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Ammonia buffer solution preparation pdf

Species	Differentiation Properties
<i>E. coli</i>	Catalase positive, most strains are oxidase positive, motile at 22-26 °C, non-motile at 37 °C, urea positive, nitrate and indole positive, most strains are sucrose positive
<i>Pseudomonas</i>	Motile at 22-26 °C, non-motile at 37 °C, urea positive, diamine positive, oxidase positive, ornithine decarboxylase negative, indole negative
<i>V. parvii</i>	Produces two antiparallel components (antiparallel), catalase positive, non-motile at 22-26 °C and 37 °C, urea positive, indole negative, ornithine decarboxylase negative

NH₃

STABILIZATION OF CORRODED ALUMINIUM

Ian D. MacLeod

Abstract—A simple washing procedure has been developed for stabilization of corroding aluminium-copper alloys. The use of an ammonia-ammonium sulphate buffer at pH 9.6 in aerated deionized water effectively removes cemented copper metal and copper corrosion products from the surface of the object whilst also removing aggressive chloride ions. This treatment has been used to stabilize a Darulamin sea-plane float which was actively corroding after 46 years' exposure to a marine environment.

1 Introduction

At the time of writing, little or no information is readily available to the conservator faced with the problem of how to treat a corroding modern alloy such as aluminium. For transport and other museums whose collections include aeroplanes and automobiles this dearth of information can be of major concern. The problem will be exacerbated as contemporary vehicles, motors, etc., are added to the collections.

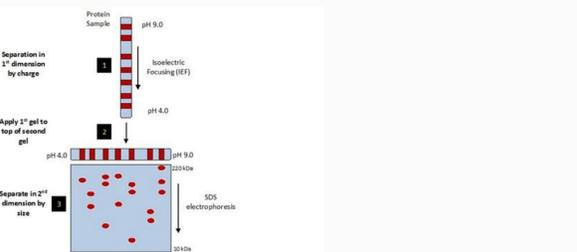
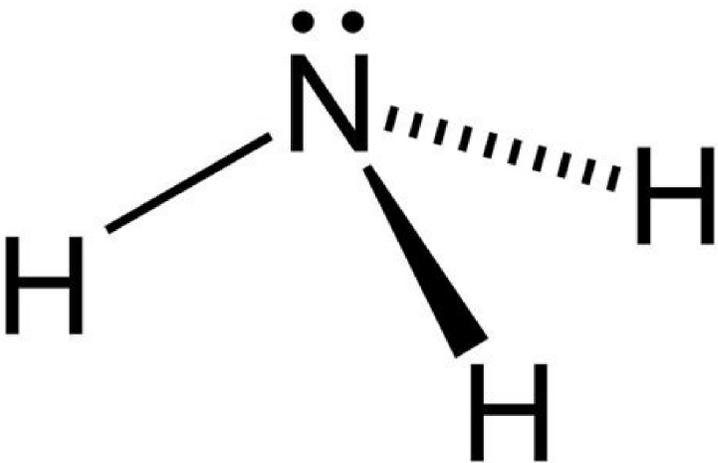
This paper describes the treatment used to stabilize a large corroded Darulamin sea-plane float from a Junkers W33 aircraft which was abandoned by H. Bettram, at Cape Bernier (14°00'S, 127°27'E) in 1932. The aircraft had been on a flight from Keeping (Timor) to Darwin when various circumstances caused it to land hundreds of miles off course in a most inhospitable part of tropical N.W. Australia [1]. The abandoned float had been used as a sailing vessel in an attempt by the crew to reach safety. The 'Atlantic' float was discovered in 1978 by members of a W.A. Museum expedition and sent to the conservation laboratories in Fremantle for treatment.

2 Description of float

When the float was uncrated at the conservation laboratories after a 3000km journey it was immediately obvious that it had suffered extensive pitting corrosion after being exposed for 46 years in a littoral location. There was evidence of continuing corrosion underneath the red-brown surface coating of pindan, a characteristic mixture of red clay soils of the Kimberley region which have been cemented together with large amounts of sand and some calcium minerals. Large areas were covered with the original tar-like paint and often the concreted sands, tar and metal corrosion products were the only materials that gave form to the float. The major corrosion products detected were aluminium hydroxide (Gibbsite, Al(OH)₃) and an unidentifiable light blue copper-containing material which had a distorted alumina structure. The major calcium mineral was gypsum, CaSO₄·2H₂O, and there were minor amounts of calcite, CaCO₃.



Figure 1 The Darulamin sea-plane float from the Junkers W33 as received in the conservation laboratories, showing extensive corrosion after 46 years' exposure to salt air.



Ammonia-ammonium chloride buffer solution preparation as per usp. Ammonia buffer solution preparation pdf. How to prepare ammonia buffer. How to prepare ammonium chloride buffer solution. Ammonia buffer solution preparation for hardness. How to make ammonia/ammonium buffer. How to make ammonia buffer. How to prepare ammonia solution.

$\text{pKa}(\text{NH}_3) = 9.25$ is sound. Working towards a solution for 0.10 molar buffers Edit 1/19/2017, noon So now the problem is that 0.1 molar buffers are needed for pH values of pH 10, 9.5, and 9.0. (I'll assume that pH 10 is 10.0 so 1 significant figure in concentration.) Even though we are going to work with pKa of $\text{pKa}(\text{NH}_4^+) = 9.25$, let's not forget that we are starting with "pure" ammonium hydroxide which ionizes as: $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$. In "pure" ammonium hydroxide not much of the ammonia will ionize to ammonium, so we can assume $[\text{OH}^-] = [\text{NH}_4^+]$ and the pH will be 11+. The contents of the flask to dilute the hydrochloric acid (0.1 mol/L) to 100 mL. Add 50 mL of potassium hydrophthalate (0.2 mol/L), Phosphate (4.8 - 8.0 pH) 8. Dissociation constants: Orthophosphoric acid: pK1=2.15; pK2=7.21; pK3=12.0 pH V, mL 4.8 4.9 5.0 5.1 5.2 5.3 5.4 5.5 5.6 0.35 0.60 0.95 1.35 1.80 2.3 3.0 3.9 4.9 pH V, mL 5.7 5.8 5.9 6.0 6.1 6.2 6.3 6.4 6.5 6.2 7.9 9.8 12.1 15.0 18.4 22.1 26.4 31.3 pH V, mL 6.6 6.7 6.8 6.9 7.0 7.1 7.2 7.3 7.4 37.1 43.0 49.2 55.2 61.2 67.0 72.6 77.7 81.8 pH V, mL 7.5 7.6 7.7 7.8 7.9 8.0 85.2 88.5 91.2 93.6 95.5 96.9 PHTHALIC (2.2 - 3.8 pH) Preparation method: Place some solution volume of HCl (0.1 mol/L) (see the table) into 200-mL volumetric flask. A 0.1 molar buffer means that given 1 liter of solution then: If 0.1 moles of a strong acid is added, the pH drops no more than 1 pH unit. Dissociation constants: o-Phthalic acid: pK1=2.93; pK2=5.41 m-Phthalic acid: pK1=3.70; pK2=4.60 p-Phthalic acid: pK1=3.54; pK2=4.46 pH V, mL 2.2 2.3 2.4 2.5 2.6 93.2 96.2 79.2 72.6 66 pH V, mL 2.7 2.8 2.9 3.0 3.1 59.4 53.0 46.8 40.8 35 pH V, mL 3.2 3.3 3.4 3.5 3.6 29.6 24.6 19.9 15.7 12.0 PHTHALIC (4.0 - 6.2 pH) Preparation method: Place some solution volume of NaOH (0.1 mol/L) (see the table) into 200-mL volumetric flask. So these buffers will guard against nearly the same amount of acid as base. at pH 9.5 there will be a bit more $[\text{NH}_3]$ than $[\text{NH}_4^+]$. You have some of the chemistry wrong. So: at pH 9.0 there will be a bit more $[\text{NH}_4^+]$ than $[\text{NH}_3]$. Using the given pKa, at pH = 9.25 then $[\text{NH}_3] = [\text{NH}_4^+]$. You could have ethanol solutions saturated with $[\text{HCl}]$ and $[\text{NH}_3]$ but that would be unusual. No problem for a computer, but it is painful by hand. Looking at your assumptions, concentrated ammonia is typically in an aqueous solution, as is concentrated hydrochloric acid. Ammonium (8 - 10 pH) 4. The volume of acetic acid is indicated in the table. Concentrated ammonia in aqueous solution is about 18 molar and is usually notated as concentrated ammonium hydroxide. ACETIC (3.8 - 6.3 pH) Preparation method: Mix the solution of acetic acid (1 mol/L) and 50 ml of solution of caustic soda (1 mol/L) in 500-mL volumetric flask. Now a buffer would be made "best" at the pKa or pKb value of a chemical so that the buffer capacity was the same for either a strong acid or a strong base. Phthalic (2.2 - 3.8 pH) 9. However if you can add additional water, then there is no unique solution and you'd need to know the buffer capacity for each of the buffers. Citric (5 - 6.6 pH) AMMONIUM (8 - 10 pH) Preparation method: Mix the solutions of ammonia (1 mol/L) and ammonium chloride (1 mol/L). So we can add various amounts of HCl and create the needed buffers. Citric (1.1 - 4.9 pH) 3. The contents of the flask to dilute the solution of KH2PO4 (0.0667 mol/L) to 100 mL. Borate (9.2 - 11 pH) 7. at pH 10.0 there will be a even more $[\text{NH}_3]$ than there was $[\text{NH}_4^+]$ at pH 9.5. The ratio of $[\text{NH}_3]$ to $[\text{NH}_4^+]$ can be calculated based on the pH using the Henderson-Halbach equation which for the ammonia/ammonium equilibrium is: $\text{pH} = 9.25 + \log(\text{NH}_3/\text{NH}_4^+)$ The problem as stated however seems incomplete. But in as concentrated solutions as concentrated as 18 molar the assumption is dicey and some correction would need to be made since the activity of the various species would be less than their actual concentrations. Since all three buffer solutions are in alkaline solution all the $[\text{HCl}]$ will be reacted according to the following reaction: $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4^+ + \text{Cl}^-$. The $[\text{Cl}^-]$ anion is a spectator anion and will not effect the pH. The contents of the flask to dilute the solution of sodium tetraborate (0.05 mol/L) to 100 mL. Dissociation constants: Orthoboric acid (H3BO3): pK1=9.15-9.24; pK2=12.74; pK3=13.80 Tetraboric acid (H2B4O7): pK1=3.74-4; pK2=7.70-9 pH V, mL 9.3 9.4 9.5 9.6 9.7 8.9 15.4 21 26.8 32.3 pH V, mL 9.8 9.9 10.0 10.1 10.2 36.3 39 41 42.7 44 pH V, mL 10.3 10.4 10.5 10.6 10.7 45.2 46.3 47.2 48.0 48.6 pH V, mL 10.8 10.9 11.0 49.1 49.5 49.9 PHOSPHATE (4.8 - 8.0 pH) Preparation method: Place some solution volume of Na2HPO4 (0.0667 mol/L) (see the table) into 100-mL volumetric flask. Dissociation constants: o-Phthalic acid: pK1=2.93; pK2=5.41 m-Phthalic acid: pK1=3.70; pK2=4.60 p-Phthalic acid: pK1=3.54; pK2=4.46 pH V, mL 4.0 4.1 4.2 4.3 4.4 4.5 0.8 4.1 7.4 11 15 19.3 pH V, mL 4.6 4.7 4.8 4.9 5.0 5.1 24.3 29.7 35.4 41.4 47.7 53.9 pH V, mL 5.2 5.3 5.4 5.5 5.6 5.7 59.9 65.7 70.9 75.6 79.7 83.1 pH V, mL 5.8 5.9 6.0 6.1 6.2 86.0 88.6 90.9 92.8 94.0 CITRIC (1.1 - 4.9 pH) Preparation method: Place some solution volume of sodium citrate (0.1 mol/L) (see the table) into 100-mL volumetric flask. The last assumption is a "gotcha" of sorts. The mix should be diluted with by distilled water to 500 ml. It seems really odd to have solutions as buffer solutions as strong as 18 molar. Flask contents to dilute with water to 200 mL. So it will take much less of a strong base like NaOH to change the pH to 11.0 than a strong acid like HCl to change the pH to 9.0. So assume that: The 0.1 molar buffer capacity means "at least 0.1 molar". I think the "exact" solution would have to be calculated iteratively. Acetic (3.8 - 6.3 pH) 5. The buffers at 9.0 and 9.5 are reasonably close to the pKa of ammonium which is 9.25. Dissociation constants: Citric acid: pK1=3.13; pK2=4.66-4.76; pK3=6.40 pH V, mL 5.0 5.1 5.2 5.3 5.4 3.6 9.7 14.9 19.6 23.7 pH V, mL 5.5 5.6 5.7 5.8 5.9 27.7 31.0 34.0 36.4 38.5 pH V, mL 6.0 6.1 6.2 6.3 6.4 40.4 42.0 43.4 44.6 45.5 pH V, mL 6.5 6.6 46.3 47.0 The contents of the flask to dilute the sodium citrate (0.1 mol/L) to 100 mL. If you really must make the buffer solution using only concentrated ammonium hydroxide and 3 molar HCl, then there is a unique solution for each buffer. I haven't worked through the solution, but I don't think there is a way to solve all the constraints directly in one pass. However there is a consideration. Borate (7.8 - 9.2 pH) 6. Phthalic (4.0 - 6.2 pH) 2. You're making 1 liter of each of the solutions You only need volumes to about 5% for the reagents. Dissociation constant: Acetic acid: pK1=4.76 pH V, mL 3.8 3.9 4.0 4.1 4.2 4.3 4.4 42.1 5 34.5 1 29.4 4 23.6 2 19.7 9 16.7 4 14.3 3 pH V, mL 4.5 4.6 4.7 4.8 4.9 5.0 5.1 12.4 1 108.9 96.8 87.2 79.5 73.4 68.6 pH V, mL 5.2 5.3 5.4 5.5 5.6 5.7 5.8 64.8 61 75.3 57.4 55.9 54.7 53.7 pH V, mL 5.9 6.0 6.1 6.2 6.3 53 52.3 51.9 51.5 51.2 BORATE (7.8 - 9.2 pH) Preparation method: Place some solution volume of Na2B4O7 (0.05 mol/L) (see the table) into 100-mL volumetric flask. 1. Dissociation constants: Orthoboric acid (H3BO3): pK1=9.15-9.24; pK2=12.74; pK3=13.80 Tetraboric acid (H2B4O7): pK1=3.74-4; pK2=7.70-9 pH V, mL 7.8 7.9 8.0 8.1 53.4 54.65 55.85 57.15 pH V, mL 8.2 8.3 8.4 8.5 58.65 60.7 62.95 65.25 pH V, mL 8.6 8.7 8.8 8.9 68 71.2 75.5 80.5 pH V, mL 9.0 9.1 9.2 85.6 91.9 98.1 BORATE (9.2 - 11 pH) Preparation method: Place some solution volume of NaOH (0.1 mol/L) (see the table) into 100-mL volumetric flask. However the buffer at pH 10.0 is a considerable distance from the pKa of ammonium. I'll explain further that the equilibrium equation $\text{pH} = \text{pKa} + \log\left(\frac{[\text{NH}_3]}{[\text{NH}_4^+]}\right)$ should really be written as the activities of the species not the concentrations. Dissociation constants: Citric acid: pK1=3.13; pK2=4.66-4.76; pK3=6.40 pH V, mL 1.1 1.2 1.3 1.4 1.5 1.6 1.7 1.8 1.9 2.0 2.1 4.8 11.1 15.9 19.3 22.2 24.6 26.5 28.2 29.5 30.6 31.7 pH V, mL 2.2 2.3 2.4 2.5 2.6 2.7 2.8 2.9 3.0 3.1 3.2 32.6 33.6 34.5 35.4 36.4 37.3 38.3 39.3 40.3 41.5 42.7 pH V, mL 3.3 3.4 3.5 3.6 3.7 3.8 3.9 4.0 4.1 4.2 4.3 44 45.4 46.8 48.4 50.1 51.9 53.8 56.0 58.5 61.1 64.3 pH V, mL 4.4 4.5 4.6 4.7 4.8 4.9 67.9 71.9 76.9 82.2 88.0 95.6 CITRIC (5 - 6.6 pH) Preparation method: Place some solution volume of NaOH (0.1 mol/L) (see the table) into 100-mL volumetric flask. So overall this is trying to avoid iterating using the quadratic equation. If 0.1 moles of a strong base is added the the pH increases no more than 1 pH unit. Below about 0.1 molar the activity and concentration are pretty equal.

Preparation of 1M Calcium chloride (CaCl₂) ... forms (CaCl₂·H₂O, CaCl₂·2H₂O, CaCl₂·4H₂O, CaCl₂·6H₂O). Any of them can be used to prepare 1 M calcium chloride stock solution. Anhydrous calcium chloride is solid white powder at room temperature. It is highly hygroscopic, therefore, ... Aqueous solution of ammonia (2) bacterial culture (1) ... What is the Hybridization of Ammonia? To understand the hybridization of ammonia we have to carefully examine the areas around Nitrogen. If we look at the atomic number of nitrogen it is 7 and if we consider its ground state it is given as 1s², 2s², 2p³. During the formation of ammonia, one 2s orbital and three 2p orbitals of nitrogen combine to form four hybrid orbitals having ... 07.06.2015 · To prepare a sample of aspirin Requirements Part 1 salicylic (2-hydroxybenzenecarboxylic) acid 100 cm³ conical flask 10 cm³ measuring cylinder ethanoic anhydride concentrated sulfuric acid in a dropping bottle 400 cm³ beaker tripod, gauze and Bunsen burner thermometer (-10 °C to 110 °C) 250

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