Acids and Bases

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Acids

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

ACIDS
Hydronium ion: $[H_3O^+] = [H^+] > 1.0 \times 10^{-7}$ M
$pH = -\log[H^+] = -\log[H_3O^+]$
$[H_3O^+] = 10^{-pH}$
$pH < 7.00$

BASES

Bases: taste bitter and feel soapy. Example NaOH

BASES
Hydroxide ion: $[OH^-] > 1.0 \times 10^{-7}$ M
$pOH = -\log [OH^-]$
$[OH^-] = 10^{\mathbf{-pOH}}$
$pH > 7.00$

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
$K_w = [H_3O^+][OH^-] = [H^+][OH^-] = 1.0 \times 10^{-14}$ $pH + pOH = 14$

Derivatives:

$pH = 14 - pOH$ and $pOH = 14 - pH$

$$[H_3O^+] = 1.0 \times 10^{-14} / [OH^-]$$

$$[OH^-] = 1.0 \times 10^{-14} / [H_3O^+]$$

Strong Acid

Strong acid: 100% ionized in H_2O. Example HCl:

$$HCl(g) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

$$pH = -\log[H_3O^+]$$

$$[H_3O^+] = [HCl]$$

Strong Base

Strong Base: 100% ionized in H_2O.

1) Example NaOH:

$$NaOH \longrightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

a) $[OH^-] = [NaOH]$

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

c) $pH + pOH = 14$ so, $pH = 14 - pOH$

2) Example $Ca(OH)_2$:

$$Ca(OH)_2 \longrightarrow Ca^{2+}_{(aq)} + \mathbf{2}\ OH^-_{(aq)}$$

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

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.

$$HA\ (aq) + H_2O\ (l) \leftrightarrow H_3O^+\ (aq) + A^-\ (aq)$$

Or:

$$HA\ (aq) \leftrightarrow H^+\ (aq) + A^-\ (aq)$$

K_a : acid-dissociation constant.
$K_a = \dfrac{[H_3O^+][A^-]}{[HA]}$ or $K_a = \dfrac{[H^+][A^-]}{[HA]}$

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:

Weak base + H_2O (l) ↔ conjugate acid + OH^- (*aq*)

NH_3 (*aq*) + H_2O (l) ↔ NH_4^+ (*aq*) + OH^- (*aq*)

K_a : acid-dissociation constant.
$$K_b = \frac{[NH_4^+]\,[OH^-]}{[NH_3]}$$

Relationship Between K_a and K_b

At 25^0C:

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^-):

$$HX(aq) \leftrightarrow H^+(aq) + X^-(aq) \qquad K_a$$

acid conjugate base

Henderson-Hasselbalch Equation
$pH = pK_a + \log \frac{[X^-]}{[HX]} = pK_a + \log \frac{[base]}{[acid]}$

$$pK_a = -\log K_a$$

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

$$pH = -\log[H_3O^+]$$

Buffered Solutions- II

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

$$HX(aq) \leftrightarrow H^+(aq) + X^-(aq) \qquad K_a$$

acid conjugate base

Henderson-Hasselbalch Equation
$pH = pK_a + \log \frac{[X^-]}{[HX]} \quad = \quad pK_a = -\log K_a$
$[X^-] = [HX]$

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-II

A buffered solution or buffer is a solution that resists a change in pH upon addition of small amounts of strong acid or strong base.

$$HX(aq) \leftrightarrow H^+(aq) + X^-(aq) \qquad K_a$$

acid conjugate base

When a small amount of OH^- is added to the buffer, the OH^- reacts with HX to produce X^- and water.

When a small amount of H^+ is added to the buffer, X^- is consumed to produce HX.

The $[X^-]/[HX]$ ratio is more or less constant, so the pH does not change significantly.

Buffered Solutions-III
Addition of Acid

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

Addition of an acid (**H^+**) to a buffered solution:

$$HX(aq) \leftrightarrow H^+(aq) + X^-(aq) \qquad K_a$$

acid conjugate base

Henderson-Hasselbalch Equation
$pH = pK_a + \log \dfrac{[\text{base} - H^+]}{[\text{acid} + H^+]} = pK_a + \log \dfrac{[X^- - H^+]}{[HX + H^+]}$

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^-**) to a buffered solution:

$$HX(aq) \leftrightarrow H^+(aq) + X^-(aq) \qquad K_a$$

acid — conjugate base

Henderson-Hasselbalch Equation
$pH = pK_a + \log \dfrac{[\text{base} + OH^-]}{[\text{acid} - OH^-]} = pK_a + \log \dfrac{[X^- + OH^-]}{[HX - OH^-]}$