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## Strong acids and bases pdf

(Back to 5 Anion page) TABLE OF STRONG ACIDS Completely ionized in Water to one (or more) Protons per acid Molecule HI H+ (aq) + I-(aq) HBr H+ (aq) + Br-(aq) HClO4 H+ (aq) + ClO4-(aq) HCl H+ (aq) + Cl-(aq) HClO3 H+ (aq) + ClO3-(aq) H2SO4 H+ (aq) + HSO4-(aq) (HSO4- is a weak acid that contributes additional protons) HNO3 H+ (aq) + NO3-(aq) TABLE OF STRONG BASES Completely ionized in Water to give one (or more) Hydroxides per Base Molecule NaOH+ OH-(aq) KOH K+(aq) + OH-(aq) LiOH Li+(aq) + OH-(aq) RbOH Rb+(aq) + OH-(aq) CsOH Cs+(aq) + OH-(aq) Ca(OH)2 Ca2+(aq) + 2OH-(aq) (but not very soluble) Ba(OH)2 Ba2+(aq) + 2OH-(aq) (but not very soluble) Sr(OH)2 Sr2+(aq) + 2OH-(aq) (but not very soluble) (For a list of common weak acids and bases, see Table 8-2 in Oxtoby) Sure and bases will neutralize each other to form liquid water and a salt. Describing the common characteristics of acids and bases, comparing the three ways to define them Key Takeaways Key Points A acid is a substance that donates protons (in the Brønsted-Lowry definition) or accepts some valence electrons to form a tape (in the Lewis definition). A base is a substance that protons can accept or donate some valence electrons to form a tire. Bases can be thought of as the chemical opposite of acids. A response between an acid and base is called a neutralization response. The strength of an acid refers to its ability or tendency to lose a proton; a strong acidity is one that dissociates completely in water. Key terms valence electron: Any of the electrons in the outer shell of an atom; able to form ties with other atoms. Lewis base: Any connection that can donate some electrons and form a coordinate coval tape. Lewis acid: Any compound that can accept some electrons and form a coordinate coval tape. Sure has long been recognized as a distinctive class of compounds whose aqueous solutions exhibit the following characteristics: A distinctive sour taste. Change the color of litmus from blue to red. Respond with certain metals to produce gaseous H2. Respond with bases to form a salt and water. Acidic solutions have a pH less than 7, with lower pH values corresponding to increasing acidity. Common examples of acids include acetic acid (in vinegar), sulfuric acid (used in car batteries), and tartaric acid (used in baking). There are three common definitions for acids: Arrhenius acid: any substances that increase the concentration of hydronium ions (H3O+) in solution. Brønsted-Lowry acid: any substance that can act as a proton donor. Lewis acid: any substance that can accept some electrons. Acid strength and strong acids The strength of an acid refers to how readily an acid will lose or donate a proton, often in solution. A stronger acid ionizes more readily, or dissociates, in a solution as a weaker acidity. The six common strong acids are: hydrochloric acid (HCl) hydrobromic acid (HBr) hydroiodic acid (HI) sulphuric acid (H2SO4; only the first proton is Strong acid) sally acid (HNO3) perchloric acid (HClO4) Each of these acids essentially ionizes 100% in solution. By definition, a strong acid is one that dissociates completely in water. In other words, one mole of the generic strong acid, HA, will deliver one mole of H+, one mole from the conjugate base, A-, with none of the unenconcene acid HA remaining in solution. By contrast, however, a weak acidity, less willing to donate his proton will only partially dissociate in solution. At equilibrium, both the acid and the conjugate base will be present, along with a significant amount of the undissociated species, HA. Factors that affect acid strength Two key factors contribute to overall strength of a acid: polarity of the molecule strength of the H-A bond These two factors are actually related. The more polar the molecule, the more the electron density within the molecule will be pulled away from the proton. The greater the partial positive charge on the proton, the weaker the H-A tape will be, and the more readily the proton will dissociate in solution. Acid strengths are also often discussed in terms of the stability of the conjuge base. Stronger acids have a larger Ka and a more negative pKa than weaker acids. Metal and acid response: Zinc reacts with water chloride acid to form hydrogen gas. Bases There are three common definitions of bases: Arrhenius base: any compound that donates a hydroxide date (OH-) in solution. Brønsted-Lowry base: any connection capable of accepting a proton. Lewis base: any connection capable of donating an electron pair. In water, basic solutions will have a pH between 7-14. Base strength and strong bases A strong base is the reverse of a strong acidity, while an acid is considered strong if it can readily donate protons, a base is considered strong if it can readily explode (that is, remove an H+ion) from other compounds. As with acids, we often talk about basic aching solutions in water, and the species that are exploded is often water itself. The common response looks like: [text]{text{A}^-} + text{(aq)} + text{H}\_2text{(O)}text{(aq)}rightarrowtext{A}text{(aq)} + text{(OH)}^-text{(aq)}[/text] Thus, deprotonated water returns hydration The concentration of hydroxide onions increases as pH increases. Most alkali metal and some alcalian earth metal hydroxides are strong bases in solution. These include: sodium hydroxide (NaOH) potassium hydroxide (KOH) lithium hydroxide (LiOH) rubidium hydroxide (RbOH) cesium hydroxide (CsOH) Calcium hydroxide [Ca(OH)2] barium hydroxide (Ba(OH)2) strontium hydroxide (Sr(OH)2) The alkali metal hydroxides dissociate completely in solution. The alcaic earth metal hydroxides are less soluble, but are still considered to be strong bases. Acid/Base Neutralization Sure and bases react with each other to yield water and a salt. For example: This response is called a neutralization response. Lewis bases and acids: A list of several Lewis bases (right) and Lewis acids (left). Sure+ Bases made easy! Part 1 - What the gate is an acid or base? – Organic Chemistry – YouTube: Ever wondered what the hen a acid or base actually is? Have you ever been super confused in high school or college chemistry? In this video I introduce to you that the gate a acid and base really forgets about the Lewis or Brønsted/Lowry definitions and then we'll go more in depth in parts 2,3, and 4. An Arrhenius acid dissociates in water to form hydrogen ions, while an Arrhenius base dissociates in water to form hydroxide ions. Remember the Arrhenius acid definition and its limitations. Key takeaways Key Points An Arrhenius acid increases the concentration of hydrogen (H+) in an aqueous solution, while an Arrhenius base increases the concentration of hydroxide (OH-) in an aqueous solution. The Arrhenius definitions of acidity and alkalinity are limited to horrible solutions and refer to the concentration of the solvent ions. The universal water acid-base definition of the Arrhenius concept is described as forming a water molecule from a proton and hydroxide ion. Therefore, in Arrhenius acid-base reactions, the reaction between an acid and a base is a neutralization response. Key terms hydrium: The hydrated hydrogen ion (SH\_3O^+S). Acid: a measure of the overall concentration of hydrogen ions in solution alkalinity: a measure of the overall concentration of hydroxide ions in solution A acidic base reaction is a chemical reaction that occurs between an acid and a base. Various concepts exist that provide alternative definitions for the relevant response mechanisms and its application to solve related problems. Despite various differences in definitions, its importance, since different methods of analysis become apparent when applied to acid base reactions for gaseous or liquid species, or when acid or base character can be somewhat less apparent. The Arrhenius definition of acid-base reactions, which were devised by Svante Arrhenius, is a development of the hydrogen theory of acids. It was used to provide a modern definition of acids and bases, and because of Arrhenius's work with Friedrich Wilhelm Ostwald in establishing the presence of onions in aqueous solution in 1884. This led to Arrhenius receiving the Nobel Prize in Chemistry in 1903. As defined by Arrhenius: An Arrhenius acid is a substance that dissociates in water to form hydrogen ions (H+). In other words, an acid increases the concentration of H+ lie in an aqueous solution. This protonation of water delivers the hydronium ion (H3O+). In modern times, H+ is used as a shorthand for H3O+ because it is now known that a bare proton (H+) does not exist as a free species in aqueous solution. A base is a substance that dissociates in water to form hydroxide (OH-) ions. In other words, a base increases the of OH- lie in an aqueous solution. Limitations of the Arrhenius Definition The Arrhenius definitions of acidity and alkalinity are limited to analytical solutions and refer to the concentration of the dissolved substances. Under this definition, pure H2SO4 or HCl dissolved in toluene is not acidic, despite the fact that both of these acids will donate a proton to toluene. In addition, under the Arrhenius definition, a solution of sodium amide (NaNH2) in liquid ammonia is not alkaline, despite the fact that the amide ion ([text]{text{NH}\_2^-}[/text]) will readily deprotonate ammonia. Thus, the Arrhenius definition can only describe acids and bases in a watery environment. Arrhenius Acid-Base Response An Arrhenius acid-base response is defined as the reaction of a proton and a hydroxide ion to form water. [text]{text{H}^+} + text{(OH)}^-rightarrowtext{H}\_2text{(O)}[/text] This is an Arrhenius acid base response simply a neutralization response. Chemistry 12.1 What are acids and bases? (Part 1 of 2) - YouTube: This introduction to acids and bases discusses their common characteristics and explains the Arrhenius definitions for acids and bases. A Brønsted-Lowry acid is any species capable of donating a proton; A Brønsted-Lowry base is any species capable of accepting a proton. Distinguish Brønsted-Lowry and Arrhenius acids. Key takeaways Key Points The formation of conjugate acids and bases is central to the Brønsted-Lowry definition of acids and bases. The conjugate base is the ion or molecule remaining after the acid has lost its proton, and the conjugate acid is the species created when the base assumes the proton. Interestingly, water is amph hysterical and can act as an acid and a base. Therefore, it can play all four roles: acid, conjugate base, acidity and base. A Brønsted-Lowry acid -base response can be defined as: acid + base [text]{rightleftharpoons}[/text] conjugate base + conjugate acid. Key terms amphiteric: To have the characteristics of both an acid and a base; is capable of both donating and accepting a proton (amphiprotic). assess acid: The species created when a base accepts a proton. conjuge base: The species left after an acid donates a proton. Originally, acids and bases were defined by Svante Arrhenius. Its original definition stated that acids were compounds that increased the concentration of hydrogen ions (H+) in solution, while bases were compounds that increased the concentration of hydroxide inns (OH-) in solutions. Problems arise with this conceptualisation because Arrhenius's definition is limited to aqueous solutions, with reference to the solution of aqueous ions, and therefore is not inclusive of acids dissolved in organic solvents. To solve this problem, Johannes Nicolaus Brønsted and Thomas Martin Lowry, in 1923, both independently proposed an alternative definition of acids and bases. In newer system, Brønsted-Lowry acids are defined as any molecule or ion that is able to donate hydrogen cation (proton, H+), while a Brønsted-Lowry base is a species with the ability to win, or accept, a hydrogen cation. A wide variety of compounds can be classified into the Brønsted-Lowry framework: mineral acids and derivatives such as sulfonate, carboxylic acids, amines, cabbage acids, and many more. Brønsted-Lowry Acid/Base Response Keep in mind that acids and bases should always respond in pairs. That's because if a connection is to act as an acid, donate its proton, then there must necessarily be a base present to accept that proton. The general scheme for a Brønsted-Lowry acid/base response can be visualized in the form: acid+ base [text]{rightleftharpoons}[/text] conjugate base + conjugate acid Here, a conjugate base is the species that is left after the Brønsted acid donates its proton. The conjugate acid is the species formed when the Brønsted base accepts a proton of the Brønsted acid. Therefore, according to the Brønsted-Lowry definition, an acid-base response is one in which a conjugate base and a conjugate acid are formed (note how it differs from the Arrhenius definition of an acid-base response, which is limited to the response of H+ with OH- to produce water). Finally, note that the response can continue in either the forward or the backward direction; In each case, the acid donates a proton to the base. Consider the response between acetic acid and water: [text]{text{H}\_3text{CCOOH}text{(aq)} + text{H}\_2text{(O)}text{(l)}rightleftharpoons text{H}\_3text{CCOO}^-text{(aq)} + text{H}\_3text{(O)}^+text{(aq)}[/text] Here serves as a Brønsted-Lowry acid, donates a proton to water, who serves as the Brønsted-Lowry base. The products include the acetateaition, which is the conjuge base formed in the response, as well as hydronium ion, which the conjugate acid was formed. Note that water is amproic; depending on the circumstances, it can act as either an acid or a base, whether donating or accepting a proton. For example, in the presence of ammonia, water will donate a proton and act as a Brønsted-Lowry acid: [text]{text{NH}\_3text{(aq)} + text{H}\_2text{(O)}text{(l)}rightleftharpoons text{(NH}\_4^+text{(aq)} + text{(OH)}^-text{(aq)}[/text] Here, ammonia is the Brønsted-Lowry base. The conjugate acid formed in the response is the ammonium ion, and the conjugate base formed is hydroxide. Chemistry 12.1 What are acids and bases? (Part 2 of 2) - YouTube: This lesson still describes acids and bases by their definition. We first look at the Brønsted-Lowry theory, and then describe Lewis acids and bases according to the Lewis Theory Water capable of acting as either an acid or a base and can undergo self-ionization. Explain the amphiteric properties of water. Key Takeaways KeyPoints The self-ionization of water can be expressed as: [text]{text{H}\_2text{(O)} + text{H}\_2text{(O)}rightleftharpoons text{(H}\_3text{(O)}^+text{(aq)} + text{(OH)}^-text{(aq)}[/text] + constant for the self-ionization of water is known as KW; it has a value of [text]{1.0times 10^{-14}}[/text]. The value of KW results in the convenient equation regarding pH with pOH: pH + pOH = 14. Important terms ionization: Any process leading to the dissociation of a neutral atom or molecule in charged particles (ions). autoprotolysis: The auto-ionization of water (or similar compounds) in which a proton (hydrogen ion) is transmitted to form a cage and an ion. Waterium: The hydrated hydrogen ion (SH\_3O^+S). Under standard conditions, water will self-ionize to a very small extent. The self-ionization of water refers to the reaction in which a water molecule donates one of its protons to a neighbouring water molecule, whether in pure water or in aqueous solution. The result is the formation of a hydroxide ion (OH-) and a hydronium ion (H3O+). The response can be written as follows: [text]{text{(H}\_2text{(O)} + text{(H}\_2text{(O)}rightleftharpoons text{(H}\_3text{(O)}^+text{(aq)} + text{(OH)}^-text{(aq)}[/text] This is an example of autorotolise (meaning self-protonizing) and it sets the amphiteric nature of water (ability to act as both an acid and a base). Autoprotolysis of water: The self-ionization of water produces hydronium and hydroxide ions in solution. The Water Ionization Constant, KW Note that the self-ionization of water is an equilibrium response: [text]{text{(H}\_2text{(O)} + text{(H}\_2text{(O)}rightleftharpoons text{(H}\_3text{(O)}^+text{(aq)} + text{(OH)}^-text{(aq)}[/text] + [text{(OH)}^-text{(aq)} + text{(O)}text{(H)}^+text{(aq)}quadquadquadtext{(K)}\_text{(W)}=1.0times 10^{-14}}[/text] Like all equilibrium responses, this response has an equilibrium constant. Because it is a special equilibrium constant, specifically for the self-ionization of water, it is indicated KW; it has a value of 1.0 x 10-14. If we write out the actual equilibrium expression for KW, we get the following: [text]{text{(K)}\_text{(W)}=text{(H}\_3^+text{(aq)}text{(OH)}^-text{(aq)}=1.0times 10^{-14}}[/text] However, because H+ and OH- are formed in a 1:1 molar relationship, we have: [text]{text{(H}\_3^+text{(aq)}=text{(OH)}^-text{(aq)}=sqrt{1.0times 10^{-14}}}=1.0times 10^{-7}}[/text] [text{(M)}] Now, note the definition of pH and pOH: [text]{text{(pH)}=-text{(log)}[text{(H}\_3^+text{(aq)})][/text] [text]{text{(pOH)}=-text{(log)}[text{(OH)}^-text{(aq)}]}[/text] If we plug in the above value in our equation for pH, we find it: [text]{text{(pH)}=-text{(log)}[1.0times 10^{-7}]}=7.0[/text] [text{(aq)}] Here we have the reason why neutral water has a pH of 7.0; this represents the condition at which the concentrations of H+ and OH- are exactly equal in solution. pH, pOH, and pKW We have already established that the equilibrium constant KW can be expressed as: [text]{text{(K)}\_text{(W)}=text{(H}\_3^+text{(aq)}text{(OH)}^-text{(aq)}[/text] If we take the negative logarithms from both sides of this equation, we get the following: [text]{-text{(log)}[text{(K)}\_text{(W)}]= -text{(log)}[text{(H}\_3^+text{(aq)}text{(OH)}^-text{(aq)}]}[/text] [text]{text{(pK)}\_text{(W)}=-text{(pH)}+text{(pOH)}[/text] however, because we know that pKW = 14, we can establish the following relationship: [text]{text{(pH)}+text{(pOH)}=14}[/text] This relationship always holds true for any aqueous solution regardless of its level of acid or alkalinity. Using this equation is a convenient way to quickly determine pH of pH and vice versa, as well as to determine watericde concentration given hydrogen concentration, or vice versa. Selfionisation of Water: Explanation of self-ionization of water and the formation of hydronium and hydroxide ions. The acid dissociation constant (Ka) is the measure of the strength of an acid in solution. Compare and contrast acid strengths using Ka and pKa values. Key takeaways Key Points A acid dissociation constant (Ka) is a quantitative measure of the strength of an acid in solution. The dissociation constant is usually written as a quote from the equilibrium concentrations (in moles/L): [text]{text{(K)}\_text{(a)}=frac{[text{(A)}^-][text{(H}\_3^+)]}{[text{(HA)}]}}[/text]. Often the Ka value is expressed using the pKa, which is equal to [text]{-text{(log)}\_10[text{(K)}\_text{(a)}]}[/text]. The greater the value of pKa, the smaller the extent of dissociation. A weak acidity has a pKa value in the approached range of -2 to 12 in water. Acids with a pKa value of less than about -2 are said to be strong acids. Important terms dissociation: Citing the process by which a connection breaks into its constituent ions in solution. Equilibrium: The state of a response in which the rates of the forward and reverse reactions are equal. The acid dissociation constant (Ka) is a quantitative measure of the strength of an acid in solution. Cache is the equilibrium constant for the following dissociation response of an acid in aqueous solution: [text]{text{(HA)}text{(aq)}rightleftharpoons text{(H}\_3^+text{(aq)} + text{(A)}^-text{(aq)}[/text] In the above response, HA (the generic acid), A- (the conjunced base of the acid), and H+ (the hydrogen ion or proton) are said to be equilibrium when their concentrations do not change over time. As with all equilibrium conventions, the value of Ka is determined by the concentrations (in moles/L) of each aqueous species in equilibrium. The Ka expression is as follows: [text]{text{(K)}\_text{(a)}=frac{[text{(H}\_3^+text{(aq)}][text{(A)}^-]}{[text{(HA)}]}}[/text] Acid dissociation constants are mostly associated with poor acids, or acids that do not completely dissociate in solution. This is because strong acids are thought to be completely in solution and therefore their Ka values are exceptionally large. Ka and pKa Due to the many orders of magnitude stretched by Ka values, a logarithmic measure of the acid dissociation constant is more commonly used in practice. The logarithmic constant (pKa) is equal to -log10(Ka). The greater the value of pKa, the smaller the extent of dissociation. A weak it's a pKa value in that approached series of -2 to 12 in water. Sure with with pKa value of less than about -2 is said to be strong acids. A strong acid is almost completely disassociated in awful solution; it is disassociated to the extent that the concentration of the undissociated acid becomes insencere. pKa values for strong acids can be estimated by theoretical means or by extrapolating methings into non-leading solvents with a smaller dissociation constant, such as acetone and dimethylsulphoxide. Acetic acid dissociation: The acetic acid partially and reversible dissociates in acetate and hydrogen ions. Acetic acid is a weak acid with an acid dissociation constant [text]{text{(K)}\_text{(a)}=1.8times 10^{-5}}[/text]. What is the pKa for aesthetic acid? [text]{text{(a)}\_text{(a)}=-text{(log)}[1.8times 10^{-5}]}=4.74[/text] pOH and Other p Scales A P scale is a negative logarithmic scale. Switch between pH and pOH scales to solve acid-based equilibrium problems. Key takeaways Key Points The p scale is a negative logarithmic scale. This allows numbers with very small units of magnitude (for example, the concentration of H+ to be converted into solution) into more convenient numbers, often within the range of -2 - 14. The most common p scales are the pH and pOH scales, which measure the concentration of hydrogen and hydroxide easts. According to the waterion product, pH + pOH = 14 for all aerial solutions. Due to the convenience of the p scale, it is used to also indicate the small dissociation constants of acids and bases, which are given by the notation pKa and pKb. Key terms dissociation: the process by which compounds are divided into smaller constituent molecules, usually reversible logarithms: for a number of SxS, the power to which a given base number must be increased to obtain x, written logbx. for example, log216 = 4 because 24 = 16 Recall the reaction for the autoionization of water: [text]{text{(H}\_2text{(O)}rightleftharpoons text{(H}\_3^+text{(aq)} + text{(OH)}^-text{(aq)}[/text] This reaction has a special equilibrium constant denoted KW, and it can be written as follows: [text]{text{(K)}\_text{(W)}=text{(H}\_3^+text{(aq)}text{(OH)}^-text{(aq)}=1.0times 10^{-14}}[/text] Because H+ and OH- dissociate in a one-to-one molar ratio, [text]{text{(H}\_3^+text{(aq)}=text{(OH)}^-text{(aq)}=sqrt{1.0times 10^{-14}}}=1.0times 10^{-7}}[/text] If we take the negative logarithm of each concentration, we get: [text]{text{(pH)}=-text{(log)}[text{(H}\_3^+text{(aq)}]}=-text{(log)}[1.0times 10^{-7}]}=7.0[/text] [text]{text{(pOH)}=-text{(log)}[text{(OH)}^-text{(aq)}]}=-text{(log)}[1.0times 10^{-7}]}=7.0[/text] Here we have the reason that neutral water has a pH of 7.0 -: this is the pH at which the concentrations of H+ and OH- are exactly equal. Lastly, we should note the following relationship: [text]{text{(pH)}+text{(pOH)}=14}[/text] This relationship will always apply to aqueous solutions. This is a quick and convenient way to find pH from pOH; hydrogen ion concentration of water-icde ion concentration, and more. The pH and pOH Scale: Relationship between p[OH] and p[H] acid, which are the lower numbers for the pH scale and higher numbers for the pOH scale; brighter blue is more basic, which are the higher numbers for the pH scale and lower numbers for the pOH scale). pKa and pKb Generics, this p-notation can be used for other scales. In acid-base chemistry, the amount by which an acid or base dissociates to form H+ or OH- ions in solution are often given in terms of their dissociation constants (Ka or Kb). However, because these values are often very small for weak acids and weak bases, the p scale is used to simplify these numbers and make them more convenient to work with. Often we'll see the notation pKa or pKb, which refers to the negative logarithms of Ka or Kb, respectively. Interactive: pH: Test the pH of things like coffee, spit and soap to determine if each is acidic, basic or neutral. Visualize the relative number of hydroxide ions and hydronium ions in solution. Switch between logarithmic and linear scales. Investigate whether changing the volume or diluted with water affects the pH. Or you can design your own liquid! pH and pOH: This lesson introduces the pH scale and discusses the relationship between pH, [H+], [OH-] and pOH. POH.

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