AcidBaseKey

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Acids and Bases

For each of the following salts, predict if a 100 mM solution would be acidic, basic, or neutral.

NaCl (neutral) KNO$_2$ (basic) sodium acetate (basic) sodium hydrogen sulfate (acidic)

Strategy: Break the salt into ions. Is the cation an acid or base? How about the anion? NaCl → Na$^+$ + Cl$^-$. Na$^+$ is not a proton donor or acceptor. Cl$^-$ is the conjugate base of a strong acid, so it is NOT a base. Neutral.

KNO$_2$ → K$^+$ + NO$_2^-$ (base)

NaC$_2$H$_3$O$_2$ → Na$^+$ + C$_2$H$_3$O$_2^-$ (base)

NaHSO$_4$ → Na$^+$ + HSO$_4^-$ (acid)

For each pair, identify which will be a more acidic solution:

10 mM HCl or 10 mM HF (strong acid)

10 mM H$_2$SO$_4$ or 10 mM HCl (diprotic)

10 mM HNO$_3$ or 20 μM HNO$_2$ (higher concentration)

Strategy: You need to consider all variables that can influence the amount of H$_3$O$^+$ that is produced: concentration, acid strength, monoprotic vs. diprotic. HCl vs. HF. Equal concentration of monoprotic acids. HCl is a strong acid and HF is a weak acid. HCl will be more acidic.

Order the following solutions by increasing acidity (lowest pH goes last). A table of pKa is attached.

pKa: 100 mM HF, 3.2; 100 mM HClO, 7.4; 100 mM HSO$_4^{-1}$, 1.99; 100 mM NH$_4^+$, 9.25

lower pKa = stronger acid

NH$_4^+$, HClO, HF, HSO$_4^-$

HCO$_3^{-1}$ can be an acid or a base. If you have a 100 mM NaHCO$_3$ solution, will it be acidic or basic?

Hint: Compare the Ka and Kb values. Is this molecule a stronger acid or base?

$HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$ $K_b = 2.2\times10^{-8}$

$HCO_3^- + H_2O \rightleftharpoons CO_3^{-2} + H_3O^+$ $K_a = 4.7\times10^{-11}$

$K_b > K_a$ so more OH$^-$ produced

Hypochlorous acid (HClO) has a pKa of 7.53. What is the Ka? What is the conjugate base? What are the pKb and Kb of the conjugate base?

$K_a = 10^{-7.53} = 2.95\times10^{-8}$ ClO$^-$ $14 - 7.53 = 6.47$

$K_b = 10^{-6.47} = 10^{-6.47} = 3.4\times10^{-7}$

Calculate the pH of each of the following solutions:

650 μM HCl ← strong acid!

6.5×10^{-4} M = [H$_3$O$^+$] $pH = -\log 6.5\times10^{-4}$

$pH = 3.18$

650 μM HF (weak acid)

$HF + H_2O \rightleftharpoons H_3O^+ + F^-$ $K_a = 6.3\times10^{-4} = \frac{x^2}{6.5\times10^{-4} - x}$

I 6.5×10^{-4} 0 0

$HF + H_2O \rightleftharpoons H_3O^+ + F^-$ $K_a = 6.3\times10^{-4} = \frac{x^2}{6.5\times10^{-4} - x}$

	HF	H_2O	H_3O^+	F^-
I	6.5×10^{-4}		0	0
C	$-x$		$+x$	$+x$
E	$6.5\times10^{-4} - x$		x	x

$0 = x^2 + 6.3\times10^{-4}x - 4.095\times10^{-7}$

$x = 3.98\times10^{-4}$ or -0.001

$[H_3O^+]$

$pH = -\log 3.98\times10^{-4}$

$pH = 3.4$

6.5 nM HF

same approach ... $x = -6.3\times10^{-4}$ or 6.499×10^{-9}

remember when $[H_3O^+] < 10^{-7}$ when an acid is added, we must account for $[H_3O^+]$ from H_2O (10^{-7})

$[H_3O^+] + [H_3O^+]_{H_2O} = [H_3O^+]_{TOTAL}$

$6.5\times10^{-9} + 10^{-7} = 1.065\times10^{-7}$

$pH = 6.97$

1 mM $Mg(OH)_2$

1mM Mg^{2+}

2mM $OH^- = 2\times10^{-3}$ M

$pOH = -\log 2\times10^{-3} = 2.699$

$pH = 14 - 2.699 = 11.3$

650 mM MgF_2

1.3 M F^-

$F^- + H_2O \rightleftharpoons HF + OH^-$ $K_b = 1.4\times10^{-11} = \frac{x^2}{1.3 - x}$

	F^-	H_2O	HF	OH^-
I	1.3		0	0
C	$-x$		$+x$	$+x$
E	$1.3 - x$		x	x

$0 = x^2 + 1.4\times10^{-11}x - 2.08\times10^{-11}$

$x = 4.56\times10^{-6}$ or -4.56×10^{-6}

$[OH^-] = 4.56\times10^{-6}$ M $pOH = 5.34$

$pH = 8.66$

For each of the following acids, determine what concentration is needed to have a pH of 5.5. Please answer in **micromolar**.

Hydrochloric acid

strong acid

$pH = 5.5 = -\log 5.5$

$[H_3O^+] = 10^{-5.5}$

$C_{HCl} = 3.16\times10^{-6}$ M $= 3.16$ mM

Ammonium chloride

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$

	NH_4^+	H_2O	NH_3	H_3O^+
I	I		0	0
C	$-x$		$+x$	$+x$
E	$I-x$		x	x

$[H_3O^+] = x = 10^{-5.5}\ M$

$K_a = 5.6\times10^{-10} = \dfrac{(10^{-5.5})^2}{I - 10^{-5.5}}$

$5.6\times10^{-10} I - 1.77\times10^{-5} = 10^{-11}$

$5.6\times10^{-10} I = 1.000177\times10^{-11}$

$I = 0.01786\ M$

$[NH_4] = 1.786\times10^{4}\ mM$