Strong and Weak Acids and Bases; pH and pOH calculations

Section 8.2

Recall from last class…

An acid is a substance that donates protons. Forming an acid solution involves a reaction with water:

$$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$$

Which can be simplified, by representing it as an ionization:

(Omit H_2O) $$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

$$\mathbf{HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}}$$

A **strong acid** is one that ionizes completely.
- equilibrium lies far to the right
- essentially goes to completion
- often represent the ionization with a single-headed arrow

K_a is large

A **weak acid** is one that does NOT ionize completely.
- ions form a dynamic equilibrium with the un-ionized form

K_a is small

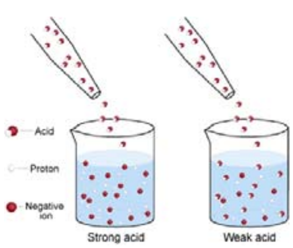
acid ionization constant, $K_a = \frac{[H^+][A^-]}{[HA]}$

Stronger acids have higher concentrations of free ions:
- higher conductivity
- higher reactivity

Which of these is vinegar (acetic acid), and which is hydrochloric acid, HCl?

Strength ≠ Concentration!

- two separate concepts
  - though both do affect amount of ions in solution
- a strong acid may either concentrated or dilute:
  - 0.1 mol/L HCl – dilute solution of a strong acid
  - 12 mol/L HCl – concentrated solution of strong acid



Similar concentration;
Different strengths

Bases can be described in the same way.

Strong bases dissociate completely (100%) in aqueous solutions.
- all hydroxides of Group 1 and 2 elements

$$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

1.0 mol 1.0 mol 1.0 mol

$$Ca(OH)_{2\,(aq)} \rightarrow Ca^{2+}_{(aq)} + 2\,OH^-_{(aq)}$$

1.0 mol 1.0 mol 2.0 mol

Weak bases only partially react with water to produce OH^- ions.

- Bronsted-Lowry bases

$$\underset{1.0\ \text{mol}}{NH_{3\,(aq)}} + H_2O_{(l)} \overset{0.42\%}{\rightleftharpoons} \underset{0.0042\ \text{mol}}{NH_4^+{}_{(aq)}} + \underset{0.0042\ \text{mol}}{OH^-{}_{(aq)}}$$

Just like with weak acids, a dynamic equilibrium is reached between the forward and reverse reactions.

For the generalized reaction of a base with water,

$$B_{(aq)} + H_2O_{(l)} \rightleftharpoons BH^+_{(aq)} + OH^-_{(aq)}$$

an equilibrium constant (the **base ionization constant**) can be calculated:

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

Base ionization constant, K_b

Value of K_b
- Infinitely large for strong bases
- Small for weak bases

Learning Checkpoint: Strong and Weak Bases

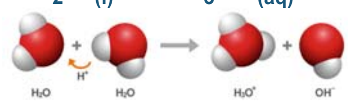
- Just like acids, bases can be classified as strong or weak, depending on how completely they react with water to produce OH^- ions.
- For the generalized reaction of a base with water,

$$B_{(aq)} + H_2O_{(l)} \rightleftharpoons BH^+_{(aq)} + OH^-_{(aq)}$$

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

Water naturally undergoes an ionization process:

$$\mathbf{2\,H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}}$$



This process is called **autoionization.**

- less than 2 water molecules in one billion will autoionize
- occurs in pure water and all aqueous solutions
- equation is often simplified:

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

For the reaction $H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$,

the equilibrium law is $K_w = [H^+]\,[OH^-]$

$= 1.0 \times 10^{-14}$ at SATP

K_w: Ion Product Constant for Water

The ion product constant for water has a value of $K_w = 1.0 \times 10^{-14}$ (SATP conditions).

In pure water, $[H^+]$ and $[OH^-]$ are produced in a 1:1 ratio, therefore $[H^+] = [OH^-]$.

Let $[H^+] = [OH^-]$ be x

$$K_w = [H^+]\,[OH^-]$$
$$1.0 \times 10^{-14} = x^2$$
$$1.0 \times 10^{-7}\ \text{mol/L} = x$$

In pure water, the concentration of $[H^+] = [OH^-] = 1.0 \times 10^{-7}$ mol/L, due to the autoionization of water.

Dissolving an acid or a base in water will alter these concentrations of $[H^+]$ or $[OH^-]$, but their product will always be constant ($K_w = 1.0 \times 10^{-14}$).

Example 1. A 0.15 mol/L solution of $HCl_{(aq)}$ is found to have a hydrogen ion concentration of 0.15 mol/L. Calculate the hydroxide ion concentration.

$$K_w = [H^+][OH^-]$$

(sub in) $1.0 \times 10^{-14} = (0.15)[OH^-]$

6.7×10^{-14} mol/L $= [OH^-]$

LeChatelier:
The concentration of OH^- goes down because because the presence of H^+ ions from the acid pushes the equilibrium to the left:

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

Example 2. Calculate the hydrogen ion concentration in a 0.025 mol/L solution of barium hydroxide, a strong base.

Strategy
1. Write the dissociation equation for this base, and determine $[OH^-]$
2. Use K_w to solve for $[H^+]$

❶ $Ba(OH)_{2\,(aq)} \rightleftharpoons Ba^{2+}_{(aq)} + 2\,OH^-_{(aq)}$

0.025 mol/L 0.050 mol/L

From the molar ratio,
$[OH^-] = 2\,[Ba(OH)_2] = 2(0.025) = 0.050$ M

❷ $K_w = [H^+][OH^-]$

$1.0 \times 10^{-14} = [H^+](0.050)$

2.0×10^{-13} mol/L $= [H^+]$

For the ionization of acetic acid in water,
$$HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$$
acid; conjugate base

$K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$

$K_a = 1.7 \times 10^{-5}$

For the reaction of the conjugate base of acetic acid ($C_2H_3O_2^-$) with water,
$$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-$$
conjugate base

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

$K_b = 5.9 \times 10^{-10}$

$$K_a K_b = (1.7 \times 10^{-5})(5.9 \times 10^{-10})$$

$$K_a K_b = 1.0 \times 10^{-14}$$

The product of the K_a of an acid and the K_b of its conjugate base is equal to the water ionization constant, $\mathbf{K_w = 1.0 \times 10^{-14}}$.

For the ionization of acetic acid in water,
$$HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$$
acid; conjugate base

$K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$

For the reaction of the conjugate base of acetic acid ($C_2H_3O_2^-$) with water,
$$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-$$
conjugate base

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

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

$$K_a K_b = [H^+][OH^-]$$

$$K_a K_b = K_w$$

The product of the K_a of an acid and the K_b of its conjugate base is equal to the water ionization constant, $\mathbf{K_w = 1.0 \times 10^{-14}}$.

Example 3.
The HPO_4^{2-} ion is a weak acid with a K_a of 1.3×10^{-13} at SATP.

a) Write the ionization equation for HPO_4^{2-} and identify its conjugate base.
b) Calculate the K_b for the conjugate base.

Solution

a) $HPO_4^{2-} \rightleftharpoons$ ________ + ________________

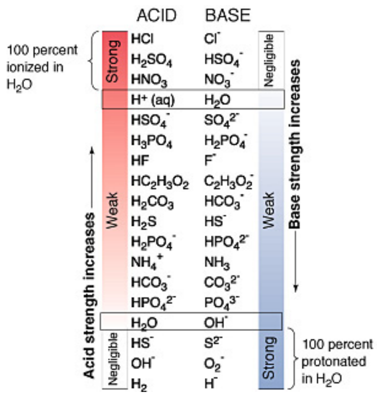
b) $K_b = \frac{K_w}{K_a}$

$$K_b = \frac{1.0 \times 10{-14}}{1.3 \times 10{-13}} = 2.4 \times 10^{-2}$$

Practice. Pg. 502 #1, 2

1. K_a for ammonium ion (NH_4^+) is 5.8×10^{-10}. Find the K_b of NH_3.
Ans: 1.7×10^{-5}

2. K_b for the fluoride ion (F^-) is 1.5×10^{-11}. Find the K_a of HF.
Ans: 6.7×10^{-4}



In general,

- If an acid is very strong, its conjugate base is very weak.
- If a base is very strong, its conjugate acid is very weak.
- Weak acids have weak conjugate bases, and vice versa.

Learning Checkpoint: K_w, K_a, and K_b

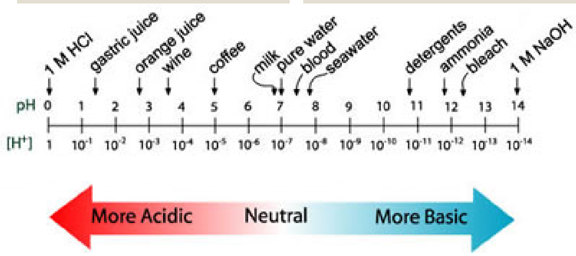
- Water spontaneously reacts with itself, undergoing autoionization: $H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$
  - For this reaction, $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ at SATP
- The value of Kw can be used to relate the Ka of an acid with the Kb of its conjugate base:

$$K_w = K_a K_b = 1.0 \times 10^{-14}$$

$$K_a = \frac{1.0\times10^{-14}}{K_b} \qquad K_b = \frac{1.0\times10^{-14}}{K_a}$$

The pH value of a solution measures the concentration of hydrogen ions.

$pH = -\log[H^+]$ $\qquad$ $[H^+] = 10^{-pH}$

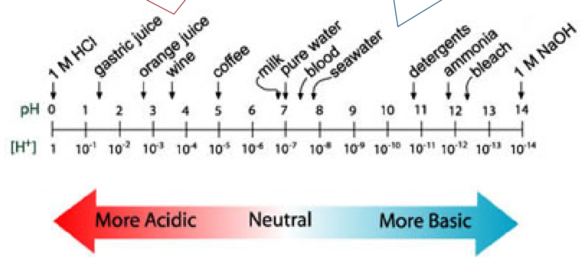


Acids have **HIGH concentrations of [H+]**

This means they have **LOW pH values**.

Bases have **LOW concentrations of [H+]**

This means they have **HIGH pH values**.



Example 4. A solution of NaOH has a $[H^+]$ of 5.2×10^{-11} mol/L. What is the pH of this solution?

The number of DECIMAL PLACES in the log must equal the number of significant figures in the original value

Solution

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

$= -\log(5.2 \times 10^{-11})$

two sig figs

$pH = 10.28$

two decimals

Check:

$[H^+] = 10^{-pH}$

$= 10^{-10.28}$

$[H^+] = 5.2 \times 10^{-11}$ mol/L

The pOH of a solution measures its concentration of hydroxide ions:

$pH = -\log[H^+]$ $\qquad$ $pOH = -\log[OH^-]$

…and interestingly enough…

For all aqueous solutions at SATP

$pH + pOH = 14.00$

See pg. 504 for the math

Example 5. The pH of the NaOH solution in Example 4 was calculated to be 10.28. What is its pOH?

Solution

pH + pOH = 14.00
pOH = 14.00 – 10.28 = 3.72

Practice. Pg. 505 #1
Calculate the pOH of a solution that has a pH of 4.

Learning Checkpoint: pH and pOH

For an aqueous solution,

- pH = - log $[H^+]$
- pOH = - log $[OH^-]$
- pH + pOH = 14.00

Homework

Pg. 504 #2, 3
Pg. 508 #1-4
Pg. 509 #1-9