

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:

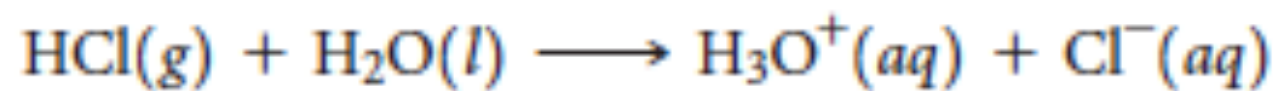
$$\text{pH} = 14 - \text{pOH} \quad \text{and} \quad \text{pOH} = 14 - \text{pH}$$

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

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

Strong Acid

Strong acid: 100% ionized in H₂O. Example HCl:



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

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

Strong Base

Strong Base: 100% ionized in H₂O.

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)₂:



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

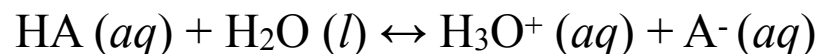
Weak Acid

Weak acids are only partially ionized in aqueous solution.

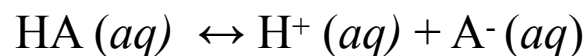
It has K_a : is called the **acid-dissociation constant**.

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:



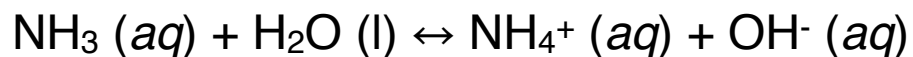
K_a : acid-dissociation constant.

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

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

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

$$pK_a + pK_b = 14.00$$

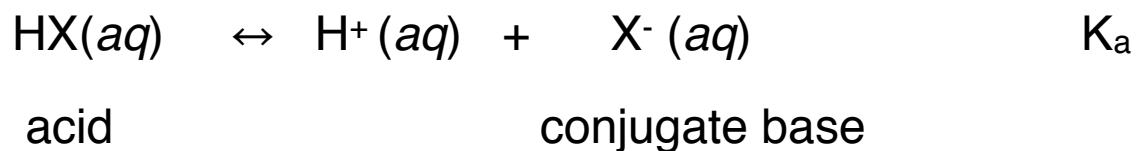
$$pK_a = -\log [K_a]$$

$$pK_b = -\log [K_b]$$

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⁻):



Henderson-Hasselbalch Equation

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

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

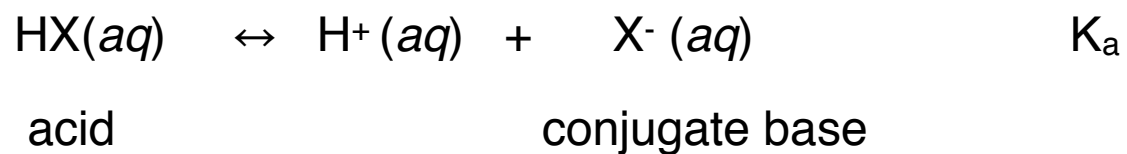
$$K_a \times K_b = 1.0 \times 10^{-14}$$

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

Buffered Solutions-III

Addition of Acid

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



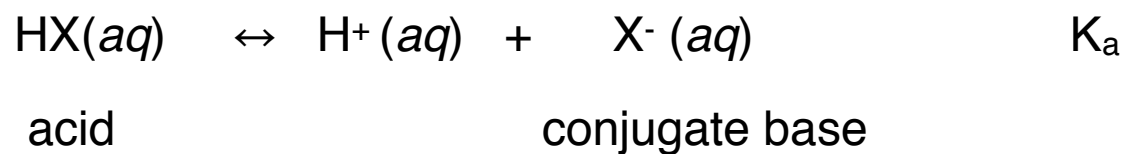
Henderson-Hasselbalch Equation

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

Buffered Solutions-III

Addition of Base

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



Henderson-Hasselbalch Equation

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