

Conductometric and volumetric study of copper sulphate

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INTRODUCTION

Significant data provided from the physicochemical properties of salts in aqueous solutions explain (solute + solute) and (solute + solvent) interactions that are crucial for the dependability of these systems and are used in some biochemical and physiological procedures in a living cell. The existence of numerous interactions makes the determination of the physical and chemical properties of organic solvents + inorganic salts solutions (partially solvated melts) over a wide concentration range very difficult. Strong ion-ion interactions are dominant in concentrated solutions. The influence of ion-dipole interactions between solvent molecules and solute ions increases with dilution. In extremely dilute solutions, a predominant weak dipole-dipole interaction is present. In diluted solutions within the range of validity of the Debye-Hückel law, most of the properties of the systems, such as conductivity, viscosity and volumetric properties, have been calculated. Because of their commercial uses in many technological areas, the ion solvation process and equilibrium involved in many concentration sections in water as well as non-aqueous media have been of great importance. The existence of ions, solvated ions, and free solvent depends on the concentration area. Hence, the investigation of volumetric properties over a wide electrolyte concentration range in different solvents is essential. The present work investigates the volumetric properties of copper sulphate in aqueous ethanol to explain the nature of solute-solvent interactions in a diluted and concentrated copper sulphate solution. Copper sulphate (CuSO_4) has been widely used for algal growth control and treatment of human copper deficiency. Ethanol, which is a colorless liquid at room temperature, is a polar solvent with a high relative permittivity of 24.4. This feature of ethanol allows the study of its volumetric properties in solutions with a high copper sulphate concentration. Ethanol is a commercial solvent that is primarily used in the pharmaceutical industry as well as in many organic syntheses. Because of the possibility of H-bond formation, ethanol solutions are often used as model systems to study the interaction of peptides and proteins. A wide range of dielectric constants (ϵ) and viscosities (η) and a high amount of hydrogen bonding at different temperatures are possible in aqueous ethanol solutions. Through determination of the conductivity, many interactions between water and ethanol can be studied by mixing them at various ratios over the mixing range. Suitable conductivity measurements offer a useful prediction of ion-solvent interactions, proton-anion and proton-solvent associations, and solvent structures. Conductive measurements began as early as the 18th century, when Andreas Baumgartner noticed that salt and mineral waters from Bad Gastein in Austria conducted electricity. As such, using conductometry to determine water purity, which is often used today to test the effectiveness of water purification systems, began in 1776. Friedrich Kohlrausch further developed conductometry in the 1860s when he applied alternating current to water, acids, and other solutions. It was also around this time when Willis Whitney, who was studying the interactions of sulfuric acid and chromium sulfate complexes, found the first conductometric endpoint. These findings culminated into potentiometric titrations and the first instrument for volumetric analysis by Robert Behrend in 1883 while titrating chloride and bromide with HgNO_3 . This development allowed for testing the solubility of salts and hydrogen ion concentration, as well as acid/base and redox titrations. Conductometry was further improved with the development of the glass electrode, which began in 1909. Acid-base titrations and redox titrations are often performed in which common indicators are used to locate the end point e.g., methyl orange, phenolphthalein for acid base titrations and starch solutions for iodometric type redox process. However, electrical conductance measurements can also be used as a tool to locate the end point.

Example: titration of an HCl solution with the strong base NaOH. As the titration progresses, the protons are neutralized to form water by the addition of NaOH. For each amount of NaOH added equivalent amount of hydrogen ions is removed. Effectively, the mobile H^+ cation is replaced by the less-mobile Na^+ ion, and the conductivity of the titrated solution as well as the measured conductance of the cell fall. This continues until the equivalence point is reached, at which one obtains a solution of sodium chloride, NaCl. If more base is added, an increase in conductivity or conductance is observed, since more ions Na^+ and OH^- are being added and the neutralization reaction no longer removes an appreciable amount of H^+ . Consequently, in the titration of a strong acid with a strong base, the conductance has a minimum at the equivalence point. This minimum can be used, instead of an indicator dye, to determine the endpoint of the titration. The conductometric titration curve is a plot of the measured conductance or conductivity values as a function of the volume of the NaOH solution added.