

Lecture 8

Acids – Bases pH Calculations

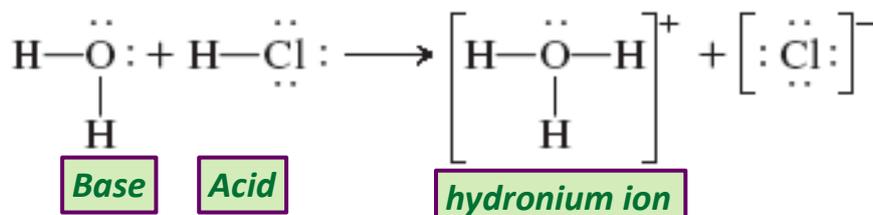
Acids and Bases

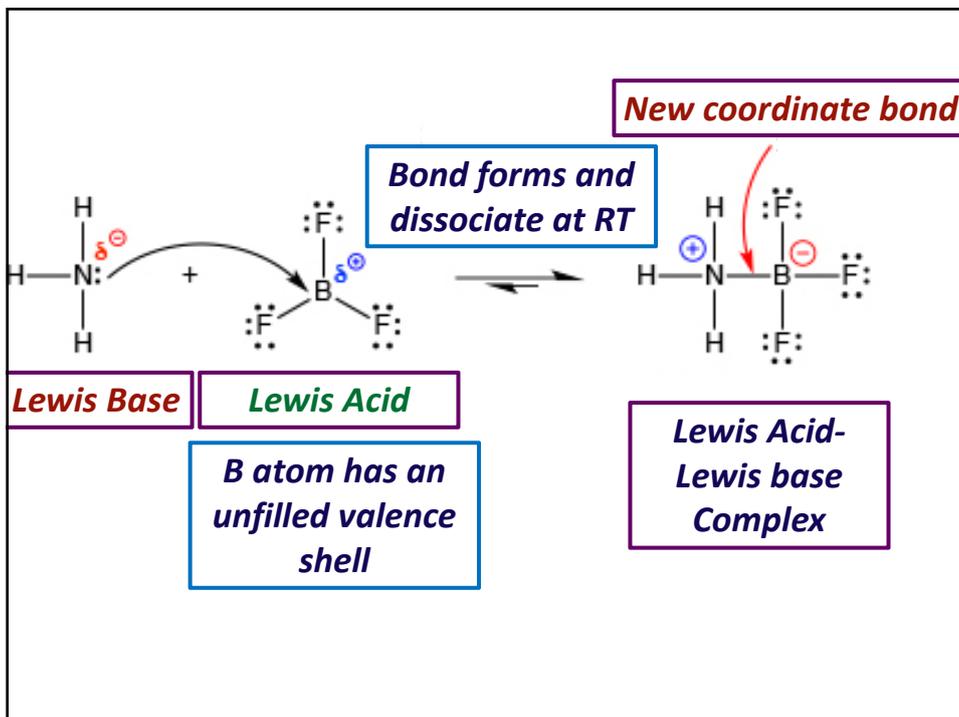
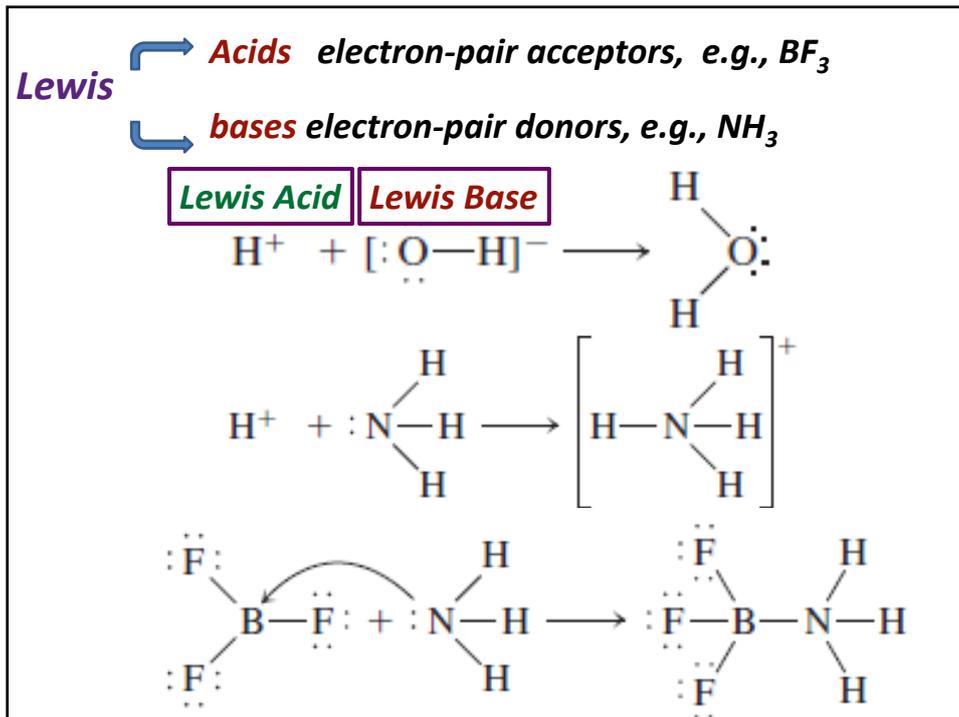
Arrhenius

- ↳ **acids** produce hydrogen ions in **aqueous** solutions, e.g., HCl, H₂SO₄
- ↳ **bases** produce hydroxide ions in **aqueous** solutions, e.g., KOH, NaOH

Brønsted–Lowry

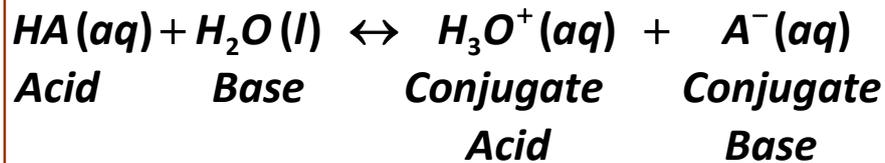
- ↳ **acids** Proton donors e.g., HCl, H₂SO₄
- ↳ **bases** Protons acceptors, e.g., ammonia NH₃





Acids and Bases		
<i>Model</i>	<i>Acid</i>	<i>Base</i>
<i>Arrhenius</i>	<i>H⁺ produce</i>	<i>OH⁻ produce</i>
<i>Brønsted–Lowry</i>	<i>H⁺ donor</i>	<i>H⁺ acceptor</i>
<i>Lewis</i>	<i>electron-pair acceptor</i>	<i>electron-pair donor</i>

Conjugate Acids and Bases		
<ul style="list-style-type: none"> When acids dissolve in water; 		
$ \begin{array}{ccccccc} HA(aq) & + & H_2O(l) & \leftrightarrow & H_3O^+(aq) & + & A^-(aq) \\ \text{Acid} & & \text{Base} & & \text{Conjugate} & & \text{Conjugate} \\ & & & & \text{Acid} & & \text{Base} \end{array} $		
<ul style="list-style-type: none"> The conjugate base is everything that remains of the acid molecule after a proton is lost. The conjugate acid is formed when the proton is transferred to the base. A conjugate acid–base pair consists of two substances related to each other by the donating and accepting of a single proton. The equation shown above contains two conjugate acid–base pairs: HA and A⁻ and H₂O and H₃O⁺. 		

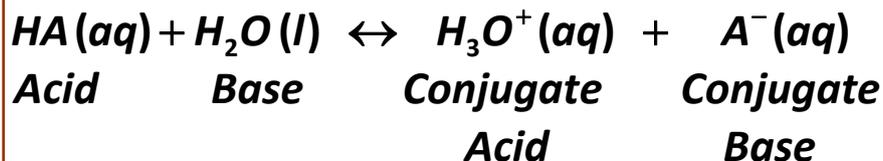


⚡ Notice the competition for the proton between the two bases H_2O and A^- .

⚡ If H_2O is a much stronger base than A^- (H_2O has a much greater affinity for H^+ than does A^-), the equilibrium position will be far to the right, i.e., most of the acid will dissolve (*dissociated*) at equilibrium.

⚡ If A^- is a much stronger base than H_2O , the equilibrium position will lie far to the left (most of the acid will remain *undissociated* as HA at equilibrium).

Acid dissociation constant, K_a

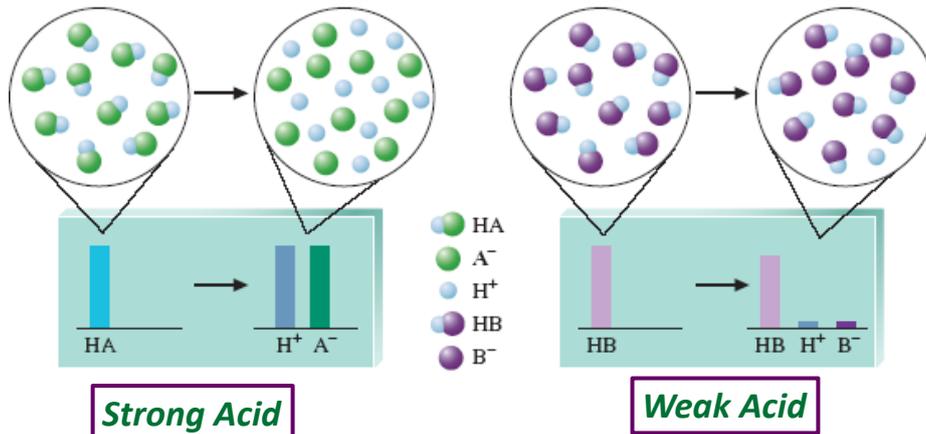


$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

▪ In a dilute solution we assume that the concentration of liquid water remains essentially constant when an acid is dissolved. Thus the term $[\text{H}_2\text{O}]$ is not included

Acid Strength

- is measured by the equilibrium position of its dissociation (ionization) reaction



The weaker the acid, the stronger its conjugate base

Strong Acids

- The dissociation equilibrium lies far to the right.
- Almost all the original HA is **dissociated (ionized)** at equilibrium.
- The dissociation produces a **weak conjugate base** that has a low affinity for a proton (is a much weaker base than water).

Weak Acids

- The dissociation equilibrium lies far to the left (i.e., **very small extent of dissociation**).
- Almost all the original HA remains undissociated (**not ionized**) at equilibrium.
- The dissociation produces a **strong conjugate base** that is much stronger base than water.

Acid Strength: Examples

Strong acids: sulfuric acid [$H_2SO_4(aq)$], hydrochloric acid [$HCl(aq)$], nitric acid [$HNO_3(aq)$], and perchloric acid [$HClO_4(aq)$].

Weak acids: phosphoric acid (H_3PO_4), nitrous acid (HNO_2), hypochlorous acid ($HOCl$), acetic acid (CH_3COOH), and benzoic acid (C_6H_5COOH).

K_a for some common monoprotic Acids

Formula	Name	Value of K_a^*
HSO_4^-	Hydrogen sulfate ion	1.2×10^{-2}
$HClO_2$	Chlorous acid	1.2×10^{-2}
$HC_2H_2ClO_2$	Monochloroacetic acid	1.35×10^{-3}
HF	Hydrofluoric acid	7.2×10^{-4}
HNO_2	Nitrous acid	4.0×10^{-4}
$HC_2H_3O_2$	Acetic acid	1.8×10^{-5}
$[Al(H_2O)_6]^{3+}$	Hydrated aluminum(III) ion	1.4×10^{-5}
$HOCl$	Hypochlorous acid	3.5×10^{-8}
HCN	Hydrocyanic acid	6.2×10^{-10}
NH_4^+	Ammonium ion	5.6×10^{-10}
$HO_6C_6H_5$	Phenol	1.6×10^{-10}

↑
Increasing acid strength

Exercise

- Arrange the following species according to their strengths as bases: H_2O , F^- , Cl^- , NO_2^- , and CN^- .

Solution

- Remember that water is a stronger base than the conjugate base of a strong acid but a weaker base than the conjugate base of a weak acid. This leads to the following order:

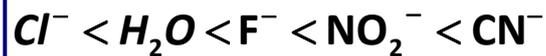


Weakest bases $\rightarrow\rightarrow\rightarrow$ Strongest bases

- Recognize that the strength of an acid is inversely related to the strength of its conjugate base.

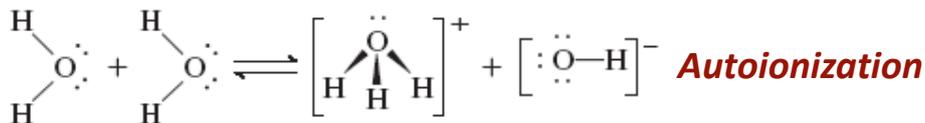


- The combined order of increasing base strength is

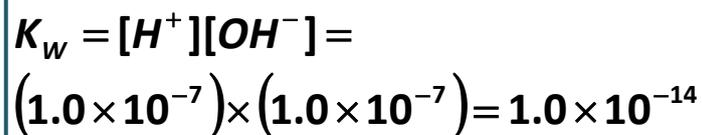
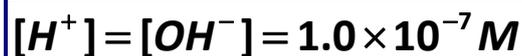


Water as an Acid and a Base

- Amphoteric substances** can behave either as an acid or as a base.



K_w : the **ion-product constant** (or the **dissociation constant for water**). Experiment shows that at 25°C in pure water,



Example

- Calculate $[H^+]$ or $[OH^-]$ as required for each of the following solutions at 25°C, and state whether the solution is neutral, acidic, or basic.

a) $1.0 \times 10^{-5} M OH^-$, b) $1.0 \times 10^{-7} M OH^-$, c) $10.0 M H^+$

- Solution

a) Basic

$$[H^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}} = 1.0 \times 10^{-9} M$$

b) Neutral

$$[H^+] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-7}} = 1.0 \times 10^{-7} M$$

c) Acidic

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H^+]} = \frac{1.0 \times 10^{-14}}{10.0} = 1.0 \times 10^{-15} M$$

The pH Scale

$$pK = -\log K$$

$$pH = -\log[H^+]$$

$$pOH = -\log[OH^-]$$

- At 25 °C $pH + pOH = 14$
- Thus for a solution where

$$[H^+] = 1.0 \times 10^{-7} M \quad pH = -(-7) = 7$$

- Calculate pH and pOH for $1.0 \times 10^{-3} M OH^-$ solution at 25°C?

$$[H^+] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-3}} = 1.0 \times 10^{-11} M$$

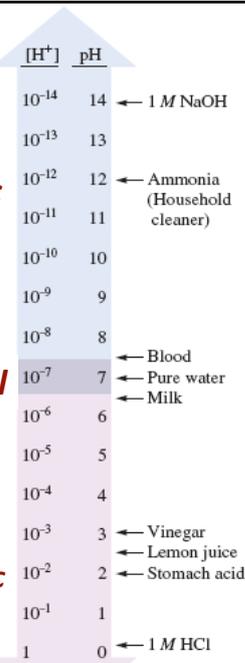
$$pH = -\log(1.0 \times 10^{-11}) = 11.00$$

$$pOH = -\log(1.0 \times 10^{-3}) = 14.00 - 11.00 = 3.00$$

Basic

Neutral

Acidic



pH of Strong Acid Solutions

- First: Write the major species in the solution.

Example: Calculate the pH of 1.0 M HCl?

Solution: List the major species: H^+ , Cl^- , and H_2O

- The major species that can produce H^+ are:

✓ the dissociation of HCl



✓ Autoionization of H_2O



- In pure H_2O at 25 °C, $[H^+] = 10^{-7}$ M.
- In 1.0 M HCl solution, the water will produce even less than 10^{-7} M H^+ (Le Châtelier's principle) (It can be ignored compared to 1.0 M H^+ from 1.0 M HCl).

$$pH = -\log [H^+] = -\log (1.0) = 0$$

Exercise

- Calculate the pH of 1.0×10^{-10} M HNO_3 ?

Solution

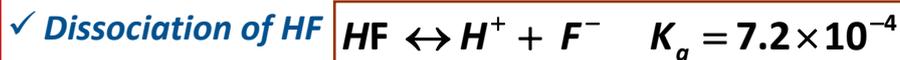
- List the major species: H^+ , NO_3^- , and H_2O
- The major species that can produce H^+ are:
 - ✓ the dissociation of HNO_3
 - ✓ Autoionization of H_2O
- In pure H_2O at 25 °C, $[H^+] = 10^{-7}$ M.
- In 1.0×10^{-10} M HNO_3 solution, the amount of HNO_3 in solution is so small that it has no effect; the only major species is H_2O .
- Thus the pH will be that of pure water, or **pH = 7.00**.

pH of Weak Acid Solutions

- Calculate the pH of a 1.00 M HF soln. ($K_a = 7.2 \times 10^{-4}$)?

Solution: List the major species: \longrightarrow HF, and H_2O

- The major species that can produce H^+ are:



- Compare K_a for HF and K_w for H_2O ,

✚ HF, *although weak*, is still a much stronger acid than water. Thus we will assume that HF will be the dominant source of H^+ . We will ignore the tiny contribution by water.

$$K_a = \frac{[H^+][F^-]}{[HF]} = 7.2 \times 10^{-4}$$

- Before dissociation of HF,

$$[HF]_0 = 1.0M, \quad [F^-]_0 = 0, \quad [H^+]_0 = 10^{-7} \approx 0$$

- Assume that *to reach equilibrium*, x mol/L HF will dissociate to produce x mol/L H^+ and x mol/L F^- .

- Equilibrium concentrations can be defined in terms of x :

$$[HF] = [HF]_0 - x = 1.0 - x,$$

$$[F^-] = [F^-]_0 + x = 0 + x = x,$$

$$[H^+] = [H^+]_0 + x \approx 0 + x = x$$

$$K_a = 7.2 \times 10^{-4} = \frac{[H^+][F^-]}{[HF]} = \frac{(x)(x)}{1.0 - x}$$

- As x is very small compared to 1.0, $1.0 - x \approx 1.0$

$$x^2 \approx (7.2 \times 10^{-4})(1.0)$$

$$x \approx \sqrt{7.2 \times 10^{-4}} = 2.7 \times 10^{-2}$$

$$x = [H^+] = 2.7 \times 10^{-2} \text{ mol / L}$$

$$pH = -\log[H^+] = -\log(2.7 \times 10^{-2}) = 1.57$$

Solving a quadratic equation of the general form

$$ax^2 + bx + c = 0$$

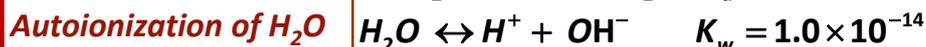
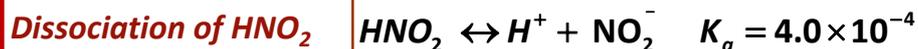
$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

The pH of Weak Acid Mixtures

Calculate the pH of a solution that contains 1.00 M HCN ($K_a = 6.2 \times 10^{-10}$) and 5.00 M HNO_2 ($K_a = 4.0 \times 10^{-4}$). Also calculate the concentration of cyanide ion (CN^-) in this solution at equilibrium?

Solution: List the major species: HCN , HNO_2 , and H_2O

▪ The major species that can produce H^+ are:



▪ Compare K_a for HCN, HNO_2 and K_w for H_2O ,

▪ HNO_2 , although weak, is still a much stronger acid than HCN and water. Thus we will assume that HNO_2 will be the dominant source of H^+ .

$$K_a = 4.0 \times 10^{-4} = \frac{[H^+][NO_2^-]}{[HNO_2]}$$

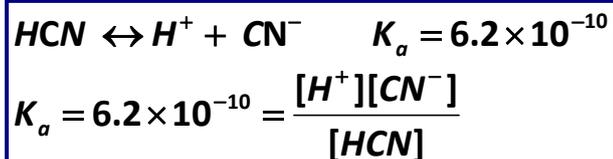
Initial Concentration (mol/L)	→	Equilibrium Concentration (mol/L)
$[HNO_2]_0 = 5.00$	x mol/L	$[HNO_2] = 5.00 - x$
$[NO_2^-]_0 = 0$	HNO_2	$[NO_2^-]_0 = x$
$[H^+]_0 \approx 0$	dissociates	$[H^+]_0 \approx x$

	HNO_2	\leftrightarrow	H^+	$+ NO_2^-$
Initial	5.00		0	0
Change	-x		+x	+x
Equilibrium	$5.00 - x$		x	x

$$K_a = 4.0 \times 10^{-4} = \frac{[H^+][NO_2^-]}{[HNO_2]} = \frac{(x)(x)}{5.0 - x} \approx \frac{x^2}{5.0}$$

$$x = 4.5 \times 10^{-2} = [H^+] \quad \text{pH} = 1.35$$

Calculation of equilibrium $[CN^-]$:



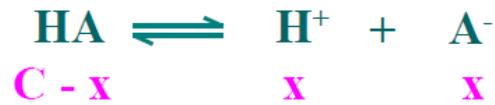
⚡ There is only one kind of H^+ in this solution. It does not matter from which acid the H^+ ions originate. $[H^+] = 4.5 \times 10^{-2} M$

⚡ Since K_a for HCN is so small, a negligible amount of HCN will dissociate.

$$K_a = 6.2 \times 10^{-10} = \frac{[H^+][CN^-]}{[HCN]} = \frac{4.5 \times 10^{-2} [CN^-]}{1.00}$$

$$[CN^-] = 1.4 \times 10^{-8} \text{ mol/L}$$

pH of Weak Acid Solutions



$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = \frac{\text{x} \cdot \text{x}}{\text{C} - \text{x}} = \frac{\text{x} \cdot \text{x}}{\text{C}}$$

$$K_a = \frac{[\text{H}^+]^2}{\text{C}}$$

$$[\text{H}^+]^2 = K_a \text{C} \qquad [\text{H}^+] = \sqrt{K_a \text{C}}$$

Percent Dissociation

$$\text{Percent dissociation} = \frac{\text{amount dissociated (mol / L)}}{\text{initial concentration (mol / L)}} \times 100 \%$$

- In a 1.00 M solution of HF, $[\text{H}^+] = 2.7 \times 10^{-2} \text{ M}$. To reach equilibrium, $2.7 \times 10^{-2} \text{ mol/L}$ of the original 1.00 M HF dissociates, so

$$\text{Percent dissociation} = \frac{2.7 \times 10^{-2} \text{ mol / L}}{1.00 \text{ mol / L}} \times 100 = 2.7 \%$$

For a given weak acid, the percent dissociation increases as the acid becomes more dilute.

Exercise

The pH of a 0.050 M weak acid is 3.00. What is the percentage ionization?

Solution

$$pH = 3.00 = -\log[H^+]$$

$$[H^+] = 0.001 \text{ mol / L}$$

$$\begin{aligned} \text{Percent dissociation} &= \frac{\text{amount dissociated (mol / L)}}{\text{initial concentration (mol / L)}} \times 100 \% \\ &= \frac{0.001 \text{ (mol / L)}}{0.050 \text{ (mol / L)}} \times 100 = 2.0 \% \end{aligned}$$

Homework

✚ Calculate the percent dissociation of acetic acid ($K_a = 1.8 \times 10^{-5}$) in 1.00 and 0.100 M $\text{HC}_2\text{H}_3\text{O}_2$?

✚ Lactic acid ($\text{HC}_3\text{H}_5\text{O}_3$) is a waste product that accumulates in muscle tissue during exertion, leading to pain and a feeling of fatigue. In a 0.100 M aqueous solution, lactic acid is 3.7% dissociated. Calculate the value of K_a for this acid?

Bases

- **Strong Bases:** are those dissociating completely when dissolved in aqueous solution, e.g., sodium hydroxide (NaOH) and potassium hydroxide (KOH).



- **Virtually** no undissociated NaOH left. Thus a 1.0 M NaOH solution really contains 1.0 M Na⁺ and 1.0 M OH⁻.

Example

- Calculate the pH of a 5.0×10^{-2} M NaOH solution?
- List the major species: **Na⁺, OH⁻, H₂O**

Although autoionization of water also produces OH⁻ ions, the pH will be dominated by the OH⁻ ions from the dissolved NaOH.

$$[\text{OH}^-] = 5.0 \times 10^{-2} \text{ mol/L}$$

$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{5.0 \times 10^{-2}} = 2.0 \times 10^{-13} \text{ mol/L}$$

$$\text{pH} = 12.70$$

- Note that this is a basic solution for which

$$[\text{OH}^-] > [\text{H}^+] \text{ and } \text{pH} > 7.0$$

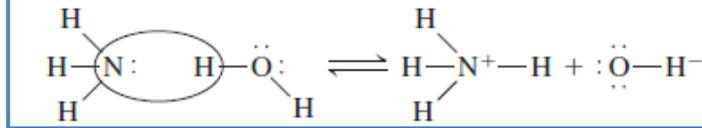
- The added OH⁻ from NaOH has shifted the water autoionization equilibrium



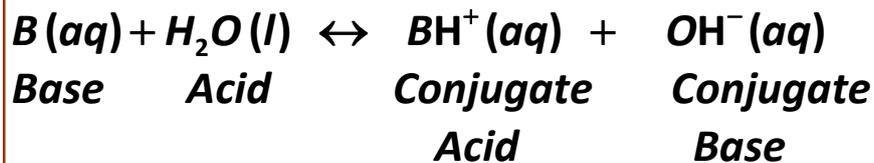
to left, significantly lowering [H⁺] compared with that in pure water.

A base does not have to contain OH^-

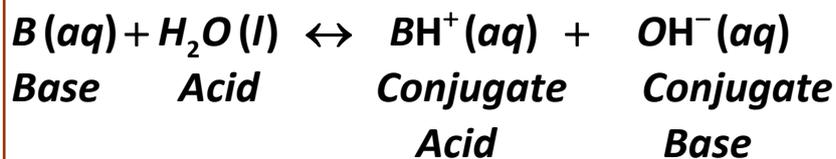
- Many bases do not contain the hydroxide ion. However, when dissolved in water $[\text{OH}^-]$ increases because of their reaction with water



- The general Reaction between a base B and water is given by



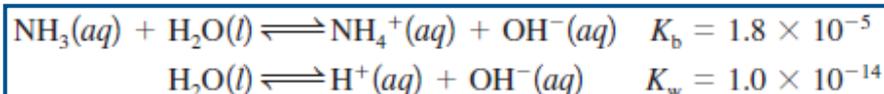
$$K_b = \frac{[BH^+][OH^-]}{[B]}$$



- Note the competition between the weak base, B , and the OH^- , which is a very strong base, for the H^+ ion. Thus their K_b values tend to be small.

Example: Calculate the pH for a 15.0 M solution of NH_3 ($K_b = 1.8 \times 10^{-5}$)?

- List the major species: NH_3 and H_2O (as K_b is very small)
- The major species that can produce OH^- are:



- The contribution from water can be neglected, since $K_b \gg K_w$

$$K_b = 1.8 \times 10^{-5} = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

Initial Concentration (mol/L)		Equilibrium Concentration (mol/L)
$[\text{NH}_3]_0 = 15.0$	x mol/L $\xrightarrow[\text{H}_2\text{O to reach equilibrium}]{\text{NH}_3 \text{ reacts with}}$	$[\text{NH}_3] = 15.0 - x$
$[\text{NH}_4^+]_0 = 0$		$[\text{NH}_4^+] = x$
$[\text{OH}^-]_0 \approx 0$		$[\text{OH}^-] = x$

$$K_b = 1.8 \times 10^{-5} = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{(x)(x)}{15.0 - x} \approx \frac{x^2}{15.0}$$

$$x \approx 1.6 \times 10^{-2}$$

$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{1.6 \times 10^{-2}} = 6.3 \times 10^{-13} \text{ M}$$

$$\text{pH} = -\log(6.3 \times 10^{-13}) = 12.20$$

Examples of weak bases

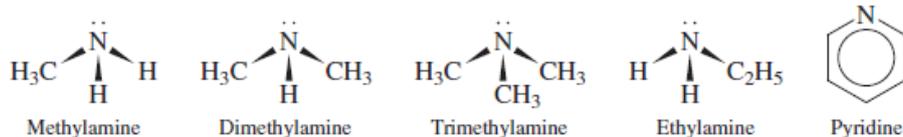


TABLE 14.3 > Values of K_b for Some Common Weak Bases

Name	Formula	Conjugate Acid	K_b
Ammonia	NH_3	NH_4^+	1.8×10^{-5}
Methylamine	CH_3NH_2	CH_3NH_3^+	4.38×10^{-4}
Ethylamine	$\text{C}_2\text{H}_5\text{NH}_2$	$\text{C}_2\text{H}_5\text{NH}_3^+$	5.6×10^{-4}
Aniline	$\text{C}_6\text{H}_5\text{NH}_2$	$\text{C}_6\text{H}_5\text{NH}_3^+$	3.8×10^{-10}
Pyridine	$\text{C}_5\text{H}_5\text{N}$	$\text{C}_5\text{H}_5\text{NH}^+$	1.7×10^{-9}

Exercise

Calculate $[H_3O^+]$, $[OH^-]$, and $[Ba^{2+}]$ in a 50.0 mL sample of 0.010 M $Ba(OH)_2$?

Solution

✚ $Ba(OH)_2$ is considered the sole source of OH^- .

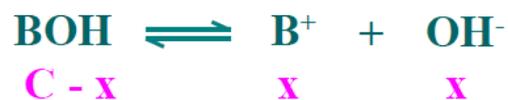


$$[Ba^{2+}] = 0.010 M \quad [OH^-] = 2 [Ba^{2+}] = 0.020 M$$

$$K_w = [H_3O^+] [OH^-] = [H_3O^+](0.020) = 1.0 \times 10^{-14}$$

$$[H_3O^+] = 5.0 \times 10^{-13} M$$

pH of Weak Bases Solutions



$$K_b = \frac{[B^+][OH^-]}{[BOH]} = \frac{x \cdot x}{C - x} = \frac{x \cdot x}{C}$$

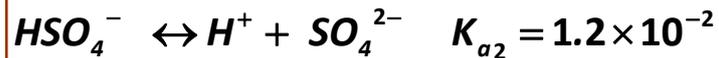
$$K_b = \frac{[OH^-]^2}{C}$$

$$[OH^-]^2 = K_b C \quad [OH^-] = \sqrt{K_b C}$$

$$pOH = -\log [OH^-] \quad pH = 14 - pOH$$

The pH of a Polyprotic Acid

- Sulfuric acid is unique among the common acids in that it is a strong acid in its first dissociation step and a weak acid in its second step:



Example : Calculate the pH of a 1.0 M H_2SO_4 solution.

Solution: The major species in the solution are

H^+ , HSO_4^- , and H_2O

$[H^+] = 1.0 + x$ (obtained from the dissociation of HSO_4^-)

- To calculate x

$$K_{a2} = 1.2 \times 10^{-2} = \frac{[H^+][SO_4^{2-}]}{[HSO_4^-]}$$

	$HSO_4^-(aq)$	\rightleftharpoons	$H^+(aq)$	+	$SO_4^{2-}(aq)$
Initial:	1.0		1.0		0
Change:	-x		+x		+x
Equilibrium:	1.0 - x		1.0 + x		x

Note that $[H^+]_0$ is not equal to zero, as it usually is for a weak acid, because the first dissociation step has already occurred.

$$K_{a2} = 1.2 \times 10^{-2} = \frac{[H^+][SO_4^{2-}]}{[HSO_4^-]} = \frac{(1.0 + x)(x)}{1.0 - x} \approx \frac{(1.0)(x)}{1.0} \approx x$$

$$[H^+] = 1.0 M \quad \text{pH} = 0.00$$

$$[H^+] = 1.0 M + x = 1.0 M + 1.2 \times 10^{-2} M \approx 1.0 M$$

Thus the dissociation of HSO_4^- does not contribute to $[H^+]$

(H) Calculate the pH of a 1.0×10^{-2} M H_2SO_4 solution ?

Acid-Base Properties of Salts

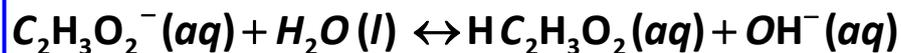
Acid-Base Equilibria

Salts Producing Neutral Solutions

- ***When a salt (ionic compound) dissolves in water, it breaks up into its ions, which move about independently, at least in dilute solutions.***
- ***Under certain conditions, these ions can behave as acids or bases.***
- ***Salts that consist of the cations of strong bases (K^+ , Na^+) and the anions of strong acids (Cl^- , NO_3^-) have no effect on $[H^+]$ when dissolved in water***
- ***Aqueous solutions of salts such as KCl , $NaCl$, $NaNO_3$, and KNO_3 are neutral (have a pH of 7).***

Salts That Produce Basic Solutions

- Salts whose *cations have neutral properties* (such as K^+ , Na^+) and whose *anions are the conjugate bases of weak acids* always produce basic aqueous solutions.
- In an aqueous solution of sodium acetate ($NaC_2H_3O_2$), The pH of this solution will be determined by the $C_2H_3O_2^-$ ion (*strong conjugate base of acetic acid*)



$$K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]}$$

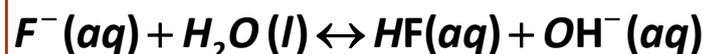
- What is the relationship between K_b and K_a of acetic acid?

$$K_a \times K_b = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} \times \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]} \\ = [H^+] \times [OH^-] = 1.0 \times 10^{-14} = K_w$$

- Calculate the pH of a 0.30 M NaF solution? The K_a value for HF is 7.2×10^{-4} .

Solution: Major species in solution are Na^+ , F^- , and H_2O

- Since HF is a weak acid, the F^- ion must have a significant affinity for protons, and the dominant reaction will be



$$K_b = \frac{[HF][OH^-]}{[F^-]} = \frac{K_w}{K_a(HF)} = \frac{1.0 \times 10^{-14}}{7.2 \times 10^{-4}} = 1.4 \times 10^{-11}$$

	$F^-(aq)$	$+ H_2O(l)$	\leftrightarrow	$HF(aq)$	$+ OH^-(aq)$
Initial	0.30	-		0	≈ 0
Change	-x	-		+x	+x
Equilibrium	$0.30 - x$	-		x	x

$$K_b = 1.4 \times 10^{-11} = \frac{[HF][OH^-]}{[F^-]} = \frac{(x)(x)}{0.3 - x} \approx \frac{x^2}{0.3}$$

$$x \approx 2.0 \times 10^{-6} = [OH^-]$$

$$pOH = 5.69$$

$$pH = 14.00 - 5.69 = 8.31$$

As expected, the solution is basic

Salts Producing Acidic Solutions

- Salts whose anion is a weak base (Cl^- , NO_3^-) and whose cation is a conjugate acid of a weak base (such as NH_4^+) produce acidic solutions
- Calculate the pH of a 0.10 M NH_4Cl solution. The K_b value for NH_3 is 1.8×10^{-5} .

Solution: Major species in solution are NH_4^+ , Cl^- , and H_2O

- Note that both NH_4^+ and H_2O can produce H^+ . The dissociation reaction for the NH_4^+ ion is



$$K_a = \frac{[NH_3][H^+]}{[NH_4^+]}$$

$$K_a(\text{for } \text{NH}_4^+) = \frac{K_w}{K_b(\text{for } \text{NH}_3)} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

Although NH_4^+ is a very weak acid, as indicated by its K_a value, it is stronger than H_2O and will dominate in the production of H^+ . Thus we will focus on the dissociation reaction of NH_4^+ to calculate the pH in this solution.

	$\text{NH}_4^+(\text{aq})$	\leftrightarrow	$\text{H}^+(\text{aq}) +$	$\text{NH}_3(\text{aq})$
Initial	0.10		≈ 0	0
Change	-x		+x	+x
Equilibrium	$0.10 - x$		x	x

$$K_a = 5.6 \times 10^{-10} = \frac{[\text{H}^+][\text{NH}_3]}{[\text{NH}_4^+]} = \frac{(x)(x)}{0.10 - x} \approx \frac{x^2}{0.10}$$

$$x \approx 7.5 \times 10^{-6} = [\text{H}^+] \quad \text{pH} = 5.13 \quad \text{As expected, the solution is acidic}$$

Salts whose cations and anions influence pH

- We can predict whether the solution will be basic, acidic, or neutral by comparing the K_a value for the acidic ion with the K_b value for the basic ion.

$K_a > K_b$	$\text{pH} < 7$	(acidic)
$K_a < K_b$	$\text{pH} > 7$	(basic)
$K_a = K_b$	$\text{pH} = 7$	(neutral)

Example

Predict whether an aqueous solution of $\text{NH}_4\text{C}_2\text{H}_3\text{O}_2$ salt will be acidic, basic, or neutral.

Solution

The ions in solution are NH_4^+ and $\text{C}_2\text{H}_3\text{O}_2^-$

$$K_a \text{ for } \text{NH}_4^+ = 5.6 \times 10^{-10}$$

$$K_b \text{ for } \text{C}_2\text{H}_3\text{O}_2^- = 5.6 \times 10^{-10}$$

$$K_a \text{ for } \text{NH}_4^+ = K_b \text{ for } \text{C}_2\text{H}_3\text{O}_2^-$$

The solution will be neutral (pH = 7).

Solutions of Acids Containing a Common Ion

Consider a solution containing a weak acid 1.0 M HF ($K_a = 7.2 \times 10^{-4}$) and 1.0 M NaF.

Recall

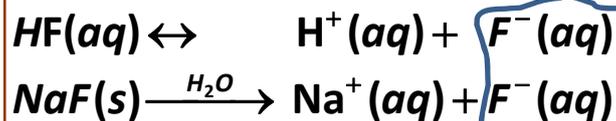


List the major species in solution: Na^+ , F^- , HF, and H_2O

The common ion in this solution is F^- .

How does F^- influence the dissociation equilibrium of HF?

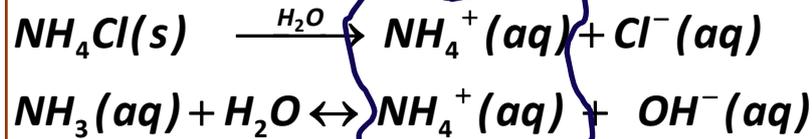
- Apply Le Châtelier's principle



The common ion effect makes a solution of NaF and HF less acidic than a solution of HF alone.

Solutions of Bases Containing a Common Ion

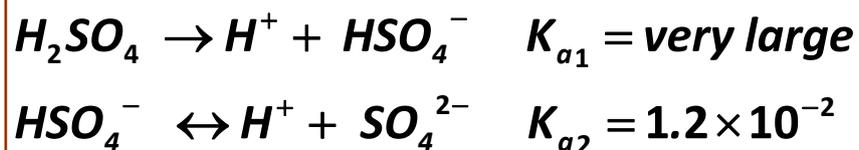
- Consider the addition of solid NH_4Cl to a 1.0 M NH_3 solution



The position of the ammonia–water equilibrium shifts to the left reducing the equilibrium concentration of OH^- ions.

Common Ion effect in Polyprotic acids

- The common ion effect is important in solutions of **polyprotic acids** (e.g., H_2SO_4 , H_3PO_4).
- The production of protons by the first dissociation step **greatly inhibits** the succeeding dissociation steps, which, of course, also produce protons, the common ion in this case

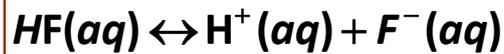


Acidic Solutions Containing Common Ions

The equilibrium concentration of H^+ in a 1.0 M HF solution is 2.7×10^{-2} M, and the percent dissociation of HF is 2.7%. Calculate $[H^+]$ and the percent dissociation of HF in a solution containing 1.0 M HF ($K_a = 7.2 \times 10^{-4}$) and 1.0 M NaF.

Solution: Major species are Na^+ , F^- , HF, and H_2O

- Note that Na^+ ions have neither acidic nor basic properties and that water is a very weak acid (or base). Therefore, the important species are HF and F^- .



$$K_a = \frac{[H^+][F^-]}{[HF]} = 7.2 \times 10^{-4}$$

Initial Conc. (mol/L)	Change (mol/L)	Eq. Conc. (mol/L)
$[HF]_0 = 1.0$ <i>From dissolved HF</i>	x mol/L HF dissociates	$[HF] = 1.0 - x$
$[F^-]_0 = 1.0$ <i>From dissolved NaF</i>		$[F^-] = 1.0 + x$
$[H^+]_0 = 0$ <i>Neglect contribution from H_2O</i>		$[H^+]_0 = x$

$$K_a = 7.2 \times 10^{-4} = \frac{[H^+][F^-]}{[HF]} = \frac{(x)(1.0 + x)}{(1.0 - x)} \approx x = [H^+]$$

$$[H^+] = 7.2 \times 10^{-4} \Rightarrow pH = 3.14$$

The percent dissociation of HF in this solution is

$$\frac{[H^+]}{[HF]_0} \times 100 = \frac{7.2 \times 10^{-4} M}{1.0 M} \times 100 = 0.072\%$$

As expected, F^- ions inhibited the dissociation of HF and position of acid dissociation equilibrium is shifted to the left

Buffered Solutions

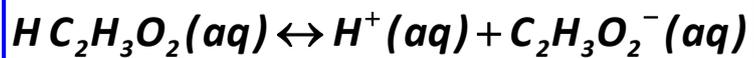
- **A buffered solution is that resisting a change in its pH when either hydroxide ions or protons are added**
- **Buffered systems are especially important in living systems, which can survive only in a relatively narrow pH range.**
- **For example, although human blood contains many buffering systems, the most important of these consists of a mixture of carbonic acid (0.0012 M) and bicarbonate ion (0.024 M). These concentrations produce a pH of 7.4 for normal blood.**
- **Because our cells are so sensitive to pH, it is important that this pH value be maintained.**
- **When reactions occur in our bodies, such as the formation of lactic acid ($HC_3H_5O_3$) when our muscles are exerted, the buffering systems must be capable of neutralizing the effects of this acid to maintain the pH at 7.4.**

A buffered solution may contain a weak acid and its salt (for example, HF and NaF) or a weak base and its salt (for example, NH_3 and NH_4Cl). By choosing the appropriate components, a solution can be buffered at virtually any pH.

Example

- A buffered solution contains 0.50 M acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$, $K_a = 1.8 \times 10^{-5}$) and 0.50 M sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$). Calculate the pH of this solution?

Solution: Major species in solution are



$$K_a = 1.8 \times 10^{-5} = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

Initial Concentration (mol/L)	Equilibrium Concentration (mol/L)
$[\text{HC}_2\text{H}_3\text{O}_2]_0 = 0.50$ $[\text{C}_2\text{H}_3\text{O}_2^-]_0 = 0.50$ $[\text{H}^+]_0 \approx 0$	$[\text{HC}_2\text{H}_3\text{O}_2] = 0.50 - x$ $[\text{C}_2\text{H}_3\text{O}_2^-] = 0.50 + x$ $[\text{H}^+] = x$

x mol/L of $\text{HC}_2\text{H}_3\text{O}_2$ dissociates to reach equilibrium

$$K_a = 1.8 \times 10^{-5} = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} = \frac{(x)(0.5 + x)}{(0.5 - x)} \approx x$$

$$[\text{H}^+] = 1.8 \times 10^{-5} \Rightarrow \text{pH} = 4.74$$

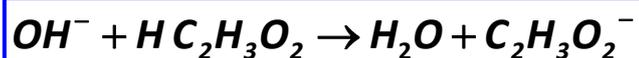
Calculate the change in pH that occurs when 0.010 mole of solid NaOH is added to 1.0 L of the abovementioned buffered solution. Compare this pH change with that which occurs when 0.010 mole of solid NaOH is added to 1.0 L water.

Solution

✚ NaOH will completely dissociate, so the major species in solution before any reaction occurs are



✚ Note that the solution contains a relatively large amount of the very strong base hydroxide ion, which has a great affinity for protons. The best source of protons is the acetic acid, and the reaction that will occur is



✚ Although acetic acid is a **weak acid**, the hydroxide ion is such a **strong base** that the reaction above will proceed essentially to **completion** (until the OH^- ions are consumed).

✚ The best approach to this problem involves two distinct steps:

- ✚ assume that the reaction goes to completion, and carry out the stoichiometric calculations, and then
- ✚ carry out the equilibrium calculations.

1. The stoichiometry problem.

	$\text{HC}_2\text{H}_3\text{O}_2$	+	OH^-	\longrightarrow	$\text{C}_2\text{H}_3\text{O}_2^-$	+	H_2O
Before	$1.0 \text{ L} \times 0.50 \text{ M}$		0.010 mol		$1.0 \text{ L} \times 0.50 \text{ M}$		
reaction:	$= 0.50 \text{ mol}$				$= 0.50 \text{ mol}$		
After	$0.50 - 0.010$		$0.010 - 0.010$		$0.50 + 0.010$		
reaction:	$= 0.49 \text{ mol}$		$= 0 \text{ mol}$		$= 0.51 \text{ mol}$		

2. The equilibrium problem. After the reaction between OH^- and $\text{HC}_2\text{H}_3\text{O}_2$ is complete, the major species in solution are



▪ The dominant equilibrium involves the dissociation of acetic acid and this will be treated as the previous example

	$\text{HC}_2\text{H}_3\text{O}_2(\text{aq})$	\rightleftharpoons	$\text{H}^+(\text{aq})$	+	$\text{C}_2\text{H}_3\text{O}_2^-(\text{aq})$
Initial:	0.49		0		0.51
Change:	$-x$		$+x$		$+x$
Equilibrium:	$0.49 - x$		x		$0.51 + x$

$$K_a = 1.8 \times 10^{-5} = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} = \frac{(x)(0.51 + x)}{0.49 - x} \approx \frac{(x)(0.51)}{0.49}$$

$$x \approx 1.7 \times 10^{-5}$$

$$[\text{H}^+] = x = 1.7 \times 10^{-5} \quad \text{and} \quad \text{pH} = 4.76$$

- The change in pH produced by the addition of 0.01 mole of OH^- to this buffered solution is then

$$\begin{array}{ccccccc} 4.76 & - & 4.74 & = & +0.02 \\ \uparrow & & \uparrow & & \\ \text{New solution} & & \text{Original solution} & & \end{array}$$

- The pH increased by 0.02 pH units.

what happens if 0.01 mole of solid NaOH is added to 1.0 L water.

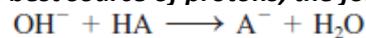
$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-2}} = 1.0 \times 10^{-12} \quad \text{pH} = 12.00$$

$$\begin{array}{ccccccc} \text{Change in pH} = & 12.00 & - & 7.00 & = & +5.00 \\ & \uparrow & & \uparrow & & \\ & \text{New solution} & & \text{Pure water} & & \end{array}$$

Buffering: How Does It Work?

Consider the addition of OH^- to a buffered solution containing relatively large quantities of a weak acid HA and its conjugate base A^- .

- HA represents the best source of protons, the following reaction occurs :



OH^- ions are not allowed to accumulate but are replaced by A^- ions

- The stability of the pH under these conditions can be understood by examining the equilibrium expression for the dissociation of HA:

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} \quad \longrightarrow \quad [\text{H}^+] = K_a \frac{[\text{HA}]}{[\text{A}^-]}$$

- The equilibrium concentration of H^+ , and thus the pH, is determined by the ratio $[\text{HA}]/[\text{A}^-]$.
- When OH^- ions are added, HA is converted to A^- , and the ratio $[\text{HA}]/[\text{A}^-]$ decreases.

- However, if the amounts of HA and A⁻ originally present are very large compared with the amount of OH⁻ added, the change in the [HA]/[A⁻] ratio will be small.

$$[H^+] = K_a \frac{[HA]}{[A^-]}$$

- In the previous example $\frac{[HA]}{[A^-]} = \frac{0.50}{0.50} = 1.0$ Initially

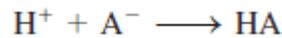
$$\frac{[HA]}{[A^-]} = \frac{0.49}{0.51} = 0.96 \quad \text{After adding 0.01 mol/L OH}^-$$

- The change in the ratio [HA]/[A⁻] is very small. Thus the [H⁺] and the pH remain essentially constant.

The essence of buffering is that [HA] and [A⁻] are large compared with the amount of OH⁻ added

- If **protons** are added to a buffered solution of a weak acid and a salt of its conjugate base:

- ✓ Because the A⁻ ion has a high affinity for H⁺, the added H⁺ ions react with A⁻ to form the weak acid:



If [A⁻] and [HA] are large compared with the [H⁺] added, little change in pH will occur

Henderson–Hasselbalch equation

- It is used to calculate the pH of buffered solutions when the ratio [HA]/[A⁻] is known.

Example : Calculate [H⁺] in a buffered solution containing 0.10 M HF (K_a = 7.2 × 10⁻⁴) and 0.30 M NaF

Solution

$$[H^+] = K_a \frac{[HA]}{[A^-]} = (7.2 \times 10^{-4}) \times \frac{0.10}{0.30} = 2.4 \times 10^{-4} \text{ mol / L}$$

$$-\log[H^+] = -\log(K_a) - \log \frac{[HA]}{[A^-]}$$

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$$pH = pK_a + \log \frac{[base]}{[acid]}$$

- For a particular buffering system (conjugate acid–base pair), all solutions that have the same ratio $[A^-]/[HA]$ will have the same pH.

System	$[A^-]/[HA]$
5.0 M $HC_2H_3O_2$ and 3.0 M $NaC_2H_3O_2$	$\frac{3.0 M}{5.0 M} = 0.60$
0.050 M $HC_2H_3O_2$ and 0.030 M $NaC_2H_3O_2$	$\frac{0.030 M}{0.050 M} = 0.60$

$$pH = pK_a + \log\left(\frac{[C_2H_3O_2^-]}{[HC_2H_3O_2]}\right) = 4.74 + \log(0.60) = 4.74 - 0.22 = 4.52$$

- In all these calculations, we are assuming the validity of the approximations

$$[A^-] = [A^-]_0 + x \approx [A^-]_0 \text{ and } [HA] = [HA]_0 - x \approx [HA]_0$$

Example

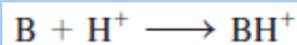
Calculate the pH of a solution containing 0.75 M lactic acid ($K_a = 1.4 \times 10^{-4}$) and 0.25 M sodium lactate?

Solution

$$pH = pK_a + \log\left(\frac{[C_3H_5O_3^-]}{[HC_3H_5O_3]}\right) = 3.85 + \log\left(\frac{0.25 M}{0.75 M}\right) = 3.38$$

Buffered solutions also can be formed from a weak base and the corresponding conjugate acid.

In these solutions, the weak base B reacts with any H^+ added:



and the conjugate acid BH^+ reacts with any added OH^- :

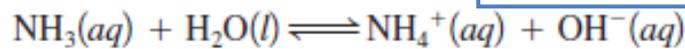


The approach needed to perform pH calculations for these systems is virtually identical to that used above.

Calculate the pH of a buffered solution containing 0.25 M NH_3 ($K_b = 1.8 \times 10^{-5}$) and 0.40 M NH_4Cl ?

- The major species in solution are NH_3 , NH_4^+ , Cl^- , and H_2O
From the dissolved NH_4Cl
- Since Cl^- is such a weak base and water is a weak acid or base, the important equilibrium is

$$K_b = 1.8 \times 10^{-5} = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$



	$\text{NH}_3(\text{aq})$	+	$\text{H}_2\text{O}(\text{l})$	\rightleftharpoons	$\text{NH}_4^+(\text{aq})$	+	$\text{OH}^-(\text{aq})$
Initial:	0.25		—		0.40		≈ 0
Change:	$-x$		—		$+x$		$+x$
Equilibrium:	$0.25 - x$		—		$0.40 + x$		x

$$K_b = 1.8 \times 10^{-5} = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{(0.40 + x)(x)}{0.25 - x} \approx \frac{(0.40)(x)}{0.25} \quad x \approx 1.1 \times 10^{-5}$$

$$[\text{OH}^-] = x = 1.1 \times 10^{-5} \quad \text{pOH} = 4.95 \quad \text{pH} = 14.00 - 4.95 = 9.05$$

Alternatively,

- We can use the dissociation equilibrium for NH_4^+ to calculate $[\text{H}^+]$ directly, that is,



$$K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

- Then, using the Henderson–Hasselbalch equation, we have

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{base}]}{[\text{acid}]}\right)$$

$$= 9.25 + \log\left(\frac{0.25 \text{ M}}{0.40 \text{ M}}\right) = 9.25 - 0.20 = 9.05$$

Adding a Strong Acid to a Buffered Solution

Example

- Calculate the pH of the solution that results when 0.10 mole of gaseous HCl is added to 1.0 L of a buffered solution containing 0.25 M NH₃ ($K_b = 1.8 \times 10^{-5}$) and 0.40 M NH₄Cl?

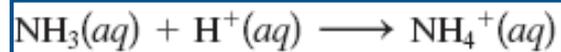
Solution

The major species in solution before any reaction occurs are



From added HCl

- ✓ H⁺ will not react with Cl⁻ to form HCl.
- ✓ NH₃ will react with H⁺ to form NH₄⁺



This Rx is assumed to go essentially to completion. Hence, we will do the stoichiometry calculations before the equilibrium calculations.

	NH ₃	+	H ⁺	→	NH ₄ ⁺
Before reaction:	(1.0 L)(0.25 M) = 0.25 mol		0.10 mol		(1.0 L)(0.40 M) = 0.40 mol
			↑ Limiting reactant		
After reaction:	0.25 - 0.10 = 0.15 mol		0		0.40 + 0.10 = 0.50 mol

The major species in solution the reaction goes to completion



$$[\text{NH}_3]_0 = \frac{0.15 \text{ mol}}{1.0 \text{ L}} = 0.15 \text{ M} \quad [\text{NH}_4^+]_0 = \frac{0.50 \text{ mol}}{1.0 \text{ L}} = 0.50 \text{ M}$$

We can use the Henderson-Hasselbalch equation, where

$$[\text{Base}] = [\text{NH}_3] \approx [\text{NH}_3]_0 = 0.15 \text{ M}$$

$$[\text{Acid}] = [\text{NH}_4^+] \approx [\text{NH}_4^+]_0 = 0.50 \text{ M}$$

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log\left(\frac{[\text{NH}_3]}{[\text{NH}_4^+]}\right) \\ &= 9.25 + \log\left(\frac{0.15 \text{ M}}{0.50 \text{ M}}\right) = 9.25 - 0.52 = 8.73 \end{aligned}$$

The addition of HCl only slightly decreases the pH, as we would expect in a buffered solution.

Buffering Capacity

- The **buffering capacity** of a buffered solution represents the amount of protons or hydroxide ions the buffer can absorb without a significant change in pH.
- A buffer containing large concentrations of buffering components will have a large buffering capacity.
- The pH of a buffered solution is determined by the ratio $[A^-]/[HA]$.
- The capacity of a buffered solution is determined by the magnitudes of $[HA]$ and $[A^-]$.

Example: Calculate the change in pH that occurs when 0.010 mole of gaseous HCl is added to 1.0 L of each of the following solutions (K_a for acetic acid = 1.8×10^{-5}):

- **Solution A:** 5.00 M $HC_2H_3O_2$ and 5.00 M $NaC_2H_3O_2$
- **Solution B:** 0.050 M $HC_2H_3O_2$ and 0.050 M $NaC_2H_3O_2$

Solution

- For both solutions the initial pH can be determined from the Henderson–Hasselbalch equation:

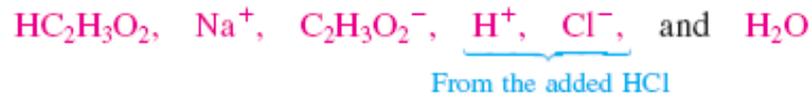
$$pH = pK_a + \log\left(\frac{[C_2H_3O_2^-]}{[HC_2H_3O_2]}\right)$$

$$[C_2H_3O_2^-] = [HC_2H_3O_2]$$

$$pH = pK_a + \log(1) = pK_a = -\log(1.8 \times 10^{-5}) = 4.74$$

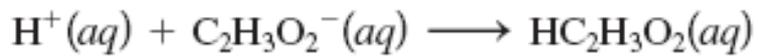
After adding HCl

The major species in solution before any reaction occurs are



✓ H^+ will not react with Cl^- to form HCl .

✓ $\text{C}_2\text{H}_3\text{O}_2^-$ will react with H^+ to form $\text{HC}_2\text{H}_3\text{O}_2$



Because $\text{HC}_2\text{H}_3\text{O}_2$ is a weak acid, the Rx is assumed to go to completion. Hence, we will do the stoichiometry calculations before the equilibrium calculations.

Solution A

	H^+	+	$\text{C}_2\text{H}_3\text{O}_2^-$	\longrightarrow	$\text{HC}_2\text{H}_3\text{O}_2$
Before reaction:	0.010 M		5.00 M		5.00 M
After reaction:	0		4.99 M		5.01 M

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log\left(\frac{[\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}\right) \\ &= 4.74 + \log\left(\frac{4.99}{5.01}\right) = 4.74 - 0.0017 = 4.74 \end{aligned}$$

**no change
in pH**

Solution B

	H^+	+	$\text{C}_2\text{H}_3\text{O}_2^-$	\longrightarrow	$\text{HC}_2\text{H}_3\text{O}_2$
Before reaction:	0.010 M		0.050 M		0.050 M
After reaction:	0		0.040 M		0.060 M

$$\begin{aligned} \text{pH} &= 4.74 + \log\left(\frac{0.040}{0.060}\right) \\ &= 4.74 - 0.18 = 4.56 \end{aligned}$$

solution A has a much higher buffering capacity than solution B.

Change in $[C_2H_3O_2^-]/[HC_2H_3O_2]$ for two solutions when 0.01 mol H^+ is added to 1.0 L of each

Solution	$\left(\frac{[C_2H_3O_2^-]}{[HC_2H_3O_2]}\right)_{orig}$	$\left(\frac{[C_2H_3O_2^-]}{[HC_2H_3O_2]}\right)_{new}$	Change	Percent Change
A	$\frac{1.00 M}{1.00 M} = 1.00$	$\frac{0.99 M}{1.01 M} = 0.98$	1.00 \rightarrow 0.98	2.00%
B	$\frac{1.00 M}{0.01 M} = 100$	$\frac{0.99 M}{0.02 M} = 49.5$	100 \rightarrow 49.5	50.5%

The optimal buffering, most resistant to change when H^+ or OH^- are added to the buffered solution, occurs when $[HA]$ is equal to $[A^-]$.

If $[A^-]/[HA] = 1$

$$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right) = pK_a + \log(1) = pK_a$$

The pK_a of the weak acid to be used in the buffer should be as close as possible to the desired pH.

Example

- A chemist needs a solution buffered at pH 4.30 and can choose from the following acids (and their sodium salts):
 - a. chloroacetic acid ($K_a = 1.35 \times 10^{-3}$)
 - b. propanoic acid ($K_a = 1.3 \times 10^{-5}$)
 - c. benzoic acid ($K_a = 6.4 \times 10^{-5}$)
 - d. hypochlorous acid ($K_a = 3.5 \times 10^{-8}$)

Calculate the ratio $[HA]/[A^-]$ required for each system to yield a pH of 4.30. Which system will work best?

Answer

➤ A pH of 4.30 corresponds to

$$[H^+] = K_a \frac{[HA]}{[A^-]}$$

$$[H^+] = 10^{-4.30} = \text{antilog}(-4.30) = 5.0 \times 10^{-5} M$$

$$[\text{H}^+] = K_a \frac{[\text{HA}]}{[\text{A}^-]}$$

Acid	$[\text{H}^+] = K_a \frac{[\text{HA}]}{[\text{A}^-]}$	$\frac{[\text{HA}]}{[\text{A}^-]}$
a. Chloroacetic	$5.0 \times 10^{-5} = 1.35 \times 10^{-3} \left(\frac{[\text{HA}]}{[\text{A}^-]} \right)$	3.7×10^{-2}
b. Propanoic	$5.0 \times 10^{-5} = 1.3 \times 10^{-5} \left(\frac{[\text{HA}]}{[\text{A}^-]} \right)$	3.8
c. Benzoic	$5.0 \times 10^{-5} = 6.4 \times 10^{-5} \left(\frac{[\text{HA}]}{[\text{A}^-]} \right)$	0.78
d. Hypochlorous	$5.0 \times 10^{-5} = 3.5 \times 10^{-8} \left(\frac{[\text{HA}]}{[\text{A}^-]} \right)$	1.4×10^3

- ✦ As $[\text{HA}]/[\text{A}^-]$ for benzoic acid is closest to 1, the system of benzoic acid and its sodium salt will be the best choice.
- ✦ The optimal buffering system has a $\text{p}K_a$ value close to the desired pH. (The $\text{p}K_a$ for benzoic acid is 4.19)

Chem 102

End of Lecture's Presentations

Best wishes