



ACIDS & BASES

INTRODUCTION: DEFINITIONS, pH, AND *K* VALUES

CHEMISTRY 165 // SPRING 2020

What is an acid?

Let's stick with the Brønsted-Lowry definition:

ACID: Proton (H^+) *donor*, or gives off free H^+

BASE: Proton (H^+) *acceptor*, or reacts with H^+ to give off OH^-

To help, consider the following examples of **acids**. Notice how the acids give off H^+ and an anion.

ACID



In general, we can express the reactivity
of **ACIDS** and **BASES** as:



What is an acid? What is a base?

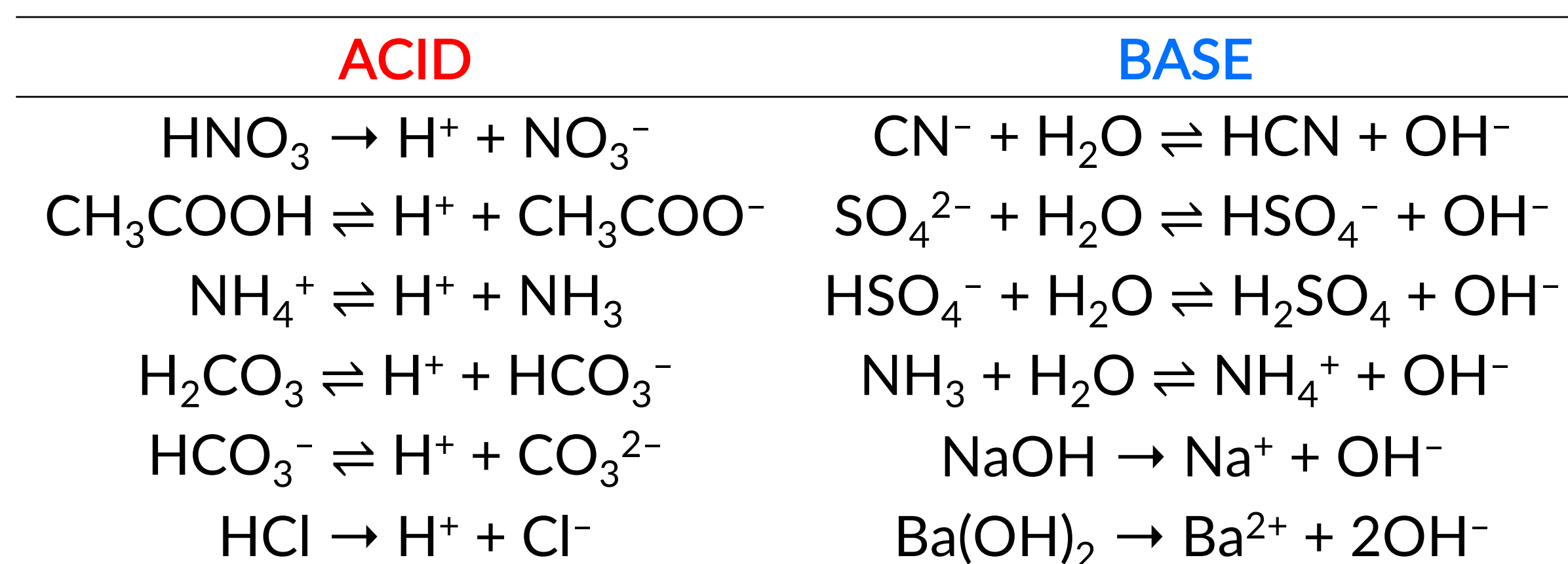
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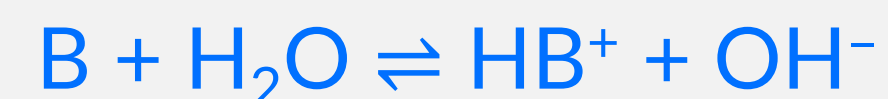
BASE: Proton (H^+) *acceptor*, or reacts with H^+ to give off OH^-

To help, consider the following examples of **acids**. Notice how the acids give off H^+ and an anion.

Now consider the following examples of **bases**, some reacting with water (H_2O). Pay close attention to how each base reacts (with water) and what each base gives off.



In general, we can express the reactivity
of **ACIDS** and **BASES** as:

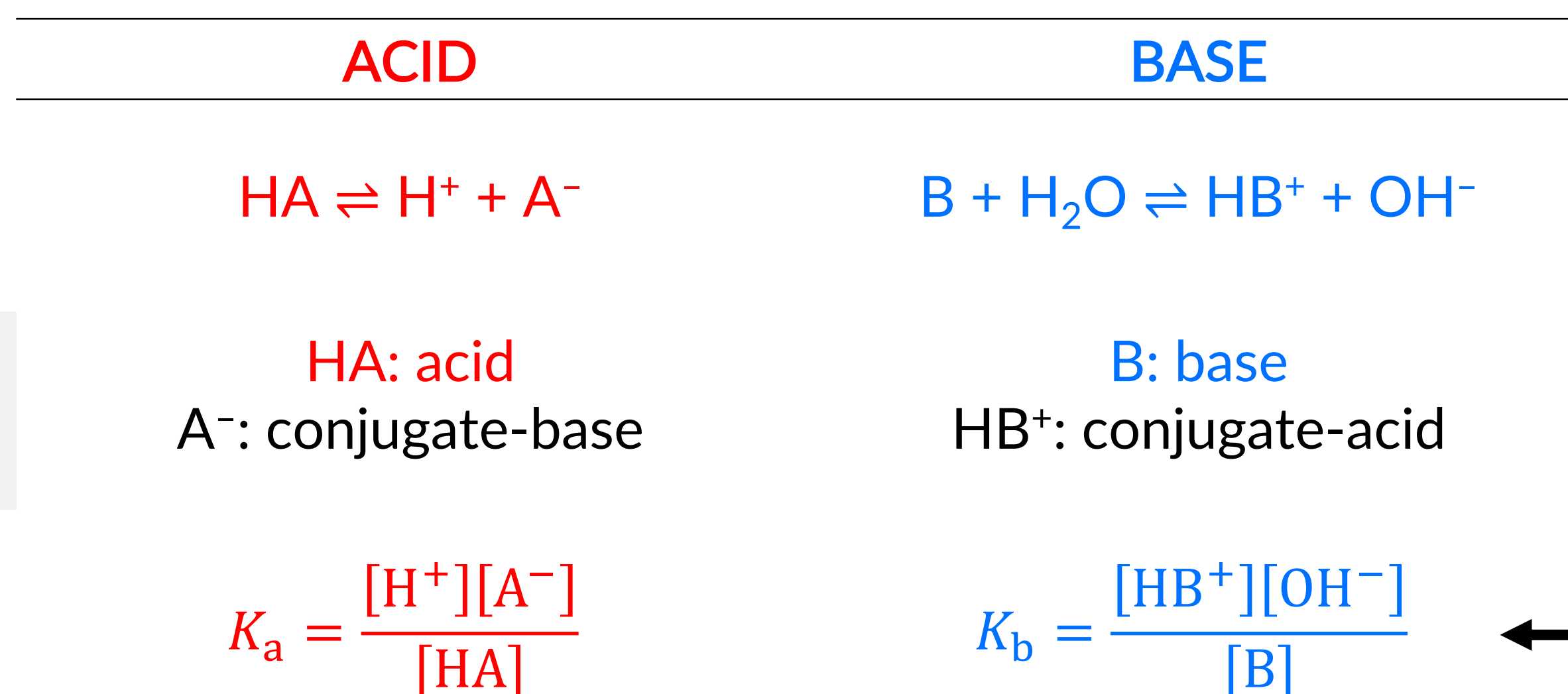


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Exercise: Go back to the previous slide and identify the conjugate-bases and conjugate-acids of each of the acids and bases, respectively.

← These are equilibrium constants!

Let's come back to this chart in a bit.

Measuring acidity or basicity: pH & pOH

Because Brønsted-Lowry theory defines acids as giving off H^+ and bases as giving off OH^- , we can formulate a metric to determine how *acidic an acid is* and how *basic a base is*.

Consider the aqueous equilibria associated with *acids* and *bases* below and their equilibrium constants: K_a and K_b .

For an *acid*, the amount of H^+ liberated, $[H^+]$, will define its acidity.

For a *base*, the amount of OH^- liberated, $[OH^-]$, will define its basicity.

Because concentrations can span large ranges (from very small to very large), we can take a logarithm of the concentrations to give us a more “user-friendly” number: *pH* and *pOH*.

ACID



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

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

BASE



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

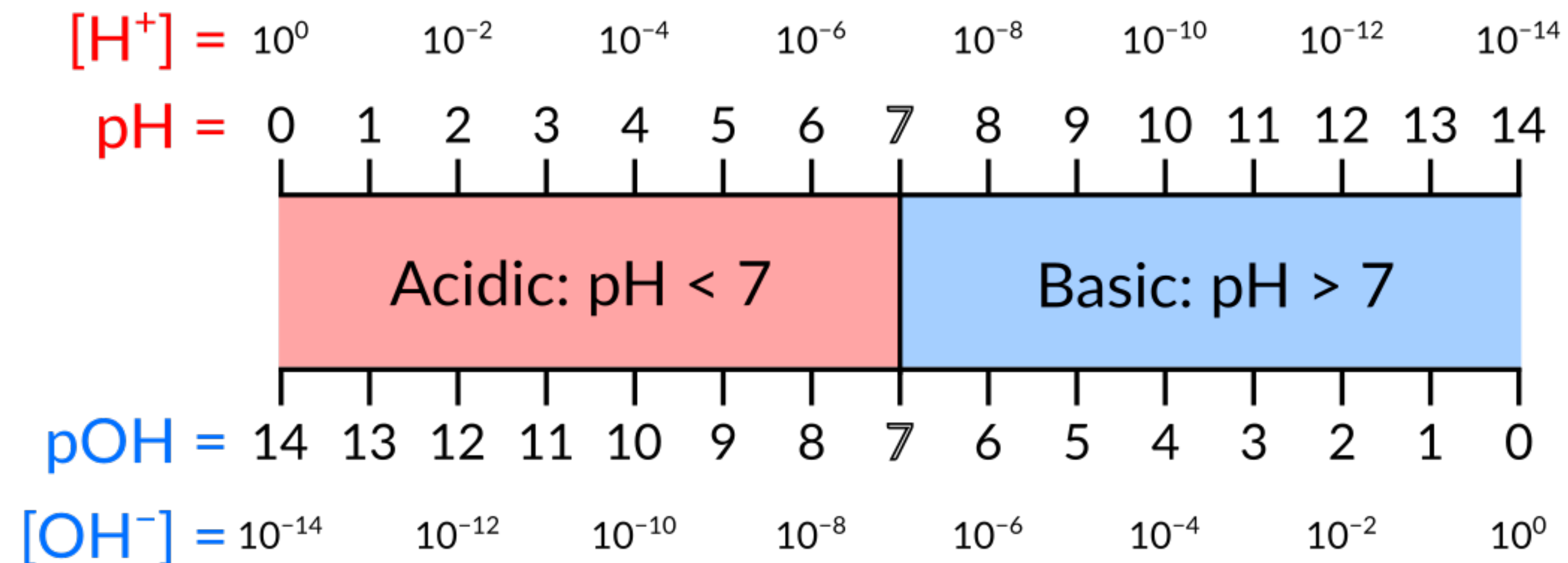
$$pOH = -\log[OH^-]$$

Measuring acidity or basicity: pH & pOH

Typically, we work in terms of pH, where acidic solutions have $\text{pH} < 7$ and basic solutions have $\text{pH} > 7$. Neutral solutions (pure water) have $\text{pH} = 7$.

The pH scale below shows how we can relate $[\text{H}^+]$ and pH (or $[\text{OH}^-]$ and pOH).

$$\begin{aligned} \text{pH} &= -\log[\text{H}^+] & \text{pOH} &= -\log[\text{OH}^-] \\ [\text{H}^+] &= 10^{-\text{pH}} & [\text{OH}^-] &= 10^{-\text{pOH}} \\ \text{pH} + \text{pOH} &= 14 \end{aligned}$$



What does K_a or K_b tell me?

Since the K_a (or K_b) value is an equilibrium constant, its magnitude tells us to what extent the acid (or base) dissociates to give off H^+ (or OH^-).

Therefore, we have the following summary table:

Strong Acid	Weak Acid
$K_a > 1$	$K_a < 1$
Complete dissociation; <u>Not</u> an equilibrium; All H^+ and A^- in solution No HA	Partial dissociation; Equilibrium lies to the left; Mostly HA in solution, some H^+ and A^-
$HA \rightarrow H^+ + A^-$	$HA \rightleftharpoons H^+ + A^-$

Strong Acids: HCl, HBr, HI, H_2SO_4 , HNO_3 , $HClO_4$

Strong Bases: LiOH, NaOH, KOH, $Ca(OH)_2$, $Ba(OH)_2$, $Sr(OH)_2$

ACID



HA: acid

A^- : conjugate base

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

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

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

BASE



B: base

HB^+ : conjugate acid

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

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

$$[OH^-] = 10^{-pOH}$$

Memorize these strong acids/base!
Everything else is weak.

PRACTICE PROBLEM 1

Calculate the pH of each of the following solutions of strong acids and strong bases.

— *answer* —

	Concentration	Acid/Base	pH
A	0.001 M	HI	
B	0.1 M	H ₂ SO ₄	
C	0.76 M	NaOH	
D	2.8×10^{-4} M	Ca(OH) ₂	

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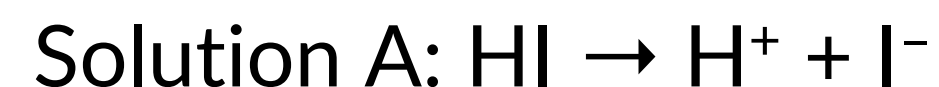
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$$[\text{H}^+] = \frac{0.001 \text{ mol HI}}{1 \text{ L}} \times \frac{1 \text{ mol H}^+}{1 \text{ mol HI}} = 0.001 \text{ M}$$

$$\text{pH} = -\log(0.001) = 3.0$$



$$[\text{OH}^-] = \frac{0.76 \text{ mol NaOH}}{1 \text{ L}} \times \frac{1 \text{ mol OH}^-}{1 \text{ mol NaOH}} = 0.76 \text{ M}$$

$$\text{pOH} = -\log(0.76) = 0.119$$

$$\text{pH} = 14 - \text{pOH} = 13.88$$

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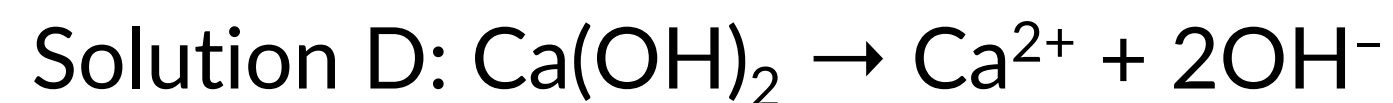


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For solutions B and D, we need to take into account the stoichiometry of the dissociation.



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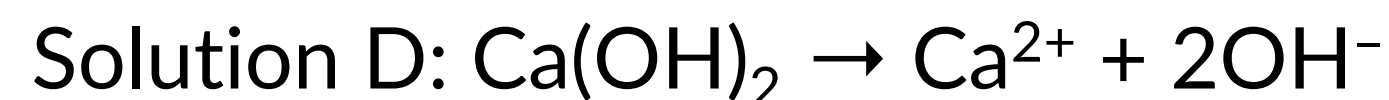
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For solutions B and D, we need to take into account the stoichiometry of the dissociation.



$$[\text{H}^+] = \frac{0.1 \text{ mol H}_2\text{SO}_4}{1 \text{ L}} \times \frac{2 \text{ mol H}^+}{1 \text{ mol H}_2\text{SO}_4} = 0.2 \text{ M}$$

$$\text{pH} = -\log(0.2) = 0.7$$



$$[\text{OH}^-] = \frac{2.8 \times 10^{-4} \text{ mol Ca(OH)}_2}{1 \text{ L}} \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ca(OH)}_2} = 5.6 \times 10^{-4} \text{ M}$$

$$\text{pOH} = -\log(5.6 \times 10^{-4}) = 3.252$$

$$\text{pH} = 14 - \text{pOH} = 10.75$$

PRACTICE PROBLEM 2

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— *answer* —

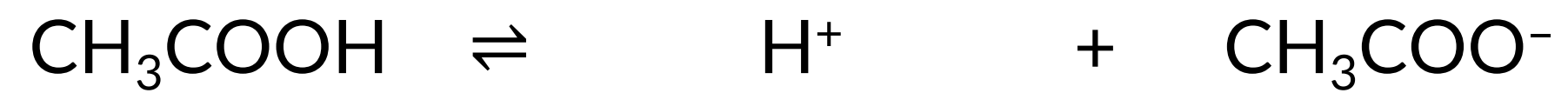
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Always begin by writing out the balanced equilibrium expression for the weak acid dissociation.



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Construct an ICE chart.

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I					
C					
E					

$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = 1.76 \times 10^{-5}$$

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C	- x		+ x		+ x
E	0.10 - x		x		x

Because the value of K_a is very small, we know that the system will undergo very little change (x value) toward the right in order to achieve equilibrium. In other words, we can take "x" to be very small/negligible and invoke the approximation that:

$$0.10 - x \approx 0.10$$

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$$1.76 \times 10^{-5} = \frac{(x)(x)}{0.10 - x}$$

$$1.76 \times 10^{-5} = \frac{x^2}{0.10}$$

$$x = 0.00133 \text{ M} = [\text{H}^+]$$

$$\text{pH} = -\log(0.00133) = 2.88$$

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$$K_w = K_a \times K_b = 1.0 \times 10^{-14}$$
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C	- x		+ x		+ x
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C	-x		+x		+x
E	0.20 - x		x		x

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$$5.5_6 \times 10^{-10} = \frac{(x)(x)}{0.20 - x}$$

$$5.5_6 \times 10^{-10} = \frac{x^2}{0.20}$$

$$x = 1.0_5 \times 10^{-5} \text{ M} = [\text{H}^+]$$

$$\text{pH} = -\log(1.0_5 \times 10^{-5}) = 4.98$$

Because the value of K_a is very small, we know that the system will undergo very little change (x value) toward the right in order to achieve equilibrium. In other words, we can take "x" to be very small/negligible and invoke the approximation that:

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Again, we can construct an ICE chart. Fill in the initial values, then the stoichiometric changes, and finally the equilibrium concentration values. **Now what?**

	HNO_2	\rightleftharpoons	H^+	+	NO_2^-
I	0.0516		0		0
C	- x		+ x		+ x
E	0.0516 - x		x		x

$$K_a = \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]}$$
$$= \frac{(x)(x)}{0.0516 - x}$$

PRACTICE PROBLEM 4

Determine the K_a of a 0.0516 M solution of HNO_2 with a $\text{pH} = 2.34$.

— answer —

HNO_2 is a weak acid, so we can start by writing the weak acid dissociation equilibrium and the associated equilibrium expression.

Again, we can construct an ICE chart. Fill in the initial values, then the stoichiometric changes, and finally the equilibrium concentration values. Now what?

They give us the pH , which tells us $[\text{H}^+]$ at equilibrium, and remember that $x = [\text{H}^+]$.

$$\begin{aligned}\text{pH} &= -\log[\text{H}^+] \\ [\text{H}^+] &= 10^{-\text{pH}} = 10^{-2.34} = 0.00457 \text{ M}\end{aligned}$$

$$\begin{aligned}K_a &= \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]} \\ &= \frac{(x)(x)}{0.0516 - x}\end{aligned}$$

	HNO_2	\rightleftharpoons	H^+	+	NO_2^-
I	0.0516		0		0
C	- x		+ x		+ x
E	0.0516 - x		x		x

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They give us the pH , which tells us $[\text{H}^+]$ at equilibrium, and remember that $x = [\text{H}^+]$.

Let's replace this value into the ICE chart.

$$\begin{aligned}\text{pH} &= -\log[\text{H}^+] \\ [\text{H}^+] &= 10^{-\text{pH}} = 10^{-2.34} = 0.0045_7 \text{ M}\end{aligned}$$

$$\begin{aligned}K_a &= \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]} \\ &= \frac{(x)(x)}{0.0516 - x}\end{aligned}$$

	HNO_2	\rightleftharpoons	H^+	+	NO_2^-
I	0.0516		0		0
C	- 0.0045 ₇		+ 0.0045 ₇		+ 0.0045 ₇
E	0.0516 - 0.0045 ₇		0.0045 ₇		0.0045 ₇

PRACTICE PROBLEM 4

Determine the K_a of a 0.0516 M solution of HNO_2 with a $\text{pH} = 2.34$.

— answer —

HNO_2 is a weak acid, so we can start by writing the weak acid dissociation equilibrium and the associated equilibrium expression.

Again, we can construct an ICE chart. Fill in the initial values, then the stoichiometric changes, and finally the equilibrium concentration values. Now what?

They give us the pH , which tells us $[\text{H}^+]$ at equilibrium, and remember that $x = [\text{H}^+]$.

Let's replace this value into the ICE chart. Finally, solve for K_a from the equilibrium concentrations.

$$\begin{aligned}\text{pH} &= -\log[\text{H}^+] \\ [\text{H}^+] &= 10^{-\text{pH}} = 10^{-2.34} = 0.00457 \text{ M}\end{aligned}$$

	HNO_2	\rightleftharpoons	H^+	+	NO_2^-
I	0.0516		0		0
C	- 0.0045 ₇		+ 0.0045 ₇		+ 0.0045 ₇
E	0.0516 - 0.0045 ₇		0.0045 ₇		0.0045 ₇

$$\begin{aligned}K_a &= \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]} \\ &= \frac{(x)(x)}{0.0516 - x} \\ &= \frac{(0.0045_7)^2}{0.0516 - 0.0045_7}\end{aligned}$$

$$K_a = 4.5 \times 10^{-4}$$