

Chapter 9 - Monoprotic Acid-Base Equilibria

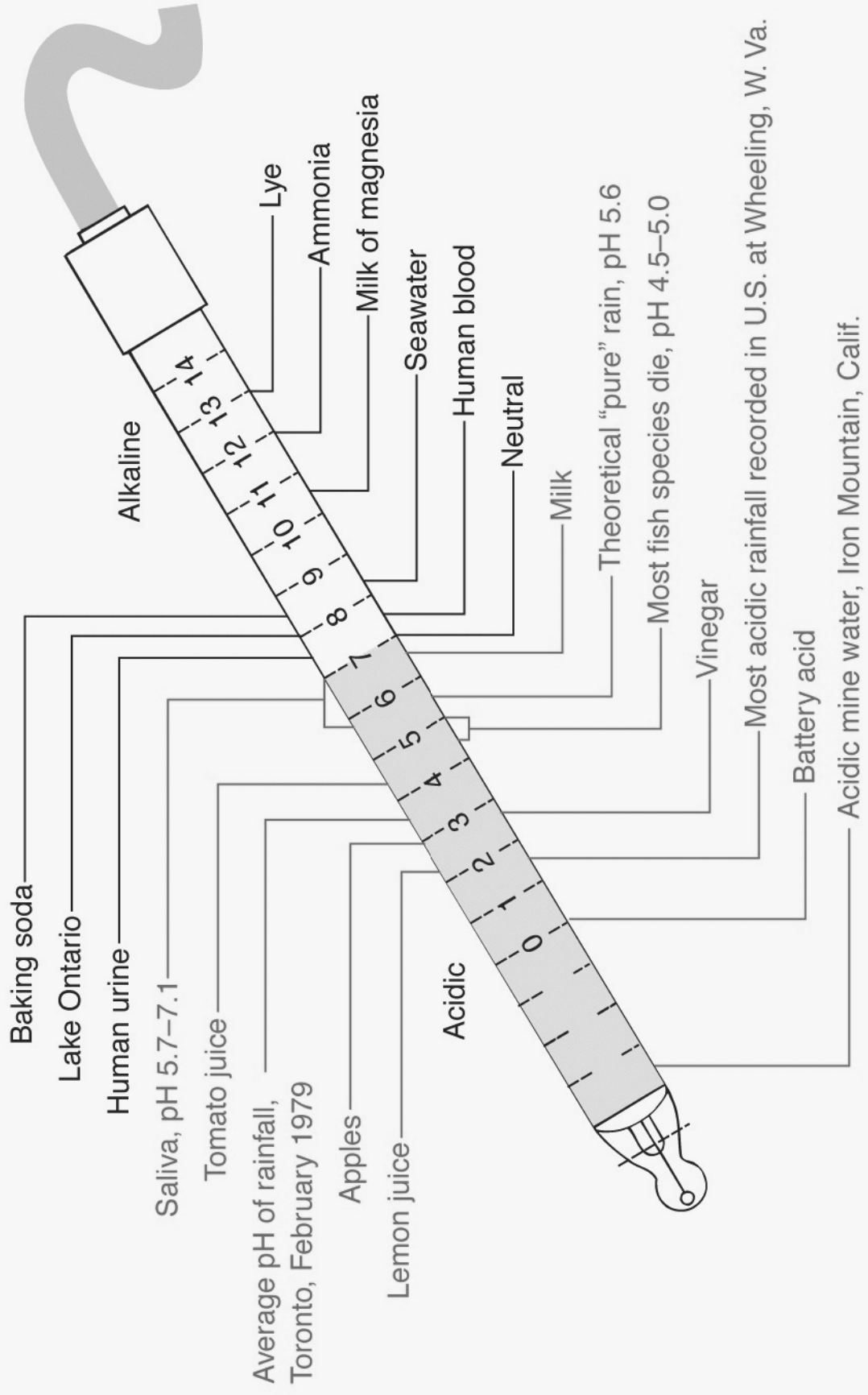
Homework: Due Friday, February 20, 2009

Problems 9-2, 9-3, 9-5, 9-8, 9-11, 9-20, 9-
27, 9-28, 9-29, 9-30, 9-33, 9-34,
9-37

ACID – BASE CHEMISTRY

Protic Acids and Bases (Bronsted-Lowry)

- Acid \Rightarrow increases $[\text{H}_3\text{O}^+]$ or $[\text{H}^+]$
- Base \Rightarrow increases $[\text{OH}^-]$
- Protic \Rightarrow transfer of H^+
- Bronsted acid \Rightarrow proton donor (like HCl)
- Bronsted base \Rightarrow proton acceptor (like NaOH)
- Salt \Rightarrow ionic solid made from reaction of acid and base
(like $\text{CaCl}_2 - \text{R}_x$ of $2\text{HCl} + \text{Ca}(\text{OH})_2 = \text{CaCl}_2 + 2\text{HOH}$)
- Conjugate \Rightarrow acid-base pair
(like acetic acid $\text{H}_3\text{C}-\text{C}(=\text{O})-\text{OH}$ and acetate ion $\text{H}_3\text{C}-\text{C}(=\text{O})-\text{O}^-$)



Equilibrium of Water



- K_w allows you to calculate $[\text{H}^+]$ and $[\text{OH}^-]$
- $[\text{H}^+]$ and $[\text{OH}^-]$ describes the acid/base content of solution
- Since K_w is small, $[\text{H}^+] [\text{OH}^-]$ will also be small.
- pH ($-\log [\text{H}^+]$) or pOH ($-\log [\text{OH}^-]$) is a useful way to report $[\text{H}^+]$ or $[\text{OH}^-]$
- $K_w = [\text{H}^+] [\text{OH}^-] = 1.0 \times 10^{-14}$ So....
$$pK_w = \text{pH} + \text{pOH} = 14$$

Acid/Base Strength

Strong acids and bases **COMPLETELY** dissociates



(No reverse arrow as no HCl present in solution)

TABLE 6-2: Strong acids and Strong bases

Be able to recognize them (memorize)

Examples: Acids – HCl, HNO₃, H₂SO₄

Bases – NaOH, KOH, NH₄OH

Table 6-2 Common strong acids and bases

Formula	Name
ACIDS	
HCl	Hydrochloric acid (hydrogen chloride)
HBr	Hydrogen bromide
HI	Hydrogen iodide
H ₂ SO ₄ ^a	Sulfuric acid
HNO ₃	Nitric acid
HClO ₄	Perchloric acid
BASES	
LiOH	Lithium hydroxide
NaOH	Sodium hydroxide
KOH	Potassium hydroxide
RbOH	Rubidium hydroxide
CsOH	Cesium hydroxide
R ₄ NOH ^b	Quaternary ammonium hydroxide

Determining pH and pOH of strong acid/base

Example:

0.1 M HCl means $[H^+] = 0.1$ or 1×10^{-1}

$$\text{pH} = -\log(1 \times 10^{-1}) = 1$$

So...	$[H^+]$	$[H^+]$	pH
	0.01	1×10^{-2}	2
	0.001	1×10^{-3}	3
	0.0001	1×10^{-4}	4

0.035 M HCl $[H^+] = 3.5 \times 10^{-2}$ M $\text{pH} = -\log(3.5 \times 10^{-2}) = 1.46$

Given pH 5.67, what is the $[H^+]$?

$$[H^+] = \text{antilog}(-5.67) \Rightarrow 10^{-\text{pH}} = 10^{-5.67} = 2.14 \times 10^{-6} \text{ M}$$

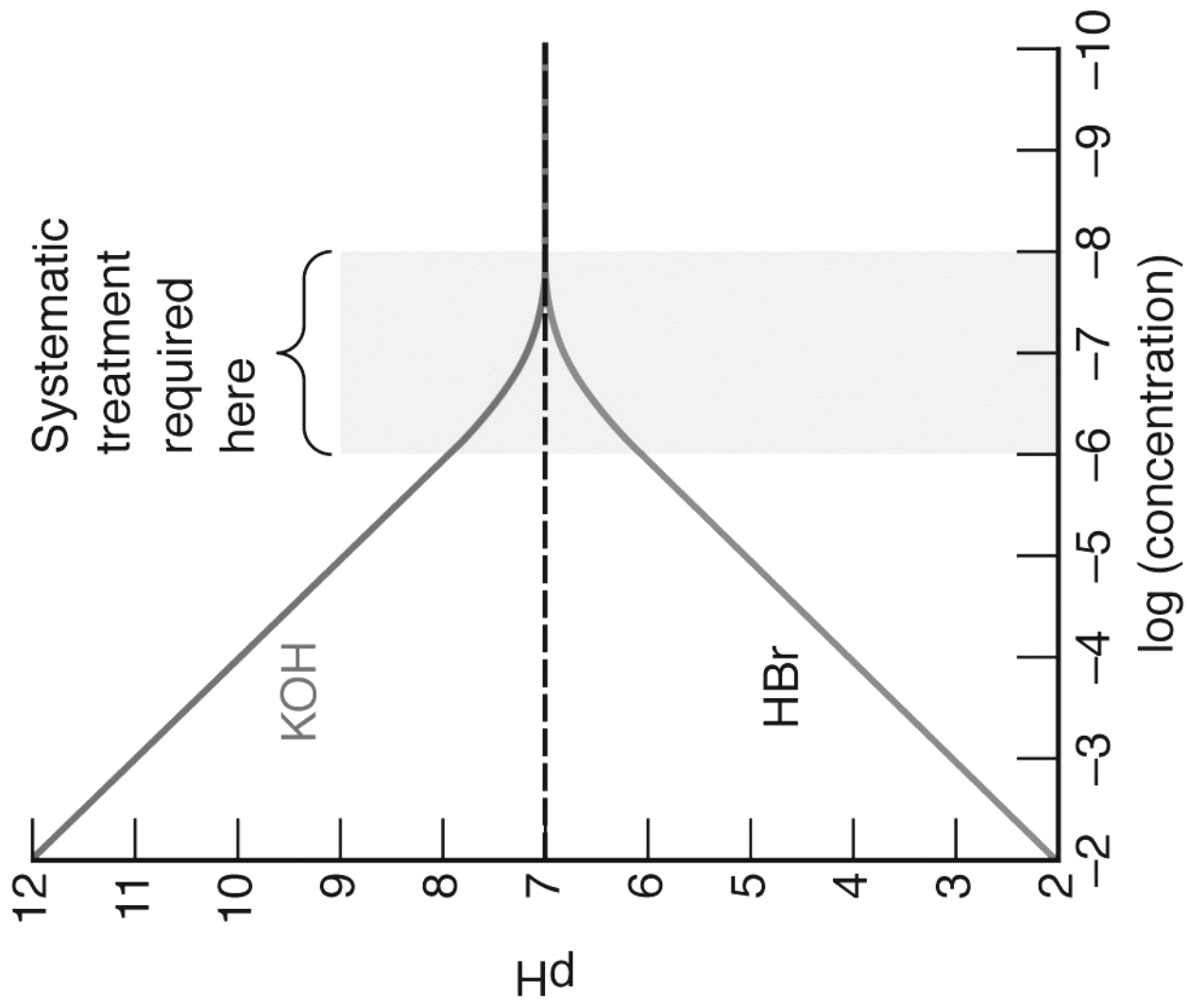
Example: What is the pH of a 0.0045 M NaOH solution?

Example: What is the $[\text{OH}^-]$ of a solution at pH 4.3?

If concentration of strong acid or strong base is very small ($< 10^{-6}$ M), then you cannot simply use the concentration of the strong acid or strong base to determine pH or pOH. The reason is that at this low concentration, the dissociation of H_2O will also contribute to the $[\text{H}^+]$ and $[\text{OH}^-]$. ($\text{H}_2\text{O} \Leftrightarrow \text{H}^+ + \text{OH}^-$)

In the case of extremely small strong acid or strong base concentrations ($< 10^{-8}$), then the main contribution of $[\text{H}^+]$ or $[\text{OH}^-]$ is from the dissociation of H_2O so the solution has a pH 7 and pOH 7.

See pg. 160 and 161 of Harris



Weak acids and bases:

PARTIAL dissociation



acid dissociation constant



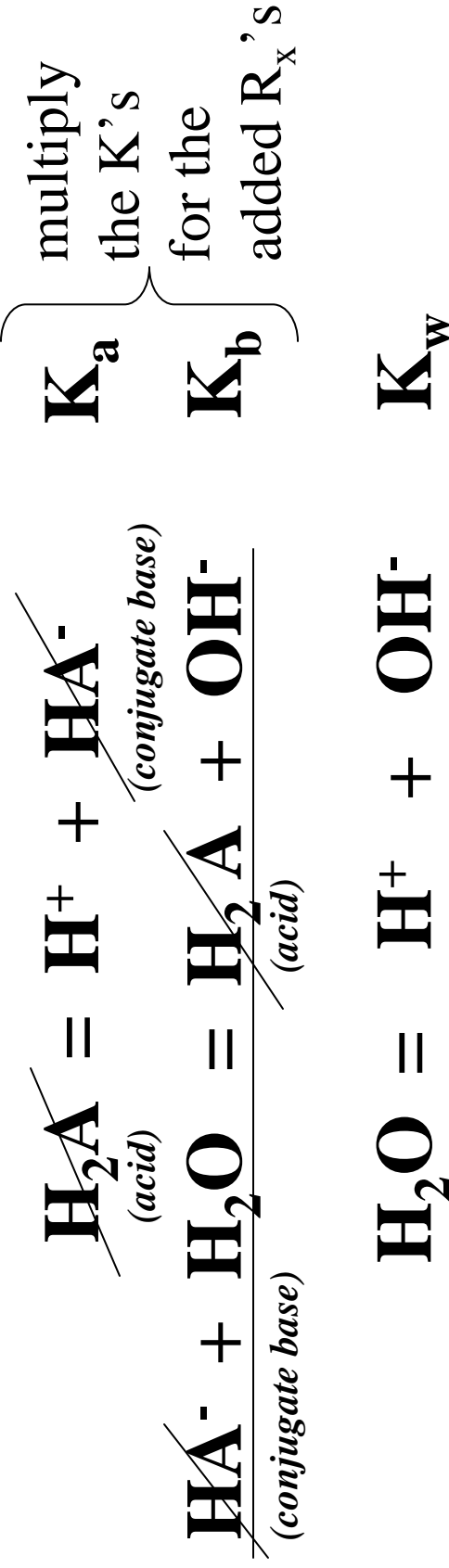
base hydrolysis constant

Relationship of K_a and K_b

The relationship between K_a and K_b for a conjugate pair is:

$$K_a \times K_b = K_w = 1 \times 10^{-14}$$

Also
$$pK_w = pK_a + pK_b$$



Find the pH of 0.050 M sodium butanoate.

$K_a = 1.52 \times 10^{-5}$ butanoic acid – $\text{CH}_3\text{CH}_2\text{CH}_2\text{CO}_2\text{H}$