CHAPTER 15 ACIDS & BASES

Acid Base Definitions - Arrhenius vs Bronsted-Lowry

Arrhenius – Aqueous solutions only

Acid: Substance contains hydrogen and dissociates to form $\text{H}_3\text{O}^+$ in water
Base: Substance contains hydroxide group and dissociates to form $\text{OH}^-$ in water

Bronsted - Lowry – Any solvent, even gas

Acid: Proton ($\text{H}^+$) donor
Base: Proton acceptor

Any substance that is an Arrhenius acid is an Bronsted-Lowry acid. Likewise, any Arrhenius base is a Bronsted-Lowry base.

Much Broader Range

$$ \text{HCl} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{Cl}^- $$
Acid Base Conjugate acid of $\text{H}_2\text{O}$ base of $\text{HCl}$

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

Label conjugate acid-base pairs and predict direction of reaction (that is, which way equilibrium lies)

$$ \text{HCN} + \text{CO}_2\text{O}^2- \leftrightarrow \text{CN}^- + \text{HCO}_3^- $$
Acid Base Acid Base
$K_a=6.2\times10^{-10}$ $K_a=4.7\times10^{-11}$

$$ \text{HF} + \text{HCO}_3^- \leftrightarrow \text{F}^- + \text{H}_2\text{CO}_3 $$
$K_a=6.8\times10^{-4}$ $K_a=4.5\times10^{-7}$

$$ \text{H}_2\text{PO}_4^- + \text{HNO}_2 \leftrightarrow \text{H}_3\text{PO}_4 + \text{NO}_2^- $$
$K_a=7.1\times10^{-4}$ $K_a=7.1\times10^{-3}$

$$ \text{NH}_4^+ + \text{C}_2\text{H}_3\text{O}_2^- \leftrightarrow \text{NH}_3 + \text{HC}_2\text{H}_3\text{O}_2 $$
$K_a=5.7\times10^{-10}$ $K_a=1.8\times10^{-5}$
The stronger the acid, the weaker is the conjugate base and vice-versa.

Dissociation of Water

Water acts as both Bronsted-Lowry acid & base

\[ \text{H}_2\text{O} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{OH}^- \]

Acid Base Acid Base

But water is pure liquid so \([\text{H}_2\text{O}]^2\) is combined with \(K_c\) yielding \(K_w\).

\[
K_c [\text{H}_2\text{O}]^2 = [\text{H}_3\text{O}^+][\text{OH}^-]
K_w = [\text{H}_3\text{O}^+][\text{OH}^-]
\]

At 25°C \(K_w = 1.0 \times 10^{-14}\)

So in neutral water \([\text{H}_3\text{O}^+][\text{OH}^-] = x\) (same)

\[
[H_3O^+][OH^-] = K_w
(x)(x) = 1.0 \times 10^{-14}
\]

\[
x^2 = 1.0 \times 10^{-14}
\]

\[
x = [\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}
\]

Definitions

Neutral => \([\text{H}_3\text{O}^+] = [\text{OH}^-]\)
Acid => \([\text{H}_3\text{O}^+] > [\text{OH}^-]\)
Base => \([\text{H}_3\text{O}^+] < [\text{OH}^-]\)

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

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

\[
\text{pK}_w = -\log K_w \quad K_w = 10^{-K_w}
\]

At 25°C, neutral \(\text{pH} = -\log(1.0 \times 10^{-7}\text{M}) = 7.00\)

\(\text{pOH} = -\log(1.0 \times 10^{-7}\text{M}) = 7.00\)

Relationships between pH and pOH

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

\[
\log[\text{H}_3\text{O}^+] + \log[\text{H}_3\text{O}^+] = \log K_w
\]

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

\[
\text{pH} + \text{pOH} = \text{pK}_w
\]

At 25°C, \(K_w = 1.0 \times 10^{-14}\)

so \(\text{pK}_w = -\log(1.0 \times 10^{-14}) = 14.00\)

\[
\text{pH} + \text{pOH} = 14.00
\]
Is pure water always 7? K values change with temperature.

At 100° C, \( K_w = 5.5 \times 10^{-13} \)

So if pH scale had been set at 100° C:

\[ [H_3O^+][OH^-] = K_w = 5.5 \times 10^{-13} \]

At 100° C, \( K_w = 5.5 \times 10^{-13} \)

So if pH scale had been set at 100° C:

\[ [H_3O^+][OH^-] = K_w = 5.5 \times 10^{-13} \]

Neutral pH \( (x)(x) = 5.5 \times 10^{-13} \)

\[ x = [H_3O^+] = [OH^-] = 7.4 \times 10^{-7}M \]

\[ pH = pOH = -\log(7.4 \times 10^{-7}M) = 6.13 \]

At 100° C, \( pH + pOH = pK_w = 12.26 \)

**Strong Acids and Bases**

Strong Acids have large \( K_a \).

Include \( HCl, HBr, HI, HNO_3, HClO_4, \) and \( H_2SO_4 \) (first \( H^+ \))

Strong bases are soluble hydroxides of groups 1A and 2A.

Include \( LiOH, NaOH, KOH, LiOH, RbOH, CsOH, Ca(OH)_2, Ba(OH)_2, \) and \( Sr(OH)_2 \)

**Calculation of pH of strong acids and bases**

What is the pH of the following solutions?

0.15M HBr, 0.20M KOH, and 0.015M Ba(OH)_2

For strong acids,

\[ [H_3O^+] = [\text{acid}] \] since 100% ionized.

**Strong acid 100%**

Note: \( HA = \text{acid} \)

\[ HA + H_2O \rightarrow H_3O^+ + A^- \]

\[ 0.15M \text{HBr} \Rightarrow [H_3O^+] = 0.15 \text{ M} \]

\[ pH = -\log[H_3O^+] = -\log(0.15) = 0.82 \]
Likewise for strong bases, 100% dissociation.

\[
\text{KOH(s)} \rightarrow \text{K}^+(aq) + \text{OH}^-(aq)
\]

0.20M KOH \Rightarrow [\text{OH}^-] = 0.20 \text{ M}

\[
p\text{OH} = -\log[\text{OH}^-] = -\log(0.20) = 0.70
\]

At 25°C, \( \text{pH} + \text{pOH} = 14.00 \)

At 25°C, \( \text{pH} + \text{pOH} = 14.00 \)
so \( \text{pH} = 14.00 - \text{pOH} = 14.00 - 0.70 = 13.30 \)

What's different about the 0.015M \( \text{Ba(OH)}_2 \) solution?

\[
[\text{OH}^-] = 2 \times 0.015 \text{ M} = 0.030 \text{ M}
\]

\[
p\text{OH} = -\log(0.030) = 1.52
\]

so \( \text{pH} = 14.00 - \text{pOH} = 14.00 - 1.52 = 12.48 \)