Acids and Bases.

1. The pKa of HNO_2 is 3.25. Calculate pKb, Ka, and Kb.

2. Write a reaction showing the neutralization of HF with KOH.

3. A flask contains 14.86 μM HNO_3.
    a. Calculate the pH, pOH, $[H_3O^+]$, and $[OH^-]$
    b. Do you expect the pH of 14.86 μM HNO_2 to be more acidic?

4. A flask contains 14.86 nM HNO_3. Calculate the pH.

5. Calculate the pH of a 148.6 μM solution of HNO_2.

6. Calculate the pH of a 148.6 μM solution of KNO_2.

7. What is the pH of 1 mM H_2SO_4? Note that the pKa of HSO_4^- is 1.99.

8. What concentration of HNO_2 has the same pH as 100 mM HBr?

9. For a 100 mL solution of 100 mM KNO_2:
    a. What volume of 1.5 M HCl is needed to completely neutralize this solution?

(1) $pK_a = 3.25$ $\quad pK_a = -\log K_a$ $\quad K_a = 10^{-3.25} = 5.62 \times 10^{-4}$

$pK_a + pK_b = 14$ $\quad pK_b = 14 - pK_a = 10.75$

$pK_b = -\log K_b$ $\quad K_b = 10^{-10.75} = 1.78 \times 10^{-11}$

(2) $HF(aq) + KOH(aq) \longrightarrow H_2O^{+} + KF(aq)$

Neutralization = producing H_2O from the reaction between an acid and a base

- boils down to a double displacement reaction

(3) 14.86 μM HNO_3 — this is a strong acid!

a. $HNO_3(aq) + H_2O(l) \longrightarrow H_3O^{+} + NO_3^{-}$

Full dissociation — meaning that the reverse reaction does NOT happen.

$$\frac{14.86\,\mu M\ HNO_3 \,|\, 1\,\mu M\ H_3O^{+} \,|\, 10^{-6}\,M}{\quad |\, 1\,\mu M\ HNO_3 \,|\, 1\,\mu M} = 1.486 \times 10^{-5}\,\mu M = [H_3O^{+}]$$

$pH = -\log 1.486 \times 10^{-5} = 4.83$

b. HNO_2 is a weak acid, so it does not dissociate completely because

NO the reaction is governed by an equilibrium constant.

- so, NOT all of the HNO_2 becomes H_3O^{+}, so the pH will be higher (less acidic)

4. $$\frac{14.86\,nM\ HNO_3 \,|\, 1\ H_3O^{+} \,|\, 10^{-9}\,M}{\quad |\, 1\ HNO_3 \,|\, 1\,nM} = 1.486 \times 10^{-8}\,M = [H_3O^{+}]$$

strong acid

- since $[H_3O^{+}] < [H_3O^{+}]$ from H_2O ($10^{-7}M$), we need to take $[H_3O^{+}]_{water}$ into account.

$[H_3O^{+}]_{total} = [H_3O^{+}]_{HNO_3} + [H_3O^{+}]_{water} = 1.486 \times 10^{-8} + 10^{-7}$

$[H_3O^{+}]_{total} = 1.1486 \times 10^{-7}\,M$ $\quad pH = -\log 1.1486 \times 10^{-7} = \boxed{6.94}$

5. 148.6 μM HNO_2 (pKa = 3.25)

$$\frac{148.6\ \mu M}{} \bigg| \frac{10^{-6}\ M}{1\ \mu M} = 1.486\times10^{-4}\ M$$

$$HNO_2 + H_2O \rightleftharpoons NO_2^- + H_3O^+$$

	HNO_2	NO_2^-	H_3O^+
I	1.486×10^{-4}	$\varnothing$	$\varnothing$
C	$-x$	$+x$	$+x$
E	$1.486\times10^{-4}-x$	x	x

$$5.62\times10^{-4} = \frac{x^2}{1.486\times10^{-4}-x} \qquad x^2 + 5.62\times10^{-4}x - 8.35\times10^{-8}$$

$$x = \frac{-5.62\times10^{-4} + \sqrt{(5.62\times10^{-4})^2 - 4(1)(-8.35\times10^{-8})}}{2} = 1.22\times10^{-4}$$

$$x = [H_3O^+] = 1.22\times10^{-4}$$

$$pH = -\log 1.22\times10^{-4} = 3.91$$

6. 148.6 μM KNO_2 ← this ionic compound breaks apart into K^+ + NO_2^- ions.

NO_2^- is a weak base! $K_b = 1.78\times10^{-11}$ (calculated in #1)

REALLY small K, So you can use the shortcut

$$NO_2^- + H_2O(l) \rightleftharpoons HNO_2(aq) + OH^-$$

	NO_2^-	HNO_2	OH^-
I	1.486×10^{-4}	$\varnothing$	$\varnothing$
C	$-x$	$+x$	$+x$
E	$1.486\times10^{-4}-x$	x	x

$x = [OH^-]$

$$1.78\times10^{-11} = \frac{x^2}{1.486\times10^{-4}-x} = \frac{x^2}{1.486\times10^{-4}}$$

$$x^2 = 2.645\times10^{-15}$$

$$x = [OH^-] = 5.14\times10^{-8}\ M$$

- Really small $[OH^-]$, so add $[OH^-]_{water}$

$$[OH^-]_{total} = 10^{-7} + 5.14\times10^{-8} = 1.514\times10^{-7}\ M$$

$$pOH = 6.82$$

$$pH = 14 - 6.82 = 7.18$$

7. H_2SO_4 is a strong acid

$$H_2SO_4 + H_2O \longrightarrow \underbrace{HSO_4^- + H_3O^+}_{\text{1 mM of each produced}}$$

HSO_4^- is a weak acid!

$$HSO_4^- + H_2O(l) \rightleftharpoons SO_4^{2-} + H_3O^+$$

$pKa = 1.99$

$K_a = 10^{-1.99}$

	HSO_4^-	SO_4^{2-}	H_3O^+
I	1×10^{-3}	$\varnothing$	$\varnothing$
C	$-x$	$+x$	$+x$
E	$1\times10^{-3}-x$	x	x

$$1.023\times10^{-2} = \frac{x^2}{1\times10^{-4}-x}$$

$$x^2 + 0.01023x - 1.023\times10^{-5} = \varnothing$$

$$x = \frac{-b+\sqrt{b^2-4ac}}{2a} = 9.990\times10^{-6} M$$

$$[H_3O^+]_{total} = \underbrace{1\times10^{-3}}_{\text{from } H_2SO_4} + 9.18\times10^{-4} = 1.918\times10^{-3} M$$

$$pH = -\log 1.918\times10^{-3} = 2.71$$

8. 100 mM HBr => 0.1 M H_3O^+ (because strong acid)

So we are looking for $[HNO_2]$ that produces 0.1 M H_3O^+ at equilibrium

$$HNO_2(aq) + H_2O(l) \rightleftharpoons NO_2^- + H_3O^+$$

$K_a = 5.62\times10^{-4}$

	HNO_2	NO_2^-	H_3O^+
I	I	$\varnothing$	$\varnothing$
C	$-x$	$+x$	$+x$
E	$I-x$	x	x

0.1 M = x

$$5.62\times10^{-4} = \frac{x^2}{I-x} = \frac{(0.1)^2}{I-0.1}$$

$I - 0.1 = 17.79$

$I = 17.89$ M

9.

$$\frac{100\ mL}{} \left| \frac{10^{-3}\ L}{1\ mL} \right| \frac{100\ mmol}{L} = 10\ mmol\ NO_2^- \quad (0.01\ mol)$$

a. to neutralize, need 10 mmol HCl

$$\frac{10\ mmol\ HCl}{} \left| \frac{10^{-3}\ mol}{1\ mmol} \right| \frac{L}{1.5\ mol\ HCl} \left| \frac{1\ mL}{10^{-3}\ L} \right. = 6.67\ mL$$