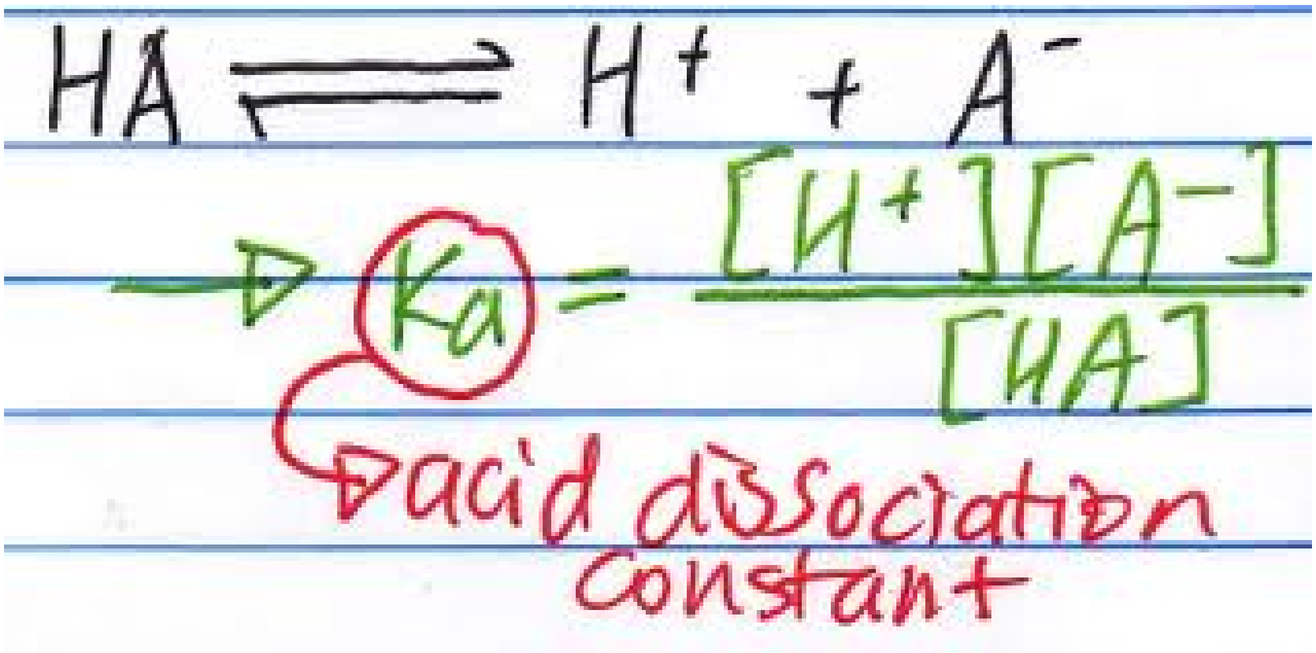
Dissociation Constants…



Outcomes:

Write the equilibrium expression (K_a or K_b) from a balanced chemical equation.

Use K_a or K_b to solve problems for pH, percent dissociation, and concentration.

Acid Dissociation Constant:

- **STRONG** acids **DISSOCIATE COMPLETELY**, so they do **NOT** reach an **EQUILIBRIUM**.
- **WEAK** acids **WILL**, however, establish an **EQUILIBRIUM**.
- We can write an **EQUILIBRIUM LAW**:

In general, for the weak acid **HA**:

$$HA + H_2O \leftrightarrow H_3O^+ + A^-$$

The equilibrium law would be:

$$K_c = \frac{[H_3O^+][A^-]}{[HA][\cancel{H_2O}]}$$

Water is a **LIQUID**, so it does **NOT** appear in equilibrium **LAW**, so we remove H_2O and replace **K_c** with ***K_a*** - the ***acid dissociation (ionization) constant:***

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

Notes on K_a:

- ***K_a reflects the equilibrium for an acid in solution.***
- ***The Larger the K_a, the more product, so the stronger the acid, the smaller the K_a, the weaker the acid.***

Acid Dissociation Constant:

Example:

HCl dissociates completely according to the equation:

$$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

If 1.0M HCl dissolves, then $[H_3O^+] = [Cl^-] = 1.0M$ (complete dissociation)

$$K_a = \frac{[H_3O^+][Cl^-]}{[HCl]}$$

$$K_a = \frac{[1.0][1.0]}{[0]}$$

K_a = very large, so HCl is a strong acid.

Base Dissociation Constant:

Same idea as **K_a**. Alkali + Alkaline earth

- **STRONG** bases dissociate **COMPLETELY**, so they do **NOT** reach an **EQUILIBRIUM**.
- **WEAK** bases **WILL**, however, establish an **EQUILIBRIUM**.
- The **HIGHER** the K_b, the **STRONGER** the base.

In general, for some weak base B:

$$B + H_2O \leftrightarrow BH^+ + OH^-$$

NaOH

Al(OH)3

The equilibrium law would be:

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

Base Dissociation Constant:

In general, for the weak base BOH:

$$BOH \leftrightarrow B^+ + OH^-$$

The equilibrium law would be:

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

Example:

For the weak base NH_3:

$$NH_{3(g)} + H_2O_{(l)} \leftrightarrow NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

The equilibrium law is:

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

The K_b for ammonia at 25°C is 1.8×10^{-5}.

Ka/Kb Problem Types:

weak → EQ

1. Calculating K_a/K_b

A 0.75M solution of Carbonic Acid dissociates partially. If at equilibrium, the $[H_3O^+] = 1.7 \times 10^{-3}$M, find K_a.

$$H_2CO_3 + H_2O_{(l)} \rightleftharpoons HCO_3^- + H_3O^+$$

	H_2CO_3	H_2O	HCO_3^-	H_3O^+
I	0.75 mol/L	X	0	0
C	-1.7×10^{-3}	X	$+1.7 \times 10^{-3}$	$+1.7 \times 10^{-3}$
E	0.7483	X	1.7×10^{-3}	1.7×10^{-3}

$$K_a = \frac{[HCO_3^-][H_3O^+]}{[H_2CO_3]}$$

$$= \frac{(1.7 \times 10^{-3})^2}{0.7483}$$

$$K_a = 3.86 \times 10^{-6}$$

Ka/Kb Problem Types:

2. Finding Concentrations of Species:

a) For a Strong Acid/Base:

Find the concentration of all species in a 0.50M HCl solution.

$$HCl \rightarrow H^+ + Cl^-$$

$$0.5\,\frac{mol}{L} \rightarrow 0.5\,\frac{mol}{L} \quad 0.5\,\frac{mol}{L}$$

Ka/Kb Problem Types:

2. Finding Concentrations of Species:

b) For a Weak Acid/Base:

Find the concentration of all species in a 0.75M solution of Acetic acid (K_a = 1.8 x 10^{-5}).

weak / ICE

$$HC_2H_3O_2 + H_2O \rightleftharpoons H_3O^+ + C_2H_3O_2^-$$

$HC_2H_3O_2$	H_2O	H_3O^+	$C_2H_3O_2^-$
$0.75 \frac{mol}{L}$	X	0	0
$-x$		$+x$	$+x$
$0.75-x$		x	x

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$1.8\times10^{-5} = \frac{(x)^2}{(0.75-x)}$$

assume x is small

$$1.8\times10^{-5} = \frac{x^2}{0.75}$$

$$\sqrt{(1.8\times10^{-5})(0.75)} = \sqrt{x^2}$$

$$x = 0.0037 \frac{mol}{L} = [H_3O^+] = [C_2H_3O_2^-]$$

$$[HC_2H_3O_2] = 0.75 - x = 0.746 \frac{mol}{L}$$

Ka/Kb Problem Types:

3. Percent Ionization:

Calculate the percent ionization of a 0.0800 M solution of hydrocyanic acid if 0.002M has ionized.

$$\% \text{ diss.} = \frac{\text{amt. ionized}}{\text{Total}} \times 100$$

$$= \frac{0.002 \frac{mol}{L}}{0.08 \frac{mol}{L}} \times 100$$

$$= 2.5\%$$

Ka/Kb Problem Types:

4. Finding K_a/K_b With Percent Dissociation:

Find K_a of HNO_2 if a 1.0M solution dissociates 2.26%.

$$HNO_2 + H_2O \rightleftharpoons H_3O^+ + NO_2^-$$

HNO_2	H_2O	H_3O^+	NO_2^-
1.0 mol/L	X	0	0
−0.0226		+0.0226	+0.0226
0.9774		0.0226	0.0226

$$1.0\,\frac{mol}{L} \times \frac{2.26\%}{100} = 0.0226\,\frac{mol}{L}$$

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]}$$

$$= \frac{(0.0226)^2}{0.9774}$$

$$K_a = 5.2 \times 10^{-4}$$

Try these ones:

1. Calculating Ka/Kb:

A 1.0 M solution of Acetic acid is only partially ionized. If at equilibrium, the $[H_3O^+] = 1.3 \times 10^{-3}$ M, find K_a.

weak EQ ICE

$$HC_2H_3O_2 + H_2O \rightleftharpoons H_3O^+ + C_2H_3O_2^-$$

$HC_2H_3O_2$	H_2O	H_3O^+	$C_2H_3O_2^-$
1.0	X	0	0
-1.3×10^{-3}		$+1.3 \times 10^{-3}$	$+1.3 \times 10^{-3}$
0.9987		1.3×10^{-3}	1.3×10^{-3}

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$= \frac{(1.3 \times 10^{-3})(1.3 \times 10^{-3})}{0.9987}$$

$$K_a = 1.69 \times 10^{-6}$$

Try these ones:

2. Calculating Concentration of Dissociated Species:

HA is a weak acid with $K_a = 7.3 \times 10^{-8}$. What is the concentration of all species (HA, H_3O^+, and A^-) if the initial [HA] = 0.50 M?

$[H_3O^+] = [A^-] = 1.9 \times 10^{-4} \frac{mol}{L}$

$[HA] \simeq 0.5 \frac{mol}{L}$

Try these ones:

3. Percent Ionization/Dissociation:

We can describe acids and bases in terms of the degree that they dissociate.

$$\text{Percent dissociation} = \frac{\text{concentration of dissociated species}}{\text{original concentration of acid or base}} \times 100\%$$

Calculate percent dissociation of a 0.100M solution of formic acid (HCH_2O_2) if the $[H_3O^+]$ = 4.21 x 10^{-3} M.

4.21%

Try these ones:

4. Finding K_a or K_b given percent dissociation:

Find K_b of the HPO_4^{2-} ion if a 0.25 M solution dissociates 0.080%

$HPO_4^{2-} + H_2O \rightleftharpoons H_2PO_4^{1-} + OH^-$ 1.6×10^{-7}