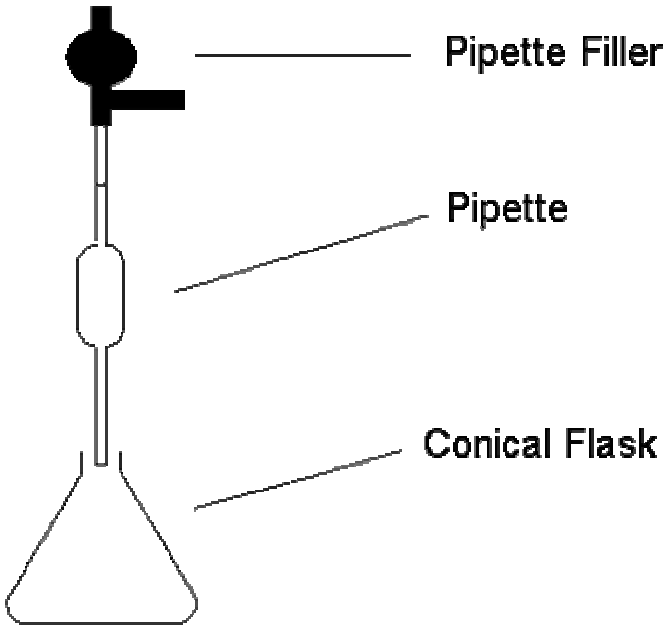
Leaving Certificate Chemistry - 1999

Question

A mass of 2.52g of ethanedioic acid crystals, $C_2O_4^{-2}.xH_2O$, was dissolved in deionised water and the solutions was made up accurately to 500 cm^3 in a volumetric flask. Some of the solutions was poured into a clean, dry beaker. A pipette, fitted with a pipette filler (see diagram), was used to transfer 25.0 cm^3 of the solution from the beaker to a conical flask and about 20 cm^3 of dilute sulphuric acid were added. A thermometer was placed in the conical flask, and the flask and contents were heated to about 70°C. The flask was removed from the heat and the thermometer was removed from the solutions. It was found by titration that the acidified solution of ethanedioic acid required 20.0 cm^3 of a 0.020 mol dm^{-3} solutions of potassium manganate(VII) for complete oxidation.



The equation for the reaction involved in the titration is

$$16H^+ + 2MnO_4^- + 5C_2O_4^{2-} \rightarrow 10CO_2 + 2Mn^{2+} + 8H_2O$$

i) In preparing the pipette for the titration it was rinsed with deionised water and then the water was removed by rinsing the pipette with the solution to be transferred. Explain why it was important to remove the water. Why was the conical flask *not* rinsed with the solution? (12)

ii) Pipette fillers are used for safety reasons. What property of ethanedioic acid makes the use of a pipette filler essential? (6)

iii) What problem was encountered when removing the thermometer from the solution in the conical flask? How could this problem have been overcome?(9)

iv) Why was it necessary to heat the contents for the conical flask immediately before carrying out the titration? Explain why it was not necessary to continue heating during the titration. (9)

v) What colour change was observed as the potassium manganate (VII) solution reacted with the acidified ethanedioic acid solution? How as the end-point of the titration identified? (12)

vi) Calculate the concentration of the ethanedioic acid solution in the volumetric flask (a) in g dm^{-3} , (b) in mol dm^{-3} . Hence find the percentage water in the ethanedioic acid crystals and the value of *x* in the formula. (18)

Solutions

i) If drops of water remain in pipette (or burette), this would dilute the solutions they are going to contain. However, if the conical flask is rinsed with the solution, drops of solution would remain which would increase the amount of the reactant.

ii) Pipette fillers are used for health safety reasons as they prevent intake of dangerous solutions through the mouth. Ethanedioic acid is, actually poisonous.

iii) Drops of the flask solution would be removed on the thermometer. If this happens calculations are inaccurate. The thermometer should be rinsed with deionised water before taking it out of the solution. (The deionised water will not alter the moles of reactant in the flask.)

iv) Heat increases the rate of a chemical reaction and also breaks bonds of the reactants. In this case the C to C bonds in $C_2O_4^{-2}$ must be broken. One of the products in this reaction is Mn^{+2} which will act as a catalyst for the reaction. The catalyst will speed up the reaction without more heat.

v) A purple to colourless change was noted. The end point occurs when one further drop won't go colourless.

vi) Parts a and b first:

a. Note that the flask solution was made up by mixing 2.52g of crystals per $500\,cm^3$ $[\frac{1}{2}\,Litre]$. This is the same concentration as $5.04 g.dm^{-3}$.

b. To change g dm^{-3} to moles dm^{-3} we divide by the molecular . But, we don't know the molecular weight as x is involved.

$$(1)\ \frac{V_1M_1}{2} = \frac{V_2M_2}{5} \Rightarrow \frac{20.(0.02)}{2} = \frac{25(M_2)}{5} \Rightarrow M_2 = 0.04$$

$$(2)\ 0.04 \text{ moles} = 5.04\text{g} \Rightarrow 1 \text{ mole} = \frac{5.04}{0.04} = 126.$$

(3) $C_2H_2O_4.xHO = 126 \Rightarrow 90 + 18x = 126 \Rightarrow x = 2$

%Water = $\frac{WtH_2O}{Total..Wt}.(100) = \frac{36}{126}.(100) = 28.6\%.$