16.3 Acid and Base Strength

Acid and Base Hydrolysis Equilibria, Ka, and Kb

Weak Acids:

$$HA_{(aq)} + H_2O_{(l)} \Leftrightarrow H_3O^+ + A^-$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

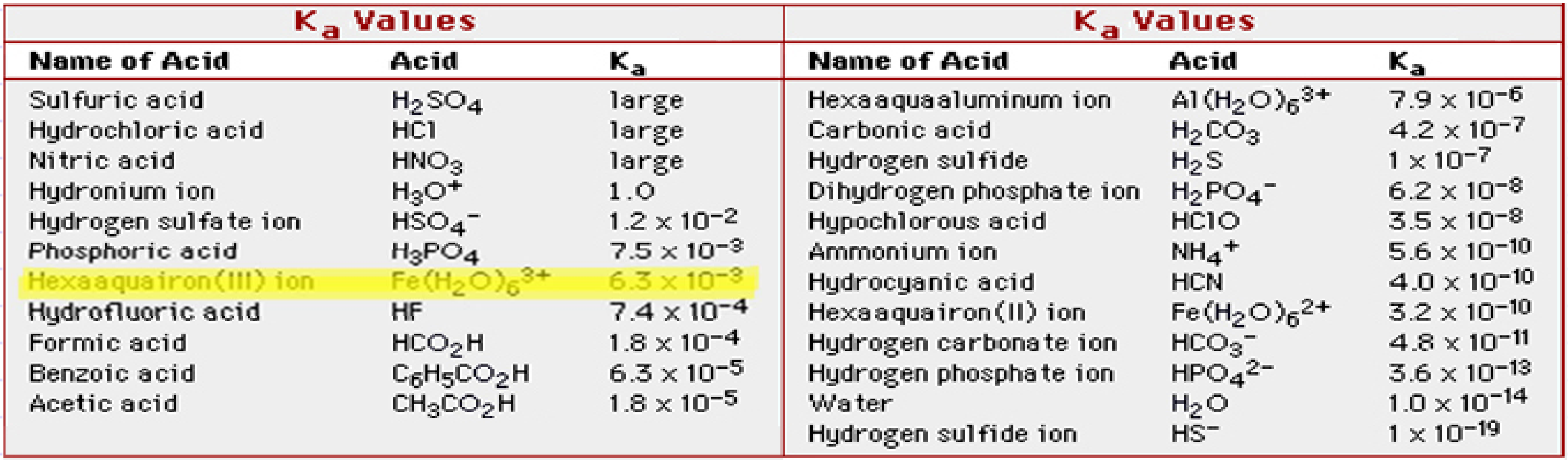
Weak Bases:

$$B_{(aq)} + H_2O_{(l)} \Leftrightarrow BH^+ + OH^-$$

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

16.3 Acid and Base Strength

Acid and Base Hydrolysis Equilibria, Ka, and Kb



K_a Values

Name of Acid	Acid	K_a
Sulfuric acid	H_2SO_4	large
Hydrochloric acid	HCl	large
Nitric acid	HNO_3	large
Hydronium ion	H_3O^+	1.0
Hydrogen sulfate ion	HSO_4^-	1.2×10^{-2}
Phosphoric acid	H_3PO_4	7.5×10^{-3}
Hexaaquairon(III) ion	$Fe(H_2O)_6^{3+}$	6.3×10^{-3}
Hydrofluoric acid	HF	7.4×10^{-4}
Formic acid	HCO_2H	1.8×10^{-4}
Benzoic acid	$C_6H_5CO_2H$	6.3×10^{-5}
Acetic acid	CH_3CO_2H	1.8×10^{-5}

K_a Values

Name of Acid	Acid	K_a
Hexaaquaaluminum ion	$Al(H_2O)_6^{3+}$	7.9×10^{-6}
Carbonic acid	H_2CO_3	4.2×10^{-7}
Hydrogen sulfide	H_2S	1×10^{-7}
Dihydrogen phosphate ion	$H_2PO_4^-$	6.2×10^{-8}
Hypochlorous acid	$HClO$	3.5×10^{-8}
Ammonium ion	NH_4^+	5.6×10^{-10}
Hydrocyanic acid	HCN	4.0×10^{-10}
Hexaaquairon(II) ion	$Fe(H_2O)_6^{2+}$	3.2×10^{-10}
Hydrogen carbonate ion	HCO_3^-	4.8×10^{-11}
Hydrogen phosphate ion	HPO_4^{2-}	3.6×10^{-13}
Water	H_2O	1.0×10^{-14}
Hydrogen sulfide ion	HS^-	1×10^{-19}

The larger the Ka, the stronger the acid.

16.3 Acid and Base Strength

Relationship Between Ka and Kb – Conjugate Acid-Base Pair

Acid : HA $\xrightarrow{-H^+}$ Conjugate Base : A^-

$$\begin{array}{lcll} HA(aq) + H_2O(l) & \Longleftrightarrow & H_3O^+ + A^- & K = K_a \\ * \quad A^- + H_2O(l) & \Longleftrightarrow & HA(aq) + OH^- & K = K_b \\ \hline H_2O(l) + H_2O(l) & \Longleftrightarrow & H_3O^+ + OH^- & K = K_aK_b \end{array}$$

$K \rightarrow K_w$

$K_aK_b = K_w$ for a conjugate acid-base pair.

* An anion acting as a base?
We will address this in more detail shortly.

16.4 Estimating the pH of Acid and Base Solutions

Strong Acid and Strong Base Solutions

What is the **pH** of an aqueous solution of **1.15x10^{-2} M hydrobromic acid**?

	HBr	+	$H_2O(l)$	=	H_3O^+	+	Br^-
I	1.15×10^{-2}				0		0
C	-1.15×10^{-2}				1.15×10^{-2}		1.15×10^{-2}
E	0				1.15×10^{-2}		1.15×10^{-2}

HBr : Strong acid ... 100%

I : Initial concentrations
C : Change in concentrations
E : Equilibrium concentrations

$$[H_3O^+] = 1.15 \times 10^{-2}$$

$$pH = -\log_{10}(1.15 \times 10^{-2})$$

$$= 1.94$$

16.4 Estimating the pH of Acid and Base Solutions

Strong Acid and Strong Base Solutions

What is the **pH** of an aqueous solution of 1.0×10^{-5} M sodium hydroxide?

pH = ?.0

a) 5
b) 6
c) 7
d) 8
e) 9

	NaOH	+	$\cancel{H_2O(l)}$	=	Na^+	+	OH^-
I	1.0×10^{-5}				0		0
C	-1.0×10^{-5}				1.0×10^{-5}		1.0×10^{-5}
E	0				1.0×10^{-5}		1.0×10^{-5}

NaOH : Strong Base → 100%

$$[OH^-] = 1.0\times10^{-5}$$

$$pOH = -\log_{10}(1.0\times10^{-5})$$

$$= 5$$

$$pH + pOH = 14 \text{ @ } 25^\circ C$$

$$pH + 5 = 14$$

$$pH = 9$$

16.4 Estimating the pH of Acid and Base Solutions

pH of a Weak Acid – Quadratic Equation

Calculate the **pH** of a 0.372 M aqueous solution of hypochlorous acid (**HClO, Ka = 3.5x10^{-8}**).

	$HClO$	+	$H_2O(l)$	=	H_3O^+	+	ClO^-
I	0.372				0		0
C	$-x$				x		x
E	$0.372-x$				x		x

$$K_a = \frac{[H_3O^+][ClO^-]}{[HClO]}$$

$$3.5\times10^{-8} = \frac{(x)(x)}{(0.372-x)}$$

$$3.5\times10^{-8}(0.372-x) = x^2$$

$$x^2 + 3.5\times10^{-8}x - 1.302\times10^{-8} = 0$$

$$x = \frac{-b \pm \sqrt{b^2-4ac}}{2a} \quad \left.\right\} \quad \begin{aligned} a &= 1 \\ b &= 3.58\times10^{-8} \\ c &= -1.302\times10^{-8} \end{aligned}$$

$$x = 1.141\times10^{-4}, \quad \underline{-1.141\times10^{-4}}$$

↳ Disregard as this solution makes no chemical sense!

$$x = 1.141\times10^{-4} = [H_3O^+]$$

$$pH = -\log_{10}(1.141\times10^{-4}) = 3.94$$

While this method is the most accurate, solving a quadriatic equation can be problematic on "bad math days!" ☺ ... as in on exam days!!

16.4 Estimating the pH of Acid and Base Solutions

pH of a Weak Acid – Approx Method

Calculate the **pH** of a 0.372 M aqueous solution of hypochlorous acid (**$HClO$, $Ka = 3.5x10^{-8}$**).

	$HClO$	+	$H_2O(l)$	=	H_3O^+	+	ClO^-
I	0.372				0		0
C	$-x$				x		x
E	$0.372 - x$				x		x

If $[HA]_i > 100\,Ka$ then :- $[HA]_i - x \approx [HA]_i$

$0.372 > 100(3.5\times10^{-8})$

then $0.372 - x \approx 0.372$

$$Ka = \frac{[H_3O^+][ClO^-]}{[HClO]}$$

$$3.5\times10^{-8} = \frac{x.x}{0.372}$$

$$x^2 = 0.372(3.5\times10^{-8})$$

$$x = \sqrt{0.372(3.5\times10^{-8})}$$
$$= 1.141\times10^{-4} = [H_3O^+]$$

$$pH = -\log_{10}(1.141\times10^{-4}) = 3.94$$

After you have done some of these you will notice :-
That as long as $[HA]_i > 100\,Ka$

$$\boxed{x = \sqrt{[HA]_i Ka}}$$

16.4 Estimating the pH of Acid and Base Solutions

pH of a Weak Base – Approx Method

Calculate the **pH** of a **0.372 M** aqueous solution of isoquinoline (C_9H_7N, $K_b = 2.5 \times 10^{-9}$)

pH = ?.0

a) 5
b) 6
c) 7
d) 8
[✓ e)] 9

	C_9H_7N	+	$H_2O(l)$	=	$C_9H_7NH^+$	+	OH^-
I	0.372				0		0
C	$-x$				x		x
E	$0.372 - x$				x		x

$0.372 > 100(2.5 \times 10^{-9})$

thus $0.372 - x \approx 0.372$

$$K_b = \frac{[C_9H_7NH^+][OH^-]}{[C_9H_7N]}$$

$$2.5 \times 10^{-9} = \frac{x \cdot x}{0.372}$$

$$x^2 = 0.372(2.5 \times 10^{-9})$$

$$x = \sqrt{0.372(2.5 \times 10^{-9})}$$

$$= 3.05 \times 10^{-5} = [OH^-]$$

$$pOH = -\log_{10}(3.05 \times 10^{-5}) = 4.52$$

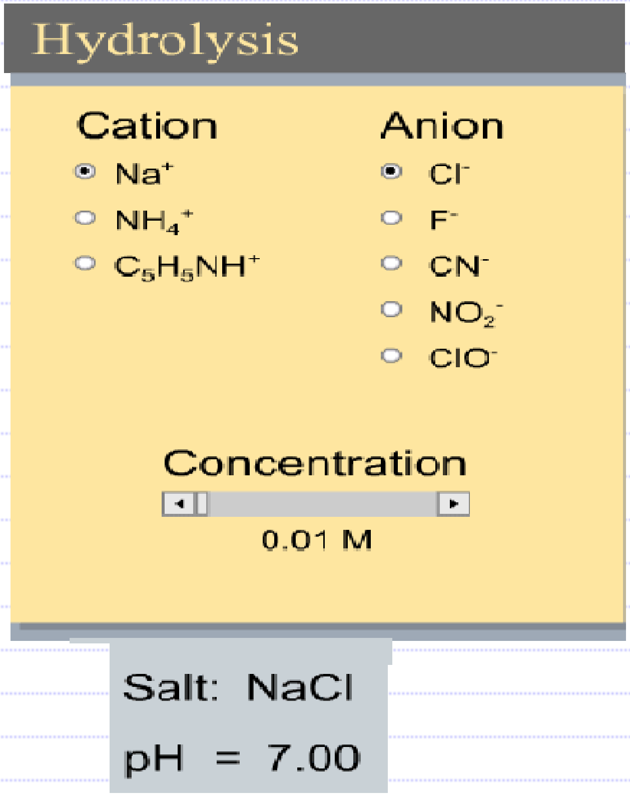
$pH + pOH = 14$ @ 25°C

$$pH = 14 - 4.52$$

$$= 9.48$$

16.5 Acid-Base Properties of Salts

Hydrolysis – Neutral Cations and Anions



Hydrolysis

Cation	Anion
● Na^+	● Cl^-
○ NH_4^+	○ F^-
○ $C_5H_5NH^+$	○ CN^-
	○ NO_2^-
	○ ClO^-

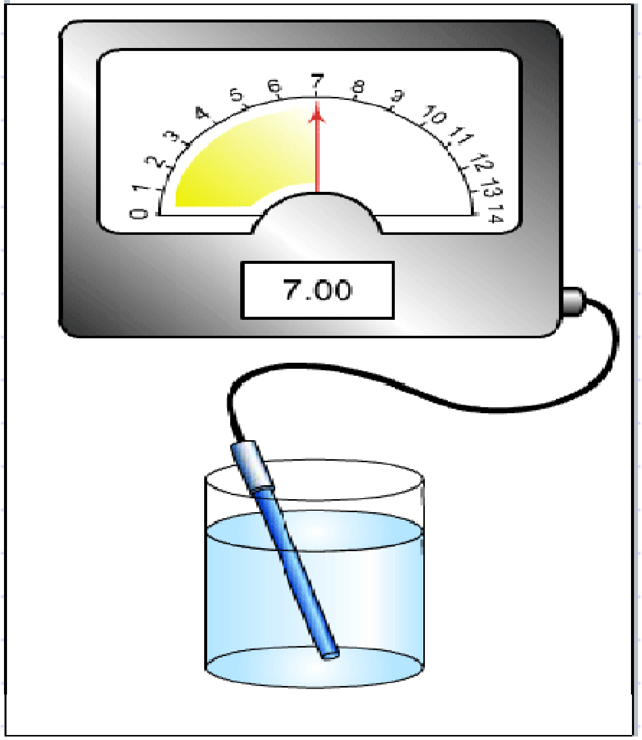
Concentration

0.01 M

Salt: NaCl

pH = 7.00

See Class Web Site.



NEUTRAL CATIONS

The 6 cations associated with the six strong bases:

Li^+	$LiOH$
Na^+	$NaOH$
K^+	KOH
Ca^{2+}	$Ca(OH)_2$
Ba^{2+}	$Ba(OH)_2$
Sr^{2+}	$Sr(OH)_2$

Non neutral cations are all potential weak acids ... ie NH_4^+

NEUTRAL ANIONS

The 6 anions associated with the six strong acids.

Cl^-	HCl
Br^-	HBr
I^-	HI
NO_3^-	HNO_3
SO_4^{2-}	H_2SO_4
ClO_4^-	$HClO_4$

Non neutral anions are all potential weak bases ... ie F^-