CHAPTER 15 ACIDS & BASES

Acid Base Definitions - Arrhenius vs Bronsted-Lowry

Arrhenius – Aqueous solutions only

Acid: Substance contains hydrogen and dissociates to form H_3O^+ in water
Base: Substance contains hydroxide group and dissociates to form OH^- in water

Bronsted-Lowry – Any solvent, even gas

Acid: Proton (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

HCl	+	H_2O	<=>	H_3O^+	+	Cl^-
Acid		Base		Conjugate acid of H_2O		Conjugate base of HCl

NH_3	+	H_2O	<=>	NH_4^+	+	OH^-

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

$\longrightarrow$

HCN	+	CO_3^{2-}	<=>	CN^-	+	HCO_3^-
Acid		Base		Base		Acid
$K_a=6.2x10^{-10}$						$K_a=4.7x10^{-11}$

$\longrightarrow$

HF	+	HCO_3^-	<=>	F^-	+	H_2CO_3
$K_a=6.8x10^{-4}$						$K_a=4.5x10^{-7}$

$\longleftarrow$

$H_2PO_4^-$	+	HNO_2	<=>	H_3PO_4	+	NO_2^-
$K_a=7.1x10^{-4}$				$K_a=7.1x10^{-3}$		

$\longleftarrow$

NH_4^+	+	$C_2H_3O_2^-$	<=>	NH_3	+	$HC_2H_3O_2$
$K_a=5.7x10^{-10}$						$K_a=1.8x10^{-5}$

The stronger the acid, the weaker is the conjugate base and vice-versa.

Dissociation of Water

Water acts as both Bronsted-Lowry acid & base

H_2O + H_2O <=> H_3O^+ + OH^-

Acid Base Acid Base

But water is pure liquid so $[H_2O]^2$ is combined with K_c yielding K_w.

$$K_c [H_2O]^2 = [H_3O^+][OH^-]$$
$$K_w = [H_3O^+][OH^-]$$

At 25°C $K_w = 1.0 \times 10^{-14}$

So in neutral water $[H_3O^+] = [OH^-] = x$ (same)

$$[H_3O^+][OH^-] = K_w$$
$$(x)(x) = 1.0 \times 10^{-14}$$
$$x^2 = 1.0 \times 10^{-14}$$
$$x = [H_3O^+] = [OH^-] = 1.0 \times 10^{-7} \text{ M}$$

Definitions

Neutral => $[H_3O^+] = [OH^-]$

Acid => $[H_3O^+] > [OH^-]$

Base => $[H_3O^+] < [OH^-]$

$pH = -\log[H_3O^+]$ $[H_3O^+] = 10^{-pH}$

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

$pK_w = -\log K_w$ $K_w = 10^{-Kw}$

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

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

Relationships between pH and pOH

$$[H_3O^+][OH^-] = K_w$$
$$\log[H_3O^+] + \log[H_3O^+] = \log K_w$$
$$-\log[H_3O^+] + -\log[H_3O^+] = -\log K_w$$
$$pH + pOH = pK_w$$

At 25°C, $K_w = 1.0 \times 10^{-14}$

so $pK_w = -\log(1.0 \times 10^{-14}) = 14.00$

$$pH + pOH = 14.00$$

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:

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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,
$[H_3O^+] = [acid]$ since 100% ionized.

Strong acid   100%
$HA + H_2O \rightarrow H_3O^+ + A^-$

Note: HA = acid

0.15M HBr => $[H_3O^+] = 0.15$ M

$pH = -\log[H_3O^+] = -\log(0.15) = 0.82$

Likewise for strong bases, 100% dissociation.

$KOH(s) \rightarrow K^+(aq) + OH^-(aq)$

0.20M KOH => $[OH^-]$ = 0.20 M

$pOH = -\log[OH^-] = -\log(0.20) = 0.70$

At 25°C, pH + pOH = 14.00

At 25°C, pH + pOH = 14.00
so pH = 14.00 – pOH = 14.00 – 0.70 = 13.30

What's different about the 0.015M $Ba(OH)_2$ solution?

$[OH^-] = 2 \times 0.015$ M = 0.030 M
pOH = –log(0.030) = 1.52
so pH = 14.00 – pOH = 14.00 – 1.52 = 12.48