

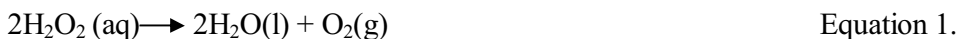
To determine the concentration of hydrogen peroxide, H₂O₂, in aqueous solution.

In this practical examination I am provided with a '100-volume' hydrogen peroxide solution. Taking the '100-volume' into consideration I have to plan two experiments that would enable me to determine the exact concentration of H₂O₂, in mol dm⁻³.

Titration:

The first experiment I shall carry out will be a redox titration; between hydrogen peroxide and potassium manganate (VII). A titration will enable the reacting volumes of H₂O₂ and KMnO₄ to be accurately determined. From this information and the stoichiometric ratio I will then be able to determine the concentration of H₂O₂.

Prior to going into detail we must first understand what a '100-volume' solution is. By definition this means that 1cm³ of H₂O₂ will decompose to produce 100cm³ of O₂ at STP. In simple terms it is just another way of indicating the strength of H₂O₂.



Consequently, we can understand that '100-volume' represents a very strong concentration of H₂O₂. Thus, before carrying out the titration the solution of H₂O₂ must be diluted.

A reasonable dilution factor for this experiment is 100; hence from a 100-volume to a 1-volume solution. As I know the strength of the original (C₁) and the new solution (C₂), I can use the below formula:

$$C_1V_1 = C_2V_2 \quad \text{Formula 1.}$$

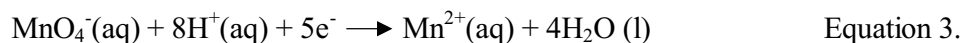
where V₁ is the original volume (dm³)
V₂ is final volume (dm³)

To make up a 1-volume H₂O₂, we would need to use a graduated pipette to transfer 0.01dm³ of the 100-volume H₂O₂ into a 1dm³ volumetric flask. As we know 3 of the 4 quantities in formula 1 we can work out V₂:

$$100(0.01) = 1(V_2) \\ \therefore V_2 = 1 \text{ dm}^3$$

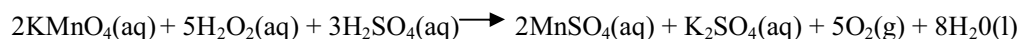
As the final volume is 1 dm³, the volumetric flask needs to be filled up to the top with distilled water.

The principle behind a redox titration is that if a solution contains a substance that can be oxidized, then the concentration of that substance can be analyzed by titrating it with a standard solution of a strong oxidizing agent. The reaction can be balanced by presuming that it occurs through two separate half-reaction.



Potassium manganate (VII) is very useful in redox titrations, due to its capability to be self indicating. However, for it to be able to function properly it must be in acid solution. Therefore, we must add excess dilute sulphuric acid to the solution.

If we combine both equation 2 and 3, we can get the full reaction between H_2O_2 and KMnO_4 . Note, that sulphuric acid is also part of the reaction.



The stoichiometric ratio provided from the above equation, along with experimental data obtained from the titration will be used to determine the concentration of H_2O_2 .

Method:

1. Measure out 25cm^3 of the diluted H_2O_2 in a graduated pipette and pour this into a conical flask. Record the volume in a data table.
2. Measure 10cm^3 of 2 molar sulphuric acid into a graduated cylinder and carefully add the acid to the conical flask.
3. Set up a clamp, boss and stand in order to fix the burette on to it.
4. Place a white tile on the bottom of the stand (underneath the conical flask) in order to make the colour change easier to recognise.
5. Fill up the burette with 0.1 molar KMnO_4 just above 0cm^3 , whilst having the tap closed. Record the molarity in a data table.
6. Open the stopcock on the burette to allow any air bubbles to escape from the tip. Close the stopcock when the liquid level in the burette is 0cm^3 . Record the initial volume, remembering to always read from bottom of meniscus.
7. Start by opening the tap of the burette slightly and gently shaking the conical flask as the KMnO_4 is poured in.
8. The KMnO_4 is a purple colour, however as it reacts with the H_2O_2 it becomes colourless. The endpoint of the titration is when all the H_2O_2 has reacted and a further drop of KMnO_4 causes the solution to remain pink/purple. This volume of KMnO_4 added should be recorded.
9. Repeat the titration until the results only differ by 0.5cm^3 .

The results obtained will be recorded in a table; below is a specimen:

	Titre 1	Titre 2	Titre 3
Molarity of KMnO_4	0.1	0.1	0.1
Initial volume of KMnO_4 solution (cm^3)	0	22.4	44.9
Final volume of KMnO_4 solution (cm^3)	22.4	44.9	67.3
Volume of KMnO_4 solution (cm^3) used	22.4	22.5	22.4
Mean volume of KMnO_4 solution (cm^3)	22.43	24.43	22.43

By using the data collected from the titration and the stoichiometric ratio, it is then possible to work out the concentration of H_2O_2 .

Gas Collection

In order to determine the concentration of H_2O_2 I will be carrying out two separate experiments; the titration explained above and a gas collection. H_2O_2 always decomposes into oxygen and water, as in equation 1.



However, the rate of decomposition of H_2O_2 is very slow without the use of a catalyst. As a result, I will be using manganese dioxide as a catalyst; this provides an alternative pathway with a lower activation energy. Thus, the rate of decomposition and production of O_2 is faster. The below arrangement will be used for the decomposition:

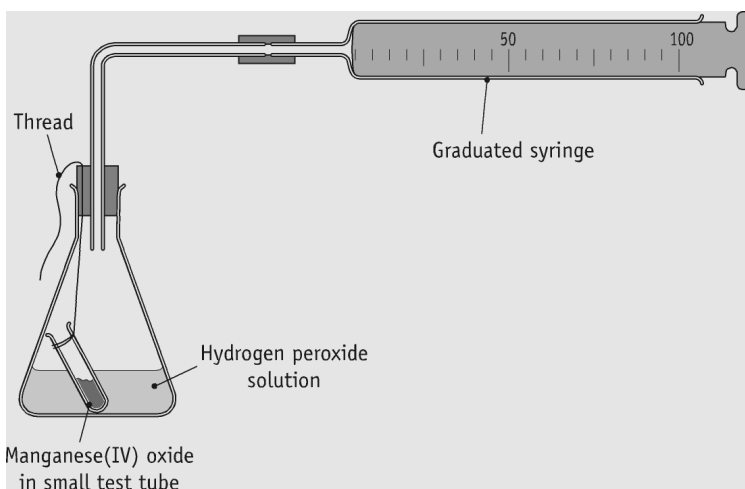


Figure 1.0

In this experiment we are provided with a 100-volume H_2O_2 . As explained earlier a 100-volume solution means that 1cm^3 of H_2O_2 decomposes to 100cm^3 of O_2 gas. Hence, 1cm^3 of H_2O_2 would fill up the graduated syringe to its maximum. As a result, I have chosen to use the 1-volume H_2O_2 which was also used in the titration.

As illustrated in figure 1.0, the manganese dioxide will be kept apart from the H_2O_2 until the start of reaction. The MnO_2 will be delivered using an inner test tube, which is turn is connected to a thread. Once the bung is attached to the conical flask the thread can be released and the reaction can proceed. The purpose of this is to increase the accuracy.

Method:

1. Measure 20cm^3 of 1-volume H_2O_2 using a measuring cylinder and add to a conical flask.
2. Weight 15g of MnO_2 powder using an electronic balance and add to a small test tube.

3. Attach a thread to the test tube; make sure it is secured tightly.
4. Place the test tube inside the conical flask using the attached thread; remember to keep part of the thread outside the conical flask.
5. Secure the bung, both to the conical flask and the gas syringe as illustrated in figure 1.0. Release the thread and measure the volume of gas collected. Repeat the experiment 3 times to get reliable results.
6. Record the volume of gas collected in both reactions in a data table:

	Trail 1	Trail 2	Trail 3
Volume of O ₂ gas collected (cm ³)	20.2	19.8	20.1
Mean volume of O ₂ gas collected (cm ³)	20.03	20.03	20.03

Below is a specimen calculation:

$$20.03 \text{ cm}^3 = 0.02003 \text{ dm}^3 \quad \text{number of moles of O}_2 = 0.02003 / 22.4 \\ = 0.000894 \quad 3\text{sf.}$$

Equation 1.0 states that H₂O₂ and O₂ are in stoichiometric ratio of, 2:1 respectively.

$$\therefore \text{ number of moles of H}_2\text{O}_2 = 0.000894 * 2 = 0.00179 \text{ 3sf.}$$

The volume of H₂O₂ used was 0.02dm³, therefore the concentration of H₂O₂ is:

$$\text{number of moles/ volume} = \text{concentration} \quad 0.00179 / 0.02 = 0.0894 \text{ mol dm}^{-3}$$

As the solution used was diluted by a factor of 100; from 100 to 1 volume, the actual concentration of 100-volume H₂O₂ is:

$$0.0894 \text{ mol dm}^{-3} * 100 = 8.94 \text{ mol dm}^{-3}$$
