Chemistry 12

Acid-Base Equilibrium V

IV 13
pg. 144-148

Name:
Date:
Block:

1. Hydrolysis

↳ Water → Salt in water

Hydrolysis

In previous Chemistry courses, you have learned about neutralization reactions where:

Acid + Base → Salt + Water

→ slightly acidic / basic

HY + XOH → XY + H_2O

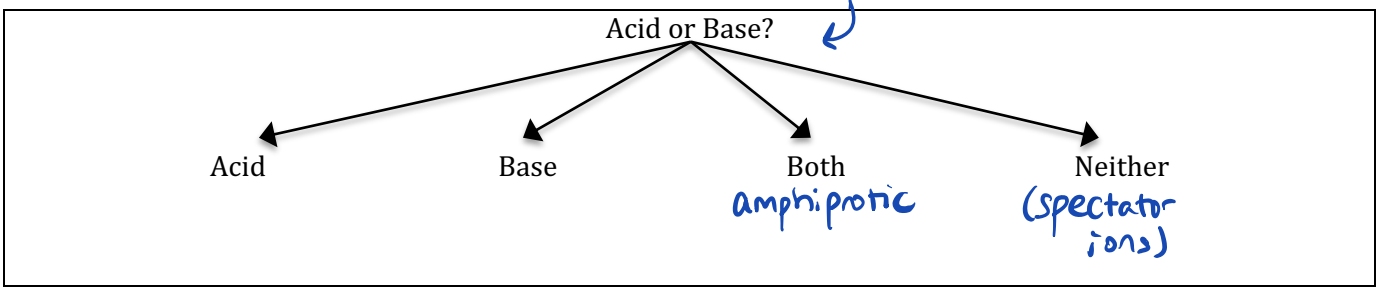
HCl + NaOH → NaCl + H_2O

The "salt" produced in neutralization reactions are actually acidic or basic. The ions that make up the salt behave as weak acids or bases.

Ionization / Dissociation Equation

XY (salt) → X^+ + Y^-

Will it be



Consider the following...

$HCl + NaOH \rightarrow NaCl + H_2O$

Spectator ions are conjugates of SA or SB
- see from data table
  - * watch out for HSO_4^-
- metals (bond w/ -OH)

$NaCl \rightarrow Na^+ + Cl^-$

Na^+: spectator ion
Cl^-: spectator ion

salt is neutral
pH = 7
- neither ion will hydrolyze (react w/ water)

$HNO_2 + NaOH \rightarrow NaNO_2 + H_2O$

$NaNO_2 \rightarrow Na^+ + NO_2^-$

Na^+: spectator ion
NO_2^-: WB — will hydrolyze

← Ionization equation

$NO_2^- + H_2O \rightleftharpoons HNO_2 + OH^-$

↑ Hydrolysis reaction

$C_6H_5OH + NH_4OH \rightarrow NH_4C_6H_5O + H_2O$

$NH_4C_6H_5O \rightarrow NH_4^+ + C_6H_5O^-$

NH_4^+: WA, $K_a = 5.6 \times 10^{-10}$ (crossed out)

$C_6H_5O^-$: WB

$K_b = \frac{K_w}{K_a(C_6H_5OH)} = \frac{1.0 \times 10^{-14}}{1.3 \times 10^{-10}} = 7.7 \times 10^{-5}$

$C_6H_5O^- + H_2O \rightleftharpoons C_6H_5OH + OH^-$

$C_6H_5COOH + Mg(OH)_2 \rightarrow Mg(C_6H_5COO)_2 + H_2O$

$Mg(C_6H_5COO)_2 \rightarrow Mg^{2+} + 2\ C_6H_5COO^-$

Mg^{2+}: spectator ion

$2\ C_6H_5COO^-$: WB

$C_6H_5COO^- + H_2O \rightleftharpoons C_6H_5COOH + OH^-$

Circle the following salts whose ions **will** hydrolyze (react with water!) when dissociated in water.

NH_4Cl Na_2CO_3 $RbClO_4$ Li_2SO_3 BaI_2 NH_4HCOO KIO_3 CsF $CaBr_2$

NH_4Cl: WA → NH_4Cl will hydrolyze as a weak acid

Na_2CO_3: WB

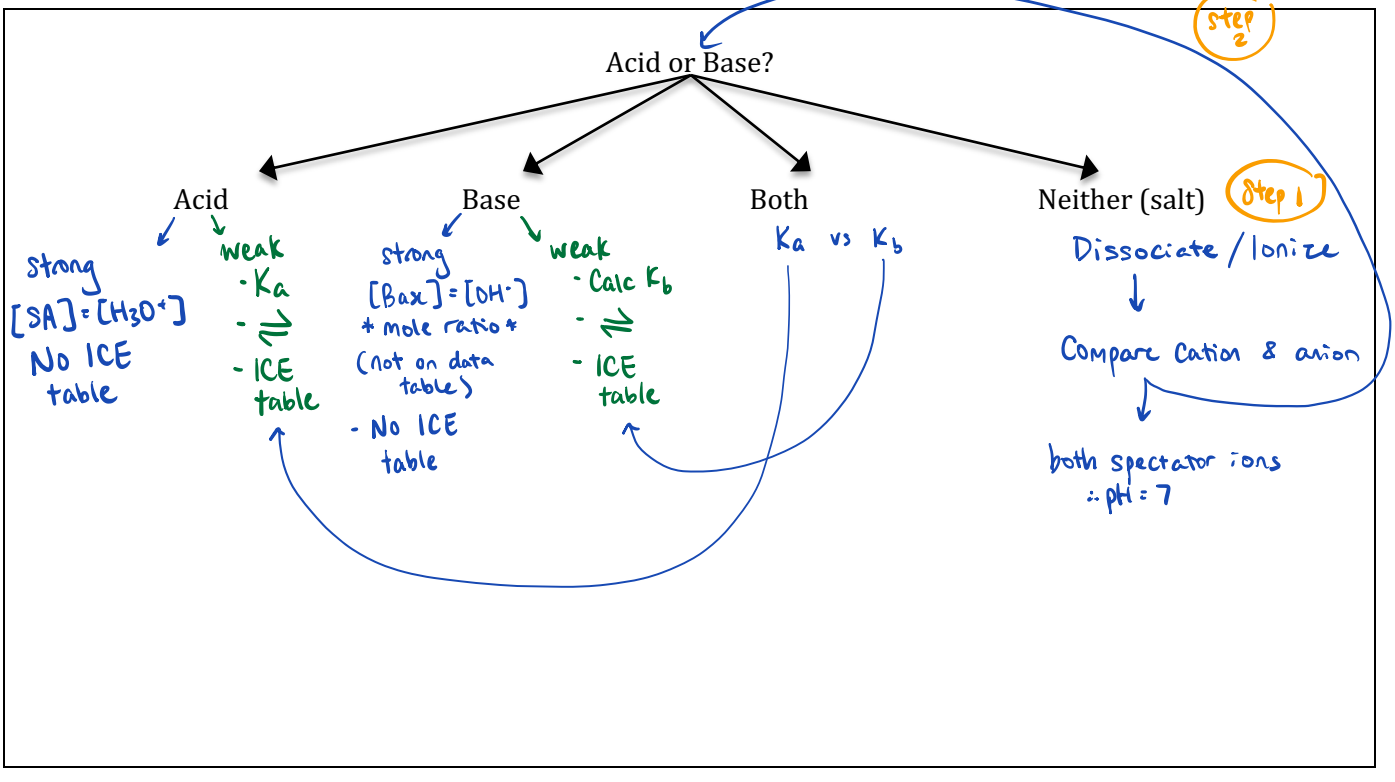
$RbClO_4$ → $RbClO_4$ will not hydrolyze

Li_2SO_3: WB

NH_4HCOO: WA, WB, $K_a > K_b$

KIO_3: WB

CsF: WB



H/W:

Decide if each of the following salts will produce an acidic, basic or neutral solution when combined with water.

		K_a		K_b	pH (A, B, or N)
1.	Na_3PO_4			4.5×10^{-2}	B
2.	KH_2PO_4 ↳amphiprotic	6.2×10^{-8}	>	1.3×10^{-12}	A
3.	Na_2CO_3			1.8×10^{-4}	B
4.	$KHSO_4$	1.2×10^{-2}			A
5.	$CaCO_3$			1.8×10^{-4}	B
6.	$NaNO_3$				N
7.	$(NH_4)_2C_2O_4$ WA ↑ ↑WB	5.6×10^{-10}	>	1.6×10^{-10}	A
8.	NH_4Cl	5.6×10^{-10}			A
9.	Na_2SO_3			1.0×10^{-7}	B
10.	$FeCl_3$				N
11.	KCH_3COO			5.6×10^{-10}	B

Order the above substances from most acidic to most basic.

Most Acidic ↑ largest K_a

4 2 8=7 6=10 11 9 5=3 1

Most Basic ↑ largest K_b

Example:

A 9.54g sample of $Mg(HCO_3)_2$ is dissolved in enough water to make 500.0 mL of solution. Calculate the pH of this solution.

- What is the concentration of $Mg(HCO_3)_2$?

$$\frac{9.54\,g}{0.5000\,L} \times \frac{1\,mol}{146.3\,g} = 0.130\,M = [Mg(HCO_3)_2]$$

- What is initial concentration of each ion? (*Hint – dissociation equation required)

$$Mg(HCO_3)_2 \rightarrow Mg^{2+} + 2\,HCO_3^-$$

$0.130\,M$ → $0.130\,M$ (spectator) + $0.260\,M$

- Which ion produced will hydrolyze?

HCO_3^- amphiprotic!

$K_a = 5.6\times10^{-11}$

$$K_b = \frac{K_w}{K_a(H_2CO_3)} = \frac{1.0\times10^{-14}}{4.3\times10^{-7}} = 2.3\times10^{-8}$$

∴ weak base

- What is the equation when it reacts with water? Make an ICE table.

$$HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$$

	HCO_3^-		H_2CO_3	OH^-
I	$0.260\,M$		$0\,M$	$0\,M$
C	$-x$		$+x$	$+x$
E	$0.260-x$		x	x

* assume $0.260 - x \cong 0.260$

$$K_b = 2.3\times10^{-8} = \frac{x^2}{0.260-x}$$

$$x = \sqrt{(0.260)(2.3\times10^{-8})}$$

$$= 7.7\times10^{-5}\,M = [OH^-]$$

- Calculate pOH and pH.

$$pOH = -\log[OH^-]$$

$$= -\log(7.7\times10^{-5})$$

$$= 4.11$$

$$pH = 14 - 4.11$$

$$= \boxed{9.89}$$

$C_1V_1 = C_2V_2$

A 200.0 mL aqueous solution of 0.50 M Na_2CO_3 is diluted to 500.0 mL. Calculate the pH of the resulting solution.

V_1 C_1 V_2

$$[Na_2CO_3] = \frac{(0.50)(200.0)}{(500.0)} = 0.20M$$

$Na_2CO_3 \rightarrow 2\,Na^+ + CO_3^{2-}$

0.20M → 0.40M (spectator) + 0.20M (WB)

$$K_b = \frac{K_w}{5.6\times10^{-11}} = 1.8\times10^{-4}$$

$$K_b = 1.8\times10^{-4} = \frac{x^2}{0.20-x}$$

	CO_3^{2-}	+ H_2O ⇌	HCO_3^-	+ OH^-
I	0.20M		0M	0M
C	−x		+x	+x
E	0.20 − x		x	x

assume $0.20 - x \cong 0.20$

$$x = \sqrt{(0.20)(1.8\times10^{-4})}$$

$$= 6.0\times10^{-3}M = [OH^-]$$

$$pOH = -\log(6.0\times10^{-3})$$

$$= 2.22$$

$$\boxed{pH = 11.78}$$

The K_b for pyridine, C_5H_5N, a weak base, is 4.7×10^{-9}. Calculate the pH of a 0.10 M solution of $C_5H_5NHNO_3$.

Salt

$C_5H_5NHNO_3 \rightarrow C_5H_5NH^+ + NO_3^-$ (spectator)

$C_5H_5NH^+$: weak acid (we know b/c we're given the K_b of a conjugate weak base)

	$C_5H_5NH^+$	+ H_2O ⇌	C_5H_5N	+ H_3O^+
I	0.10M		0M	0M
C	−x		+x	+x
E	0.10 − x		x	x

$$K_a = \frac{K_w}{K_b\,(\text{conjugate base})} = \frac{1.0\times10^{-14}}{4.7\times10^{-9}} = 2.1\times10^{-6} = \frac{x^2}{0.10-x}$$

assume $0.10 - x \cong 0.10$

$$x = \sqrt{(0.10)(2.1\times10^{-6})}$$

$$= 4.6\times10^{-4} = [H_3O^+]$$

$$pH = -\log(4.6\times10^{-4})$$

$$= \boxed{3.34}$$

Complete: Hebden Pg. 148 #69, 70, 73