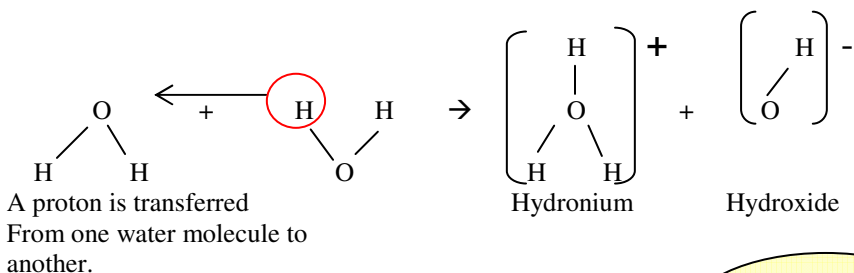


Ionization of Water

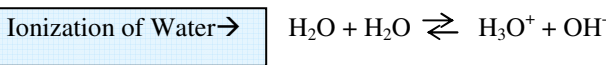
DEMONSTRATION OF CONDUCTIVITY OF TAP WATER AND DISTILLED WATER

- Pure distilled water still has a small conductivity. Why?
- There are a few ions present.
- Almost all the pure water is H₂O molecules.
- But every once in a while, this happens:

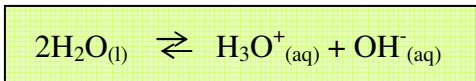


An equilibrium
is established

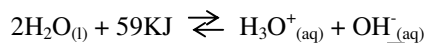
Equation:



Or



Process is Endothermic

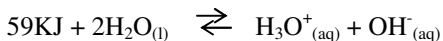


All water or aqueous solutions
contain these.

In neutral water	$[H_3O^+] = [OH^-]$	} Know these!!
In acidic solutions	$[H_3O^+] > [OH^-]$	
In basic solutions	$[OH^-] > [H_3O^+]$	

Write the Keq Expression for this equilibrium:

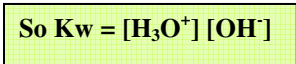
Keq =



Keq = $[H_3O^+][OH^-]$ (liquid water left out)

Given a special name for ionization of water—called Kw

Ionization
constant
for **water**



Always true at any temp!

Since reaction is *endothermic*: $59\text{KJ} + 2\text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{OH}^-_{(aq)}$

At higher temps _____ are favoured and Kw is _____ er.

At lower temps _____ are favoured and Kw is _____ er.

At 25°C (only) Kw = 1.00×10^{-14}

← Know this!!

For
comparison

At 10°C Kw = 0.295×10^{-14} (smaller)

At 60°C Kw = 9.55×10^{-14} (larger)

So

Always: $[\text{H}_3\text{O}^+][\text{OH}^-] = K_w$

At 25°C only: $[\text{H}_3\text{O}^+][\text{OH}^-] = 1.00 \times 10^{-14}$

$[\text{H}_3\text{O}^+]$ & $[\text{OH}^-]$ in Neutral Water

At 25°C (NOTE: Assume Temp = 25°C unless otherwise noted)

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

and $[\text{H}_3\text{O}^+] = [\text{OH}^-]$ if water is **neutral**. (If “water” is mentioned in a problem, it can be assumed to be NEUTRAL unless otherwise stated!)

(substitute. $[\text{H}_3\text{O}^+]$ for $[\text{OH}^-]$)

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

$$[\text{H}_3\text{O}^+]^2 = 1.00 \times 10^{-14}$$

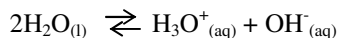
$$[\text{H}_3\text{O}^+] = \sqrt{1.00 \times 10^{-14}} = \underline{1.00 \times 10^{-7}} \text{ M}$$

Also $[\text{OH}^-] = [\text{H}_3\text{O}^+] = \underline{1.00 \times 10^{-7}} \text{ M}$

At Higher Temp

Given: Kw at 60°C = 9.55×10^{-14}

Calculate $[\text{H}_3\text{O}^+]$ & $[\text{OH}^-]$ at 60°C

[H₃O⁺] & [OH⁻] in Acids and Bases

Add acid, H₃O⁺ increases, so equilibrium shifts LEFT and [OH⁻] decreases

All acids produce
H₃O⁺ in water

Add base, [OH⁻] increases, so the equilibrium shifts LEFT and [H₃O⁺] decreases.

Finding [H₃O⁺] and [OH⁻] in Acids and Bases

At 25^oC

Eg.) Find the [OH⁻] in 0.0100 M HCl

In a Strong Acid ([H₃O⁺] = [acid])

$$[\text{H}_3\text{O}^+] = 0.0100 \text{ M}$$

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

$$[\text{OH}^-] = \frac{1.00 \times 10^{-14}}{[\text{H}_3\text{O}^+]} = \frac{1.00 \times 10^{-14}}{1.00 \times 10^{-2}} = 1.00 \times 10^{-12} \text{ M}$$

[OH⁻] is less in
an acid than in
neutral water.

Find [H₃O⁺] in 0.300 M NaOH.

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

$$[\text{H}_3\text{O}^+] = \frac{1.00 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.00 \times 10^{-14}}{0.300} = 3.33 \times 10^{-14}$$

[H₃O⁺] is very low
in a strong fairly
concentrated base

Find [H₃O⁺] in 0.020 M Ba(OH)₂

In a STRONG BASE, [OH⁻] = [base] x # of OH's

$$[\text{OH}^-] = ? (\quad) \text{ M}$$

$$[\text{H}_3\text{O}^+] = \frac{1.00 \times 10^{-14}}{(\quad)} = \underline{\hspace{2cm}} \text{ M}$$

At Other Temps

- you'd be given K_w

eg.) K_w at $60^\circ\text{C} = 9.55 \times 10^{-14}$

Calculate $[\text{OH}^-]$ in 0.00600 M HNO_3 at 60°C .

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

SA

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 9.55 \times 10^{-14}$$

$$[\text{OH}^-] = \frac{9.55 \times 10^{-14}}{0.00600} = 1.59 \times 10^{-11} \text{ M}$$

Rd. pg. 126-127 Ex. 28-30 pg. 127 of SW.

pH

-Shorthand method of showing acidity (or basicity, alkalinity)

If $[\text{H}_3\text{O}^+] = 0.10 \text{ M}$ ($1.0 \times 10^{-1} \text{ M}$) $\text{pH} = 1.00$

$[\text{H}_3\text{O}^+] = 0.00010 \text{ M}$ ($1.0 \times 10^{-4} \text{ M}$) $\text{pH} = 4.00$

Definition of pH

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+] \quad (\text{assume } \log = \log_{10})$$

If $[\text{H}_3\text{O}^+] = 1.0 \times 10^{-7}$

$$\text{pH} = -\log (1.0 \times 10^{-7})$$

Regular Scientific Calculator. Enter: $1 \rightarrow \text{EXP} \rightarrow 7 \rightarrow +/- \rightarrow \text{LOG} \rightarrow +/-$ and the answer should be 7

For DAL (Sharp) calc. Enter: $+/- \rightarrow \log \rightarrow 1 \rightarrow \text{Exp} \rightarrow +/- \rightarrow 7 \rightarrow =$ and the answer should be 7

For a TI 83 Enter $(-) \rightarrow \text{LOG} \rightarrow 1 \rightarrow 2^{\text{nd}} \rightarrow \text{EE} \rightarrow (-) \rightarrow 7 \rightarrow \text{ENTER}$ and the answer should be 7

Use the (-) button at the bottom right of the white pad. NOT the “—“ on the far right

NOTE: If you are using a DAL or a TI 83 calculator and the number you want to find the pH of is the answer to an ongoing calculation, leave the answer to your calculation in the calculator and press $(-)$ or $+/- \rightarrow \log \rightarrow 2^{\text{nd}} \rightarrow \text{ANS} \rightarrow \text{ENTER}$ (or =).

Practice finding pH's on your own calculator. You will be doing many of these in the rest of this unit and it's important that you can do it quickly and easily and accurately!

Find the pH of 0.030 M HCl

2 SD's

SA. So $[H_3O^+] = [\text{acid}]$

$$[H_3O^+] = 0.030 \text{ M}$$

$$\text{pH} = -\log(0.030) = 1.522878745$$

How to round off??

Sig. Digits in pH start at decimal point!!!

1.52287..... so pH = 1.52

Start counting here.

No units for pH

Find the pH of neutral water at 25⁰ C

$$[H_3O^+] = 1.00 \times 10^{-7}$$

$$\text{pH} = 7.000$$

Find the pH of 0.00100 M NaOH at 25⁰ C

Basic solution

$$[H_3O^+] = \frac{1.00 \times 10^{-14}}{0.00100} = 1.00 \times 10^{-11} \text{ M}$$

[OH]

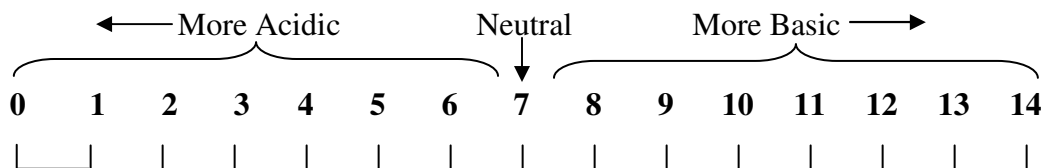
so pH = 11.000

3 sig. digits

At 25 °C

In **neutral** water pH = 7.0
 In **acid** solution pH < 7.0
 In **basic** solution pH > 7.0

pH Scale (@ 25°C)



Calculate pH of 12.0 M HCl

Very concentrated Acid !

Answer _____

Another example: Calculate the pH of 15.0 M NaOH:

$$[\text{H}_3\text{O}^+] = \frac{1.00 \times 10^{-14}}{15.0} = 6.67 \times 10^{-16} \text{ M}$$

$$\text{pH} = -\log (6.67 \times 10^{-16})$$

$$\text{pH} = 15.176$$

Very concentrated Base

Converting pH to $[\text{H}_3\text{O}^+]$

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$-\text{pH} = \log [\text{H}_3\text{O}^+]$$

$$\text{antilog} (-\text{pH}) = [\text{H}_3\text{O}^+]$$

or

$$[\text{H}_3\text{O}^+] = \text{antilog} (-\text{pH})$$

3 SD's

Put in (-) before
you take the
antilog

eg.) If pH = 11.612 , find $[\text{H}_3\text{O}^+]$

$$[\text{H}_3\text{O}^+] = \text{antilog} (-11.612)$$

For regular Scientific Calculator:

Enter: 11.612 \rightarrow +/- \rightarrow 2nd \rightarrow log The calculator answer should be $2.443430553 \times 10^{-12}$

The original pH had 3 SD's, so the answer must also have 3 SD's (Remember the answer is NOT a pH, so digits to the LEFT of the decimal point **are** also significant!) . Remember that concentration also has a **unit**!. So the answer would be reported as:

$$[\text{H}_3\text{O}^+] = 2.44 \times 10^{-12} \text{ M}$$

3 SD's

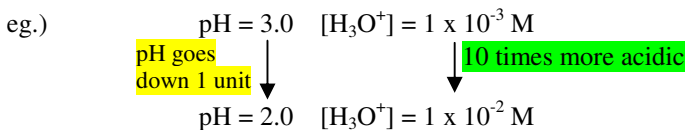
unit

For a DAL or TI 83 enter: 2nd \rightarrow log \rightarrow (-) (or +/-) \rightarrow 11.612 \rightarrow ENTER (or =)

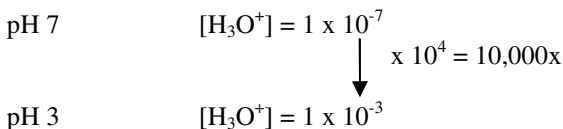
If pH = 3.924 calculate $[\text{H}_3\text{O}^+]$

Logarithmic Nature of pH

A change of 1 pH unit → a factor of 10 in $[\text{H}_3\text{O}^+]$ or (acidity)



How many times more acidic is pH 3 than pH 7?



or taking **antilog of difference in pH** $7 - 3 = 4$
 antilog 4 = $10^4 = 10,000$ times
 (remember lower pH more acidic)

Natural rainwater pH ~ 6

Extremely acidic acid rain pH ~ 3 diff = 3 & antilog (3) = 10^3 (1,000)

So, the acid rain is 1000 times more acidic than natural rain water!

pOH

$$\text{pOH} = -\log [\text{OH}^-]$$

And $[\text{OH}^-] = \text{antilog} (-\text{pOH})$

Calculate the pOH of 0.0020 M KOH

$$[\text{OH}^-] = 2.0 \times 10^{-3} \text{ M}$$

$$\text{pOH} = -\log (2.0 \times 10^{-3}) = 2.70$$

Find the pH of the same solution:

$$[\text{OH}^-] = 2.0 \times 10^{-3} \text{ M}$$

$$[\text{H}_3\text{O}^+] = \frac{1.00 \times 10^{-14}}{2.0 \times 10^{-3}} = 5.0 \times 10^{-12}$$

$$\text{pH} = 11.30$$

Notice: **pH + pOH = 14.00**

From Math: If $a \times b = c$
 Then: $\log_a + \log_b = \log_c$

Eg.)

10	x	100	=	1000
Log(10)	+	log(100)	=	log(1000)
1	+	2	=	3

So since

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

$$\log[\text{H}_3\text{O}^+] + \log[\text{OH}^-] = \log(K_w)$$

or make everything *negative*

$$-\log[\text{H}_3\text{O}^+] + -\log[\text{OH}^-] = -\log K_w$$

pH + pOH = pK_w	(relation)
----------------------------------	------------

True at all temperatures

where **pK_w = -log K_w** (definition of pK_w)

Specifically at 25⁰C

$$K_w = 1.00 \times 10^{-14}$$

$$pK_w = -\log(1.00 \times 10^{-14})$$

$$\mathbf{pK_w = 14.000}$$

Only at 25⁰C

so at 25⁰C

pH + pOH = 14.000

At 25⁰C If pH = 4.00 pOH = 10.00
 Or: If pH = 2.963 pOH = 11.037

eg.) Find the pH of 5.00 x 10⁻⁴ M LiOH (25⁰C)

plan: [OH⁻] → pOH → pH

$$[\text{OH}^-] = 5.00 \times 10^{-4} \rightarrow \text{pOH} = 3.301 \rightarrow \text{pH} = 14.000 - 3.301 = \mathbf{10.699}$$

eg.) Find the pOH of 0.0300 M HBr (25⁰C)

$$[\text{H}_3\text{O}^+] = 0.0300 \text{ M (HBr is a strong acid)}$$

$$\text{pH} = 1.523$$

$$\text{pOH} = 14.000 - 1.523$$

$$\mathbf{\text{pOH} = 12.477}$$

See pOH scale & pH scale. Pg. 140 S.W.

When not at 25°CEg.) At 60°C $K_w = 9.55 \times 10^{-14}$

Find the pH of neutral water at 60°C.

One way: Calculate pK_w

$$pK_w = -\log K_w = -\log (9.55 \times 10^{-14})$$

At 60°C $pK_w = 13.020$ For **neutral** water $pH = pOH$ ($[H_3O^+] = [OH^-]$)

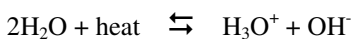
$$pH + pOH = pK_w \text{ (substitute pH for pOH)}$$

$$pH + pH = 13.020$$

$$2pH = 13.020 \rightarrow \text{so } pH = \frac{13.020}{2} = \mathbf{6.510}$$

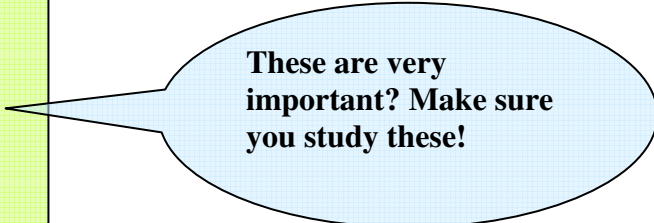
Is pH always 7.00 in neutral water? _____

At higher temp:



$$[H_3O^+] > 1.0 \times 10^{-7} \quad \text{so } pH < 7$$

$$[OH^-] > 1.0 \times 10^{-7} \quad \text{so } pOH < 7$$

Summary:**In neutral water $pH = pOH$ at any temp.** **pH & $pOH = 7.00$ at 25°C only.****At lower temps pH and pOH are > 7** **At higher temps pH and pOH are < 7** **At any temp: $pH + pOH = pK_w$** **At 25°C: $pH + pOH = 14.000$**


These are very important? Make sure you study these!

Read p. 134-141 in SW.

Do ex. 49-53 + 55-57 (p. 139-141 S.W.)

Do Worksheet 4-3 pH and pOH Calculations

K_a & K_b for Weak acids and BasesReminder: pH of SA's

$$[H_3O^+] = [\text{acid}] \quad \text{strong means 100\% ionized}$$

so, to find pH of 0.100 M HCl

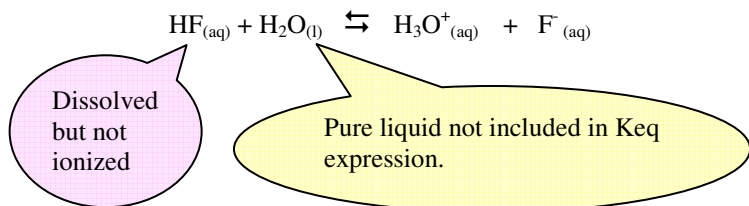
$$[H_3O^+] = 0.100 \text{ M}$$

$$\mathbf{pH = 1.000}$$

For weak acids $[H_3O^+] \ll [\text{acid}]$

Eg.) What is pH of 0.10 M HF?

Look at equilibrium for **Weak Acid HF**



$$K_{eq} = \frac{[H_3O^+][F^-]}{[HF]} \quad \text{for WA's } K_{eq} \text{ is called } \mathbf{K_a} \text{ (acid ionization constant)}$$

- see acid table for list of K_a 's.

higher $K_a \rightarrow$ stronger acid

lower $K_a \rightarrow$ weaker acid

For SA's (eg. HCl) $K_a = \frac{[H_3O^+][Cl^-]}{[HCl]}$ = called "very large!"

Essentially **zero**
molecular HCl

-Discuss Relative Strengths of Oxyacids

Calculations Using K_a (Used for Weak Acids)

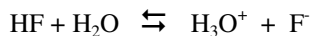
$[H_3O^+]$ from K_a (pH from K_a)

1. **$[H_3O^+]$ from K_a and Original concentration (C_o)**

eg.) Find the $[H_3O^+]$ in 0.10 M HF

WA

1. Write out equilibrium equation for *ionization*



2. Ice table

$$\text{HF} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{F}^-$$

[I]	0.10		0	0
[C]	-x		+x	+x
[E]	0.10 - x		x	x

3. Ka expression: $K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]}$

$$K_a = \frac{(x)(x)}{0.10 - x}$$

To avoid quadratic assume x is insignificant compared to 0.10. This can be confirmed later.

In Chem. 12 with weak acids, always use this assumption - Make sure you state it!

4. Substitute

$$K_a = \frac{x^2}{0.10}$$

(Assume $0.10 - x \cong 0.10$)

You **must** state this assumption here!

5. Solve for x ($[\text{H}_3\text{O}^+]$)

$$K_a = \frac{x^2}{0.100}$$

$$\text{So } x^2 = 0.10 K_a$$

$$[\text{H}_3\text{O}^+] = x = \sqrt{0.10 K_a}$$

$$= \sqrt{0.10 (3.5 \times 10^{-4})}$$

Ka from Acid Table

$$[\text{H}_3\text{O}^+] = 5.9 \times 10^{-3} \text{ M}$$

~ Check assumption (we see that this is quite small compared to 0.10)

Do ex. 74 & 75 Pg. 152

After Questions 74 & 75

Short cut for **multiple choice Only!!**

For WA $[\text{H}_3\text{O}^+] = \sqrt{C_o K_a}$

“root” a “coka”
(beer) (cola)

C_o is Original
Acid
Concentration

NOTE: In W.R. questions, full solution **must** be shown including the assumption!

Eg.) Find pH of 2.0 M acetic acid (Multiple Choice Question)

1. First $[\text{H}_3\text{O}^+] = \sqrt{C_o K_a}$

$$= \sqrt{2.0 (1.8 \times 10^{-5})}$$

$$[\text{H}_3\text{O}^+] = 6.0 \times 10^{-3} \text{ M}$$

Ka from Acid Table

2. Find pH = $-\log (6.0 \times 10^{-3})$

$$\text{pH} = 2.22$$

2 SD's. The “2.0 M” was 2 SD's and the Ka was 2 SD's

Try using this for a SA!!

NOTE: Ions which act as acids can come from compounds.

Eg.) See table ~ ammonium ion NH_4^+

- can be found in NH_4NO_3 , NH_4Cl , NH_4Br , etc....
 ferric (hexaquoiron) Fe^{3+} ($\text{Fe}(\text{H}_2\text{O})_6^{3+}$) could be found in $\text{Fe}(\text{H}_2\text{O})_6\text{Br}_3$ (also called FeBr_3)
 or $\text{Fe}(\text{H}_2\text{O})_6(\text{NO}_3)_3$ (also called $\text{Fe}(\text{NO}_3)_3$)

Aluminum (hexaquoaluminum) Al^{3+} ($\text{Al}(\text{H}_2\text{O})_6^{3+}$) could be found in $\text{Al}(\text{H}_2\text{O})_6\text{Cl}_3$ (also called AlCl_3)

Do ex. 79 & 81, Pg. 152

More Ka Calculations:

Ka from pH

NOTE: the pH is to 3 SD's
 so your final answer cannot
 have more than 3 SD's.

Eg.) a 0.350 M Solution of the weak acid HA has a pH of 1.620. Find the K_a of HA.

1. First convert pH to $[\text{H}_3\text{O}^+]$

$$\begin{aligned} [\text{H}_3\text{O}^+] &= \text{antilog}(-\text{pH}) \\ &= \text{antilog}(-1.620) \\ [\text{H}_3\text{O}^+] &= 2.399 \times 10^{-2} \text{ M} \end{aligned}$$

This is the
 $[\text{H}_3\text{O}^+]$ at
 equilibrium

2. Write out equilibrium equation for ionization. Make an ICE table:

	HA	+	H_2O	\rightleftharpoons	H_3O^+	+	A^-
[I]	0.350				0		0
[C]							
[E]					2.399×10^{-2}		

Now, you can see that the change in concentration [C] of $[\text{H}_3\text{O}^+]$ is $+ 2.399 \times 10^{-2}$ M and using the mole ratios (mole bridges) in the balanced equation, you can figure out the [C]'s for the A^- and the HA:

	HA	+	H_2O	\rightleftharpoons	H_3O^+	+	A^-
[I]	0.350				0		0
[C]	-2.399×10^{-2}				$+ 2.399 \times 10^{-2}$		$+ 2.399 \times 10^{-2}$
[E]					2.399×10^{-2}		

Now, we can figure out the equilibrium concentrations of HA and A^- . There are no "x"s in the table so we don't need to make any assumptions. It is best to use your calculator to figure out the equilibrium [HA], because the [C] may or **MAY NOT** be insignificant. Using a calculator $0.350 - 2.399 \times 10^{-2} = 0.32601$. Don't round it off too much here. I would keep it in a memory in my calculator. BUT BECAUSE THE "0.350" IS 3 DECIMAL PLACES AND YOU ARE SUBTRACTING, THE [E] OF "HA" CANNOT HAVE MORE THAN 3 DECIMAL PLACES (although you should use 0.32601 in your calculator) JUST REMEMBER THAT IN THE NEXT CALCULATION, THE 3 DECIMAL PLACES IN THE ICE TABLE TRANSLATES TO 3 SD'S, SO YOUR FINAL ANSWER CANNOT HAVE MORE THAN 3 SD'S.

$$-2.399 \times 10^{-2}M \quad + 2.399 \times 10^{-2}M \quad + 2.399 \times 10^{-2}M$$

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

[I]	0.350		0	0
[C]	-2.399×10^{-2}		$+2.399 \times 10^{-2}$	$+2.399 \times 10^{-2}$
[E]	0.326		2.399×10^{-2}	2.399×10^{-2}

3. Write Ka expression & substitute values.

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{(2.399 \times 10^{-2})^2}{0.326} = 1.7653 \times 10^{-3} \text{ and expressing in 3SD's, the answer is:}$$

$$K_a = 1.77 \times 10^{-3}$$

For those that want a short-cut for multiple choice:

Ka from $[H_3O^+]$:

$$K_a = \frac{[H_3O^+]^2}{(C_o - [H_3O^+])}$$

Don't forget this

OK for multiple choice ONLY!

Don't rearrange $[H_3O^+] = \sqrt{C_o K_a}$

Was derived using an assumption which may NOT be valid!

Do Ex. 77 & 80 on p. 152 SW

To Calculate C_o (conc. of acid needed) form pH & K_a

Eg. Find the concentration of HCOOH needed to form a solution with pH = 2.69

- First change pH to $[H_3O^+]$
 $[H_3O^+] = \text{antilog}(-\text{pH})$
 $= \text{antilog}(-2.69)$
 $[H_3O^+] = 2.0417 \times 10^{-3} \text{ M}$ (notice that the given pH limits us to 2SD's, but keep more in your calculations.)
- Write out ionization equilibrium with an ICE TABLE. You can insert 2.0417×10^{-3} for equilibrium $[H_3O^+]$. And since our **unknown** is the initial [HCOOH], we put in an " C_o " for the [I] of HCOOH:

$$HCOOH + H_2O \rightleftharpoons H_3O^+ + HCOO^-$$

[I]	C_o		0	0
[C]				
[E]			2.0417×10^{-3}	

Now we can see that the change in concentration [C] of $[H_3O^+]$ is “+ 2.0417 x 10⁻³” and the [C] of $HCOO^-$ will be the same. The [C] of $HCOOH$ will be “- 2.0417 x 10⁻³”

	HCOOH	+	H ₂ O	↔	H ₃ O ⁺	+	HCOO ⁻
[I]	C _o				0		0
[C]	- 2.0417 x 10 ⁻³				+ 2.0417 x 10 ⁻³		+ 2.0417 x 10 ⁻³
[E]					2.0417 x 10 ⁻³		

We can now calculate the equilibrium concentrations [E] of $HCOOH$ and $HCOO^-$.

	HCOOH	+	H ₂ O	↔	H ₃ O ⁺	+	HCOO ⁻
[I]	C _o				0		0
[C]	- 2.0417 x 10 ⁻³				+ 2.0417 x 10 ⁻³		+ 2.0417 x 10 ⁻³
[E]	C _o - 2.0417 x 10 ⁻³				2.0417 x 10 ⁻³		2.0417 x 10 ⁻³

The next step will be to write the K_a expression and substitute the equilibrium concentrations in:

3. Write K_a expression. Substitute equilibrium concentrations in. Find K_a for $HCOOH$ on the acid table:

$$K_a = \frac{[H_3O^+][HCOO^-]}{[HCOOH]}$$

Find K_a on Acid Table

$$1.8 \times 10^{-4} = \frac{(2.0417 \times 10^{-3})^2}{(C_o - 2.0417 \times 10^{-3})}$$

Now we can solve for C_o (the original concentration of the acid):

$$C_o - 2.0417 \times 10^{-3} = \frac{(2.0417 \times 10^{-3})^2}{1.8 \times 10^{-4}}$$

$$C_o - 2.0417 \times 10^{-3} = 2.3159 \times 10^{-2}$$

$$C_o = 2.3159 \times 10^{-2} + 2.0417 \times 10^{-3}$$

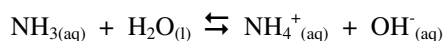
$$C_o = 2.52 \times 10^{-2} \text{ M}$$

$$C_o = 2.5 \times 10^{-2} \text{ M or } 0.025 \text{ M (remember, we are restricted to 2 SD's)}$$

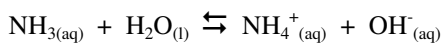
Do Ex. 76 & 78 on p. 152 of SW.

Now For Bases

Base ionization $NH_3 \rightarrow$ very common weak base. It partially ionizes in water to form NH_4^+ and OH^- :



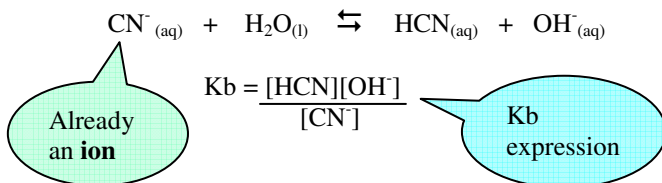
Equilibrium constant – called **base ionization constant (K_b)**



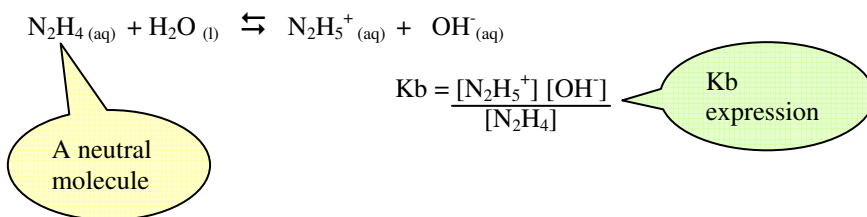
$$\text{Kb expression: } \text{Kb} = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

NOTE: Ions can also act as a weak bases. The reaction of an **ion** with water to form OH^- is called **base hydrolysis**. Equilibrium constant is still called Kb.

Eg.) **Hydrolysis of CN^-**



Ionization of N_2H_4 (weak base)



Do Ex 32 on p. 128 of SW.

How to Find Kb using Acid Table

(not shown directly)

Derivation



$$\text{Kb}_{(\text{F}^-)} = \frac{[\text{HF}][\text{OH}^-]}{[\text{F}^-]}$$



$$\text{Ka}_{(\text{HF})} = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]}$$



Conj. Acid

Conj. base

$$\text{Ka}_{[\text{HF}]} \times \text{Kb}_{[\text{F}^-]} = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} \times \frac{[\text{HF}][\text{OH}^-]}{[\text{F}^-]} = [\text{H}_3\text{O}^+][\text{OH}^-] \text{ (notice that } [\text{HF}] \text{ and } [\text{F}^-] \text{ will cancel.)}$$

or $\text{Ka}_{(\text{HF})} \times \text{Kb}_{(\text{F}^-)} = \text{Kw}$

$$K_{a(\text{HF})} \times K_{b(\text{F}^-)} = K_w$$

$$\text{Or } K_{b(\text{F}^-)} = \frac{K_w}{K_{a(\text{HF})}}$$

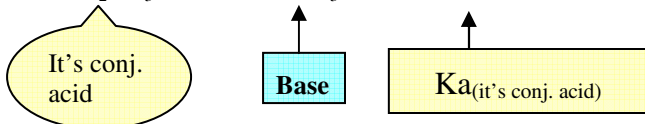
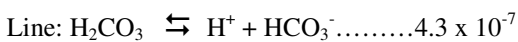
In general:

$$K_{b(\text{weak base})} = \frac{K_w}{K_{a(\text{it's conj. acid})}}$$

Using Acid Table:

1. Find base on **right** side (if amphiprotic -locate base on **right** side only)
2. It's conjugate acid will be across from it on the left side.
3. The K_a of it's conjugate acid is on the **far right** of the same line.
4. Use equation: $K_{b(\text{base})} = \frac{K_w}{K_{a(\text{conj. acid})}}$

Eg.) Calculate the K_b for HCO_3^- : (find HCO_3^- on **RIGHT** SIDE)



$$K_{b(\text{HCO}_3^-)} = \frac{K_w}{K_{a(\text{H}_2\text{CO}_3)}} = \frac{1.0 \times 10^{-14}}{4.3 \times 10^{-7}} = 2.3 \times 10^{-8}$$

Find K_b of SO_3^{2-}

Similarly : If K_b (base) given

$$K_{a(\text{weak acid})} = \frac{K_w}{K_{b(\text{it's conj. Base})}}$$

Eg.) The K_b for base B^- is 2.73×10^{-9}
Find the K_a for the acid HB

3 sig. Digs.

$$K_{a(\text{HB})} = \frac{K_w}{K_{b(\text{B}^-)}} = \frac{1.00 \times 10^{-14}}{2.73 \times 10^{-9}} = 3.66 \times 10^{-6}$$

NOTES:

- Table only states K_a values. For questions like this K_b will have to be calculated if not given.
- All K_a 's on table are 2 SD's—limits any calculation using them to 2 SD's maximum.
- The larger the K_b , the “stronger” the weak base - the more OH^- produced.
- The smaller the K_a of an acid, the larger the K_b of its conjugate Base. (*Weaker acids have stronger conjugate bases*)

Calculations Involving K_b

Given: [Base] & K_b → Find [OH⁻] → or find pOH → or find pH

eg.) Find [OH⁻] in a 0.20 M solution of KNO_2 (*this is a salt, so it must be dissociated into it's ions first*)

Dissociation of $\text{KNO}_2 \rightarrow \text{K}^+ + \text{NO}_2^-$

Neutral
spectator
(discard)

Weak base

$$\begin{aligned}
 &1. \text{ Find } K_b \text{ of } \text{NO}_2^- \\
 &K_b(\text{NO}_2^-) = \frac{K_w}{K_a(\text{HNO}_2)} \\
 &= \frac{1.0 \times 10^{-14}}{4.6 \times 10^{-4}} \\
 &K_b(\text{NO}_2^-) = 2.174 \times 10^{-11}
 \end{aligned}$$

2. Hydrolysis (if ion) or ionization (if molecule) equation followed by an ICE table:

	NO_2^-	+ H_2O	\rightleftharpoons	HNO_2	+	OH^-
[I]	0.20			0		0
[C]						
[E]						

We can let “x” equal the moles/L of NO_2^- which hydrolyze as the reaction reaches equilibrium. Using the mole ratios in the equation and calculating [E]’s we get:

	NO_2^-	+ H_2O	\rightleftharpoons	HNO_2	+	OH^-
[I]	0.20			0		0
[C]	-x			+x		+x
[E]	0.20 - x			x		x

3. K_b Expression:

$$K_b = \frac{[\text{HNO}_2][\text{OH}^-]}{[\text{NO}_2^-]} \quad \text{at equilibrium}$$

On the next page, we substitute the [E]’s into the K_b expression (*Don't forget, we are in “baseland” not “acidland” now!*)

Substitute: $K_b = \frac{[\text{HNO}_2][\text{OH}^-]}{[\text{NO}_2^-]}$ or $K_b = \frac{x^2}{(0.20 - x)}$

Original
conc. of
 NO_2^- (Co)

Assume that
 $0.20 - x \cong 0.20$

Approximation:

$$K_b \cong \frac{x^2}{0.20}$$

$$x^2 = 0.20 K_b$$

$$[\text{OH}^-] = \sqrt{0.20 K_b}$$

$$[\text{OH}^-] = \sqrt{0.20 (2.174 \times 10^{-11})}$$

$$[\text{OH}^-] = 2.085 \times 10^{-6} \text{ M}$$

$$[\text{OH}^-] = 2.1 \times 10^{-6} \text{ M} \quad \text{Both the } 0.20 \text{ M and the } K_a \text{ used in calculating } K_b \text{ limit us to } 2 \text{ SD's.}$$

NOTE: In many problems, finding $[\text{OH}^-]$ is only the first step. Often you have to find the pH. Then you would use the process $[\text{OH}^-] \rightarrow \text{pOH} \rightarrow \text{pH}$

Kb From pH and Concentration

Ex.) At a certain temp, a 0.20 M solution of K_2SO_3 has a pH of 10.25. Calculate the K_b of SO_3^{2-} at this temp.

- Identify SO_3^{2-} as a weak base (When K_2SO_3 is dissociated, it yields K^+ (a spectator) and SO_3^{2-})
- $[\text{OH}^-]$ can be obtained from pH ($\text{pH} \rightarrow \text{pOH} \rightarrow [\text{OH}^-]$)

$$\text{pH} = 10.25 \text{ so } \text{pOH} = 14.00 - 10.25 = 3.75$$

$$[\text{OH}^-] = \text{antilog}(-\text{pOH}) = \text{antilog}(-3.75) = 1.778 \times 10^{-4} \text{ M}$$

- Write hydrolysis equation and an ICE table. (It is called *hydrolysis* this time because SO_3^{2-} is an ion.) We know the *initial* $[\text{SO}_3^{2-}]$ is 0.20M and the *equilibrium* $[\text{OH}^-]$ is $1.778 \times 10^{-4} \text{ M}$:

	SO_3^{2-}	+ H_2O	\rightleftharpoons	HSO_3^-	+	OH^-
[I]	0.20			0		0
[C]						
[E]						1.778×10^{-4}

So, from this, the change in conc. [C] of OH^- is “+ 1.778×10^{-4} ” and using the coefficient ratios we can insert the [C]’s for the other species and calculate the equilibrium concentrations [E]’s:

	SO_3^{2-}	+ H_2O	\rightleftharpoons	HSO_3^-	+	OH^-
[I]	0.20			0		0
[C]	-1.778×10^{-4}			$+1.778 \times 10^{-4}$		$+1.778 \times 10^{-4}$
[E]	0.1998			1.778×10^{-4}		1.778×10^{-4}

4. Now we write the Kb expression and substitute the values from the [E]'s in our ICE table:

$$K_b = \frac{[\text{HSO}_3^-][\text{OH}^-]}{[\text{SO}_3^{2-}]} = \frac{(1.778 \times 10^{-4})^2}{(0.199822)} = 1.5825 \times 10^{-7}$$

We round it to 1.6×10^{-7} because the 0.20M and the pH given both had 2 SD's

So:

Kb of SO_3^{2-} at the temperature given is $= 1.6 \times 10^{-7}$

Short Cut for Multiple Choice:

$$K_b = \frac{[\text{OH}^-]^2}{(C_0 - [\text{OH}^-])}$$

Like Weak acids:

$$K_a = \frac{[\text{H}_3\text{O}^+]^2}{(C_0 - [\text{H}_3\text{O}^+])}$$

Rd. pg. 128 – 129 & 152-153

Do Ex. 84, 87, 88, & 89 on p. 153 of SW

- Do worksheet on Ka, Kb calculations