16.5 Acid-Base Properties of Salts

Acid–Base Properties of Salts

An aqueous solution of **NaClO** is expected to be:

a) Acidic
b) Basic ✓
c) Neutral

$NaClO_{(aq)} = Na^+ + ClO^-$

- Na^+ → Neutral cation
- ClO^- → Non neutral anion … weak base
  - $ClO^- + H_2O_{(l)} \rightleftharpoons HClO_{(aq)} + OH^-$

16.5 Acid-Base Properties of Salts

Acid–Base Properties of Salts

An aqueous solution of **ammonium nitrite** is expected to be:

Ka HNO_2 = 4.5×10^{-4} Kb NH_3 = 1.8×10^{-5}

a) Acidic ✓
b) Basic
c) Neutral

$NH_4NO_2(aq) = NH_4^+ + NO_2^-$

- NH_4^+ → Non neutral cation ... weak acid
- NO_2^- → Non neutral anion ... weak base

$NH_4^+ + H_2O(l) \Leftrightarrow NH_3(aq) + H_3O^+$; Ka = ?

Acid (NH_4^+); Conjugate base (NH_3)

$K_a K_b = 1 \times 10^{-14}$ @ 25°C

$$K_a = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

$NO_2^- + H_2O(l) \Leftrightarrow HNO_2(aq) + OH^-$; Kb = ?

Base (NO_2^-); Conjugate acid (HNO_2)

$K_b K_a = 1 \times 10^{-14}$ @ 25°C

$$K_b = \frac{1 \times 10^{-14}}{4.5 \times 10^{-4}} = 2.2 \times 10^{-11}$$

Ka for NH_4^+ > Kb for NO_2^-

NH_4^+ wins hence solution acidic

16.5 Acid-Base Properties of Salts

Determining pH of a Salt Solution

What is the **pH** of an **0.432M** aqueous solution of **NaCN**
Ka HCN = 4.0×10^{-10}.

$$NaCN_{(aq)} = Na^+ + CN^-$$

	CN^-	+	$H_2O(l)$	=	HCN	+	OH^-
I	0.432				0		0
C	$-x$				x		x
E	$0.432-x$				x		x

CN^- (weak base) $\xrightarrow{+H^+}$ HCN (its conjugate acid)

$$K_a K_b = 1\times10^{-14} \text{ @ } 25°C$$

$$K_b = \frac{1\times10^{-14}}{4.0\times10^{-10}} = 2.5\times10^{-5}$$

$[CN^-] > 100(2.5\times10^{-5})$

Thus $0.432 - x \approx 0.432$

$$K_b = \frac{[HCN][OH^-]}{[CN^-]}$$

$$2.5\times10^{-5} = \frac{x \cdot x}{0.432}$$

$$x^2 = 0.432(2.5\times10^{-5})$$

$$x = \sqrt{0.432(2.5\times10^{-5})}$$

$$= 3.29\times10^{-3} = [OH^-]$$

$$pOH = -\log_{10}(3.29\times10^{-3}) = 2.48$$

$$pH + pOH = 14 \text{ @ } 25°C$$

$$pH = 14 - 2.48$$

$$= 11.52$$

17.1 Acid-Base Reactions

Types of Reactions

	Reaction	Example	pH at Eq	Extent
1.	Strong acid + strong base	$HCl(aq) + NaOH(aq) \longrightarrow H_2O(\ell) + NaCl(aq)$	7	100%
2.	Strong acid + weak base	$HCl(aq) + NH_3(aq) \longrightarrow NH_4Cl(aq)$	< 7	100%
3.	Strong base + weak acid	$NaOH(aq) + HClO(aq) \longrightarrow H_2O(\ell) + NaClO(aq)$	> 7	~100%
4.	Weak acid + weak base	$HClO(aq) + NH_3(aq) \rightleftarrows NH_4ClO(aq)$	Depends on Ka Vs Kb	?

SA : Strong Acid. SB : Strong Base WA : Weak Acid WB : Weak Base

	Reaction	Salt Produced		pH
1	SA + SB	$NaCl(aq) = Na^+ + Cl^-$	Cation & Anion neutral	7
2.	SA + WB	$NH_4Cl(aq) = NH_4^+ + Cl^-$	NH_4^+ is a weak acidic cation. $NH_4^+ + H_2O(l) \Leftrightarrow NH_3(aq) + H_3O^+$	< 7
3.	WA + SB	$NaClO(aq) = Na^+ + ClO^-$	ClO^- is a weak basic anion. $ClO^- + H_2O(l) \Leftrightarrow HClO(aq) + OH^-$	> 7
4.	WA + WB	$NH_4ClO(aq) = NH_4^+ + ClO^-$	NH_4^+ is a weak acidic cation. $NH_4^+ + H_2O(l) \Leftrightarrow NH_3(aq) + H_3O^+$ ClO^- is a weak basic anion $ClO^- + H_2O(l) \Leftrightarrow HClO(aq) + OH^-$	?*

* pH depends on which K is larger, Ka or Kb

17.1 Acid-Base Reactions

Strong Acid/Strong Base Reactions

$HCl(aq) + NaOH(aq) \overset{?}{=} NaCl(aq) + H_2O(l)$

Net Ionic Equation:

$HCl(aq) + NaOH(aq) = NaCl(aq) + H_2O(l)$

$H_3O^+ + Cl^- + Na^+ + OH^- = Na^+ + Cl^- + H_2O(l) + H_2O(l)$ *

$H_3O^+ + OH^- = 2\,H_2O(l)$

* As we wrote H_3O^+ on reactant side, had to add $H_2O(l)$ to product side so as to keep the chemical equation balanced.

$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+ + OH^-$; $K = K_w$

$H_3O^+ + OH^- = 2\,H_2O(l)$; $K = \frac{1}{K_w}$

$= \frac{1}{1 \times 10^{-14}}$

$= 1 \times 10^{14}$

$K \ggg 1$: Very product favored, goes 100%.
Thus = rather than $\rightleftharpoons$

17.1 Acid-Base Reactions

Strong Acid/Weak Base Reactions

$HCl(aq) + NH_3(aq) \overset{?}{=} NH_4Cl(aq)$

NET IONIC EQUATION:

$HCl(aq) + NH_3(aq) = NH_4Cl(aq)$

$H_3O^+ + Cl^- + NH_3(aq) = NH_4^+ + Cl^- + H_2O(l)$*

$H_3O^+ + NH_3(aq) = NH_4^+ + H_2O(l)$

* See previous slide as to why we had to add $H_2O(l)$ to the product side.

$NH_4^+ + H_2O(l) \Leftrightarrow NH_3(aq) + H_3O^+$: $K_a = 5.6 \times 10^{-10}$

$H_3O^+ + NH_3(aq) \Leftrightarrow NH_4^+ + H_2O(l)$: $K = \frac{1}{5.6 \times 10^{-10}} = 1.8 \times 10^9$

$K \ggg 1$, very product favored, essentially 100%

17.1 Acid-Base Reactions
Weak Acid/Weak Base

$HClO(aq) + NH_3(aq) \overset{?}{\Leftrightarrow} NH_4ClO(aq)$

NET IONIC EQUATION:

$HClO(aq) + NH_3(aq) \Leftrightarrow NH_4^+ + ClO^-$

$HClO(aq) + H_2O(l) \Leftrightarrow H_3O^+ + ClO^-$ $K = 3.5 \times 10^{-8}$

$NH_3(aq) + H_2O(l) \Leftrightarrow NH_4^+ + OH^-$ $K = 1.8 \times 10^{-5}$

$HClO(aq) + NH_3(aq) + 2H_2O(l) \Leftrightarrow NH_4^+ + ClO^- + H_3O^+ + OH^-$ $K = 6.3 \times 10^{-13}$

$H_3O^+ + OH^- \Leftrightarrow 2\,H_2O(l)$ $K^* = 1.0 \times 10^{14}$

$HClO(aq) + NH_3(aq) \Leftrightarrow NH_4^+ + ClO^-$ $K = 63$

$K^* = 1/K_w$

$K \approx 1$: Significant quantities of reactants and products present at equilibrium.

Curious about the pH at the equivalence point?

Hydrolysis of NH_4ClO

$NH_4ClO(aq) = NH_4^+ + ClO^-$

a) NH_4^+ is an acidic cation.

b) ClO^- is a basic anion.

a) $NH_4^+ + H_2O(l) \Leftrightarrow NH_3(aq) + H_3O^+$: $K_a = 5.6 \times 10^{-10}$

b) $ClO^- + H_2O(l) \Leftrightarrow HClO(aq) + OH^-$: $K_b = 2.9 \times 10^{-7}$

$K_b > K_a$

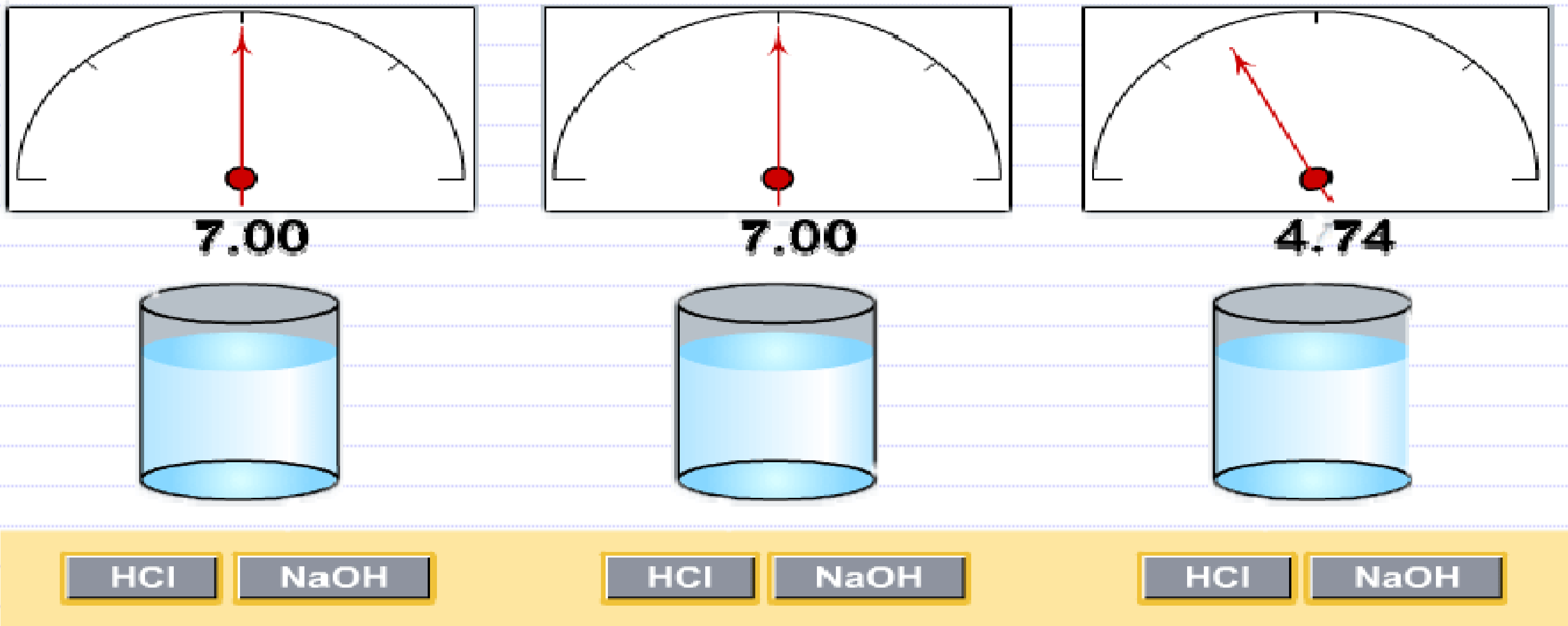
$pH > 7$.

17.2 Buffers

What Constitutes a Buffer and why are they Special

Buffer Solutions

See Class Web Site



100 mL H_2O	100 mL 1.0 M NaCl	100 mL 1.0 M CH_3CO_2H(aq) 1.0 M $NaCH_3CO_2$(aq)

Weak acid (-H^+)
Its conjugate base

17.2 Buffers

How do they Resist Drastic pH Change – A Summary

1. BUFFER : Buffer Acid + Buffer Base

   Buffer Acid: A Weak Acid

   Buffer Base: Its Conjugate Base

   (Weak Acid $-H^+$ → Its Conjugate Base)

2. How does a buffer resist drastic pH changes when it comes in contact : H_3O^+ or OH^-

   a) H_3O^+ : H_3O^+ + Buffer Base = $H_2O(l)$ + Buffer Acid.

   b) OH^- : OH^- + Buffer Acid = $H_2O(l)$ + Buffer Base.