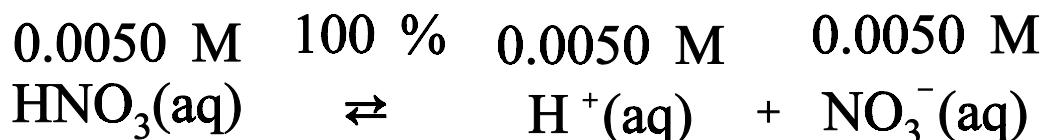


Calculate the pH of 0.0050 M HNO₃(aq).

Since HNO₃ is a strong acid, HNO₃ is completely ionized, and the [H⁺(aq)] equals the molarity of the acid.



$$\text{pH} = -\log [\text{H}^+(\text{aq})]$$

$$\text{pH} = -\log (0.0050 \text{ M}) = 2.30$$

What is the [H⁺(aq)] of a solution with a pH of 2.74?

$$[\text{H}^+(\text{aq})] = 10^{-\text{pH}}$$

$$[\text{H}^+(\text{aq})] = 10^{-2.74} = 1.8 \times 10^{-3} \text{ M}$$

Find the pH of a 0.0025 M NaOH solution.

Since NaOH is a strong base, it is . 100% dissociated. Therefore, the $[\text{OH}^-(\text{aq})]$ equals the molarity of the NaOH solution.

$$\text{pOH} = -\log [\text{OH}^-(\text{aq})]$$

$$\text{pOH} = -\log (0.0025 \text{ M}) = 2.60$$

$$\text{pH} + \text{pOH} = 14$$

$$\text{pH} + 2.60 = 14$$

$$\text{pH} = 14 - 2.60 = 11.40$$

What is the $[\text{OH}^-(\text{aq})]$ of a solution with a pH of 12.70?

$$\text{pH} + \text{pOH} = 14$$

$$12.70 + \text{pOH} = 14$$

$$\text{pOH} = 14 - 12.70 = 1.30$$

$$[\text{OH}^-(\text{aq})] = 10^{-\text{pOH}}$$

$$[\text{OH}^-(\text{aq})] = 10^{-1.30} = 0.05 \text{ M}$$

Calculate the percent ionization of a 0.0040 M solution of a weak acid with a pH of 2.52.

$$[\text{H}^+(\text{aq})] = 10^{-\text{pH}}$$

$$[\text{H}^+(\text{aq})] = 10^{-2.52} = 0.0030 \text{ M}$$

$$\text{percent ionization} = \frac{[\text{H}^+(\text{aq})]}{\text{Molarity}} \times 100$$

$$\% \text{ ionization} = \frac{0.0030 \text{ M}}{0.0040 \text{ M}} \times 100 = 75 \%$$

What is the pH of a 0.0060 M weak acid solution that is 3.0 % ionized?

$$[\text{H}^+] = 0.0060 \text{ M} \times 0.030 = 1.8 \times 10^{-4} \text{ M}$$

$$\text{pH} = -\log [\text{H}^+(\text{aq})]$$

$$\text{pH} = -\log (1.8 \times 10^{-4} \text{ M}) = 3.74$$

Calculate the K_a of a 0.0080 M acid which is 5.0 % ionized.



$$[\text{H}^+(\text{aq})] = 0.0080 \text{ M} \times 0.05 = 0.00040 \text{ M}$$

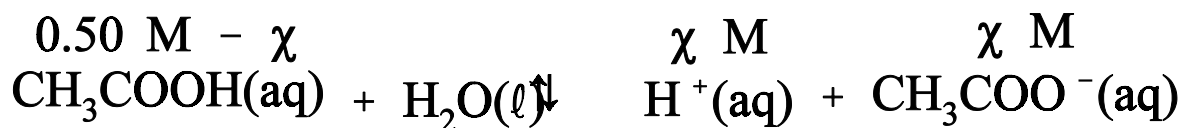
$$[\text{B}^-(\text{aq})] = 0.0080 \text{ M} \times 0.05 = 0.00040 \text{ M}$$

$$[\text{HB(aq)}] = 0.0080 \text{ M} \times 0.95 = 0.0076 \text{ M}$$

$$K_a = \frac{[\text{H}^+(\text{aq})][\text{B}^-(\text{aq})]}{[\text{HB(aq)}]}$$

$$K_a = \frac{(0.00040 \text{ M})(0.00040 \text{ M})}{0.0076 \text{ M}} = 2.1 \times 10^{-5}$$

Calculate the hydrogen ion concentration, $[H^+(aq)]$, in 0.5 M acetic acid. What is the pH of the solution? $K_a = 1.8 \times 10^{-5}$.



$$K_a = \frac{[H^+(\text{aq})][\text{CH}_3\text{COO}^-(\text{aq})]}{[\text{CH}_3\text{COOH(aq)}]} = \frac{(x)(x)}{(0.5 - x)} = 1.8 \times 10^{-5}$$

$$x^2 = 1.8 \times 10^{-5}(0.5) - 1.8 \times 10^{-5}x$$

since $x \ll 0.5 \text{ M}$, then $1.8 \times 10^{-5}x \approx 0$

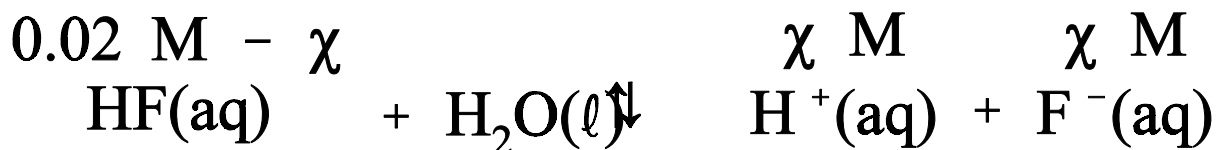
$$x^2 = 1.8 \times 10^{-5}(0.5) = 9.0 \times 10^{-6}$$

$$x = [H^+(\text{aq})] = 3.0 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\log[H^+(\text{aq})]$$

$$\text{pH} = -\log(3.0 \times 10^{-3} \text{ M}) = 2.52$$

Calculate the percent ionization of a 0.02 M HF(aq) solution. $K_a = 3.5 \times 10^{-4}$.



$$K_a = \frac{[\text{H}^+(\text{aq})][\text{F}^-(\text{aq})]}{[\text{HF(aq)}]} = \frac{(\chi)(\chi)}{(0.02 - \chi)} = 3.5 \times 10^{-4}$$

$$x^2 = 3.5 \times 10^{-4}(0.02) - 3.5 \times 10^{-4}x$$

since $x \ll 0.02 \text{ M}$, then $3.5 \times 10^{-4}x \approx 0$

$$x^2 = 3.5 \times 10^{-4}(0.02) = 7.0 \times 10^{-6}$$

$$x = [\text{H}^+(\text{aq})] = 2.6 \times 10^{-3} \text{ M}$$

$$\text{percent ionization} = \frac{[\text{H}^+(\text{aq})]}{\text{Molarity}} \times 100$$

$$\% \text{ ionization} = \frac{0.0026 \text{ M}}{0.020 \text{ M}} \times 100 = 13 \%$$

Acids ionize in water to form the hydronium ion, H_3O^+ , and the base ion, $\text{B}^-(\text{aq})$.



The K_a becomes

$$K_a = \frac{[\text{H}_3\text{O}^+(\text{aq})][\text{B}^-(\text{aq})]}{[\text{HB}(\text{aq})]}$$

Bases hydrolyze in water to form a weak acid and hydroxide ions, $\text{OH}^-(\text{aq})$.



The K_b becomes

$$K_b = \frac{[\text{HB}(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{B}^-(\text{aq})]}$$

What is the relationship between K_a and K_b ?

$$K_a \times K_b = K_w \text{ (} 1.00 \times 10^{-14} \text{ at } 25^\circ\text{C)}$$

$$K_w = \frac{[\text{H}_3\text{O}^+(\text{aq})][\text{B}^-(\text{aq})]}{[\text{HB}(\text{aq})]} \times \frac{[\text{HB}(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{B}^-(\text{aq})]}$$

$$K_w = [\text{H}_3\text{O}^+(\text{aq})][\text{OH}^-(\text{aq})] = 1.00 \times 10^{-14}$$

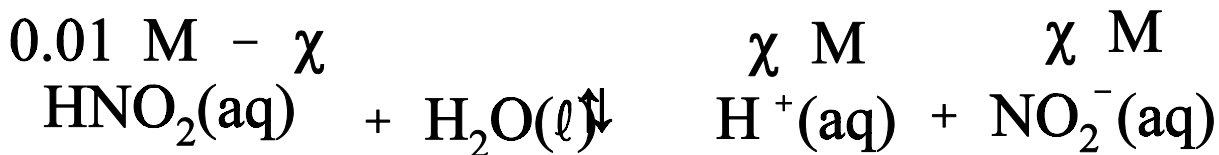
What is the K_b for the acetate ion, if the K_a for acetic acid is 1.8×10^{-5} ?

$$K_a \times K_b = 1.0 \times 10^{-14}$$

$$1.8 \times 10^{-5} \times K_b = 1.0 \times 10^{-14}$$

$$K_b = 5.6 \times 10^{-10}$$

What is the pH of a 0.01 M nitrous acid solution, $\text{HNO}_2(\text{aq})$? $K_a = 5.1 \times 10^{-4}$.



$$K_a = \frac{[\text{H}^+(\text{aq})][\text{NO}_2^-(\text{aq})]}{[\text{HNO}_2(\text{aq})]} = \frac{(x)(x)}{(0.01 - x)} = 5.1 \times 10^{-4}$$

$$x^2 = 5.1 \times 10^{-4}(0.01) - 5.1 \times 10^{-4}x$$

since $x \ll 0.01 \text{ M}$, then $5.1 \times 10^{-4}x \approx 0$

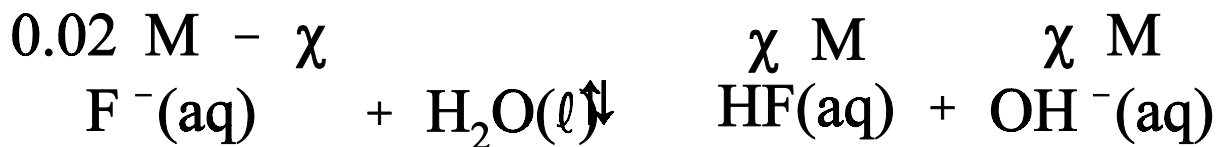
$$x^2 = 5.1 \times 10^{-6}$$

$$x = [\text{H}^+(\text{aq})] = 0.0023 \text{ M}$$

$$\text{pH} = -\log[\text{H}^+(\text{aq})]$$

$$\text{pH} = -\log(0.0023 \text{ M}) = 2.65$$

What would be the pH of a 0.02 M NaF solution? $K_a(\text{HF}) = 3.5 \times 10^{-4}$.



$$K_a \times K_b = 1.0 \times 10^{-14}$$

$$K_b = 2.86 \times 10^{-11}$$

$$K_b = \frac{[\text{HF}(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{F}^-(\text{aq})]} = 2.86 \times 10^{-11}$$

$$K_b = \frac{(\chi)(\chi)}{(0.02 - \chi)} = 2.86 \times 10^{-11}$$

$$x^2 = 2.86 \times 10^{-11}(0.02) - 2.86 \times 10^{-11}x$$

since $x \ll 0.02 \text{ M}$, then $2.86 \times 10^{-11}x \approx 0$

$$x^2 = 5.72 \times 10^{-13}$$

$$x = 7.6 \times 10^{-7} \text{ M}$$

$$\text{pOH} = -\log(7.6 \times 10^{-7}) = 6.12$$

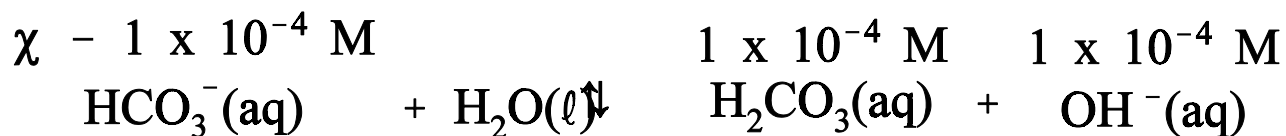
$$\text{pH} + \text{pOH} = 14$$

$$\text{pH} = 14 - \text{pOH} = 7.88$$

What must be the molarity of a NaHCO_3 solution if its pH is 10.00. $K_b = 2.33 \times 10^{-8}$.

$$\text{pH} + \text{pOH} = 14 \qquad [\text{OH}^-(\text{aq})] = 10^{-\text{pOH}}$$

$$\begin{aligned} \text{pOH} &= 14 - 10.00 = 4.00 \\ \text{pOH} &= 10^{-4.00} = 1.0 \times 10^{-4} \text{ M} \end{aligned}$$



$$K_b = \frac{[\text{H}_2\text{CO}_3(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{HCO}_3^-(\text{aq})]} = 2.33 \times 10^{-8}$$

$$K_b = \frac{(1.0 \times 10^{-4} \text{ M})(1.0 \times 10^{-4} \text{ M})}{(\chi - 1.0 \times 10^{-4} \text{ M})} = 2.33 \times 10^{-8}$$

$$1.0 \times 10^{-8} = 2.33 \times 10^{-8} \chi - 2.33 \times 10^{-12}$$

$$\chi = \frac{1.0 \times 10^{-8}}{2.33 \times 10^{-8}} = 0.43 \text{ M NaHCO}_3$$

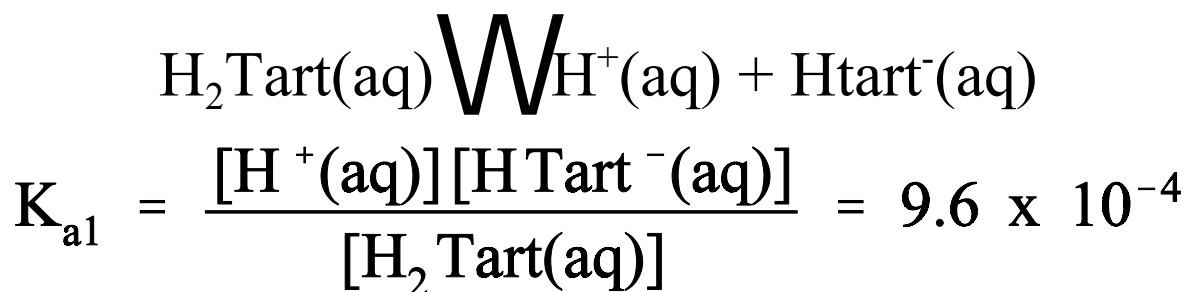
What is the pH and tartrate ion concentration in a 0.20 M solution of tartaric acid, $\text{H}_2\text{C}_4\text{H}_4\text{O}_6(\text{aq})$, or H_2Tart .

$$K_{a1} = 9.6 \times 10^{-4}$$

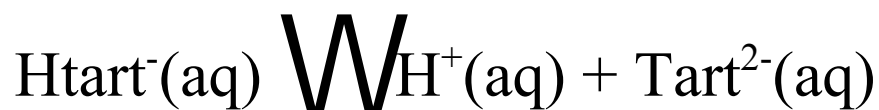
$$K_{a2} = 2.9 \times 10^{-5}$$

Tartaric acid, a diprotic acid (two ionizable hydrogen), ionizes in two steps:

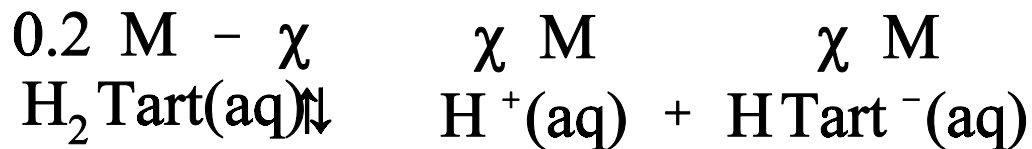
Step #1:



Step #2:



$$K_{a2} = \frac{[H^+(aq)][Tart^{2-}(aq)]}{[HTart^-(aq)]} = 2.9 \times 10^{-5}$$



$$K_a = \frac{[\text{H}^+(\text{aq})][\text{HTart}^-(\text{aq})]}{[\text{H}_2\text{Tart(aq)}]} = 9.6 \times 10^{-4}$$

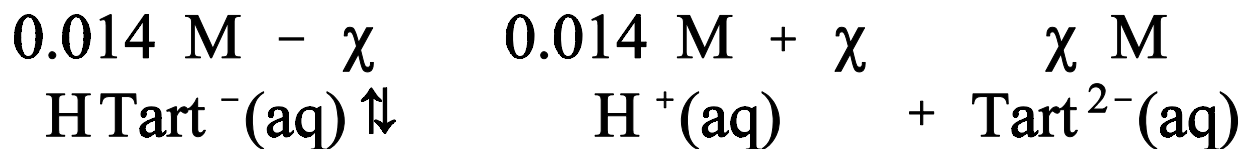
$$K_a = \frac{(\chi)(\chi)}{(0.2 - \chi)} = 9.6 \times 10^{-4}$$

$$x^2 = 9.6 \times 10^{-4}(0.2) - 9.6 \times 10^{-4}x$$

since $x \ll 0.2 \text{ M}$, then $9.6 \times 10^{-4}x \approx 0$

$$x^2 = 1.92 \times 10^{-4}$$

$$x = [\text{H}^+(\text{aq})] = 0.014 \text{ M}$$



$$K_a = \frac{[\text{H}^+(\text{aq})][\text{Tart}^{2-}(\text{aq})]}{[\text{HTart}^-(\text{aq})]} = 2.9 \times 10^{-5}$$

$$K_a = \frac{(0.014 + x)(x)}{(0.014 - x)} = 2.9 \times 10^{-5}$$

since $x \ll 0.014$, then $0.014 + x \approx 0.014$ and $0.014 - x \approx 0.014$

$$\text{therefore, } x = 2.9 \times 10^{-5} \text{ M}$$

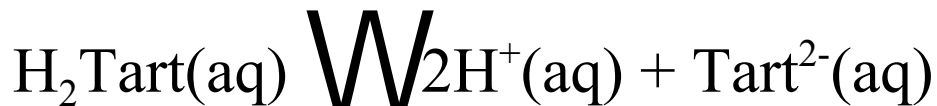
$$[\text{H}^+(\text{aq})] = 0.014 \text{ M} + 2.9 \times 10^{-5} \text{ M}$$

$$[\text{H}^+(\text{aq})] \approx 0.014 \text{ M}$$

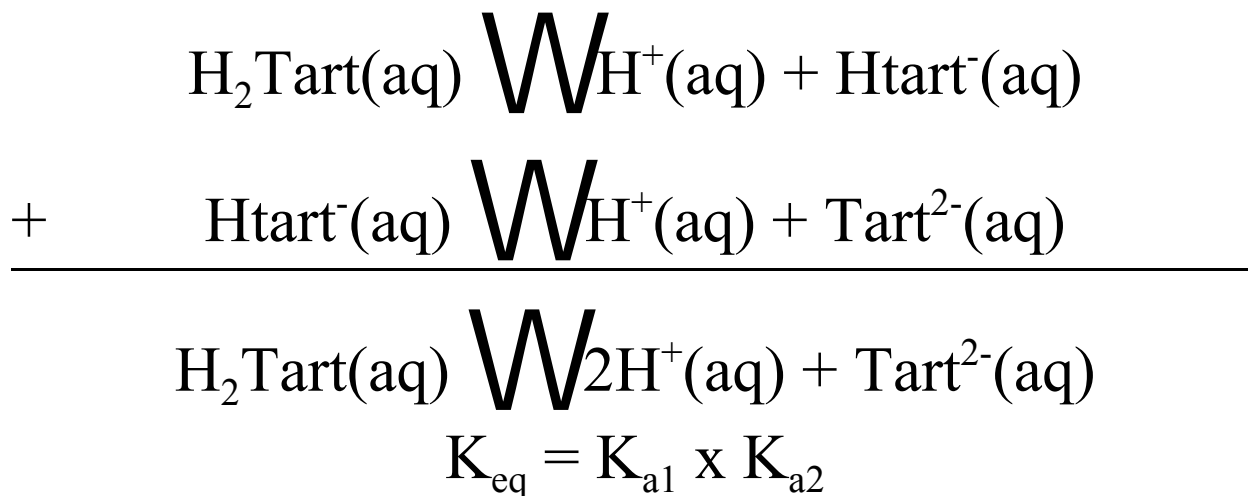
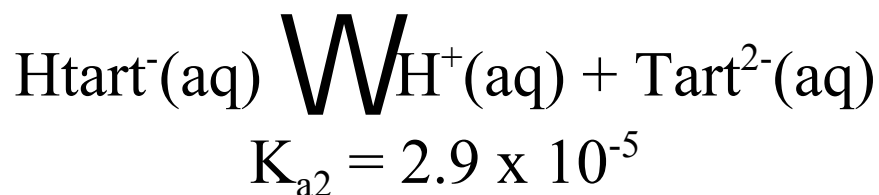
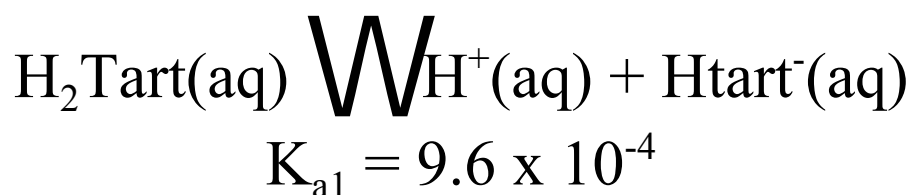
$$\text{pH} = -\log [\text{H}^+] = -\log (0.014) = 1.86$$

$$[\text{Tart}^{2-}] = 2.9 \times 10^{-5} \text{ M}$$

Calculate the ionization constant, K_a , for the reaction:



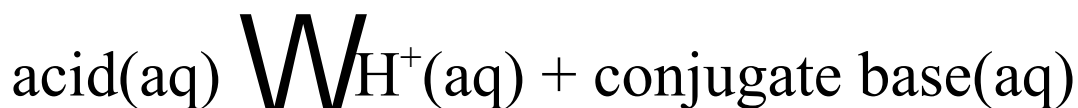
Tartaric acid ionizes in two steps:



$$K_{\text{eq}} = (9.6 \times 10^{-4})(2.9 \times 10^{-5}) = 2.73 \times 10^{-8}$$

Henderson-Hasselbalch Equation

$$\text{pH} = \text{pK}_a + \log \frac{[\text{base}]}{[\text{acid}]}$$



$$K_a = \frac{[\text{H}^+][\text{conjugate base}]}{[\text{acid}]}$$

$$K_a = [\text{H}^+] \times \frac{[\text{conjugate base}]}{[\text{acid}]}$$

$$-\log \left(K_a = [\text{H}^+] \times \frac{[\text{conjugate base}]}{[\text{acid}]} \right)$$

$$\text{pK}_a = \text{pH} - \log \frac{[\text{conjugate base}]}{[\text{acid}]}$$

$$\text{pH} = \text{pK}_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]}$$

Another form of the Henderson-Hasselbalch equation is:



$$K_b = \frac{[HB(aq)][OH^-(aq)]}{[B^-(aq)]}$$

$$K_b = [OH^-(aq)] \times \frac{[HB(aq)]}{[B^-(aq)]}$$

$$-\log \left(K_b = [OH^-(aq)] \times \frac{[HB(aq)]}{[B^-(aq)]} \right)$$

$$pK_b = pOH - \log \frac{[\text{acid}]}{[\text{base}]}$$

$$pOH = pK_b + \log \frac{[\text{acid}]}{[\text{base}]}$$

What is the pH of a solution containing equal volumes of 0.20 M HF(aq) and 0.20 M NaF(aq)? $K_a = 3.5 \times 10^{-4}$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$\text{p}K_a = -\log K_a = -\log (3.5 \times 10^{-4}) = 3.45$$

$$\text{pH} = 3.45 + \log \frac{0.20 \text{ M}}{0.20 \text{ M}}$$

$$\text{pH} = 3.45 + \log 1 = 3.45 + 0 = 3.45$$

What is the pH of a solution which is 0.10 M acetic acid and 0.10 M sodium acetate? $K_a = 1.8 \times 10^{-5}$

$$\text{pH} = \text{pK}_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$\text{pK}_a = -\log K_a = -\log (1.8 \times 10^{-5}) = 4.74$$

$$\text{pH} = 4.74 + \log \frac{0.10 \text{ M}}{0.10 \text{ M}}$$

$$\text{pH} = 4.74 + \log 1 = 4.74 + 0 = 4.74$$

What is the pH of a solution which is 0.10 M acetic acid and 0.010 M sodium acetate? $K_a = 1.8 \times 10^{-5}$

$$\text{pH} = 4.74 + \log \frac{0.010 \text{ M}}{0.10 \text{ M}}$$

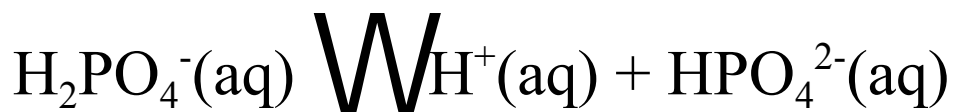
$$\text{pH} = 4.74 + \log 0.1 = 4.74 - 1 = 3.74$$

What is the pH of a solution which is 0.010 M acetic acid and 0.10 M sodium acetate? $K_a = 1.8 \times 10^{-5}$

$$\text{pH} = 4.74 + \log \frac{0.10 \text{ M}}{0.010 \text{ M}}$$

$$\text{pH} = 4.74 + \log 10 = 4.74 + 1 = 5.74$$

What is the pH of a solution containing 2.84 grams of Na_2HPO_4 and 1.20 grams of NaH_2PO_4 in 100 ml of solution?



$$K_a = 6.3 \times 10^{-8}$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$\text{p}K_a = -\log K_a = -\log (6.3 \times 10^{-8}) = 7.20$$

$$\frac{2.84 \text{ g. Na}_2\text{HPO}_4}{142 \text{ g/mol}} = 0.020 \text{ mol}, \frac{0.020 \text{ mol}}{0.10 \text{ l}} = 0.20 \text{ M}$$

$$\frac{1.20 \text{ g. NaH}_2\text{PO}_4}{120 \text{ g/mol}} = 0.010 \text{ mol}, \frac{0.010 \text{ mol}}{0.10 \text{ l}} = 0.10 \text{ M}$$

$$\text{pH} = 7.20 + \log \frac{0.20 \text{ M}}{0.10 \text{ M}}$$

$$\text{pH} = 7.20 + \log 2 = 7.20 + 0.30 = 7.50$$

Calculate the pH of a mixture containing 40.0 ml of 0.20 M KOH and 60.0 ml of 0.15 M HBr.



40.0 ml x 0.20 M = 8.0 mmol KOH (strong base)

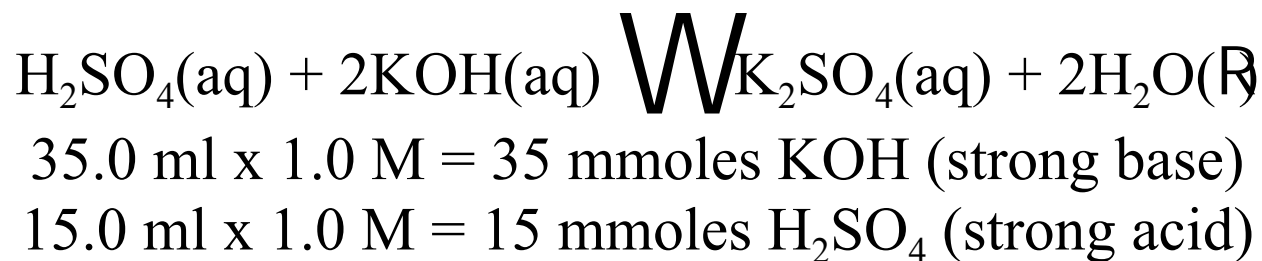
60.0 ml x 0.15 M = 9.0 mmol HBr (strong acid)

	HBr	KOH	KBr	H ₂ O
Start	9.0 mmol	8.0 mmol	0 mmol	-
React	8.0 mmol	8.0 mmol	-	-
At equilibrium	1.0 mmol	0 mmol	8.0 mmol	8.0 mmol

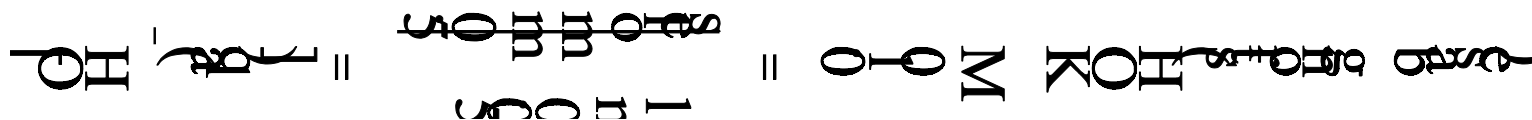


$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (0.01 \text{ M}) = 2.00$$

35.0 ml of 1.0 M KOH is mixed with 15.0 ml of 1.0 M H₂SO₄, What is the pH of the resulting solution?



	H ₂ SO ₄	2KOH	K ₂ SO ₄	H ₂ O
Start	15 mmol	35 mmol	-	-
React	15 mmol	30 mmol	-	-
At Equilibrium	0 mmol	5 mmol	15 mmol	30 mmol



$$\text{pOH} = -\log[\text{OH}^-(\text{aq})] = -\log(0.10 \text{ M}) = 1.00$$

$$\text{pH} = 14 - \text{pOH} = 14 - 1.00 = 13.00$$

Calculate the pH of a mixture containing 150. ml of 0.40 M HF(aq) and 50.0 ml of 1.0 M LiOH(aq). $K_a = 3.5 \times 10^{-4}$



150. ml x 0.40 M = 60 mmol HF (weak acid)

50.0 ml x 1.0 M = 50 mmol LiOH (strong base)

$$\text{p}K_a = -\log K_a = -\log (3.5 \times 10^{-4}) = 3.46$$

	HF	LiOH	LiF	H ₂ O
Start	60 mmol	50 mmol	0 mmol	0 mmol
React	50 mmol	50 mmol	-	-
At equilibrium	10 mmol	0 mmol	50 mmol	50 mmol

$$\text{pH} = \text{p}K_a + \log \frac{[\text{LiF}]}{[\text{HF}]} = 3.46 + \log \frac{50 \text{ mmol}}{10 \text{ mmol}} = 3.46 + 0.70 = 4.16$$

75.0 ml of 0.40 M NH₃(aq) is mixed with 40.0 ml of 0.50 M HNO₃(aq). Calculate the pH of the mixture. K_a = 5.6 x 10⁻¹⁰



75.0 ml x 0.40 M = 30.0 mmol NH₃ (weak base)

40.0 ml x 0.50 M = 20.0 mmol HNO₃ (strong acid)

$$\text{pK}_a = -\log K_a = -\log (5.6 \times 10^{-10}) = 9.25$$

	NH ₃	HNO ₃	NH ₄ NO ₃	H ₂ O
Start	30 mmol	20 mmol	0 mmol	0 mmol
React	20 mmol	20 mmol	-	-
At equilibrium	10 mmol	0 mmol	20 mmol	20 mmol

$$\text{pH} = \text{pK}_a + \log \frac{[\text{NH}_3]}{[\text{NH}_4^+]}$$

75.0 ml of 0.40 M NH₃(aq) is mixed with 40.0 ml of 0.50 M HNO₃(aq). Calculate the pH of the mixture. K_b = 1.8 x 10⁻⁵



75.0 ml x 0.40 M = 30.0 mmol NH₃ (weak base)
 40.0 ml x 0.50 M = 20.0 mmol HNO₃ (strong acid)

$$\text{p}K_b = -\log K_b = -\log (1.8 \times 10^{-5}) = 4.74$$

	NH ₃	HNO ₃	NH ₄ NO ₃	H ₂ O
Start	30 mmol	20 mmol	0 mmol	0 mmol
React	20 mmol	20 mmol	-	-
At equilibrium	10 mmol	0 mmol	20 mmol	20 mmol

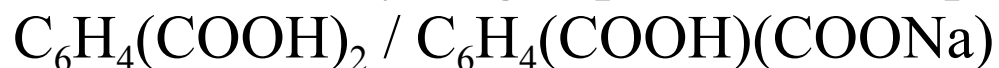
$$\text{pH} = 14 - \text{pOH} = 14 - (\text{p}K_b + \log \frac{[\text{NH}_4^+]}{[\text{NH}_3]}) = 14 - (4.74 + \log \frac{20}{10}) = 14 - 5.04 = 8.96$$

Buffers

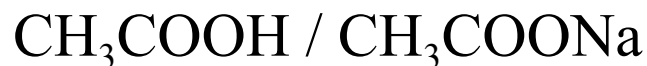
Buffer solutions are solutions that resist a change in pH when hydronium ions or hydroxide ions are added. Buffers usually consist of a mixture of an acid and its conjugate base.

Some common buffer systems are:

Phthalic acid/hydrogen phthalate ion, $pK_a = 2.89$



Acetic acid/acetate ion, $pK_a = 4.74$



Dihydrogen phosphate ion/ hydrogen phosphate ion,



Hydrogen phosphate ion/phosphate ion, $pK_a = 12.44$

