



## Ka in chemistry

Learning goals: Know the relationship between acidic or basic strength and size \(K\_a\), \(pK\_a\) and \(pK\_b\). To understand the balancing effect. The equilibrium constant size for an ionizing reaction can be used to determine the relative strength of acids. For example, the general equation for ionization of weak acid in water, where HA is the mother acid and A- is its conjugate base, is as follows:  $[HA_{(aq)}+H_2O_{(l)} rightleftharpoons H_3O^+_{(aq)}+A^-_{(aq)} label{16.5.2}]$  As we have already stated, the water concentration is basically constant for all reactions in an aqueous solution, so  $([H_2O])$  in the  $(\sqrt{161.5.2})$  equation can be incorporated into the new quantity, acidization constant:  $(K_a)$ , also called acid dissociation constant:  $(K_a)$ , also in equation \(\ref{16.5.3}\). However, keep in mind that free \(H^+\) does not exist in aqueous solutions and that proton is transferred to \(H\_2O\) in all acid ionization reactions that make up hydronium ions, \(H\_3O^+\). The larger \(K\_a\), the stronger the acid and the higher the concentration \(H^+\) at balance. Like all equilibrium constants, acid-base ionization constants are actually measured in terms of activities \(H^+\) or \(OH^-\), thus becoming inconsistent. The \(K\_a\) values for a series of common acids are listed in table \(\PageIndex{1}\). Table \(\PageIndex{1}\). Table \(\PageIndex{1}\). Table \(\PageIndex{1}\).  $(K_b) (pK_b) * The number in parentheses indicates the ionization step specified for polyprotic acid. hydroionic acid ((HI)) (2.3\times 10^{2}) -2.0 ((HSO_4^-)) (1 \ times 10^{2}) -2.0 ((HSO_4^-)) (1 \ times 10^{2}) -2.0 ((HSO_4^-)) (1 \ times 10^{2}) 16.0 acid nitric ((HNO_3)) (2.3\times 10^{1}) -1.37 ((NO_3^-)) (4.3\times 10^{-16}) (1 - 1.3$ 15.37 hydro ion \(H\_3O^+\) \(1.1.1 0\) 0.00 \(H\_2O\) \(1.0\times 10^{-14}\) 14.00 sulphuric acid (2)\* \(HSO\_4^-\) \(1.0\times 1.0\times  $(NO2^{-}) (1.8 \ times 10^{-11}) 10.75 \ uric acid ((C_6H_5CO_2+)) (1.7 \ times 10^{-5}) 4.76 ((CH_3CO_2+)) (5.8 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 9.80 \ Acetic Acid ((C_6H_5CO_2+)) (1.7 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 9.80 \ Acetic Acid ((C_6H_5CO_2+)) (1.7 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.7 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 9.80 \ Acetic Acid ((C_6H_5CO_2+)) (1.7 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 9.80 \ Acetic Acid ((C_6H_5CO_2+)) (1.7 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 9.80 \ Acetic Acid ((C_6H_5CO_2+)) (1.6 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 9.80 \ Acetic Acid ((C_6H_5CO_2+)) (1.6 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) (1.6 \ times 10^{-11}) 10.25 \ benzoic acid ((C_6H_5CO_2+)) 10.25 \ benzoic acid ((C_6H$  $k_{-5}$  4.79 amonium ion \(NH\_4^+) \(5.9\times 10^{-5}\) 5.23\(C\_5H\_5N\) \(1.7\times 10^{-9}\) 8.77 chloruric acid \(HOCl\) \(4.0 \times 10^{-5}\) 4.79 amonium ion \(NH\_4^+) \(5.6\times 10^{-7}\) 9.25 \(NH\_3\) \(1.8 \ times 10^{-5}\) 4.75 water \(H\_20\) \(1.0 \times 10^{-14}\) 14.00 \(OH^-\) \(1.00\) 0.00 acetylene \(C\_2H\_2\) \(1\times 10^{-23}\) 26, 0 \(HC\_2^-\) \(1\times 10^{2-1}\) -21.0 Weak bases react with water to form a hydroxide ion, as shown in the following general equation, where B is the parent of the parent by (1.1) \(1.1) base and BH+ is its conjugated acid: \[B\_{(aq)}+H\_2O\_{(l)} \rightleftharpoons BH^+\_{(aq)}+OH^-\_{(aq)}+ an equilibrium constant expression; instead, it is part of \(K\_b\). The larger \(K\_b\), the stronger the base and the higher the concentration \(OH^-\) in balance. \(K\_b\), \(PageIndex{2}\): Values for a number of common weak bases (B) and their cyjugated acids (BH+) Basic \(B\) \(K\_b\) \(PK\_b\) \(PK\_b\) \(BH^+\) \(K\_a\) \(PA\_a\) \(PA\_a\) \(PA\_a\) \(PA\_a\) \(2.1x \1 times 10^{-13}\) 12.32 dimethylamine \((CH\_3)\_2NH\) \(5.4\times 10^{-4}) \) 3.27 \  $((CH_3)_2NH_2^+) (1.9 \times 10^{-11}) 10.73 \text{ methylamine} ((CH_3NH_2)) (4.6 \times 10^{-11}) 10.66 \text{ trimethylamine} ((CH_3)_3NH^+) (2.2 \times 10^{-11}) 10.66 \text{ trimethylamine} ((CH_3)_3NH^+) (1.6 \times 10^{-11}) 10.66 \text{ trimethylamine} ((CH_3)_3NH^+) (2.2 \times 10^{-11}) 10.66 \text{ trimethylam$  $(C_{6H_{5NH_{3}}) = 10^{-9}) = 10^{-9}) = 10^{-9}) = 10^{-19}) =$ Take, for example, the ionisation of hydrocyanic acid (\(HCN\)) in water to form an acid solution and the reaction of \(CN^-\) with water to solution:  $||HCN_{(aq)}+H_2O_{(l)}|$  hightleftharpoons  $H^+_{(aq)}+H_2O_{(l)}$  hightleftharpoons  $H^+_{(aq)}+H_2O_{(l)}$  hightleftharpoons  $H^+_{(aq)}+H_2O_{(l)}$  hightleftharpoons  $H^+_{(aq)}+H_2O_{(l)}$  hightleftharpoons  $H^+_{(aq)}+H_2O_{(l)}$  hightleftharpoons  $H^+_{(aq)}$  hightleftharpo of HCN is as follows: \[K\_a=\dfrac{[H^+][CN^-]}{[HCN]} \label{16.5.8}] The corresponding expression for the reaction of cyanide with water is as follows: \[K\_b=\dfrac{[OH^-]}{[HCN]} \label{16.5.8}] The corresponding expression for the reaction of cyanide with water is as follows: \[K\_b=\dfrac{[OH^-]}{[HCN]} \label{16.5.8}] The corresponding expression for the reaction of cyanide with water is as follows: \[K\_b=\dfrac{[OH^-]}{[HCN]} \label{16.5.8}] The corresponding expression for the reaction of cyanide with water is as follows: \[K\_b=\dfrac{[OH^-]}{[HCN]} \label{16.5.8}] The corresponding expression for the reaction of cyanide with water is as follows: \[K\_b=\dfrac{[OH^-]}{[HCN]} \label{16.5.8}] The corresponding expression for the reaction of cyanide with water is as follows: \[K\_b=\dfrac{[OH^-]}{[HCN]} \label{16.5.8}] The corresponding expression for the reaction of cyanide with water is as follows: \[K\_b=\dfrac{[OH^-]}{[HCN]} \label{16.5.8}] The corresponding expression for the reaction of cyanide with water is as follows: \[K\_b=\dfrac{[OH^-]}{[HCN]} \label{16.5.8}] The corresponding expression for the reaction of cyanide with water is as follows: \[K\_b=\dfrac{[OH^-]}{[HCN]} \label{16.5.8}] The corresponding expression for the reaction of cyanide with water is as follows: \[K\_b=\dfrac{[OH^-]}{[HCN]} \label{16.5.8}] The corresponding expression for the reaction of cyanide with water is as follows: \[K\_b=\dfrac{[OH^-]}{[HCN]} \label{16.5.8} \label{16  $H^+_{(aq)}+(cancel{CN^-_{(aq)}}) (K_a=[H^+](cancel{[HCN]}) (K_b=[OH^-](cancel{[HCN]}) (K_b=[OH^-](aq))) (K_b=[OH^-](cancel{[HCN]}) (K_b=[OH^-](cancel{[HCN})) (K_b=[OH^-](cancel{[HCN$ (K\_a)) and \(K\_b)) is the equation for water autoionization, and the product of the two equilibrium constants is \(K\_w): \[K\_aK\_b = K\_w \label{16.5.10}\] So if we know that either \(K\_a) for acid or \(K\_b) for its conjugate base, we can calculate an additional equilibrium constant for any pair of conjugated acid and base. As with \(pH), \(pOH) and pKw, we can use negative logarithms to avoid exponential writing of acid and base ionization constants, by defining \(pK\_a\) as follows:  $[pKa = -\log_{10}K_b + \log_{16.5.13}] \[K_b = 10^{-pK_b} \]$ E2. Due to the use of negative logarithms, smaller values \(pK\_a\) orrespond to larger acidic ionization constants and thus stronger acids. For example, nitrous acid (\(HNO\_2\)), with \(pK\_a\) 9.21. Conversely, the smaller values \(pK\_b\) correspond to the larger basic ionization constants and thus the stronger the bases. Figure \(\PageIndex{1}\): Relative strengths of some common conjugated acid and base pairs. The strongest principles are in the upper right corner. Conjugate base of strong acid is very weak base and conversely conjugated acid of strong base is very weak acid. Relative forces of some common acids and their conjugate are graphically displayed in the image \(\PageIndex{1}\). Conjugated acid and base or to the increasing values of \(pK\_b\). Strong acids are common in the lower left corner of the {2} \(\PageIndex{2}\) image; at the top right are the most common strong bases. Note the inverse relationship between the strength of the parent acid and the force of the conjugate base. Thus, the conjugate base of a strong acid is very weak, and the conjugate base of a very weak acid is a strong principle. Conjugate base of strong acid is weak base and vice versa. The relative forces of acids and bases can be used to predict the direction of acid-base balance always favors the side with weaker acid and base, according to a single rule: acid-base balance always favors the side with weaker acid and base, according to a single rule: acid-base balance always favors the side with weaker acid and base, according to the following arrows: \[\text{stronger acid + stronger acid + base} \ce{ <=&gt;&gt;} \text{weaker acid + weaker base} \] In acidobasic reaction, the proton always reacts with a stronger base. For example, hydrochloric acid is a strong acid that ionizes essentially completely in a dilute aqueous solution to form \(H\_3O^+\) and \(Cl^-\); only negligible amounts of \(HCl\) molecules remain undisciplined. Therefore, the ionization balance lies practically to the right, represented by a single arrow: \[HCl\_{(aq)} + H\_2O\_{(l)} \rightarrow H\_3O^+\_{(aq)} \label{16.5.17}] By contrast, acetic acid is weak base. Aqueous acetic acid solutions therefore contain mainly acetic acid molecules in balance with a small concentration of \(H\_3O^+) and octate ions, and the ionization balance lies far to the left, represented by the following arrows: \[ \ce{ CH\_3CO\_2H\_{(aq)} + H\_2O\_{(1)} <&lt;=&gt; H\_3O^+- }\] Similarly, in response to water ammo, hydroxide ion is a strong base and ammonia is a weak base , while the ammonium ion is a stronger acid than water. Therefore, this balance also lies on the left: \[H\_2O\_{(1)} + NH\_{3(aq)} \ce{ <&lt;=&gt;} NH^+\_{4(aq)} + OH^-\_{{(aq)}} All acidobelizeria prefer a side with weaker acid and base. Proton is therefore tied to a stronger base. Example \(\PageIndex{1}\): Butyrate and Dimethylammonium Calculate ions \(K\_b\) and \(pK\_b\) butyrate ion (\(CH\_3CH\_2CO\_2^-\)). \(pK\_a\) butyric acid at 25 °C is 4,83. Butyric acid is responsible for the smell of yellowed butter. Calculate \(K\_a\) and \(pK\_a\) dimethylamine (\(CH\_3)\_2NH)) is \(5.4 \times 10^{-4}\) at 25°C. Because: \(pK\_a\) and \(K\_b\) Have requested: matching \(K\_b\) and \(pK\_b\), \(K\_a\) and \(K\_b\) dimethylamine (\(CH\_3)\_2NH)) is \(5.4 \times 10^{-4}\).  $(K_a)$  and  $(K_b)$  are linked, as shown in equations  $(ref{16.5.13})$  and  $(K_b)$  are linked, as shown in the  $(K_a)$  and  $(pK_a)$  and  $(pK_b)$  for acid and its conjugate base are related, as shown in equations  $(K_a)$  and  $(pK_a)$  and  $(PK_b)$  for acid and its conjugate base are related, as shown in the  $(K_a)$  and  $(PK_a)$  and  $(PK_b)$  for acid and its conjugate base are related, as shown in equations  $(K_b)$  and  $(PK_b)$  and  $(PK_b)$  for acid and its conjugate base are related, as shown in the  $(K_b)$  and  $(PK_b)$  and \(pK\_b\). Solution: We are given \(pK\_a\) for butyrate acid and asked to calculate \(K\_b\) and \(pK\_b\) for its conjugate base, butyrate ion. Since the quoted value \(pK\_a\) + \(pK\_b) = pKw = 14.00. Replacement \(pK\_a\) and Solution \(pK\_b\), \[4.83+pK\_b=14.00\] \  $pK_b=14.00-4.83 = 0 (pK_b = -\log K_b), (K_b) is (10^{-9.17} = 6.8 \times 10^{-10}).$  In this case, we got  $(K_b)$  for its conjugated acid, dimethylamonium ion. Because the initial quantity is  $(K_b)$  instead of  $(pK_b)$ , we can use equation  $(ref{16.5.10})$ :  $(K_aK_b = K_w)$ . Replacing \(K\_b\) and \(K\_w\) at 25°C and solutions for \(K\_a\), \[K\_a(5.4\times 10^{-14})] \] [K\_a=1.9 \times 10^{-11}] Because \(pK\_a\), we have \(pK\_b\) to \(pK\_b\) to get the same answer: \[pK\_b=-\log(5.4\times 10^{-14})] \] [K\_a=1.9 \times 10^{-11}] Because \(pK\_a\), we have \(pK\_b\) to get the same answer: \[pK\_b=-\log(5.4\times 10^{-14})] \] [K\_a=1.9 \times 10^{-11}] Because \(pK\_a\), we have \(pK\_b\) to \(pK\_b\) to get the same answer: \[pK\_b=-\log(5.4\times 10^{-14})] \] [K\_a=1.9 \times 10^{-11}] Because \(pK\_a\) = -log \((K\_a\), we have \(pK\_b\) to get the same answer: \[pK\_b=-\log(5.4\times 10^{-14})] \] [K\_a=1.9 \times 10^{-11}] Because \(pK\_a\) = -log \((K\_a\), we have \(pK\_b\) to get the same answer: \[pK\_b=-\log(5.4\times 10^{-14})] \] [K\_a=1.9 \times 10^{-11}] Because \(pK\_a\) = -log \((K\_a\), we have \(pK\_b-11)] Because \(pK\_b-11)] = 10.72 \). [pK\_a=10.73\] \[pK\_b K\_a=10^{-pK\_a}=10^{-pK\_a}=10^{-pK\_a}=10^{-11}] If we received any of these four amounts for acid or base (\(K\_a\), \(pK\_a\), or \(pK\_b\)), or \(pK\_a\), \(pK\_a\), or \(pK\_b\), or \(pK\_b), or \(pK\_b), or \(pK\_b), or \(pK\_b), or \(pK\_b), or \(pK\_b(a), (pK\_b), or \(pK\_b(a), (pK\_b(a), (pK\_ in tired muscles. Its \(pK\_a\) is 3.86 at 25°C. Calculate \(K\_a\) for lactic acid and \(pK\_b\) and \(HNO\_3\) lie above hydronio, which means that (K\_a = 1.4 \times 10^{-4}\) for lactic acid; \(pK\_b\) = 10.14 and \(K\_b = 7.2 \times 10^{-4}\) for lactic ion. Answer \(K\_a = 1.4 \times 10^{-4}\) for lactic acid; \(pK\_b\) = 10.14 and \(PK\_b = 7.2 \times 10^{-11}\) for lactic ion. the \(pK\_a\) values are less than zero and are stronger acids than the ion \(H\_3O^+\). Recall from Chapter 4 that acid proton in virtually all oxoacids is associated with one of the oxygen atoms of oxoanione. Nitric acid should therefore be correctly spelled as \(HONO\_2\). Unfortunately, however, the formulas of oxokys are almost always written with hydrogen on the left side and oxygen on the right, giving \(HNO\_3\) a place. In fact, all six common strong acids we first encountered in Chapter 4 have values \(PK\_a\), which means that they are more likely to loses the proton than the \(H\_3O^+\) ion. Conjugated bases of these strong acids, on the other hand, are weaker bases than water. As a result, the protontransfer balance for these strong acids lies far to the right, and the addition of any of the common strong acids to the water results in essentially a schiometric reaction of \(H\_3O ^+\) ions and conjugated base acid. Although \(K\_a\) for \(HI\) is about 108 greater than \(K\_a\) for \(HNO\_3\), the response of either \(HI\) or \ (HNO\_3) with water provides essentially a stichiometric solution \(H\_3O^+\), regardless of the identity of the strong acid. This phenomenon is called a balancing effect: all species that are stronger acid than conjugated water acid (\(H\_3O^+)) are aligned to the strength \(H\_3O^+\) in an aqueous solution, because \(H\_3O^+\) is the strongest acid that can exist in balance with water. As a result, it is not possible to distinguish between strong acidity such as HI and HNO3 in an aqueous solution and an alternative approach must be used to determine their relative acidic forces. One method is to use a solvent such as anhydronic acetic acid. Since acetic acid is a stronger acid than water, there must also be a weaker base, with less tendency to take proton than \(HNO\_3\) in acetic acid indicate that HI is completely dissociated, but \(HNO\_3\) is only partially dissociated and acts as a weak acid in this solvent. This result clearly tells us that HI is a stronger acid than \(HNO\_3\). The relative order of acidic forces and approximate values \(K\_a\) and \(pK\_a\) for strong acids at the top of the table \(PageIndex{1}\) have been determined by measurements such as this and various unumulated solvents. In aqueous solutions \(H\_3O^+\) is the strongest acid, and \(OH^-\) is the strongest base that can exist in balance with \(H\_2O\). The balancing effect also applies to strong base solutions: In an aqueous solution, any base is stronger than OH- compensated by oh- because OH- is the strongest base that can exist in balance with \(H\_2O\), \(NaOCH\_3\) (sodium methoxide) and \(NaNH\_2\) (sodamide or sodium amide), whose anions are conjugated bases of species that would lie underwater in the table \(\PageIndex{2}\), are all strong bases, that respond essentially completely (and often violently) with water, accepting proton for solution \(OH^-\) and corresponding cation: \[K\_2O\_{{(s)}+H\_2O\_{(l)} \rightarrow  $2OH^-_{(aq)}+2K^+_{(aq)} \be {16.5.18}] \be {16.5.18}} \be {16.5.18} \be {16.5.18} \be {16.5.19}} \be {16.5.19} \be {16.5.19} \be {16.5.19}} Other examples you may encounter are potassium hydride (\(KH\)) and organometic compounds such as methylhimonity (\(CH_3Li\)). As you$ have learned, polyprotic acids such as \(H\_2SO\_4\), \(H\_3PO\_4\) and \(H\_2CO\_3\) contain more than one ionisable proton, and protons are gradually lost. A fully toned species is always the strongest acid because it is easier to remove a proton from a neutral molecule than from a of subsequent protons and increases accordingly \(pK a\). Consider \(H 2SO 4\), for example: \[HSO^- {4 (ag)} \ce{ <=&gt;&gt;} SO^{2-} {4(ag)}+H^+ {{(ag)} \ce{ &lt;=&gt;&gt;} SO^{2-} {4(ag)}+H^+ {{(ag)} \ce{ &lt;=&gt;} SO^{2-} {4(ag)}+H^+ {{(ag)} \ce{ &lt;=&gt;} SO^{ (HSO 4^-)) and \(SO 4^{2-})) are present in the balance. For polyprotic acid, the strength of the acid decreases, and \(pK a\) increases the sequential loss of each proton. Hydrogen sulphate ion (\(HSO 4^-)) is a conjugate base of \(H 2SO 4\) and conjugate base of \(H 2SO 4\) a whether the second reactant is a stronger acid or a stronger base. By contrast, the sulphate ion (\(SO\_4^{2-}\)) is a polyprotic base that is able to receive two protons sequentially: \[SO^{{2-}\_{4}(aq)} + H\_2O\_{{(aq)}} \[HSO^{{-}\_{4}(aq)} + H\_2O\_{{(aq)}} \[HSO^{{-}\_{4}(aq)} + H\_2O\_{{(aq)}} \[HSO^{{2-}\_{4}(aq)} + H\_2O\_{{(aq)}} \[HSO^{{2-}\_{4}(aq)} + H\_2O\_{{(aq)}} \] (a) = 0  $\{16.6\}\$  Like any other conjugate acid pair, the forces of conjugated acids and acids are joined  $(pK_a) + (pK_b) = pKw$ . Take, for example,  $(PK_a) + (PK_b) = pKw$ . Take, for example,  $(PK_a) + (PK_b) = pKw$ . Take, for example,  $(PK_a) + (PK_b) = pKw$ . Take, for example,  $(PK_a) + (PK_b) = pKw$ . Take, for example,  $(PK_a) + (PK_b) = pKw$ . Take, for example,  $(PK_a) + (PK_b) = pKw$ . Take, for example,  $(PK_a) + (PK_b) = pKw$ . Take, for example,  $(PK_a) + (PK_b) = pKw$ . Take, for example,  $(PK_b) = p$ is a relatively weak base, while  $(OH^-)$  is a strong base, so the balance shown in the equation  $(/ref{16.6})$  lies to the left. The ion  $(HSO_4^- = 14 - (-2.0) = 16)$ , which is consistent with what we expect for the conjugate base of strong acid. Example  $(/PageIndex{2})$  Predict whether the balance for each reaction lies to the left or right as written. \(NH^+\_{4(aq)}+PO^{2-}\_{4(aq)}) \(CH\_3CH\_2CO^-\_{2(aq)}) +HPO^{2-}\_{4(aq)}) Because: balanced chemical equation Asked: equilibrium position Strategy: In each reaction, determine the pairs between the ecjugate base. Then look at the \(\PageIndex{1}) and \(PageIndex{2}) tables and the picture determine which is the stronger acid and base. Balance always prefers the formation of a weaker acid-base pair. Solution: The base acid pairs are \(NH 4^+/NH 3) and \(HPO 4^{2}-)/PO 4^{3}-). According to tables \ (\PageIndex{1}) and \(PageIndex{2}), \(NH\_4^+\) is stronger acid (\(pK\_a = 9.25)) than \(HPO\_4^{2-}) (pKa = 12.32) and \(PO\_4^{3-}). The balance will therefore lie on the right, prioritising the formation of a weaker acid and base pair: \[ \[ \underset{text{stronger acid}}{NH^+\_{4(aq)}} + (QA\_{3-}). The balance will therefore lie on the right, prioritising the formation of a weaker acid and base pair: \[ \[ \underset{text{stronger acid}}{NH^+\_{4(aq)}} + (QA\_{3-}). The balance will therefore lie on the right, prioritising the formation of a weaker acid and base pair: \[ \[ \underset{text{stronger acid}}{NH^+\_{4(aq)}} + (QA\_{3-}). The balance will therefore lie on the right, prioritising the formation of a weaker acid and base pair: \[ \[ \underset{text{stronger acid}}{NH^+\_{4(aq)}} + (QA\_{3-}). The balance will therefore lie on the right, prioritising the formation of a weaker acid and base pair: \[ \[ \underset{text{stronger acid}}{NH^+\_{4(aq)}} + (QA\_{3-}). The balance will therefore lie on the right, prioritising the formation of a weaker acid and base pair: \[ \[ \underset{text{stronger acid}}{NH^+\_{4(aq)}} + (QA\_{3-}). The balance will therefore lie on the right, prioritising the formation of a weaker acid and base pair: \[ \[ \underset{text{stronger acid}}{NH^+\_{4(aq)}} + (QA\_{3-}). The balance will therefore lie on the right, prioritising the formation of a weaker acid and base pair: \[ \[ \underset{text{stronger acid}}{NH^+\_{4(aq)}} + (QA\_{3-}). The balance will therefore lie on the right, prioritising the formation of a weaker acid and base pair: \[ \[ \underset{text{stronger acid}}{NH^+\_{4(aq)}} + (QA\_{3-}). The balance will therefore lie on the right. The prioritising the formation of a weaker acid and base pair: \[ \[ \underset{text{stronger acid}}{NH^+\_{4(aq)}} + (QA\_{3-}). The prioritising the formation of a weaker acid and base pair: \[ \[ \underset{text{stronger acid}}{NH^+\_{4(aq)}} + (QA\_{3-}). The prioritising the formation of a weaker acid and base pair: \[ \[ \underset{text{stronger acid}}{N  $(CH_3CH_2CO_2^-)$  and  $(HCN/CN^-)$ . According to the  $(PageIndex{1})$  table, HCN is a weak acid (pKa = 9.21) and  $(CN^-)$  is a moderately weak acid (pKa = 9.21) and  $(CN^-)$ base (pKb = 4.79). However, propionic acid (\(CH 3CH 2CO 2H\)) is not listed in table \(\PageIndex{1}\). In a situation like this, the best approach is to look for a similar compound whose acid-base properties are listed. For example, propionic acid and acetic acid are identical except for groups attached to the carbon atom of carboxylic acid (\ (\ce{-CH\_3})), so we can expect both compounds to have similar acid-base properties. In particular, we would expect \(pK\_a\) propionic acid is 4.87, compared to 4.76 for acetic acid, making propionic acid a slightly weaker acid than acetic acid.) Propionic acid should therefore be significantly stronger in acid than \(HCN\). Because stronger acid forms a weaker conjugate base, we predict that cyanide will lie on the right and prefers to produce a weaker pair of acid and base: \[ \[ subset{\text{stronger acid}}(CH\_3CH\_2CO\_2H\_{{(aq)}} +  $\$ (HCO^- {2(aq)}+HSO ^-\rightleftharpoons HCO 2H {(aq)}+SO^{2-} {4(aq)}) Answer to left answer b left Acid-base reactions always contain two conjugated pairs of acid and base. Each acid and base. Each acid and each base has an associated ionization constant corresponding to its acidic or basic strength. Two species, which differ only by protona, form a conjugate acidbase pair. The equilibrium constant size for the ionization reaction can be used to determine strong acids and acids. For an aqueous solution of weak acid, the dissociation constant (\(K\_a\)). Similarly, the equilibrium constant for a weak base reaction with water is the basic ionization constant (\(K\_b\)). For any pair of conjugated acid and base\(K\_aK\_b = K\_w\). Smaller values \(pK\_a\) correspond to larger acidic ionization constants and thus stronger acids. Conversely, the smaller values \(pK\_b) correspond to the larger basic ionization constants and thus the stronger the bases. At 25°C\(pK\_a + pK\_b = 14.00\). Acid-base reactions always move in a direction that produces a weaker pair of acid and base. In an aqueous solution, there can be no acid stronger than \(H\_3O^+\) and no base stronger than \(OH^-\), resulting in a phenomenon known as a balancing effect. Polyprotic acids (and bases) lose (and acquire) protons gradually, with fully protoned species being the strongest acid and fully deprotonated species the strongest base. Key equations Acid ionization constant:  $|K a=K[H 20]=\frac{1}{(K a, b, conjugated acid and base pair: |K aK b = K w]} Definition (pK a): |pKa$ = -\log {10}K a \] \[K a=10^{-pK a} \Definition \(pK b\) : \[pK b = -\log {10}K b \] \[K b=10^{-pK b} \] Relationship between \(pK a\) and \(pK b\) conjugated acid and base pairs: \[pK a + pK b = pK w \] \[pK a + pK b = 14.00 \; \text{at 25°C} \] Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook Textbooks

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