

Acids and Bases.

1. The pK_a of HNO_2 is 3.25. Calculate pK_b , K_a , and K_b .
2. Write a reaction showing the neutralization of HF with KOH .
3. A flask contains $14.86 \mu\text{M HNO}_3$.
 - a. Calculate the pH , pOH , $[\text{H}_3\text{O}^+]$, and $[\text{OH}^-]$
 - b. Do you expect the pH of $14.86 \mu\text{M HNO}_2$ to be more acidic?
4. A flask contains 14.86 nM HNO_3 . Calculate the pH .
5. Calculate the pH of a $148.6 \mu\text{M}$ solution of HNO_2 .
6. Calculate the pH of a $148.6 \mu\text{M}$ solution of KNO_2 .
7. What is the pH of $1 \text{ mM H}_2\text{SO}_4$? Note that the pK_a 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?

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

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

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



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

- boils down to a double displacement reaction



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

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

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

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)

$$4. \quad 14.86 \text{ nM} \text{ } HNO_2 \bigg| \frac{1 \text{ } H_3O^+}{1 \text{ } HNO_2} \bigg| \frac{10^{-9} \text{ M}}{1 \text{ nM}} = 1.486 \times 10^{-8} \text{ M} = [H_3O^+]$$

strong acid

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

$$[H_3O^+]_{\text{total}} = [H_3O^+]_{HNO_2} + [H_3O^+]_{\text{water}} = 1.486 \times 10^{-8} + 10^{-7}$$

$$[H_3O^+]_{\text{total}} = 1.1486 \times 10^{-7} \text{ M}$$

$$pH = -\log 1.1486 \times 10^{-7} = \boxed{6.94}$$

5. 148.6 μM HNO₂ (pK_a = 3.25)

$$\frac{148.6 \mu\text{M}}{1 \text{ mM}} \cdot 10^{-6} \text{ M} = 1.486 \times 10^{-4} \text{ M}$$



I	1.486 × 10 ⁻⁴	0	0
C	-x	+x	+x
E	1.486 × 10 ⁻⁴ - x	x	x

$$5.62 \times 10^{-4} = \frac{x^2}{1.486 \times 10^{-4} - x} \quad x^2 + 5.62 \times 10^{-4}x - 8.35 \times 10^{-8}$$

$$x = \frac{-5.62 \times 10^{-4} + \sqrt{(5.62 \times 10^{-4})^2 - 4(1)(-8.35 \times 10^{-8})}}{2} = 1.22 \times 10^{-4}$$

$$x = [\text{H}_3\text{O}^+] = 1.22 \times 10^{-4}$$

$$\text{pH} = -\log 1.22 \times 10^{-4} = 3.91$$

6. 148.6 μM KNO₂ ← this ionic compound breaks apart into K⁺ + NO₂⁻ ions.

NO₂⁻ is a weak base!

$$K_b = 1.78 \times 10^{-11} \text{ (calculated in \#1)}$$

← REALLY small K,
So you can use the
shortcut



I	1.486 × 10 ⁻⁴	0	0
C	-x	+x	+x
E	1.486 × 10 ⁻⁴ - x	x	x

$$\therefore x = [\text{OH}^-]$$

$$1.78 \times 10^{-11} = \frac{x^2}{1.486 \times 10^{-4} - x} = \frac{x^2}{1.486 \times 10^{-4}}$$

$$x^2 = 2.645 \times 10^{-15}$$

$$x = [\text{OH}^-] = 5.14 \times 10^{-8} \text{ M}$$

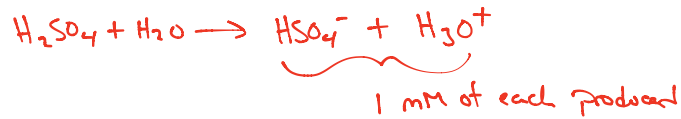
- Really small [OH⁻], so add [OH⁻]_{water}

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

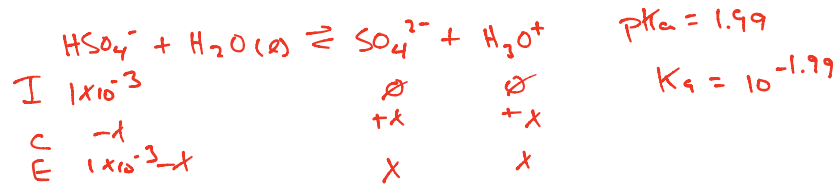
$$\text{pOH} = 6.82$$

$$\text{pH} = 14 - 6.82 = 7.18$$

7. H_2SO_4 is a strong acid



HSO_4^- is a weak acid!



$$1.023 \times 10^{-2} = \frac{x^2}{1 \times 10^{-4} - x}$$

$$x^2 + 0.01023x - 1.023 \times 10^{-5} = 0$$

$$x = \frac{-b + \sqrt{b^2 - 4ac}}{2a} = 9.99 \times 10^{-6} \text{ M}$$

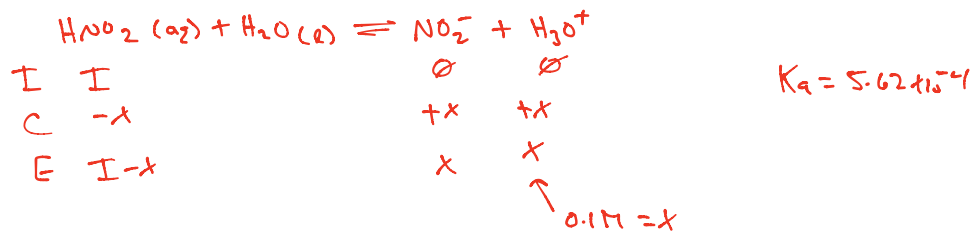
from H_2SO_4

$$[\text{H}_3\text{O}^+]_{\text{total}} = 1 \times 10^{-3} + 9.99 \times 10^{-6} = 1.918 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\log 1.918 \times 10^{-3} = 2.71$$

8. 100 mM $\text{HBr} \Rightarrow 0.1 \text{ M } \text{H}_3\text{O}^+$ (because strong acid)

So we are looking for $[\text{HNO}_2]$ that produces $0.1 \text{ M } \text{H}_3\text{O}^+$ at equilibrium



$$5.62 \times 10^{-4} = \frac{x^2}{I - x} = \frac{(0.1)^2}{I - 0.1}$$

$$I - 0.1 = 17.79$$

$$I = 17.89 \text{ M}$$

9.

$$\frac{100 \text{ mL}}{1 \text{ mL}} \times \frac{10^{-3} \text{ L}}{\text{L}} \times \frac{100 \text{ mmol}}{\text{L}} = 10 \text{ mmol NO}_2^- \quad (0.01 \text{ mol})$$

a. to neutralize, need 10 mmol HCl

$$\frac{10 \text{ mmol HCl}}{1 \text{ mmol}} \times \frac{10^{-3} \text{ mol}}{1.5 \text{ mol HCl}} \times \frac{\text{L}}{10^{-3} \text{ L}} = 6.67 \text{ mL}$$