

# Acids and Bases

By

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## Acids

Acids: taste sour and cause certain dyes to change color. Example HCl.

### ACIDS

Hydronium ion:  $[\text{H}_3\text{O}^+] = [\text{H}^+] > 1.0 \times 10^{-7} \text{ M}$

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

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

$$\text{pH} < 7.00$$

# BASES

Bases: taste bitter and feel soapy. Example NaOH

## BASES

Hydroxide ion:  $[\text{OH}^-] > 1.0 \times 10^{-7} \text{ M}$

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

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

$$\text{pH} > 7.00$$

## Relationship Between ACIDS and BASES

$K_w$  is called the ion-product constant.

At 25°C the ion-product of water is:

### ACIDS and BASES

$$K_w = [\text{H}_3\text{O}^+] [\text{OH}^-] = [\text{H}^+] [\text{OH}^-] = 1.0 \times 10^{-14}$$

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

Derivatives:

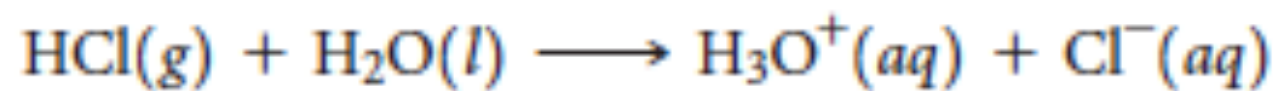
$$\text{pH} = 14 - \text{pOH} \quad \text{and} \quad \text{pOH} = 14 - \text{pH}$$

$$[\text{H}_3\text{O}^+] = 1.0 \times 10^{-14} / [\text{OH}^-]$$

$$[\text{OH}^-] = 1.0 \times 10^{-14} / [\text{H}_3\text{O}^+]$$

## Strong Acid

Strong acid: 100% ionized in H<sub>2</sub>O. Example HCl:



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

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

## Strong Base

Strong Base: 100% ionized in H<sub>2</sub>O.

1) Example NaOH:



a)  $[\text{OH}^-] = [\text{NaOH}]$

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

c)  $\text{pH} + \text{pOH} = 14$  so,  $\text{pH} = 14 - \text{pOH}$

2) Example Ca(OH)<sub>2</sub>:



a)  $[\text{OH}^-] = 2 [\text{Ca(OH)}_2]$

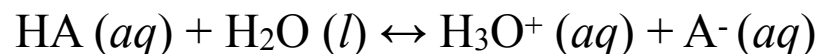
## Weak Acid

Weak acids are only partially ionized in aqueous solution.

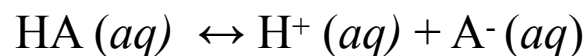
It has  $K_a$ : is called the **acid-dissociation constant**.

The larger the  $K_a$  the stronger the acid.

There is a mixture of ions and un-ionized acid in solution. Weak acids are in equilibrium.



**Or:**



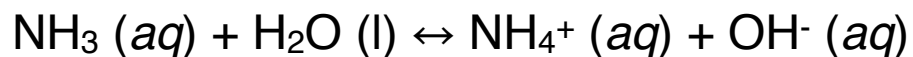
**$K_a$  : acid-dissociation constant.**

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} \quad \text{or} \quad K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

## Weak Bases

Weak bases remove protons from substance. It has  $K_b$ : is called the **base-dissociation constant**. The larger the  $K_a$  the stronger the base.

There is an equilibrium between the base and the resulting ions:



**$K_a$  : acid-dissociation constant.**

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



## Relationship Between $K_a$ and $K_b$

At 25°C:

### Relationship Between $K_a$ and $K_b$

$$K_a \times K_b = 1.0 \times 10^{-14}$$

$$pK_a + pK_b = 14.00$$

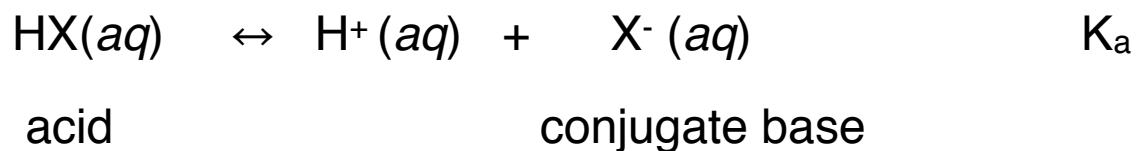
$$pK_a = -\log K_a$$

$$pK_b = -\log K_b$$

The larger  $K_a$  (and the smaller  $pK_a$ ), the smaller  $K_b$  (and the larger  $pK_b$ ).  
The stronger the acid, the weaker its conjugate base and vice versa.

## Buffered Solutions- I

A buffer consists of a mixture of a weak acid (HX) and its conjugate base (X<sup>-</sup>):



### Henderson-Hasselbalch Equation

$$\text{pH} = \text{p}K_a + \log \frac{[\text{X}^-]}{[\text{HX}]} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

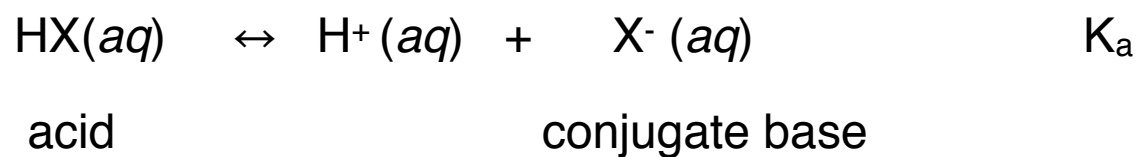
$$\text{p}K_a = -\log K_a$$

$$K_a \times K_b = 1.0 \times 10^{-14}$$

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

## Buffered Solutions- II

A buffer contains approximately equal amount of a weak acid (HX) and its conjugate base (X<sup>-</sup>). (not necessarily equal amount but a substantial concentration of each)



### Henderson-Hasselbalch Equation

$$\text{pH} = \text{p}K_a + \log \frac{[\text{X}^-]}{[\text{HX}]} = \text{p}K_a = -\log K_a$$

$$[\text{X}^-] = [\text{HX}]$$

### Buffer Capacity:

The greater the concentration of the two buffer components, the greater the ability of the buffer to resist the changes in pH (the greater the capacity). Lowering the concentration will lower the capacity of the buffer.

### Buffer Efficiency:

The efficiency of the buffer is the greatest when the concentrations of the two components are equal, but this condition is not necessary for the buffer to work.

$$[X^-] = [HX] \quad \text{so} \quad [X^-] / [HX] = 1$$

### Selecting the Buffer:

The  $pK_a$  of the weak acid to be used in buffer should be as close as possible to the desired pH.



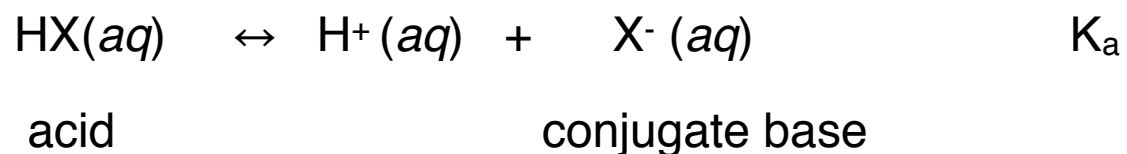


## Buffered Solutions-III

### Addition of Base

Adding more of the buffer base will make the buffer more resistant to pH change when adding more acid.

Addition of a base (**OH<sup>-</sup>**) to a buffered solution:



#### Henderson-Hasselbalch Equation

$$\text{pH} = \text{pK}_a + \log \frac{[\text{base} + \text{OH}^-]}{[\text{acid} - \text{OH}^-]} = \text{pK}_a + \log \frac{[\text{X}^- + \text{OH}^-]}{[\text{HX} - \text{OH}^-]}$$