

## Acids and Bases

- Acids generate  $H^+$  in water
  - Bases generate  $OH^-$  in water
- Aqueous equilibria**

Syllabus Learning Outcomes : 6, 7

pH is proportional to  $[H^+]$  (acidity)  
in solution

Low pH: high  $[H^+]$       High pH: low  $[H^+]$

Acidic solution	pH < 7
Neutral	pH = 7
Basic solution	pH > 7

### Calculate pH from $[H^+]$

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

In a neutral water solution,

$$[H^+] = [OH^-] = 1.00 \times 10^{-7} \text{ M at } 25^\circ \text{C}$$

$$pH = -\log [H^+] = -\log (1.00 \times 10^{-7})$$

$$= - [0 + (-7)]$$

$$= 7, \text{ the pH of a neutral solution}$$

### Calculate pH from $[H^+]$

If the  $[H^+]$  of juice is  $1.6 \times 10^{-3} \text{ M}$ ,  
the pH is \_\_\_\_?

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

$$pH = -\log(1.6 \times 10^{-3})$$

$$pH = 2.80$$

(acidic)

### Calculate $[H^+]$ from pH

If the pH of juice is 3.12, it is \_\_\_\_\_.

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

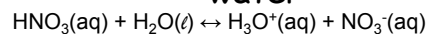
$$\log [H^+] = -pH$$

Take antilog ( $10^x$ ) and get

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

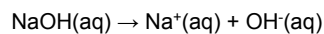
$$[H^+] = 10^{-3.12} = 7.6 \times 10^{-4} \text{ M}$$

### Strong acids ionize 100% in water



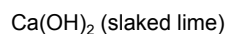
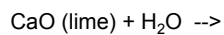
6 Strong acids:  $HNO_3$ ,  $HCl$ ,  $HBr$ ,  $HI$ ,  $H_2SO_4$ ,  $HClO_4$

### Strong bases ionize 100% in water



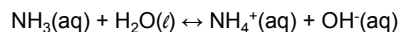
Strong bases: LiOH, NaOH, KOH, Ba(OH)<sub>2</sub>

Other common strong bases  
are CaO and Ca(OH)<sub>2</sub>.



### Weak bases ionize ~5% or less in water

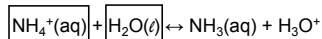
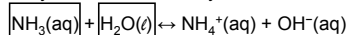
Ammonia, NH<sub>3</sub>, is a well known weak base



### Bronsted-Lowry acid-base theory is

- ACIDS DONATE H<sup>+</sup> IONS
- BASES ACCEPT H<sup>+</sup> IONS

Identify the Bronsted-Lowry acids and bases:

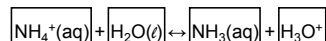
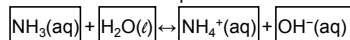


Conjugate acid: base that gained H<sup>+</sup>

Conjugate base: acid that lost H<sup>+</sup>

- Find conjugate acid-base pairs across the reaction arrow.

Identify the Bronsted-Lowry conjugate acid-base pairs



### Conjugate acid-base pairs

Acid	Conjugate Base
acetic, CH <sub>3</sub> CO <sub>2</sub> H	CH <sub>3</sub> CO <sub>2</sub> <sup>-</sup> , acetate
ammonium, NH <sub>4</sub> <sup>+</sup>	NH <sub>3</sub> , ammonia
bicarbonate, HCO <sub>3</sub> <sup>-</sup>	CO <sub>3</sub> <sup>2-</sup> , carbonate

Weak acids (or bases) ionize < 5%.

### Water (H<sub>2</sub>O)

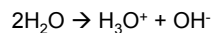
can be an acid and a base. (amphoteric/amphiprotic)

can autoionize

has an autoionization constant, Kw



### Water autoionizes



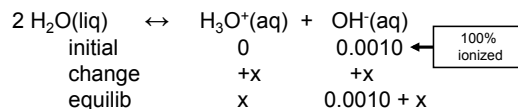
$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.00 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

In a neutral solution  $[\text{H}_3\text{O}^+] = [\text{OH}^-]$

$$\text{so } K_w = [\text{H}_3\text{O}^+]^2 = [\text{OH}^-]^2$$

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

Calculate  $[\text{H}_3\text{O}^+]$  and  $[\text{OH}^-]$  from adding 0.0010 mol of NaOH to make 1.0L of solution.



$$K_w = (x)(0.0010 + x)$$

Because  $x \ll 0.0010 \text{ M}$ , assume  $0.0010 + x = 0.0010$

$$[\text{OH}^-] = 0.0010 \text{ M} \quad (\text{Is } 100^\circ\text{K} < \text{Co?})$$

$$[\text{H}_3\text{O}^+] = K_w / 0.0010 = 1.0 \times 10^{-11} \text{ M}$$

The solution is acidic / basic

### Be able to convert among concentration scales

	pH	$[\text{H}_3\text{O}^+]$	$[\text{OH}^-]$	pOH
Basic	14	$10^{-14}$	1	0
Neutral	7	$10^{-7}$	$10^{-7}$	7
Acidic	0	1	$10^{-14}$	14

$$\text{pH} + \text{pOH} = 14$$

In general  $\text{pX} = -\log X$

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

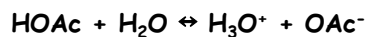
$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.00 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

Take the log of both sides, giving

$$-\log(10^{-14}) = -\log[\text{H}_3\text{O}^+] + (-\log[\text{OH}^-])$$

$$\text{p}K_w = 14 = \text{pH} + \text{pOH}$$

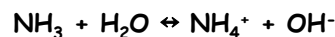
### Define $K_a$ for (weak) acids



$$K = K_a = \frac{[\text{H}_3\text{O}^+][\text{OAc}^-]}{[\text{HOAc}]} = 1.8 \times 10^{-5}$$

- (K is  $K_a$  for acid)
- $K_a$  for weak acids is  $\ll 1$
- $K_a$  can be looked up (see Table 17.3, p770)
- pH of weak acids is 2-7

### Define $K_b$ for (weak) bases



$$K = K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = 1.8 \times 10^{-5}$$

- (K is  $K_b$  for base)
- $K_b$  for weak bases is  $\ll 1$
- $K_b$  can be looked up (see Table 17.3, p770)
- pH of weak bases is 7-12

## Table of Ka and Kb values

Substance	Ka	Substance	Kb
HNO <sub>3</sub>	large	NH <sub>2</sub> <sup>-</sup>	large
H <sub>3</sub> O <sup>+</sup>	55.6	S <sup>2-</sup>	1.0x10 <sup>5</sup>
HF	6.6x10 <sup>-4</sup>	OH <sup>-</sup>	55.6
HOAc	1.8x10 <sup>-5</sup>	NH <sub>3</sub>	1.8x10 <sup>-5</sup>
HCN	6.2x10 <sup>-10</sup>	HCO <sub>3</sub> <sup>-</sup>	2.4x10 <sup>-8</sup>
H <sub>2</sub> O	1.8x10 <sup>-16</sup>	H <sub>2</sub> O	1.8x10 <sup>-16</sup>
HS <sup>-</sup>	1.0x10 <sup>-19</sup>	NO <sub>3</sub> <sup>-</sup>	small
NH <sub>3</sub>	small		

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## Relation of K<sub>a</sub>, K<sub>b</sub>, [H<sub>3</sub>O<sup>+</sup>], pH and K<sub>w</sub>

Larger K<sub>a</sub> means stronger acid, higher [H<sub>3</sub>O<sup>+</sup>], and lower pH

Larger K<sub>b</sub> means stronger base, higher [OH<sup>-</sup>], and higher pH

K<sub>a</sub> for an acid times K<sub>b</sub> for its conjugate base equals K<sub>w</sub>:  
K<sub>a</sub> x K<sub>b</sub> = K<sub>w</sub>

## Predict spontaneous direction of acid-base reactions

stronger acid + stronger base ↔  
weaker conjugate acid +  
weaker conjugate base

Reactions always go from the side with the stronger acid-base pair (larger K<sub>a</sub> and K<sub>b</sub>) to side with the weaker acid-base pair (smaller K<sub>a</sub> and K<sub>b</sub>). (See Table 17.3)

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## For strong acids and bases, equilibrium lies to products

HNO<sub>3</sub>(aq) + H<sub>2</sub>O(l) ↔ H<sub>3</sub>O<sup>+</sup>(aq) + NO<sub>3</sub><sup>-</sup>(aq)  
str. acid + str. base → wk. acid + wk. base

K is large →

HOAc(l) + NaOH(aq) ↔ H<sub>2</sub>O(l) + OAc<sup>-</sup>(aq)  
Str. acid + str. base → wk. acid + wk. base

K is large →

Equilibrium lies to the side with the weaker acid-base pair.

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## Look for the stronger acid or base

Acetic acid, a weak acid, is 0.42% ionized when [HOAc] = 1.0 M.

HOAc + H<sub>2</sub>O ↔ H<sub>3</sub>O<sup>+</sup> + OAc<sup>-</sup>  
Weak acid base strong acid strong base

Weak acid means small [H<sub>3</sub>O<sup>+</sup>] and that:

1. H<sub>3</sub>O<sup>+</sup> is a stronger acid than HOAc
2. OAc<sup>-</sup> is a stronger base than H<sub>2</sub>O
3. K is small

Table 17.3

← K is small

## Mixing strong acid & strong base gives neutral products (equimolar)

HCl + NaOH ↔ H<sub>2</sub>O + NaCl

H<sup>+</sup> + Cl<sup>-</sup> + Na<sup>+</sup> + OH<sup>-</sup> ↔ H<sub>2</sub>O + Na<sup>+</sup> + Cl<sup>-</sup>

*Net ionic equation*

H<sub>3</sub>O<sup>+</sup>(aq) + OH<sup>-</sup>(aq) ↔ 2H<sub>2</sub>O(l)

K = 1/K<sub>w</sub> = 1 x 10<sup>14</sup> → K is large

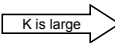
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**Mixing weak acid & strong base  
gives basic products** (equimolar)

$$\text{HOAc(aq)} + \text{OH}^{\text{-}}(\text{aq}) \leftrightarrow \text{H}_2\text{O}(\ell) + \text{OAc}^{\text{-}}$$

$$K = 1/K_b = 1/5.6 \times 10^{-10} = 1.8 \times 10^9$$

K is large 

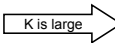
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**Mixing strong acid & weak base  
gives acidic products** (equimolar)

$$\text{H}_3\text{O}^{\text{+}} + \text{NH}_3 \leftrightarrow \text{H}_2\text{O} + \text{NH}_4^{\text{+}}$$

$$K = 1/K_a = 1/5.6 \times 10^{-10} = 1.8 \times 10^9$$

K is large 

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**Mixing weak acid & weak base,  
pH depends on Ka and Kb** (equimolar)

$$\text{HOAc} + \text{NH}_3 \leftrightarrow \text{OAc}^{\text{-}} + \text{NH}_4^{\text{+}}$$

$K_a(\text{HOAc}) = K_b(\text{NH}_3)$ , products neutral

$$\text{HF} + \text{NH}_3 \leftrightarrow \text{F}^{\text{-}} + \text{NH}_4^{\text{+}}$$

$K_a(\text{HF}) > K_b(\text{NH}_3)$ , products acidic

$$\text{HCN} + \text{NH}_3 \leftrightarrow \text{CN}^{\text{-}} + \text{NH}_4^{\text{+}}$$

$K_b(\text{NH}_3) > K_a(\text{HCN})$ , products basic

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- pH of an acetic acid solution depends on [HOAc].
- We could measure pH
- Next, we calculate pH

$$\text{HOAc} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^{\text{+}} + \text{OAc}^{\text{-}}$$

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Calculate  $[\text{HOAc}]_E$ ,  $[\text{H}_3\text{O}^{\text{+}}]_E$ ,  $[\text{OAc}^{\text{-}}]_E$ ,  
and pH for 1.00M HOAc

**Step 1. Put [ ] in ICE table.**

	[HOAc]	[H <sub>3</sub> O <sup>+</sup> ]	[OAc <sup>-</sup> ]
<b>initial</b>			
<b>change</b>			
<b>equilib</b>			

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Calculate  $[\text{HOAc}]_E$ ,  $[\text{H}_3\text{O}^{\text{+}}]_E$ ,  $[\text{OAc}^{\text{-}}]_E$ ,  
and pH for 1.00M HOAc

**Step 1. Put [ ] in ICE table.**

	[HOAc]	[H <sub>3</sub> O <sup>+</sup> ]	[OAc <sup>-</sup> ]
<b>initial</b>	1.00	0	0
<b>change</b>	-x	+x	+x
<b>equilib</b>	1.00-x	x	x

Check if  $100 \cdot K < C_0$ ,  $C_0$  is initial concentration

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31  
Calculate  $[\text{HOAc}]_E$ ,  $[\text{H}_3\text{O}^+]_E$ ,  $[\text{OAc}^-]_E$ ,  
and pH for 1.00M HOAc

Step 2. Write Ka.

$$K_a = 1.8 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+][\text{OAc}^-]}{[\text{HOAc}]}$$

$$1.8 \times 10^{-5} = \frac{x^2}{(1.00-x)}$$

Solve using quadratic equation or a  
simplifying assumption ( $100 \cdot K < C_0$ ).

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32  
Calculate  $[\text{HOAc}]_E$ ,  $[\text{H}_3\text{O}^+]_E$ ,  $[\text{OAc}^-]_E$ ,  
and pH for 1.00M HOAc

Step 3. Simplify and solve.

$$1.8 \times 10^{-5} = \frac{x^2}{(1.00-x)}$$

Because  $100 \cdot 1.8 \times 10^{-5} < 1$ ,  $1-x \approx 1$

$$1.8 \times 10^{-5} = \frac{x^2}{1.00}$$

$$x = 4.24 \times 10^{-3}$$

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33  
Calculate  $[\text{HOAc}]_E$ ,  $[\text{H}_3\text{O}^+]_E$ ,  $[\text{OAc}^-]_E$ ,  
and pH for 1.00M HOAc

Step 4. Calculate concentrations.

$$[\text{H}_3\text{O}^+] = [\text{OAc}^-] = x = 4.24 \times 10^{-3} \text{M}$$

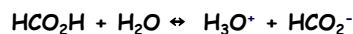
$$\text{pH} = -\log(4.24 \times 10^{-3}) = 2.37$$

$$[\text{HOAc}] = 1.00 - 0.004 \approx 1.00 \text{M}$$

Assumption was valid! ( $100 \cdot K < C_0$ )

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34  
Calculate the pH of 0.0010 M  
formic acid,  $\text{HCO}_2\text{H}$ ,  $K_a = 1.8 \times 10^{-4}$



$100K < C_0$  is not valid, making assumption gives

$$[\text{H}_3\text{O}^+] = 4.2 \times 10^{-4} \text{M}, \text{pH} = 3.37$$

Exact solution by quadratic equation gives

$$[\text{H}_3\text{O}^+] = [\text{HCO}_2^-] = 3.4 \times 10^{-4} \text{M}$$

$$[\text{HCO}_2\text{H}] = 0.0010 - 3.4 \times 10^{-4} = 0.0007 \text{M}$$

$$\text{pH} = 3.47$$

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35  
Calculate the pH of 0.010M  
 $\text{NH}_3$ ,  $K_b = 1.8 \times 10^{-5}$



Step 1. Put [ ] in ICE table

	$[\text{NH}_3]$	$[\text{NH}_4^+]$	$[\text{OH}^-]$
initial			
change			
equilib			

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36  
Calculate the pH of 0.010M  
 $\text{NH}_3$ ,  $K_b = 1.8 \times 10^{-5}$



Step 1. Put [ ] in ICE table

	$[\text{NH}_3]$	$[\text{NH}_4^+]$	$[\text{OH}^-]$
initial	0.010	0	0
change	$0.010-x$	$x$	$x$
equilib	$0.010-x$	$x$	$x$

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37  
Calculate the pH of 0.010M  
 $\text{NH}_3$ ,  $K_b=1.8 \times 10^{-5}$

Step 2. Solve  $K_b$  expression

$$K_b = 1.8 \times 10^{-5} = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

$$1.8 \times 10^{-5} = \frac{x^2}{(0.010 - x)}$$

$$100 \times 1.8 \times 10^{-5} < 0.010, \text{ so}$$

$$1.8 \times 10^{-5} = \frac{x^2}{0.010}$$

38  
Calculate the pH of 0.010M  
 $\text{NH}_3$ ,  $K_b=1.8 \times 10^{-5}$

Step 3. Simplify and solve

$$1.8 \times 10^{-5} = \frac{x^2}{0.010}$$

$$x = 4.2 \times 10^{-4}$$

$$[\text{OH}^-] = 4.2 \times 10^{-4}$$

$$\text{pOH} = 3.38, \text{ so } \text{pH} = 14 - 3.38 = 10.62$$