

8.4: Acid Strength and the Acid Dissociation Constant (K_a)

Skills to Develop

- To know the relationship between acid or base strength and the magnitude of K_a , K_b , pK_a , and pK_b .
- To understand the leveling effect.

The magnitude of the equilibrium constant for an ionization reaction can be used to determine the relative strengths of acids and bases. For example, the general equation for the ionization of a weak acid in water, where HA is the parent acid and A^- is its conjugate base, is as follows:



The equilibrium constant for this dissociation is as follows:

$$K = \frac{[H_3O^+][A^-]}{[H_2O][HA]} \quad (8.4.2)$$

As we noted earlier, the equilibrium constant expression (Law of Mass Action) is really a ratio of activities. The activities of the solutes can be replaced with their molar concentration without too much error if the solutions are dilute enough. However, the activity of water is essentially constant for all reactions in dilute aqueous solutions, and is assigned the value of "1" by definition. Thus, the $[H_2O]$ in Equation 8.4.2 is actually a misapplication of the approximation carried out for solutes. The proper sequence of approximations and substitutions is

$$K = K_a = \frac{(a_{H_3O^+})(a_{A^-})}{(a_{HA})(a_{H_2O})} \approx \frac{[H_3O^+][A^-]}{[HA][1]} = \frac{[H_3O^+][A^-]}{[HA]} \quad (8.4.3)$$

Again, for simplicity, H_3O^+ can be written as H^+ in Equation 8.4.3. Keep in mind, though, that free H^+ does not exist in aqueous solutions and that a proton is transferred to H_2O in all acid ionization reactions to form **hydronium ions**, H_3O^+ . The larger the K_a , the stronger the acid and the higher the H^+ concentration at equilibrium. As with all equilibrium constants, acid-base ionization constants are unitless because each of the activities is divided by the standard activity with the same units. The values of K_a for a number of common acids are given in Table 8.4.1.

Table 8.4.1: Values of K_a , pK_a , K_b , and pK_b for Selected Acids (HA and Their Conjugate Bases (A^-))

Acid	HA	K_a	pK_a	A^-	K_b	pK_b
hydroiodic acid	HI	2×10^9	-9.3	I^-	5.5×10^{-24}	23.26
sulfuric acid (1)*	H_2SO_4	1×10^2	-2.0	HSO_4^-	1×10^{-16}	16.0
nitric acid	HNO_3	2.3×10^1	-1.37	NO_3^-	4.3×10^{-16}	15.37
hydronium ion	H_3O^+	1.0	0.00	H_2O	1.0×10^{-14}	14.00
sulfuric acid (2)*	HSO_4^-	1.0×10^{-2}	1.99	SO_4^{2-}	9.8×10^{-13}	12.01
hydrofluoric acid	HF	6.3×10^{-4}	3.20	F^-	1.6×10^{-11}	10.80
nitrous acid	HNO_2	5.6×10^{-4}	3.25	NO_2^-	1.8×10^{-11}	10.75
formic acid	HCO_2H	1.78×10^{-4}	3.750	HCO_2^-	5.6×10^{-11}	10.25
benzoic acid	$C_6H_5CO_2H$	6.3×10^{-5}	4.20	$C_6H_5CO_2^-$	1.6×10^{-10}	9.80
acetic acid	CH_3CO_2H	1.7×10^{-5}	4.76	$CH_3CO_2^-$	5.8×10^{-10}	9.24
pyridinium ion	$C_5H_5NH^+$	5.9×10^{-6}	5.23	C_5H_5N	1.7×10^{-9}	8.77
hypochlorous acid	HOCl	4.0×10^{-8}	7.40	OCl^-	2.5×10^{-7}	6.60
hydrocyanic acid	HCN	6.2×10^{-10}	9.21	CN^-	1.6×10^{-5}	4.79
ammonium ion	NH_4^+	5.6×10^{-10}	9.25	NH_3	1.8×10^{-5}	4.75

*The number in parentheses indicates the ionization step referred to for a polyprotic acid.

Acid	HA	K_a	pK_a	A^-	K_b	pK_b
water	H_2O	1.0×10^{-14}	14.00	OH^-	1.00	0.00
acetylene	C_2H_2	1×10^{-26}	26.0	HC_2^-	1×10^{12}	-12.0
ammonia	NH_3	1×10^{-35}	35.0	NH_2^-	1×10^{21}	-21.0

*The number in parentheses indicates the ionization step referred to for a polyprotic acid.

Weak bases react with water to produce the hydroxide ion, as shown in the following general equation, where B is the parent base and BH^+ is its conjugate acid:



The equilibrium constant for this reaction is the base ionization constant (K_b), also called the base dissociation constant:

$$K = K_b = \frac{[BH^+][OH^-]}{[B]} \quad (8.4.5)$$

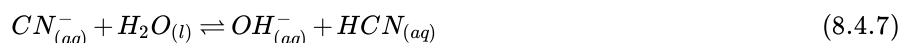
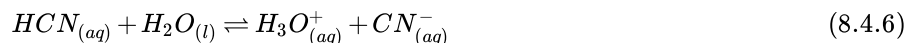
Once again, the activity of water is defined as a value of "1" because it is the solvent. The larger the K_b , the stronger the base and the higher the OH^- concentration at equilibrium. The values of K_b for a number of common weak bases are given in Table 8.4.2.

Table 8.4.2: Values of K_b , pK_b , K_a , and pK_a for Selected Weak Bases (B) and Their Conjugate Acids (BH^+)

Base	B	K_b	pK_b	BH^+	K_a	pK_a
hydroxide ion	OH^-	1.0	0.00*	H_2O	1.0×10^{-14}	14.00
phosphate ion	PO_4^{3-}	2.1×10^{-2}	1.68	HPO_4^{2-}	4.8×10^{-13}	12.32
dimethylamine	$(CH_3)_2NH$	5.4×10^{-4}	3.27	$(CH_3)_2NH_2^+$	1.9×10^{-11}	10.73
methylamine	CH_3NH_2	4.6×10^{-4}	3.34	$CH_3NH_3^+$	2.2×10^{-11}	10.66
trimethylamine	$(CH_3)_3N$	6.3×10^{-5}	4.20	$(CH_3)_3NH^+$	1.6×10^{-10}	9.80
ammonia	NH_3	1.8×10^{-5}	4.75	NH_4^+	5.6×10^{-10}	9.25
pyridine	C_5H_5N	1.7×10^{-9}	8.77	$C_5H_5NH^+$	5.9×10^{-6}	5.23
aniline	$C_6H_5NH_2$	7.4×10^{-10}	9.13	$C_6H_5NH_3^+$	1.3×10^{-5}	4.87
water	H_2O	1.0×10^{-14}	14.00	H_3O^+	1.0*	0.00

*As in Table 8.4.1.

There is a simple relationship between the magnitude of K_a for an acid and K_b for its conjugate base. Consider, for example, the ionization of hydrocyanic acid (HCN) in water to produce an acidic solution, and the reaction of CN^- with water to produce a basic solution:



The equilibrium constant expression for the ionization of HCN is as follows:

$$K_a = \frac{[H_3O^+][CN^-]}{[HCN]} \quad (8.4.8)$$

The corresponding expression for the reaction of cyanide with water is as follows:

$$K_b = \frac{[OH^-][HCN]}{[CN^-]} \quad (8.4.9)$$

If we add Equations 8.4.6 and 8.4.7, we obtain the following:

Reaction	Equilibrium Constants
$H_2O_{(l)} + HCN_{(aq)} \rightleftharpoons H_3O^+_{(aq)} + CN^-_{(aq)}$	$K_a = [H_3O^+] [CN^-] / [HCN]$
$CN^-_{(aq)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + HCN_{(aq)}$	$K_b = [OH^-] [HCN] / [CN^-]$
$2H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$	$K = K_a \times K_b = [H_3O^+][OH^-]$

In this case, the sum of the reactions described by K_a and K_b is the equation for the autoionization of water, and the product of the two equilibrium constants is K_w :

$$K_a K_b = K_w \quad (8.4.10)$$

Thus if we know either K_a for an acid or K_b for its conjugate base, we can calculate the other equilibrium constant for any conjugate acid–base pair.

We can use negative logarithms to avoid exponential notation in writing acid and base ionization constants, by defining pK_a as follows:

$$pK_a = -\log_{10} K_a \quad (8.4.11)$$

$$K_a = 10^{-pK_a} \quad (8.4.12)$$

and pK_b as

$$pK_b = -\log_{10} K_b \quad (8.4.13)$$

$$K_b = 10^{-pK_b} \quad (8.4.14)$$

Similarly, Equation 8.4.10, which expresses the relationship between K_a and K_b , can be written in logarithmic form as follows:

$$pK_a + pK_b = pK_w \quad (8.4.15)$$

At 25 °C, this becomes

$$pK_a + pK_b = 14.00 \quad (8.4.16)$$

The values of pK_a and pK_b are given for several common acids and bases in Tables 8.4.1 and 8.4.2, respectively, and a more extensive set of data is provided in Tables E1 and E2. Because of the use of negative logarithms, smaller values of pK_a correspond to larger acid ionization constants and hence stronger acids. For example, nitrous acid (HNO_2), with a pK_a of 3.25, is about a million times stronger acid than hydrocyanic acid (HCN), with a pK_a of 9.21. Conversely, smaller values of pK_b correspond to larger base ionization constants and hence stronger bases.

ACID		BASE	
negligible	OH^-	O^{2-}	strong
	HS^-	S^{2-}	
weak	H_2O	OH^-	weak
	HPO_4^{2-}	PO_4^{3-}	
	HCO_3^-	CO_3^{2-}	
	NH_4^+	NH_3	
	HCN	CN^-	
	H_2PO_4^-	HPO_4^{2-}	
	HSO_3^-	SO_3^{2-}	
	H_2S	HS^-	
	H_2CO_3	HCO_3^-	
	$\text{C}_5\text{H}_5\text{NH}^+$	$\text{C}_5\text{H}_5\text{N}$	
	$\text{CH}_3\text{CO}_2\text{H}$	CH_3CO_2^-	
	HF	F^-	
	H_3PO_4	H_2PO_4^-	
	H_2SO_3	HSO_3^-	
	HSO_4^-	SO_4^{2-}	
strong	H_3O^+	H_2O	negligible
	HNO_3	NO_3^-	
	H_2SO_4	HSO_4^-	
	HCl	Cl^-	
	HBr	Br^-	

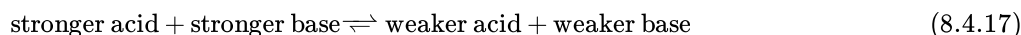
← Relative acid strength increasing Relative base strength increasing →

Figure 8.4.1: The Relative Strengths of Some Common Conjugate Acid–Base Pairs. The strongest acids are at the bottom left, and the strongest bases are at the top right. The conjugate base of a strong acid is a very weak base, and, conversely, the conjugate acid of a strong base is a very weak acid.

The relative strengths of some common acids and their conjugate bases are shown graphically in Figure 8.4.1. The conjugate acid–base pairs are listed in order (from top to bottom) of increasing acid strength, which corresponds to decreasing values of pK_a . This order corresponds to decreasing strength of the conjugate base or increasing values of pK_b . At the bottom left of Figure 8.4.2 are the common strong acids; at the top right are the most common strong bases. Notice the inverse relationship between the strength of the parent acid and the strength of the conjugate base. Thus the conjugate base of a strong acid is a very weak base, and the conjugate base of a very weak acid is a strong base.

The conjugate base of a strong acid is a weak base and vice versa.

We can use the relative strengths of acids and bases to predict the direction of an acid–base reaction by following a single rule: an acid–base equilibrium always favors the side with the weaker acid and base, as indicated by these arrows:

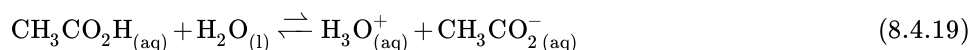


In an acid–base reaction, the proton always reacts with the stronger base.

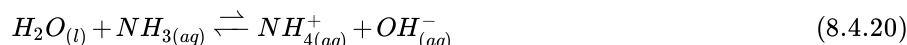
For example, hydrochloric acid is a strong acid that ionizes essentially completely in dilute aqueous solution to produce H_3O^+ and Cl^- ; only negligible amounts of HCl molecules remain undissociated. Hence the ionization equilibrium lies virtually all the way to the right, as represented by a single arrow:



In contrast, acetic acid is a weak acid, and water is a weak base. Consequently, aqueous solutions of acetic acid contain mostly acetic acid molecules in equilibrium with a small concentration of H_3O^+ and acetate ions, and the ionization equilibrium lies far to the left, as represented by these arrows:



Similarly, in the reaction of ammonia with water, the hydroxide ion is a strong base, and ammonia is a weak base, whereas the ammonium ion is a stronger acid than water. Hence this equilibrium also lies to the left:



All acid–base equilibria favor the side with the weaker acid and base. Thus the proton is bound to the stronger base.

Example 8.4.1: Butyrate and Dimethylammonium Ions

- Calculate K_b and pK_b of the butyrate ion ($\text{CH}_3\text{CH}_2\text{CH}_2\text{CO}_2^-$). The pK_a of butyric acid at 25°C is 4.83. Butyric acid is responsible for the foul smell of rancid butter.
- Calculate K_a and pK_a of the dimethylammonium ion ($(\text{CH}_3)_2\text{NH}_2^+$). The base ionization constant K_b of dimethylamine ($(\text{CH}_3)_2\text{NH}$) is 5.4×10^{-4} at 25°C.

Given: pK_a and K_b

Asked for: corresponding K_b and pK_b , K_a and pK_a

Strategy:

The constants K_a and K_b are related as shown in Equation 8.4.10. The pK_a and pK_b for an acid and its conjugate base are related as shown in Equations 8.4.15 and 8.4.16. Use the relationships $pK = -\log K$ and $K = 10^{-pK}$ (Equations 8.4.11 and 8.4.13) to convert between K_a and pK_a or K_b and pK_b .

Solution:

We are given the pK_a for butyric acid and asked to calculate the K_b and the pK_b for its conjugate base, the butyrate ion. Because the pK_a value cited is for a temperature of 25°C, we can use Equation 8.4.16: $pK_a + pK_b = pK_w = 14.00$. Substituting the pK_a and solving for the pK_b ,

$$4.83 + pK_b = 14.00 \quad (8.4.21)$$

$$pK_b = 14.00 - 4.83 = 9.17 \quad (8.4.22)$$

Because $pK_b = -\log K_b$, K_b is $10^{-9.17} = 6.8 \times 10^{-10}$.

In this case, we are given K_b for a base (dimethylamine) and asked to calculate K_a and pK_a for its conjugate acid, the dimethylammonium ion. Because the initial quantity given is K_b rather than pK_b , we can use Equation 8.4.10: $K_a K_b = K_w$. Substituting the values of K_b and K_w at 25°C and solving for K_a ,

$$K_a(5.4 \times 10^{-4}) = 1.01 \times 10^{-14} \quad (8.4.23)$$

$$K_a = 1.9 \times 10^{-11} \quad (8.4.24)$$

Because $pK_a = -\log K_a$, we have $pK_a = -\log(1.9 \times 10^{-11}) = 10.72$. We could also have converted K_b to pK_b to obtain the same answer:

$$pK_b = -\log(5.4 \times 10^{-4}) = 3.27 \quad (8.4.25)$$

$$pK_a + pK_b = 14.00 \quad (8.4.26)$$

$$pK_a = 10.73 \quad (8.4.27)$$

$$K_a = 10^{-pK_a} = 10^{-10.73} = 1.9 \times 10^{-11} \quad (8.4.28)$$

If we are given any one of these four quantities for an acid or a base (K_a , pK_a , K_b , or pK_b), we can calculate the other three.

Exercise 8.4.1: Lactic Acid

Lactic acid ($\text{CH}_3\text{CH}(\text{OH})\text{CO}_2\text{H}$) is responsible for the pungent taste and smell of sour milk; it is also thought to produce soreness in fatigued muscles. Its pK_a is 3.86 at 25°C. Calculate K_a for lactic acid and pK_b and K_b for the lactate ion.

Answer:

- $K_a = 1.4 \times 10^{-4}$ for lactic acid;
- $pK_b = 10.14$ and
- $K_b = 7.2 \times 10^{-11}$ for the lactate ion

Solutions of Strong Acids and Bases: The Leveling Effect

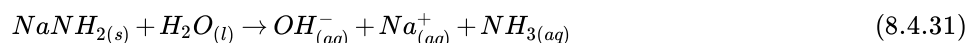
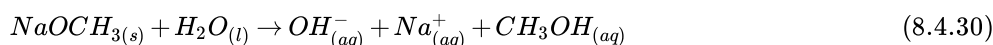
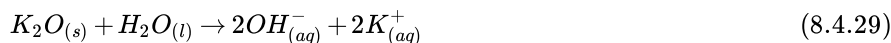
You will notice in Table 8.4.1 that acids like H_2SO_4 and HNO_3 lie above the hydronium ion, meaning that they have pK_a values less than zero and are stronger acids than the H_3O^+ ion. Recall from Chapter 4 that the acidic proton in virtually all oxoacids is bonded to one of the oxygen atoms of the oxoanion. Thus nitric acid should properly be written as $HONO_2$. Unfortunately, however, the formulas of oxoacids are almost always written with hydrogen on the left and oxygen on the right, giving HNO_3 instead. In fact, all six of the common strong acids that we first encountered in Chapter 4 have pK_a values less than zero, which means that they have a greater tendency to lose a proton than does the H_3O^+ ion. Conversely, the conjugate bases of these strong acids are weaker bases than water. Consequently, the proton-transfer equilibria for these strong acids lie far to the right, and adding any of the common strong acids to water results in an essentially stoichiometric reaction of the acid with water to form a solution of the H_3O^+ ion and the conjugate base of the acid.

Although K_a for HI is about 108 greater than K_a for HNO_3 , the reaction of either HI or HNO_3 with water gives an essentially stoichiometric solution of H_3O^+ and I^- or NO_3^- . In fact, a 0.1 M aqueous solution of any strong acid actually contains 0.1 M H_3O^+ , regardless of the identity of the strong acid. This phenomenon is called the leveling effect: any species that is a stronger acid than the conjugate acid of water (H_3O^+) is leveled to the strength of H_3O^+ in aqueous solution because H_3O^+ is the strongest acid that can exist in equilibrium with water. Consequently, it is impossible to distinguish between the strengths of acids such as HI and HNO_3 in aqueous solution, and an alternative approach must be used to determine their relative acid strengths.

One method is to use a solvent such as anhydrous acetic acid. Because acetic acid is a stronger acid than water, it must also be a weaker base, with a lesser tendency to accept a proton than H_2O . Measurements of the conductivity of 0.1 M solutions of both HI and HNO_3 in acetic acid show that HI is completely dissociated, but HNO_3 is only partially dissociated and behaves like a weak acid in this solvent. This result clearly tells us that HI is a stronger acid than HNO_3 . The relative order of acid strengths and approximate K_a and pK_a values for the strong acids at the top of Table 8.4.1 were determined using measurements like this and different nonaqueous solvents.

In aqueous solutions, H_3O^+ is the strongest acid and OH^- is the strongest base that can exist in equilibrium with H_2O .

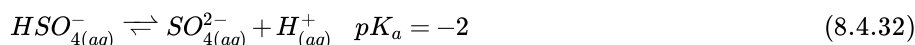
The leveling effect applies to solutions of strong bases as well: In aqueous solution, any base stronger than OH^- is leveled to the strength of OH^- because OH^- is the strongest base that can exist in equilibrium with water. Salts such as K_2O , $NaOCH_3$ (sodium methoxide), and $NaNH_2$ (sodamide, or sodium amide), whose anions are the conjugate bases of species that would lie below water in Table 8.4.2, are all strong bases that react essentially completely (and often violently) with water, accepting a proton to give a solution of OH^- and the corresponding cation:



Other examples that you may encounter are potassium hydride (KH) and organometallic compounds such as methyl lithium (CH_3Li).

Polyprotic Acids and Bases

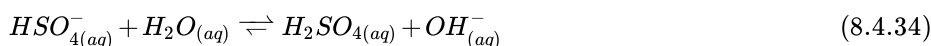
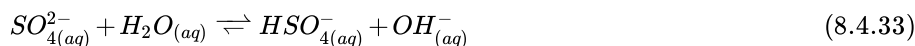
As you learned, polyprotic acids such as H_2SO_4 , H_3PO_4 , and H_2CO_3 contain more than one ionizable proton, and the protons are lost in a stepwise manner. The fully protonated species is always the strongest acid because it is easier to remove a proton from a neutral molecule than from a negatively charged ion. Thus acid strength decreases with the loss of subsequent protons, and, correspondingly, the pK_a increases. Consider H_2SO_4 , for example:



The equilibrium in the first reaction lies far to the right, consistent with H_2SO_4 being a strong acid. In contrast, in the second reaction, appreciable quantities of both HSO_4^- and SO_4^{2-} are present at equilibrium.

For a polyprotic acid, acid strength decreases and the pK_a increases with the sequential loss of each proton.

The hydrogen sulfate ion (HSO_4^-) is both the conjugate base of H_2SO_4 and the conjugate acid of SO_4^{2-} . Just like water, HSO_4^- can therefore act as either an acid or a base, depending on whether the other reactant is a stronger acid or a stronger base. Conversely, the sulfate ion (SO_4^{2-}) is a polyprotic base that is capable of accepting two protons in a stepwise manner:



Like any other conjugate acid–base pair, the strengths of the conjugate acids and bases are related by $pK_a + pK_b = pK_w$. Consider, for example, the HSO_4^-/SO_4^{2-} conjugate acid–base pair. From Table 8.4.1, we see that the pK_a of HSO_4^- is 1.99. Hence the pK_b of SO_4^{2-} is $14.00 - 1.99 = 12.01$. Thus sulfate is a rather weak base, whereas OH^- is a strong base, so the equilibrium shown in Equation 8.4.34 lies to the left. The HSO_4^- ion is also a very weak base (pK_a of $H_2SO_4 = 2.0$, pK_b of $HSO_4^- = 14 - (-2.0) = 16$), which is consistent with what we expect for the conjugate base of a strong acid.

Example 8.4.2

Predict whether the equilibrium for each reaction lies to the left or the right as written.

- $NH_{4(aq)}^+ + PO_{4(aq)}^{3-} \rightleftharpoons NH_{3(aq)} + HPO_{4(aq)}^{2-}$
- $CH_3CH_2CO_2H_{(aq)} + CN_{(aq)}^- \rightleftharpoons CH_3CH_2CO_2^-(aq) + HCN_{(aq)}$

Given: balanced chemical equation

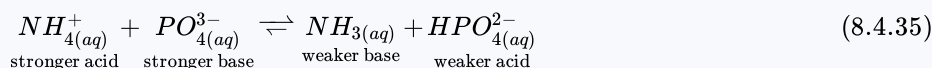
Asked for: equilibrium position

Strategy:

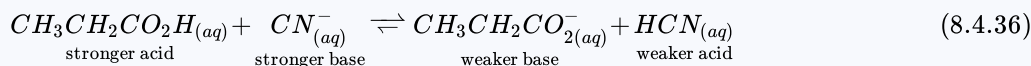
Identify the conjugate acid–base pairs in each reaction. Then refer to Tables 8.4.1 and 8.4.2 and Figure 8.4.2 to determine which is the stronger acid and base. Equilibrium **always** favors the formation of the weaker acid–base pair.

Solution:

The conjugate acid–base pairs are NH_4^+/NH_3 and HPO_4^{2-}/PO_4^{3-} . According to Tables 8.4.1 and 8.4.2, NH_4^+ is a stronger acid ($pK_a = 9.25$) than HPO_4^{2-} ($pK_a = 12.32$), and PO_4^{3-} is a stronger base ($pK_b = 1.68$) than NH_3 ($pK_b = 4.75$). The equilibrium will therefore lie to the right, favoring the formation of the weaker acid–base pair:



The conjugate acid–base pairs are $CH_3CH_2CO_2H/CH_3CH_2CO_2^-$ and HCN/CN^- . According to Table 8.4.1, HCN is a weak acid ($pK_a = 9.21$) and CN^- is a moderately weak base ($pK_b = 4.79$). Propionic acid ($CH_3CH_2CO_2H$) is not listed in Table 8.4.1, however. 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 the groups attached to the carbon atom of the carboxylic acid ($-CH_2CH_3$ versus $-CH_3$), so we might expect the two compounds to have similar acid–base properties. In particular, we would expect the pK_a of propionic acid to be similar in magnitude to the pK_a of acetic acid. (In fact, the pK_a of propionic acid is 4.87, compared to 4.76 for acetic acid, which makes propionic acid a slightly weaker acid than acetic acid.) Thus propionic acid should be a significantly stronger acid than HCN. Because the stronger acid forms the weaker conjugate base, we predict that cyanide will be a stronger base than propionate. The equilibrium will therefore lie to the right, favoring the formation of the weaker acid–base pair:



Exercise 8.4.1

Predict whether the equilibrium for each reaction lies to the left or the right as written.

- a. $H_2O_{(l)} + HS_{(aq)}^- \rightleftharpoons OH_{(aq)}^- + H_2S_{(aq)}$
 b. $HCO_{2(aq)}^- + HSO_{4(aq)}^- \rightleftharpoons HCO_{2H(aq)} + SO_{4(aq)}^{2-}$

Answer a

left

Answer b

left

Summary

Acid–base reactions always contain two conjugate acid–base pairs. Each acid and each base has an associated ionization constant that corresponds to its acid or base strength. Two species that differ by only a proton constitute a conjugate acid–base pair. The magnitude of the equilibrium constant for an ionization reaction can be used to determine the relative strengths of acids and bases. For an aqueous solution of a weak acid, the dissociation constant is called the acid ionization constant (K_a). Similarly, the equilibrium constant for the reaction of a weak base with water is the base ionization constant (K_b). For any conjugate acid–base pair, $K_a K_b = K_w$. Smaller values of pK_a correspond to larger acid ionization constants and hence stronger acids. Conversely, smaller values of pK_b correspond to larger base ionization constants and hence stronger bases. At 25°C, $pK_a + pK_b = 14.00$. Acid–base reactions always proceed in the direction that produces the weaker acid–base pair. No acid stronger than H_3O^+ and no base stronger than OH^- can exist in aqueous solution, leading to the phenomenon known as the leveling effect. Polyprotic acids (and bases) lose (and gain) protons in a stepwise manner, with the fully protonated species being the strongest acid and the fully deprotonated species the strongest base.

Key Equations

- Acid ionization constant:

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} \quad (8.4.37)$$

- Base ionization constant:

$$K_b = \frac{[BH^+][OH^-]}{[B]} \quad (8.4.38)$$

- Relationship between K_a and K_b of a conjugate acid–base pair:

$$K_a K_b = K_w \quad (8.4.39)$$

- Definition of pK_a :

$$pK_a = -\log_{10} K_a \quad (8.4.40)$$

$$K_a = 10^{-pK_a} \quad (8.4.41)$$

- Definition of pK_b :

$$pK_b = -\log_{10} K_b \quad (8.4.42)$$

$$K_b = 10^{-pK_b} \quad (8.4.43)$$

- Relationship between pK_a and pK_b of a conjugate acid–base pair:

$$pK_a + pK_b = pK_w \quad (8.4.44)$$

$$pK_a + pK_b = 14.00 \text{ at } 25^\circ\text{C} \quad (8.4.45)$$

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