

## Acid-Base Chemistry and the Equilibrium Law: $K_a$



Today, you will:

- define  $K_w$ ,  $K_a$ ,  $K_b$  and use these to determine pH, pOH,  $[H_3O^+]$  and  $[OH^-]$  of acidic and basic solutions

### The Acid Dissociation Constant: $K_a$

Recall our application of equilibrium theory to the dissociation of water:



$$K_w = [\text{H}_3\text{O}^+] [\text{OH}^-] = 1.00 \times 10^{-14}$$

We can also apply this idea to acids dissociating.

For example, the dissociation of  $\text{HI}_{(aq)}$  in water:



If we were to apply the equilibrium law to this reaction, it would look like this:

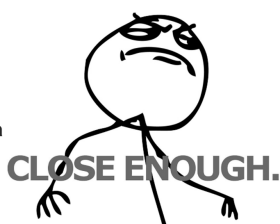
$$K_a = \frac{[\text{H}_3\text{O}^+] [\text{I}^-]}{[\text{HI}]}$$

Where:

$K_a$  = acid ionization constant

## \*Important Note\*

The concentration of water is omitted from the equilibrium equation, even though water will partly dissociate to make a the  $[H_3O^+]$  a little larger. Because of this, any  $K_a$  you calculate will never be 100% accurate, but they are good enough for our purposes.



## $K_a$ 's of some common acids

The  $K_a$ 's of some common acids are found on your Acid-Base table.

Relative Strengths of Acids And Bases at 298.15 K

Acid Name	Acid Formula	Conjugate Base Formula	$K_a$
perchloric acid	$HClO_4(aq)$	$ClO_4^-(aq)$	very large
hydroiodic acid	$HI(aq)$	$I^-(aq)$	very large
hydrobromic acid	$HBr(aq)$	$Br^-(aq)$	very large
hydrochloric acid	$HCl(aq)$	$Cl^-(aq)$	very large
sulfuric acid	$H_2SO_4(aq)$	$HSO_4^-(aq)$	very large
nitric acid	$HNO_3(aq)$	$NO_3^-(aq)$	very large
hydronium ion	$H_3O^+(aq)$	$H_2O(l)$	1
oxalic acid	$HOOC-COOH(aq)$	$HOOC-COO^-(aq)$	$5.6 \times 10^{-2}$
sulfurous acid ( $SO_2 + H_2O$ )	$H_2SO_3(aq)$	$HSO_3^-(aq)$	$1.4 \times 10^{-2}$
hydrogen sulfate ion	$HSO_4^-(aq)$	$SO_4^{2-}(aq)$	$1.0 \times 10^{-2}$
phosphoric acid	$H_3PO_4(aq)$	$H_2PO_4^-(aq)$	$6.9 \times 10^{-3}$
nitrous acid	$HNO_2(aq)$	$NO_2^-(aq)$	$5.6 \times 10^{-3}$
citric acid	$H_3C_6H_5O_7(aq)$	$H_2C_6H_5O_7^-(aq)$	$7.4 \times 10^{-4}$
hydrofluoric acid	$HF(aq)$	$F^-(aq)$	$6.3 \times 10^{-4}$

\*Note: the  $K_a$  of the strong acids is just "very large", because any strong acid dissociates >99%. So the  $[strong\ acid] = [H_3O^+]$ .

These  $K_a$  values only really apply to the weak acids.

There are two main calculations using  $K_a$  values with weak acids:

1. Use the [acid], [conjugate base] and  $[H_3O^+]$  at equilibrium to calculate a  $K_a$ .
2. Use the concentrations of some species and the  $K_a$  to work out the concentrations of other species in solution.

ex) The pH of a 1.00 mol/L solution of acetic acid is measured to be 2.38 at SATP. What is the  $K_a$  of the acid?

Step 1: Write out eqn and set up equilibrium law.



$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

Step 2: Determine the  $[H_3O^+]$  from the pH.

$$[H_3O^+] = 10^{-2.38} \\ = 0.0042 \text{ mol/L}$$

3. Because all equilibrium equations deal with [ ]'s at equilibrium, set up a quick ICE table to determine all [ ]'s at equilibrium.

	$CH_3COOH$	$H_3O^+$	$CH_3COO^-$
I	1.00	0	0
C	-0.0042	+0.0042	+0.0042
E	1.00*	0.0042	0.0042

\*Note that  $1.00 - 0.0042$ , when using our seldom talked about rules of subtraction, rounds to 1.00. We can basically forget about this subtraction step if the initial [acid] is at least 1000x larger than the  $K_a$  (as a rule of thumb).



4. Fill in the [ ]'s and calculate  $K_a$ .

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = \frac{[0.0042][0.0042]}{[1.00]} = \underline{1.7 \times 10^{-5}}$$

ex) What is the pH for a 0.200 mol/L aqueous solution of methanoic acid?

\*Hint: Assume that the initial  $[HCOOH] = \text{final } [HCOOH]$ , like in the previous examples and use the  $K$  value from the data booklet.

Ans:

ex) A student measures the pH of a 0.25 mol/L solution of carbonic acid to be 3.48. Calculate the  $K_a$  for this acid.

Ans:

ex) Predict the pH for a 0.500 mol/L aqueous solution of hydrocyanic acid.

Ans:

Homework: Read Lab Procedure on page 694-695 for tomorrow.