Diprotic acid example

17. Acid and Base Equilibrium Imagine a world where acids play a crucial role in everyday reactions. Diprotic acids are fascinating compounds that can donate two protons (H^+ ions) in solution, making them essential in various chemical processes. From biological systems to industrial applications, understanding these acids opens up a realm of possibilities.In this article, you'll explore some common examples of diprotic acids like sulfuric acid and carbonic acid. These substances not only demonstrate unique properties but also highlight their significance in both nature and technology. Have you ever wondered how these acids influence your favorite beverages or even the environment? Diprotic acids can donate two protons (H^+ ions) in aqueous solutions. This characteristic makes them significant in various chemical reactions. Here are some common examples:Sulfuric Acid (H_2SO_4): Commonly used in batteries and industrial processes. It dissociates into H^+ and HSO_4^-, then further into SO_4^{2-}.Carbonic Acid (H_2CO_3): Formed when carbon dioxide dissolves in water. It plays a crucial role in maintaining pH levels in blood and natural waters.Oxalic Acid ($C_2H_2O_4$): Found in many plants like spinach and rhubarb. It's known for its ability to chelate metal ions.Phosphoric Acid (H_3PO_4): Used widely in fertilizers and food flavoring. It dissociates into three stages, with the first two being diprotic.Each of these acids has unique properties that contribute to their roles in both nature and industry. Understanding how they function helps you appreciate their importance more fully.Diprotic acids exhibit unique properties due to their ability to donate two protons (H^+ ions) in solution. Understanding these properties is essential for grasping their behavior in various chemical contexts.Diprotic acids have two distinct acid dissociation constants, denoted as K_1 and K_2. These constants quantify the strength of each proton donation:K_1 represents the equilibrium constant for the first ionization step.K_2 indicates the equilibrium constant for the second ionization step.For example, sulfuric acid (H_2SO_4) has a very high K_1 value, indicating strong ionization in its first step. However, K_2 is significantly lower, showing that the second proton is less readily donated. This difference impacts how diprotic acids behave in different environments.The pH of a solution containing a diprotic acid reflects its concentration and degree of dissociation. Generally, a higher concentration leads to a lower pH, indicating stronger acidity. In aqueous solutions:A strong diprotic acid like sulfuric acid can yield a very low pH due to complete ionization.A weak diprotic acid like carbonic acid results in a higher pH since it only partially donates its protons.Knowing how these acids influence pH helps you understand their roles in biological systems and industrial applications. For instance, maintaining blood pH levels relies on carbonic acid's balance between HCO_3^- and CO_2.Diprotic acids play a vital role in various chemical processes. Understanding their examples helps grasp their significance.Sulfuric acid is one of the strongest diprotic acids. It donates two protons in its dissociation process, making it highly effective in industrial applications. You might find sulfuric acid used in:Batteries: Essential for lead-acid batteries.Fertilizers: Serves as a key component in producing phosphate fertilizers.Chemical manufacturing: Used to synthesize various chemicals.Its first dissociation constant (K_1) is very high, indicating strong ionization. The second dissociation constant (K_2), however, is lower and reflects less ionization.Citric acid is another notable diprotic acid found naturally in citrus fruits. It's widely recognized for its role as a food preservative and flavoring agent. In addition to culinary uses, citric acid serves functions such as:pH regulation: Maintains acidity levels in food products.Cleaning agents: Effective at removing mineral deposits due to its chelating properties.Biological processes: Plays a part in the Krebs cycle within living organisms.With K_1 and K_2 values that indicate moderate strength, citric acid showcases versatility across different applications while remaining safe for consumption.Diprotic acids play crucial roles in various applications across industries and biological systems. Their ability to donate two protons makes them valuable in numerous processes.Sulfuric acid is a prime example, commonly used in battery production, fertilizer manufacturing, and chemical synthesis. Its strong acidity enhances its effectiveness in these applications. Another notable diprotic acid is phosphoric acid, utilized extensively in food flavoring and as an essential ingredient in fertilizers due to its nutrient content. Additionally, citric acid, found naturally in citrus fruits, acts as a preservative and pH regulator in food products.Sulfuric Acid: Battery production, fertilizer manufacturing.Phosphoric Acid: Food flavoring, fertilizer formulation.Citric Acid: Food preservation, pH regulation.In biological contexts, diprotic acids like carbonic acid are vital for maintaining blood pH levels. It forms when carbon dioxide dissolves in water and plays a significant role in respiration and metabolic processes. Furthermore, oxalic acid, present in many plants such as spinach, contributes to calcium absorption but can also bind with minerals forming insoluble complexes.Carbonic Acid: Blood pH regulation.Oxalic Acid: Metal ion chelation; impacts mineral absorption.Understanding the diverse applications of diprotic acids highlights their importance not only industrially but also within living organisms. Sulfuric acid is an example of a diprotic acid.A diprotic acid is an acid that can donate two hydrogen ions (H+) or protons per molecule in an aqueous solution. Another name for a diprotic acid is a dibasic acid. A diprotic acid is a type of polyprotic acid, which is an acid able to donate more than one proton per molecule. In contrast, a monoprotic acid only donates one proton or hydrogen in water.Examples of diprotic acids include sulfuric acid (H2SO4), carbonic acid (H2CO3), chromic acid (H2CrO4), hydrogen sulfide (H2S), and oxalic acid (H2C2O4).Students generally assume a diprotic acid always loses both of its protons or hydrogen ions. However, this is not the case because the ease of losing the first and second protons usually is very different. The first acid dissociation constant Ka is always larger than the second one. In other words, it's always easier for a diprotic acid to lose its first proton than to lose its second proton.For example, sulfuric acid loses its first proton so easily (Ka > 1) that it acts as a strong acid, forming the hydrogen sulfate anion, HSO4–. H2SO4(aq) + H2O(l) → H3O+(aq) + HSO4–(aq) [Ka1 = 1 x 103]The Ka for losing the second proton is much lower, so only around 10% of sulfuric acid molecules in a 1M solution go on to form the sulfate anion (SO42-).HSO4–(aq) + H2O(l) ↔ H3O+(aq) + SO42-(aq) [Ka2 = 1.2 x 10-2]In practice, sulfuric acid fully deprotonates or loses both hydrogen atoms when it reacts with a base, like ammonia.Titration is used to calculate acid dissociation constants. For a monoprotic acid, the pH of the point of a titration curve halfway between the start of the curve and the equivalence point is used to find the pKa value using the Henderson-Hasselbalch equation:pH = pKa + log ([base]/[acid]pH = pKa + log(1)pH = pKaFor a diprotic acid, you can find the first acid dissociation constant the same way as for a monoprotic acid. The second acid dissociation constant is the point halfway between the first equivalence point and the second equivalence point.The titration curve of a diprotic acid, like oxalic acid, has two equivalence points. (JWSchmidt)Ebbing, Darrell; Gammon, Steven D. (January 1, 2016). General Chemistry. Cengage Learning. ISBN 9781305887299.Jameson, Reginald F. (1978). "Assignment of the proton-association constants for 3-(3,4-dihydroxyphenyl)alanine (L-dopa)". Journal of the Chemical Society. Dalton Transactions. 0 (1): 43–45. doi:10.1039/DT9780000043Petrucci R.H., Harwood, R.S.; Herring, F.G. (2002). General Chemistry (8th ed.). Prentice-Hall. ISBN 0-13-014329-4.Skoog, D.A; West, D.M.; Holler, J.F.; Crouch, S.R. (2004). Fundamentals of Analytical Chemistry (8th ed.). Thomson Brooks/Cole. ISBN 0-03-035523-0.Related Posts Sulfuric acid is an example of a diprotic acid.A diprotic acid is an acid that can donate two hydrogen ions (H+) or protons per molecule in an aqueous solution. Another name for a diprotic acid is a dibasic acid. A diprotic acid is a type of polyprotic acid, which is an acid able to donate more than one proton per molecule. In contrast, a monoprotic acid only donates one proton or hydrogen in water.Examples of diprotic acids include sulfuric acid (H2SO4), carbonic acid (H2CO3), chromic acid (H2CrO4), hydrogen sulfide (H2S), and oxalic acid (H2C2O4).Students generally assume a diprotic acid always loses both of its protons or hydrogen ions. However, this is not the case because the ease of losing the first and second protons usually is very different. The first acid dissociation constant Ka is always larger than the second one. In other words, it's always easier for a diprotic acid to lose its first proton than to lose its second proton.For example, sulfuric acid loses its first proton so easily (Ka > 1) that it acts as a strong acid, forming the hydrogen sulfate anion, HSO4–. H2SO4(aq) + H2O(l) → H3O+(aq) + HSO4–(aq) [Ka1 = 1 x 103]The Ka for losing the second proton is much lower, so only around 10% of sulfuric acid molecules in a 1M solution go on to form the sulfate anion (SO42-).HSO4–(aq) + H2O(l) ↔ H3O+(aq) + SO42-(aq) [Ka2 = 1.2 x 10-2]In practice, sulfuric acid fully deprotonates or loses both hydrogen atoms when it reacts with a base, like ammonia.Titration is used to calculate acid dissociation constants. 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Prentice-Hall. ISBN 0-13-014329-4.Skoog, D.A; West, D.M.; Holler, J.F.; Crouch, S.R. (2004). Fundamentals of Analytical Chemistry (8th ed.). Thomson Brooks/Cole. ISBN 0-03-035523-0.Related Posts The concept of proton donation is central to understanding the behavior of acids in chemistry. In simple terms, an acid is a substance that can donate hydrogen ions (H+), commonly referred to as protons, to another substance. Diprotic and triprotic acids, as the names imply, refer to acids that can donate two or three protons, respectively.This donation of protons occurs through a process called ionization. For diprotic acids, this means that they can undergo two distinct ionization steps, each releasing a proton into solution. Similarly, triprotic acids undergo three ionization steps. In aqueous solutions, the ease of proton donation affects the acid's strength and its pH, which is a measure of acidity. For example, in the ionization of sulfuric acid (a diprotic acid), the first proton is released more readily than the second, indicating a difference in ionization steps and the relative strengths of the resulting ions. Diprotic and Triprotic Acids and Bases Diprotic Acids Diprotic Bases Triprotic Acids Diprotic Acids The acid equilibrium problems discussed so far have focused on a family of compounds known as monoprotic acids. Each of these acids has a single H+ ion, or proton, it can donate when it acts as a Brnsted acid. Hydrochloric acid (HCl), acetic acid (CH3CO2H or HOAc), nitric acid (HNO3), and benzoic acid (C6H5CO2H) are all monoprotic acids. Several important acids can be classified as polyprotic acids, which can lose more than one H+ ion when they act as Brnsted acids. Diprotic acids, such as sulfuric acid (H2SO4), carbonic acid (H2CO3), hydrogen sulfide (H2S), chromic acid (H2CrO4), and oxalic acid (H2C2O4) have two acidic hydrogen atoms. Triprotic acids, such as phosphoric acid (H3PO4) and citric acid (C6H8O7), have three. There is usually a large difference in the ease with which these acids lose the first and second (or second and third) protons. When sulfuric acid is classified as a strong acid, students often assume that it loses both of its protons when it reacts with water. That isn't a legitimate assumption. Sulfuric acid is a strong acid because Ka for the loss of the first proton is much larger than 1. We therefore assume that essentially all the H2SO4 molecules in an aqueous solution lose the first proton to form the HSO4-, or hydrogen sulfate, ion. H2SO4(aq) + H2O(l) H3O+(aq) + HSO4-(aq) Ka1 = 1 x 103 But Ka for the loss of the second proton is only 10-2 and only 10% of the H2SO4 molecules in a 1 M solution lose a second proton. HSO4-(aq) + H2O(l) H3O+(aq) + SO42-(aq) Ka2 = 1.2 x 10-2 H2SO4 only loses both H+ ions when it reacts with a base, such as ammonia. The table below gives values of Ka for some common polyprotic acids. The large difference between the values of Ka for the sequential loss of protons by a polyprotic acid is important because it means we can assume that these acids dissociate one step at a timean assumption known as stepwise dissociation. Acid-Dissociation Equilibrium Constants for Common Polyprotic Acids Acid Ka1 Ka2 Ka3 sulfuric acid (H2SO4) 1.0 x 103 1.2 x 10-2 chromic acid (H2CrO4) 9.6 3.2 x 10-7 oxalic acid (H2C2O4) 5.4 x 10-2 5.4 x 10-5 sulfurous acid (H2SO3) 1.7 x 10-2 6.4 x 10-8 phosphoric acid (H3PO4) 7.1 x 10-3 6.3 x 10-8 4.2 x 10-13 glycine (C2H6NO2) 4.5 x 10-3 2.5 x 10-10 citric acid (C6H8O7) 7.5 x 10-4 1.7 x 10-5 4.0 x 10-7 carbonic acid (H2CO3) 4.5 x 10-7 4.7 x 10-11 hydrogen sulfide (H2S) 1.0 x 10-7 1.3 x 10-13 Let's look at the consequence of the assumption that polyprotic acids lose protons one step at a time by examining the chemistry of a saturated solution of H2S in water. Hydrogen sulfide is the foul-smelling gas that gives rotten eggs their unpleasant odor. It is an excellent source of the S2- ion, however, and is therefore commonly used in introductory chemistry laboratories. H2S is a weak acid that dissociates in steps. Some of the H2S molecules lose a proton in the first step to form the HS-, or hydrogen sulfide, ion. First step: H2S(aq) + H2O(l) H3O+(aq) + HS-(aq) A small fraction of the HS- ions formed in this reaction then go on to lose another H+ ion in a second step. Second step: HS-(aq) + H2O(l) H3O+(aq) + S2-(aq) Since there are two steps in this reaction, we can write two equilibrium constant expressions. Although each of these equations contains three terms, there are only four unknowns[H3O+], [H2S], [HS-], and [S2-] because the [H3O+] and [HS-] terms appear in both equations. The [H3O+] term represents the total H3O+ ion concentration from both steps and therefore must have the same value in both equations. Similarly, the [HS-] term, which represents the balance between the HS- ions formed in the first step and the HS- ions consumed in the second step, must have the same value for both equations. Four equations are needed to solve for four unknowns. We already have two equations: the Ka1 and Ka2 expressions. We are going to have to find either two more equations or a pair of assumptions that can generate two equations. We can base one assumption on the fact that the value of Ka1 for this acid is almost a million times larger than the value of Ka2. Ka1 >> Ka2 This means that only a small fraction of the HS- ions formed in the first step go on to dissociate in the second step. If this is true, most of the H3O+ ions in this solution come from the dissociation of H2S, and most of the HS- ions formed in this reaction PSS remain in solution. As a result, we can assume that the H3O+ and HS- ion concentrations are more or less equal. First assumption: [H3O+][HS-] We need one more equation, and therefore one more assumption. Note that H2S is a weak acid (Ka1 = 1.0 x 10-7, Ka2 = 1.3 x 10-13). Thus, we can assume that most of the H2S that dissolves in water will still be present when the solution reaches equilibrium. In other words, we can assume that the equilibrium concentration of H2S is approximately equal to the initial concentration. Second assumption: [H2S] CH2S We now have four equations in four unknowns. [H3O+] [HS-] [H2S] CH2S Since there is always a unique solution to four equations in four unknowns, we are now ready to calculate the H3O+, H2S, HS-, and S2- concentrations at equilibrium in a saturated solution of H2S in water. All we need to know is that a saturated solution of H2S in water has an initial concentration of about 0.10 M. Because Ka1 is so much larger than Ka2 for this acid, we can work with the equilibrium expression for the first step without worrying about the second step for the moment. We therefore start with the expression for Ka1 for this acid. We then invoke one of our assumptions. [H2S] CH2S 0.10 M Substituting this approximation into the Ka1 expression gives the following equation. We then invoke the other assumption. [H3O+] [HS-]C Substituting this approximation into the Ka1 expression gives the following result. We now solve this approximate equation for C. C 1.0 x 10-4 If our two assumptions are valid, we are three-fourths of the way to our goal. We know the H2S, H3O+, and HS- concentrations. [H2S] 0.10 M [H3O+] [HS-] 1.0 x 10-4 M Having extracted the values of three unknowns from the first equilibrium expression, we turn to the second equilibrium expression. Substituting the known values of the H3O+ and HS- ion concentrations into this expression gives the following equation. Because the equilibrium concentrations of the H3O+ and HS- ions are more or less the same, the S2- ion concentration at equilibrium is approximately equal to the value of Ka2 for this acid. [S2-] 1.3 x 10-13 M It is now time to check our assumptions. Is the dissociation of H2S small compared with the initial concentration? Yes. The HS- and H3O+ ion concentrations obtained from this calculation are 1.0 x 10-4 M, which is 0.1% of the initial concentration of H2S. The following assumption is therefore valid. [H2S] CH2S 0.10 M Is the difference between the S2- and HS- ion concentrations large enough to allow us to assume that essentially all of the H3O+ ions at equilibrium are formed in the first step and that essentially all of the HS- ions formed in this step remain in solution? Yes. The S2- ion concentration obtained from this calculation is 109 times smaller than the HS- ion concentration. Thus, our other assumption is also valid. [H3O+] [HS-] We can therefore summarize the concentrations of the various components of this equilibrium as follows. [H2S] 0.10 M [H3O+] [HS-] 1.0 x 10-4 M [S2-] 1.3 x 10-13 M Diprotic Bases The techniques we have used with diprotic acids can be extended to diprotic bases. The only challenge is calculating the values of Kb for the base. Example: Let's calculate the H2CO3, HCO3-, CO32-, and OH- concentrations at equilibrium in a solution that is initially 0.10 M in Na2CO3. (H2CO3: Ka1 = 4.5 x 10-7; Ka2 = 4.7 x 10-11) Because it is a salt, sodium carbonate dissociates into its ions when it dissolves in water. H2O Na2CO3(aq) 2 Na+(aq) + CO32-(aq) The carbonate ion then acts as a base toward water, picking up a pair of protons (one at a time) to form the bicarbonate ion, HCO3- ion, and then eventually carbonic acid, H2CO3. CO32-(aq) + H2O(l) HCO3-(aq) + OH-(aq) Kb1 = ? HCO3-(aq) + H2O(l) H2CO3(aq) + OH-(aq) Kb2 = ? The first step in solving this problem involves determining the values of Kb1 and Kb2 for the carbonate ion. We start by comparing the Kb expressions for the carbonate ion with the Ka expressions for carbonic acid. The expressions for Kb1 and Ka2 have something in commonthey both depend on the concentrations of the HCO3- and CO32- ions. The expressions for Kb2 and Ka1 also have something in commonthey both depend on the HCO3- and H2CO3 concentrations. We can therefore calculate Kb1 from Ka2 and Kb2 from Ka1. We start by multiplying the top and bottom of the Ka1 expression by the OH- ion concentration to introduce the [OH-] term. We then group terms in this equation as follows. The first term in this equation is the inverse of the Kb2 expression, and the second term is the Kw expression. Rearranging this equation gives the following result. Ka1Kb2 = Kw Similarly, we can multiply the top and bottom of the Ka2 expression by the OH-ion concentration. Collecting terms gives the following equation. The first term in this equation is the inverse of Kb1, and the second term is Kw. This equation can therefore be rearranged as follows. Ka2Kb1 = Kw We can now calculate the values of Kb1 and Kb2 for the carbonate ion. We are finally ready to do the calculations. We start with the Kb1 expression because the CO32- ion is the strongest base in this solution and therefore the best source of the OH- ion. The difference between Kb1 and Kb2 for the carbonate ion is large enough to suggest that most of the OH- ions come from this step and most of the HCO3- ions formed in this reaction remain in solution. [OH-] [HCO3-] C The value of Kb1 is small enough to assume that C is small compared with the initial concentration of the carbonate ion. If this is true, the concentration of the CO32- ion at equilibrium will be roughly equal to the initial concentration of Na2CO3. [CO32-] CNa2CO3 Substituting this information into the Kb1 expression gives the following result. This approximate equation can now be solved for C. C 0.0046 M We then use this value of C to calculate the equilibrium concentrations of the OH-, HCO3-, and CO32- ions. [CO32-] = 0.10 - C 0.095 M [OH-] [HCO3-] C 0.0046 M We now turn to the Kb2 expression. Substituting what we know about the OH- and HCO3- ion concentrations into this equation gives the following result. According to this equation, the H2CO3 concentration at equilibrium is approximately equal to Kb2 for the carbonate ion. [H2CO3] 2.2 x 10-8 M Summarizing the results of our calculations allows us to test the assumptions made generating these results. [CO32-] 0.095 M [OH-] [HCO3-] 4.6 x 10-3 M [H2CO3] 2.2 x 10-8 M All of our assumptions are valid. The extent of the reaction between the CO32- ion and water to give the HCO3- ion is less than 5% of the initial concentration of Na2CO3. Furthermore, most of the OH- ion comes from the first step, and most of the HCO3- ion formed in this step remains in solution. Triprotic Acids Our techniques for working diprotic acid or diprotic base equilibrium problems can be applied to triprotic acids and bases as well. To illustrate this, let's calculate the H3O+, H3PO4, H2PO4-, HPO42-, and PO43- concentrations at equilibrium in a 0.10 M H3PO4 solution, for which Ka1 = 7.1 x 10-3, Ka2 = 6.3 x 10-8, and Ka3 = 4.2 x 10-13. Let's assume that this acid dissociates by steps and analyze the first stepthe most extensive reaction. We now assume that the difference between Ka1 and Ka2 is large enough that most of the H3O+ ions come from this first step and most of the H2PO4- ions formed in this step remain in solution. [H3O+] [H2PO4-]C Substituting this assumption into the Ka1 expression gives the following equation. The assumption that is small compared with the initial concentration of the acid fails in this problem. But we don't really need this assumption because we can use the quadratic formula or successive approximations to solve the equation. Either way, we obtain the same answer. C = 0.023 M We can then use this value of C to obtain the following information. [H3PO4] 0.10 - C 0.077 M [H3O+] [H2PO4-] 0.023 M We now turn to the second strongest acid in this solution. Substituting what we know about the H3O+ and H2PO4- ion concentrations into this expression gives the following equation. If our assumptions so far are correct, the HPO42- ion concentration in this solution is equal to Ka2. [HPO42-] 6.3 x 10-8 We have only one more equation, the equilibrium expression for the weakest acid in the solution. Substituting what we know about the concentrations of the H3O+ and HPO42- ions into this expression gives the following equation. This equation can be solved for the phosphate ion concentration at equilibrium. [PO43-]1.2 x 10-18 M Summarizing the results of the calculations helps us check the assumptions made along the way. [H3PO4] 0.077 M [H3O+] [H2PO4-] 0.023 M [HPO42-] 6.3 x 10-8 M [PO43-] 1.2 x 10-18 M The only approximation used in working this problem was the assumption that the acid dissociates one step at a time. Is the difference between the concentrations of the H2PO4- and HPO42- ions large enough to justify the assumption that essentially all of the H3O+ ions come from the first step? Yes. Is it large enough to justify the assumption that essentially all of the H2PO4- formed in the first step remains in solution? Yes. You may never encounter an example of a polyprotic acid for which the difference between successive values of Ka are too small to allow us to assume stepwise dissociation. This assumption works even when we might expect it to fail.