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
$pOH = -\log [OH^-]$
$[OH^-] = 10^{-pOH}$
$pH > 7.00$

Relationship Between ACIDS and BASES

K_w is called the ion-product constant.

At 25^oC 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 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

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

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

acid conjugate base

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

Buffered Solutions-III
Addition of Base

Addition of a base (**OH^-**) to a buffered solution:

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

acid conjugate base

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