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Iodometry lab report discussion

Iodine, known as iodine titration, is a method of volumetric chemical analysis, redox titration, where the appearance or disappearance of elemental iodine marks the endpoint. Note that iodine involves indirect titration of iodine liberated by reaction by analysis, and iodimetry is directly titration using iodine as a titrate. Redox titration using sodium thiosulphate, Na₂S₂O₃ (usually) as a reducing active substance is known as iodometric titration as it is used specifically to any iodine. Iodometric titration is a general method for determining the concentration of the oxidised agent in the solution. In iodometric titration, a starch solution is used as an indicator as it can absorb I₂, which is released. This absorption will cause the solution to change colour from deep blue to light yellow when titrated with a standardised solution of thiosulphate. This points to the end point of titration. Iodometry is usually used to analyse the concentration of oxidation in water samples, such as oxygen saturation in ecological studies or active chlorine when analyzing water in the pool. Color iodometric titration of the mixture before (left) and for (right) endpoint Basic principles Dilute solutions containing iod-starch complex. Using starch as an indicator can help create a sharper color change at the end point (Dark Blue – Colorless). The above color can be seen before reaching the end point. An excess but known amount of iodine is added to the known quantity of the sample, which is then oxidised to iodine by the oxidising agent. Iod is dissolved in a solution containing iodid to give triiodide ions, which have a dark brown color. The triiodide ion solution is then titrated against the standard solution of the thiosulphate to re-use with the starch indicator: I₃⁻ + 2 e⁻ + 3 I⁻ (E_o = + 0.5355 V) Together with reduced thiosulphate potential:[1] S₄O₆²⁻ + 2 e⁻ + 2 S₂O₃²⁻ (E_o = + 0.08 V) The general reaction is as is as this: I₃⁻ + 2 S₂O₃²⁻ → S₄O₆²⁻ + 3 I⁻ (E_{reduction} = + 0.4555 V) For simplicity, equations must be written u meaning aquatic molecular iodine , but not triiod ion, How iod ion u analysis of the ratio of mole does not tional reaction. The disappearance of deep blue is due to the degradation of iod-starch clatrare indicates the end point. The reduction used does not necessarily have to be a thiosulphate; stannous chloride, sulphides, sulphides, arsenic(III) and antimoinite(III) salts are commonly used alternatives. [2] at higher pH (> 8). At low pH, the following reaction may occur with the thiosulphate: S₂O₃²⁻ + 2 H⁺ → SO₂ + S + H₂O Some reactions involving certain reductases are reversible at a certain pH and therefore the pH of the sample solution should be carefully adjusted before the analysis is carried out. For example, reaction: H₃AsO₃ + I₂ + H₂O → H₃AsO₄ + 2 H⁺ + 2 I⁻ is reversible at pH < 4. Iodine volatility is also a source of error for this can be effectively prevented by ensuring that excess iodine and cooling of the titration of the mixture are present. Strong light, nitrite and copper ions catalyse the conversion of iodine into iodine, so they should be removed before adding iodine to the sample. In long-term titrations, it is recommended to add dry ice to the titration mixture to avoid air from the erlenmeyer flask to prevent air oxidation of iodine in iodine. The standard iodate solution is prepared from potassium iodate and potassium iodide, but this su both primary standards): IO₃⁻ + 8 I⁻ + 6 H⁺ → 3 I₃⁻ + 3 H₂O iodine u organic sodium, e.g. diethyl ether and uljitetrachloride, sodium thiosulphate dissolved [necessary explanations] Iod standard solution closed in the ampoule for iodometric analysis of iodoty applications iometry in many versions is extremely useful in volumetric analysis. Examples include the determination of copper(II), chlorate, hydrogen peroxide, i dissolved oxygen: 2 Cu²⁺ + 4 I⁻ → 2 CuI + I₂ 6 H⁺ + ClO₃⁻ + 6 I⁻ → 3 I₂ + Cl⁻ + 3 H₂O 2 H⁺ + H₂O₂ + 2 I⁻ → I₂ + 2 H₂O 2 H₂O + 4 Mn(OH)₂ + O₂ → 4 Mn(OH)₃ 2 Mn³⁺ + 2 I⁻ → I₂ + 2 Mn²⁺ Available chlorine refers to chlorine , liberated from the action of deflated acid on hypochlorite. Iodometry is often employed to determine the active amount of hypochlorite in bleach responsible for bleaching. In this method, an excess but known quantity of iodine is added to the known quantity of sample, at which only the active (electrophilic) can oxidize iodine to iodine. The iodine content and thus the active chlorine content can be determined by iometrics. [3] Determination of arsenic(V) compounds is the reverse standardisation of iodine solution with sodium arsenite. When added to the view, a known i excess amount of iodine is added: As₂O₅ + 4 H⁺ + 4 I⁻ - As₂O₃ + 2 I₂ + 2 H₂O For the analysis of the antimone(V) compound, a certain moist acid is added to melt the antimonium(III) product. [2] Determination of hydrogen sulphites and hydrogen sulphides reduce iod in acidic medium to iodide. When a known amount of sample is added to the diluted but superfluous amount of the standard iodine solution, Sulphuric acid i sulphites present quantitative iogu iod: SO₃²⁻ + I₂ + H₂O → SO₄²⁻ + 2 H⁺ + 2 I⁻ -HSO₃⁻ + I₂ + H₂O → SO₄²⁻ + 3 H⁺ + 2 I⁻ Determination of sulphide i hydrogen sulphide Although the sulphide u sample content may directly defer how it is described for sulphites, the results are often poor and inaccurate. Better, an alternative method is available sa higher accuracy, which involves the addition of excess, or known amounts of standard sodium arsenite solution in a sample in which arsenic tsulphide is precipitating; As₂O₃ + 3 H₂S → As₂S₃ + 3 H₂O Excess arsenide arsenide is then determined to titration u according to the standard iothe solution of starch. For best results, the sulphide solution should be With a sulphide concentration not exceeding 0,01 M.[2] Determination of hexacyanoferrate(II) When iodine is added to the hexacyanoferrate(II) solution, there is the following balance: 2 [Fe(CN)₆]³⁻ + 2 I⁻ + 2 [Fe(CN)₆]⁴⁻ + I₂ Under a strong acid solution, the upper balance lies far to the right side, or is revitalised in an almost neutral solution. This is difficult as the analysis of hexacyanoferrate (III) is difficult as iodid and mysulfate decay in a highly acidic medium. For the start of the reaction to completion, an excess amount of zinc salt may be added to the reaction mixture containing potassium ions, 2 [Fe(CN)₆]³⁻ + 2 I⁻ + 2 K⁺ + 2 Zn²⁺ → 2 K₂Zn[Fe(CN)₆] + I₂ Precipitation is caused to a little acid medium. Thus, the problem of degradation of iodide i thiosulfate u jaja kiseononn medium, a hexacyanoferrate(III) is customary iodometry iodometry. [2] References ^ Lide, David R., ed. (2006). CRC Manual of Chemistry and Physics (87th Hour). Boca Raton, FL: CRC Press. ISBN 0-8493-0487-3. i a b c d Mandham, J.; Denney, R.C.; Barnes, J.D.; Thomas, M. J. K. (2000). Vogel's Quantitative Chemical Analysis (6th ed.). New York: Prentice Hall, ISBN 0-582-22628-7 ^ Chlorine after ioty. National Environmental Methods Index. U.S. Geological Survey. [permanent dead connection] Retrieved from 1. Page 1 of 10 KWAME NKURAMAH UNIVERSITY OF SCIENCE AND TECHNOLOGY DEPARTMENT OF CHEMY YEAR TWO (CHEM 270) TITLE: IODOMETRY NAME: OPOKU ERNEST EMAIL: ernest.opoku@gasp.knust.edu.gh EXPERIMENT: I.2.2.3 DATE: 28TH JANUARY, 2014 2. PAGE 2 OF 10 TITLE: IODOMETRY OBJECTIVES AND OBJECTIVES: 1. To determine the amount of active ingredient, hypochlorite in two white households, by titration of iodine released by a reaction of hypochlorite with iodine. 2. To determine which bleach is best to buy. INTRODUCTION Iodine is a quantitative analysis of the oxidative agent solution by adding iodine, which reacts to the formation of iodine, which is then titrated. Iodometric titration is a method of volumetric analysis, redox titration, where the appearance or disappearance of elemental iodine marks the endpoint. Iodine ation involves indirect titration of iodine liberated by reaction by analysis, and iodography involves direct titration using iodine as a titrate. Iodine can be used as an oxidative agent in a number of anti-oxidation reduction titrations and iodine can be used as a reduction in other oxidation-reducing titration: I₂ + 2 e⁻ = 2 I⁻ (1) If the standard iodine solution is used as a titrant for oxidative analyte, the technique is iodetry. If excess iodine is used to quantify the chemical species while forming iodine, and if iodine is then titrate with a thiosulphate, it is a technique of iodine. Iodometry is an example of indirect previous titration reaction. The use of iodine as a titrant suffers from two major disadvantages. First, iod is not particularly soluble in water, secondly, iod is a bit volatile. Therefore, significant amounts of iodine solution are extracted from the solution. Both deficiencies are overcome by adding iodide (I⁻) iod (I₂) to solutions. In the presence of iodide iod reacts to the form of triiodide (I₃), which is very soluble and not volatile. I₂ + I⁻ = I₃ (2) The main chemical species present in these solutions is triiodide. Reducing triiodine to iodine is an analogous to reducing iodine. I₃ + 2 e⁻ = 3 I⁻ (3) Triiod reacts with thiosulphate to give iodid and tetraionate. 3. Page 3 of 10 I₂ + 2S₂O₃²⁻ + S₄O₆²⁻ (4) Diluted triiodide solutions are yellow, more concentrated solutions are brown and even more concentrated solutions are purple. Iodide solutions are colourless. If all other components of the solution are colourless, it is possible to detect the endpoint of titration involving triiodid without using the indicator. Detecting endpoints is much easier, but with an indicator. The indicator normally selected for titration involving iodide (triiodid) is starch. Starch forms a dark blue complex with iodine. The endpoint of iodometrics corresponds to a sudden change of color to blue. Also, the endpoint in iodometry corresponds to the sudden loss of blue due to the complex. Potato starch, rather than cornstarch, is preferred to produce an indicator solution, since the change in colour due to the starch of the complex is more sasteed at the end point. In iodimime, starch is added only after the colour has started to fade due to triiodide, i.e. close to the end point, as starch can be destroyed in the presence of excess triiodide. In the first part of this experiment, the sodium thiosulphate solution is prepared and standardised with the primary standard potassium iodate. Iodate (IO₃⁻) reacts with excess iodide in acid solution for triiodide yield, which is then titrated with a standardised solution of thiosulphate. Iodometric determination of hypochlorite in commercial bleach Product Most whites for liquid linen and in some solid bleaches the active ingredient is hypochlorite (OCl⁻). Liquid bleaching usually contains sodium hypochlorite, and bleaching is due to the strong oxidation properties of OCl⁻ which are also exploited in this iodometric determination of hypochlorite. In the experiment, iodid is used to reduce hypochlorite in bleach. The reaction in acid solution is administered by chloride and triiodide. OCl⁻ + 3 I⁻ + 2 H⁺ = Cl⁻ + I₃ + H₂O (5) Triiod, which is formulated in this reaction, is titrated with a standard solution of thiosulphate prepared earlier. 2 S₂O₃²⁻ + I₃⁻ = S₄O₆²⁻ + 3 I⁻ (6) Thiosulphate + i Tetrationate The endpoint of titration is located with a starch indicator solution that was prepared earlier. EQUIPMENT 1.0.1N Sodium thiosulphate 2. Potassium iodid 3.2M sulphuric acid 4. Page 4 of 10 4. 0.1M KMnO₄ 5. Starch indicator 6. White Magic Bleaching Solution (GHC3.50) 7. Janola bleaching solution (GHC 5.0) 8.10ml pipette 9.3 Stand-off flask 10. Electronic Balance 11. Burette 12. FLOW PROCEDURE 1. Saturation of sodium thiosulphate Observation experiment 2g KI + 10ml KMnO₄ Purple colored 2g KI + 10ml KMnO₄ + 5ml 2M H₂SO₄, followed by titration with sodium thiosulphate. Color is still. color changes are observed until there is a pale yellow color, the change of color has remained for a while. 1 ml of freshly prepared iodine was added to the pale yellow solution, which was formed by a blue black solution indicating the presence of iodine. The blue solution was titr versus a drop of sodium thiosulphate. The blue black color disappeared to form a colorless solution. The procedure was repeated for two more attempts. 2. Analysis of the bleaching solution. Janola Experiment Observation 1g KI was measured in a conical flask and then 20ml distilled water added to them. The re-prepared solution was added 1ml of bleach (janola) + 10ml 2M H₂SO₄. An orange solution freed from iod was then titrated against the solution of the thiosulphate. A colour change from orange to pale yellow 1ml starch solution was added to the pale yellow solution indicating the presence of iodine. The blue solution was titr versus a drop of sodium thiosulphate. Blue black color has disappeared to form a colorless solution the process is repeated for two more experiments. 5. Page 5 of 10 WHITE MAGIC Experiment Observation 1g KI was measured in a cone flask and then 20ml distilled water added k. 1ml of bleach solution (white magic) has been added to the reconstituted solution. The freed iod was titrated against the solution of thiosulfate. The color changes from orange to pale yellow. 1ml of starch solution was added to the pale yellow solution Blue black solution indicating the presence of iodine. Blue solution was titrated against a drop of sodium thiosulphate Blue black color has disappeared to form a colorless solution The process is repeated for two more experiments. TABLE OF RESULTS STANDARDISATION THIOSULFAT TITRATION SODIUM TITRATION 1 2 3 Final volume/ml 10.80 10.85 10.80 Initial volume/ml 0.00 0.00 0.00 Titrant/ml 10.80 10.85 10.80 Average titre = (10.80 + 10.85 +10.80) = 10.82ml 3 JANOLA TITRATION 1 2 3 Konačna comunion/ml 1.50 2.80 4.10 Initial volume/ml 0. 0.00 1.50 2.80 Titra value/ml 1.50 1.30 1.30 Average titre= (1.30 + 1.30)= 1.30ml 2 WHITE MAGIC Titration 1 2 3 Final volume/ml 6.4 5 6.65 6.89 Initial volume/ml 6.10 6.45 6.65 Titra value/ml 0.35 0.20 0.24 Average titre= (0.20 + 0.24) = 0.22ml 6. Page 6 of 10 2 CALCULATIONS AND EVALUATION OF DATA FROM THE REACTION EQUATION 2MnO₄⁻ + 16H⁺ + 10I⁻ → 2Mn²⁺ + 5I₂ + 8H₂O n(MnO₄⁻) = 2 n(I₂) n(I₂)= 5 × n(MnO₄⁻) concentration × volume n(MnO₄⁻)= 0.1mol/L × 0.01L = 0.001 mol(I₂) = 5 × 0.001mol= 2.5×10⁻³ mol 2 from equation I₂ + 2S₂O₄²⁻ + S₄O₆²⁻ n(S₂O₄²⁻) = 2 × 2×10⁻³ mol = 0.005mol from titration average titra = 10.82ml = 0.01082L concentration = mole [S₂O₄²⁻] = 0.005mol = 0.462M 0.01082L JANOLA From I₂ + 2S₂O₃²⁻ + S₄O₆²⁻ 6 7 . Page 7 of 10 n(S₂O₃²⁻) that reacted with the liberated I₂ = conc. × average titre volume n(S₂O₃²⁻) that reacted with the liberated I₂ = 0.462 M × 0.00130L = 6.006 × 10⁻⁴ mol but n(I₂) = ½ × n(S₂O₃²⁻) = ½ × 6.006 × 10⁻⁴ = 3.003 × 10⁻⁴ mol hence n(I₂) liberated= 3.003 × 10⁻⁴ mol from the equation ClO⁻ + 2I⁻ + 2H⁺ + Cl⁻ + I₂ + H₂O n(I₂) liberated = n(ClO⁻) hence n(ClO⁻) = 3.003 × 10⁻⁴ mol since the volume of bleach taken = 1ml = 0.001L then [ClO⁻]= (3.003×10⁻⁴) = 0.3003M 0.001L Mass(ClO⁻) = mole (ClO⁻) × molar mass of (ClO⁻) Molar mass(ClO⁻) = (35.5 + 16) = 51.5g/mol m (ClO⁻) in Janola= 3.003 × 10⁻⁴ × 51.5 = 0.0155g WHITE MAGIC n(S₂O₃²⁻) that reacted with the liberated I₂ = 0.462×0.22×10⁻³ = 1.0164 × 10⁻⁴ mol but n(I₂) = ½ × n(S₂O₃²⁻) = ½ × 1.0164×10⁻⁴ = 5.082 × 10⁻⁵ mol since n(I₂)= n(ClO⁻) , n(ClO⁻) = 5.082 × 10⁻⁵ mol because the volume of bleach is used = 1ml = 0.001M 8. Page 8 out of 10 thus [ClO⁻]= (5.082 × 10⁻⁵ = 0.05082M 0.001L mass (ClO⁻) = mole (ClO⁻) × molar mass (ClO⁻) molar mass (ClO⁻) = (ClO) 35.5 + 16) = 51.5g/mol m (ClO⁻) u belu magic = 5.082×10⁻⁵ × 51.5= 2.61723 × 10⁻³ g 1. Janola If the cost is 5.00 Ga a unit in cedis per gram = GHC 5.00 = GHC 322.58p/g 0.0155g = GHC 322.58p/g ii. White Magic If the cost of GHC is 3.50p A unit × cedis per gram = GHC 3.50p = GHC 1337.29p/g 2.61723 × 10⁻³ g = GHC 1337.29p/g Based on my results, Janola is best for buying between two. DISCUSSION From the experiment carried out above, the reaction between the measured gram of potassium iodide and 10 ml potassium permanganate, the purple colour of permanganate was decorated with iodine. 5ml of sulphuric acid has been added to the permanganate-iodide solution to prevent the degradation of ions in the solution and also creates a medium to continue the solution. 2MnO₄⁻ + 16H⁺ + 10I⁻ → 2Mn²⁺ + 5I₂ + 8H₂O The solution was then titrated into standard sodium thiosulphate and the titration continues with colour changes until a pale yellow colour is observed, indicating the presence of iodine in the solution. The solution was allowed for two minutes, this was made to see whether the pale yellow colour would stand, 1ml of starch solution was added to the light yellow solution and the colour was blue black, confirming the presence of iodine. A few drops of thiosulfate changes the solution colorless. The extent of the titration was recorded and two attempts were carried out. The average titre was determined for three experiments. 9. Page 9 of 10 Iodide content of two different bleaches (Janola and White Magic) was taken 1 ml of each sample the different conical flask, the measured amount of potassium iodide is dissolved in a specified volume of water. The measured bleach samples were transferred to the iodine solution giving the yellow solution. This was also titrated against sodium thiosulphate until a pale yellow solution was obtained. 1ml of starch solution was added to a pale yellow solution that turns blue black. A drop or two of the heosulfate turns the solution into colourless. Janola was observed to have higher titre values and white magic had less titre. I₂ + 2S₂O₃²⁻ + S₄O₆²⁻ 2. The ratio of mole permanganate and iodine together with their stoichiometric ratio was used to find mole iodine in the calculation using the concentration and volume of permanganate. If we are aware of a mole of iodine which is in a ratio of one to two (1:2) to a thiosulphate, the mole of the thiosulphate is twice the mole of iodine. Therefore, the concentration of thiosulphate was calculated using volume and whiting. With a mole of iodine that is in a relationship (1:1) with perchlorate? The mole and the concentration of hypochlorate were calculated and the masses of hypochlorates were calculated in both janola and white magic. These masses are used to determine the price of each of the bleach. ClO⁻ + 2I⁻ + 2H⁺ + Cl⁻ + I₂ + H₂O. CONCLUSION From the above experiment and calculation it can be concluded that the active components of these two bleach are hypochlorite. They found that the mass of hypochlorite in janola and white magic was 15.5mg and 2.86172mg. Their price tags were right because of the difference in quantity. However, it is recommended that Janola is the best to buy. PRECAUTIONS 1. Gases produced during the dissolution of the intererobic samples are toxic. Therefore, all reactions involving the development of gas were carried out in smoke. 2. The glasses were worn to prevent any solution from being in contact with the eye during boiling. 3. Gloves have been carried to ensure that no solution is in contact with the body. REFERENCES 1. Modern Inorganic Chemistry, Second Edition by William L. Jolly, 2010, pages 46 to 49. Page 10 out of 10 2. Journal of Solid State Chemistry, 2012, pages 4 to 7, retrieved from . 3. Concise inorganic chemistry, Fifth edition of J.D Lee, 2007, pages 202-204 and 951. 4. KNUST Chemistry Guide for the second year, CHEM 269 & 270, page 32-34. 34.