuSO_4 in solution to precipitate the 1.98 mol of hydroxide.
5. A. 58.8 B. +2

CALCULATIONS Set 19

1. A. (i) $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$
(ii) $\text{Br}_2 + 2\text{e}^- \rightarrow 2\text{Br}^-$
B. 1.56 g C. 3.14 amp
2. A. 85.81% B. 9.779% C. $\text{Pb}_3\text{Cl}_2\text{O}_2$ D. $\text{Pb}^{2+}, \text{Cl}^-, \text{O}^{2-}$
3. $m(\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}) = 35.71$
Extra $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} = 14.3 \text{ g}$
4. A. $2\text{MnO}_4^- + 5\text{H}_2\text{O}_2 + 6\text{H}^+ = 2\text{Mn}^{2+} + 5\text{CO}_2 + 8\text{H}_2\text{O}$
B. 27.96 mL C. 0.898 mol L^{-1}
5. A. 10.34%
B. (i) 0.002753 mol (ii) 81.75% (iii) 73.30%
C. 16.4%

CALCULATIONS Set 20

1. $\text{C}_6\text{H}_8\text{O}_7$
2. A. 17.67 mL B. 0.02170 mol C. 63.9 D. 127.8
E. 191.7 F. 192.1
Any structural formula with molecular formula $\text{C}_6\text{H}_8\text{O}_7$ and 3 $-\text{COOH}$ groups. (Citric acid has 3 $-\text{COOH}$ groups and 1 $-\text{OH}$ group)
3. A. 1 mol As produces 1 mol H_3AsO_4 produces 1mol $\text{H}_3\text{AsO}_3 + \text{I}_2$ which reacts with 2 mol of $\text{S}_2\text{O}_3^{2-}$
E. 16.0%
4. A. $\text{C}_3\text{H}_8 + 5\text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$
B. 41.0 L C. 5.88 mL
5. A. 22710 s 51.79 g B. 6
C. Cr [It has atomic mass of 52.00, well within experimental error of 51.8 and has an oxidation state +3]
[The original grayish-blue salt then is $\text{CrCl}_3 \cdot 6\text{H}_2\text{O}$ or $\text{Cr}(\text{H}_2\text{O})_6\text{Cl}_3$]

CALCULATIONS Set 21

- A. $C_3H_6O_2$ B. 73 g C. $C_3H_6O_2$ D. methylmethanoate
- A. 1 mol As_2O_3 produces 2 mol AsO_3^{3-} which produces 2 mol H_3AsO_3 which reacts with 4 mol of Ce^{4+}

B. $[Ce^{4+}] = 0.1080 \text{ mol L}^{-1} = [Ce(SO_4)_2]$
- A. 6.359 g B. 456 g/kg
- A. (i) 114.3 g (ii) 21.40 L B. 77%
- A. 246.9 g B. 0.02264 mol C. 4 D. $Cu(NH_3)_4SO_4 \cdot H_2O$

CALCULATIONS Set 22

- A. $C_7H_6O_3$ B. 140 g C. 140 g

D. 1-hydroxybenzoic acid or 2-hydroxybenzoic acid or 3-hydroxybenzoic acid
- A. 0.211 g B. 198 g C. Gold
- A. $Mg + 2H^+ \rightarrow H_2 + Mg^{2+}$ B. H^+ is limiting reagent

C. (i) 0.04382 g (ii) No H^+ remaining (iii) 0.0612 L (iv) $0.0500 \text{ mol L}^{-1}$
- A. $C_2O_4^{2-} (aq) + 2MnO_4^- (aq) + 16H^+ (aq) \rightarrow 10CO_2(aq) + 2Mn^{2+}(aq) + 8H_2O(l)$

B. $1.284 \times 10^{-5} \text{ mol}$ C. $5.15 \times 10^{-3} \text{ g}$
- A. $As \rightarrow Na_3AsO_4 \rightarrow AsO_4^{3-} \rightarrow H_2AsO_4^- \rightarrow Ag_3AsO_4 \rightarrow Ag^+$

1 mol 1mol 1mol 1mol 1mol 3mol

Ag^+ reacts with SCN^- in a ratio of 3mol : 3mol or 1mol : 1 mol

B. 97.2%

CALCULATIONS Set 23

- A. $MgCl_2 \cdot KCl \cdot 6H_2O$ (i.e. $x=1, y=1, z=6$) B. 0.105 mol
- A. 93.5% B. $3.37 \times 10^5 \text{ L}$ C. $1.06 \times 10^5 \text{ s}$
- A. 8.50 mol L^{-1} B. $2H_2O_2 (O = -1) \rightarrow 2H_2O (O = -2) + O_2 (O = 0)$
- A. 0 as $10^0 = 1$ B. 5.04 L
- A. (i) 6.24 g (ii) 0.0520 mol B. 9.89 g

CALCULATIONS Set 24

- A. 2.089 g (4sig figs) B. 1.996 g (4 sig figs)

C. 1.458 g (theoretical yield) 53.8% (actual yield)

D. Some water of crystallization may be lost or the product may be impure.

E. No, it is not necessary to weigh the sodium carbonate accurately as long as the sodium carbonate is in excess, so that all the Cu^{2+} is precipitated as $CuCO_3$.
- A. +1.54 V B. $5.373 \times 10^{-3} \text{ mol}$ C. 0.733 g D. 24.1 cents

E. The conditions were not standard i.e. the concentrations of solutions were not 1.00 mol L^{-1} .
- A. (i) 0.5 mol iodine (ii) 2.0 mol thiosulfate (iii) 1.0 mol thiosulfate

B. 15.04 mL; ignore readings 1 & 2 as they are discrepant (not consistent with the others)

C. 8.053%
- A. $9.85 \times 10^{-4} \text{ mol}$ B. $5.62 \times 10^{-4} \text{ mol}$ C. 0.473%

D. Nitrogen is not lost because the acid keeps the N in as NH_4^+ ion or the nitrogen is combined in a compound.
- A. $Cu_3(CO_3)_2(OH)_2$

B. $3 \times Cu^{2+} = +6$; $2 \times CO_3^{2-} = -4$ and $2 \times OH^- = -2$ giving a total -6 maintaining charge neutrality.