1 Sulfamic acid is a white solid used by plumbers as a limescale remover.

(a) Sulfamic acid contains 14.42% by mass of nitrogen, 3.09% hydrogen and 33.06% sulfur. The remainder is oxygen.

(i) Calculate the empirical formula of sulfamic acid.

(3)

(ii) The molar mass of sulfamic acid is 97.1 g mol^{-1}. Use this information to deduce the molecular formula of sulfamic acid.

(1)

(b) A solution of sulfamic acid contains hydrogen ions. The hydrogen ions react with magnesium to produce hydrogen gas. In an experiment, a solution containing 5.5×10^{-3} moles of sulfamic acid was reacted with excess magnesium. The volume of hydrogen produced was 66 cm^3, measured at room temperature and pressure.

(i) Draw a labelled diagram of the apparatus you would use to carry out this experiment, showing how you would collect the hydrogen produced and measure its volume.

(2)

(ii) Calculate the number of moles of hydrogen, H_2, produced in this reaction.

[The molar volume of a gas is 24 $dm^3 mol^{-1}$ at room temperature and pressure]

(1)

(iii) Show that the data confirms that each mole of sulfamic acid produces one mole of hydrogen ions in solution.

(2)

(c) Plumbers use sulfamic acid powder for descaling large items such as boilers. Sulfamic acid acts as a descaler because the hydrogen ions react with carbonate ions in limescale.

(i) Write an ionic equation for the reaction of hydrogen ions with carbonate ions. State symbols are **not** required.

(1)

(ii) Suggest ONE reason why sulfamic acid is considered less hazardous than hydrochloric acid as a descaler.

(1)

(Total for Question = 11 marks)

2 The concentration of iodine in solution can be measured by titration with sodium thiosulfate solution.

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

(a) Name a suitable indicator which could be used for this titration.

(1)

(b) The amount of sulfur dioxide in the atmosphere can be measured by passing a known volume of air through iodine solution. Sulfur dioxide converts iodine to iodide ions.

$$SO_2(g) + I_2(aq) + 2H_2O(l) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2I^-(aq)$$

In an experiment, 100 m^3 of air were passed through 100 cm^3 of iodine, concentration 0.0100 mol dm^{-3}. The remaining iodine was titrated with sodium thiosulfate solution and reacted with 12.60 cm^3 of sodium thiosulfate, concentration 0.100 mol dm^{-3}.

(i) How many moles of iodine were present in the solution of the iodine at the start of the experiment?

(1)

(ii) How many moles of iodine remained in the solution at the end of the experiment?

(2)

(iii) Calculate the number of moles of iodine which **reacted** with the sulfur dioxide, and hence the number of moles of sulfur dioxide in 100 m^3 of air.

(2)

(iv) The European Commission recommend exposure to sulfur dioxide in air should be less than 350 micrograms (350×10^{-6} g) per cubic metre.

Calculate whether the sulfur dioxide in this sample of air was within this limit. One mole of sulfur dioxide has mass 64.1 g.

(2)

(c) Explain whether the changes below would or would not improve the experimental procedure for measuring the concentration of sulfur dioxide in air used in (b).

(i) The 100 cm^3 of iodine was divided into 25 cm^3 samples before titration.

(1)

(ii) The concentration of sodium thiosulfate used to titrate the iodine was changed from 0.100 mol dm^{-3} to 0.050 mol dm^{-3}.

(2)

(iii) 150 m^3 of air was passed through the iodine. The solutions used were of the same concentrations as in the original experiment.

(2)

(Total for Question = 13 marks)