

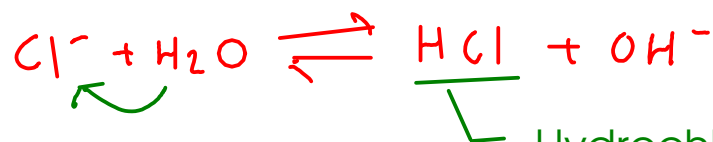
0.100 M NaCl, Find pH



Check these ions to see if they are acidic, basic, or neutral

Na^+ : Cannot be B-L acid (no protons), also not likely to be B-L base; positively charged.

Cl^- : Cannot be B-L acid (no protons). What about a base?



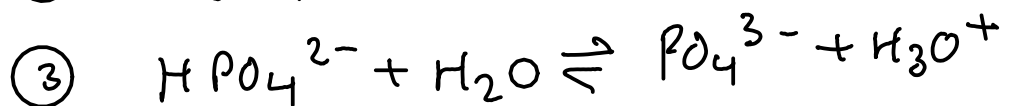
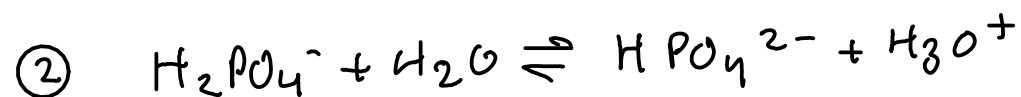
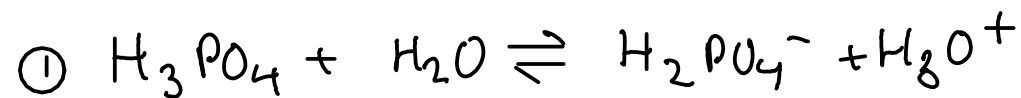
Hydrochloric acid is a STRONG ACID, meaning that it doesn't stay together in water. The reaction we wrote above is not likely to occur. Chloride ion, therefore, is a neutral ion.

Since neither ion affects the water equilibrium, the pH of this solution should be 7.00 (same as distilled water)

POLYPROTIC ACIDS

Find pH of 0,10 M H_3PO_4

... what's special about phosphoric acid?



Phosphoric acid has THREE acidic protons!

$$K_{a1} = 6,9 \times 10^{-3}$$

$$K_{a2} = 6,2 \times 10^{-8}$$

$$K_{a3} = 4,8 \times 10^{-13}$$

The first dissociation is dominant here, and for simple calculations of phosphoric acid in water, we will simply use the first ionization and ignore the other two.

Remember: This is a weak acid. It exists in water mostly as undissociated phosphoric acid molecules.

Solving the equilibrium of phosphoric acid's first proton:



$$K_a = 6.9 \times 10^{-3} = \frac{[\text{H}_2\text{PO}_4^-][\text{H}_3\text{O}^+]}{[\text{H}_3\text{PO}_4]}$$

Species	[Initial]	Δ	[Equilibrium]
H_2PO_4^-	0	+ x	x
H_3O^+	0	+ x	x
H_3PO_4	0.10	- x	0.10 - x

$$\frac{x^2}{0.10 - x} = 6.9 \times 10^{-3}$$

This time, we should solve the quadratic. The $x \ll 0.10$ might not be safe in this problem. (K_a isn't 1000x smaller than 0.10)

$$x^2 = (6.9 \times 10^{-4}) - (6.9 \times 10^{-3})x$$

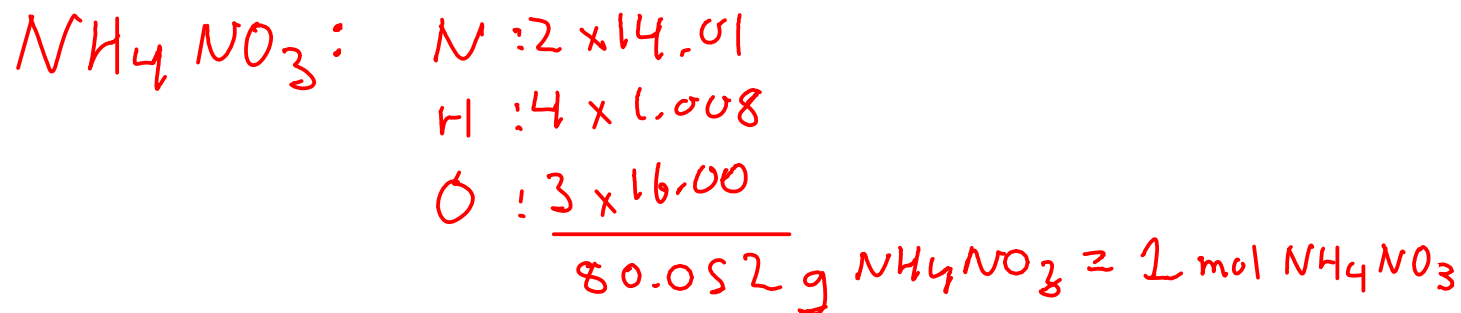
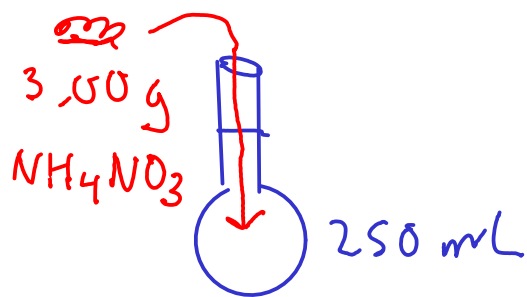
$$x^2 + (6.9 \times 10^{-3})x - (6.9 \times 10^{-4}) = 0$$

$$a=1, b=6.9 \times 10^{-3}, c=-6.9 \times 10^{-4}$$

$$x = 0.0230434426 \text{ OR } -0.0299434426$$

$$\text{"} [\text{H}_3\text{O}^+], \text{ so } \boxed{\text{pH} = 1.64}$$

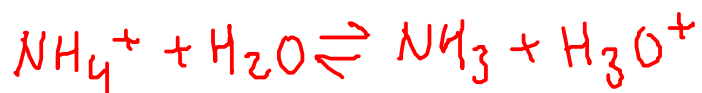
Find the pH of a solution prepared by dissolving 3.00 g of ammonium nitrate solid into enough water to make 250. mL of solution.



Find out what kind of salt we have: acidic, basic, or neutral:



So, we solve the equilibrium between ammonium ion and water



$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$

K_a :

$$K_{b, \text{NH}_3} = 1.8 \times 10^{-5} \quad (\text{p}A-14, \text{Ebbing})$$

$$(K_{a, \text{NH}_4^+})(K_{b, \text{NH}_3}) = 1.0 \times 10^{-14}$$

$$\text{So, } K_{a, \text{NH}_4^+} = 5.56 \times 10^{-10}$$

$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = 5.56 \times 10^{-10}$$

Species	[Initial]	Δ	[Equilibrium]
H_3O^+	0	+x	x
NH_3	0	+x	x
NH_4^+	0.14990	-x	0.14990 - x

To set this problem up further, we need to determine the initial ammonium ion concentration.

$$3.00 \text{ g } \text{NH}_4\text{NO}_3 \times \frac{\text{mol } \text{NH}_4\text{NO}_3}{80.052 \text{ g } \text{NH}_4\text{NO}_3} \times \frac{1}{0.250 \text{ L}} = 0.14990 \text{ M}$$

Solve the equilibrium...

$$\frac{x^2}{0.14990 - x} = 5.56 \times 10^{-10}$$

$$\downarrow x \ll 0.14990$$

$$\frac{x^2}{0.14990} = 5.56 \times 10^{-10}$$

$$x = 9.13 \times 10^{-6} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = 5.04$$

BUFFERS

- resist pH change caused by either the addition of strong acid/base OR by dilution

Made in one of two ways:

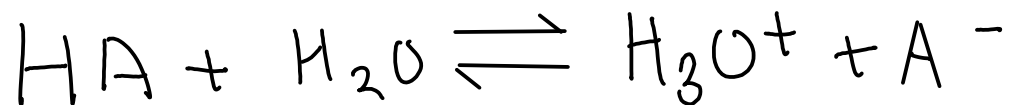
- ① Make a mixture of a weak acid and its conjugate base (as the SALT)
 $\text{HC}_2\text{H}_3\text{O}_2$ $\text{NaC}_2\text{H}_3\text{O}_2$
- ② Make a mixture of a weak base and its conjugate acid (as the SALT)
 NH_3 NH_4Cl

For a weak acid, you would:



- Add HA (weak acid)
- Add a salt containing A^- (example: NaA)

- This solution actually contains an acid and a base at equilibrium, with a significant concentration of BOTH.
- The acid in the buffer can neutralize bases, while the base can neutralize acids.



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

① Take log of both sides

② Multiply by -1

③ Rearrange, solving for pH

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right) \quad \left| \begin{array}{l} \text{Henderson-} \\ \text{Hasselbalch} \\ \text{Equation} \end{array} \right.$$

$[\text{A}^-]$... from the salt

$[\text{HA}]$... from the weak acid

- We ASSUME that the initial concentrations of both the acid and its conjugate are equal to the equilibrium concentrations. Valid IF there are significant amounts of both species initially.

$$\text{pH} = \text{p}K_{a, \text{acidic}} + \log \left(\frac{[\text{basic species}]}{[\text{acidic species}]} \right) \quad \left| \begin{array}{l} \text{Henderson-} \\ \text{Hasselbalch} \\ \text{Equation} \end{array} \right.$$

ex: acidic buffer



$$\text{pH} = \text{p}K_{a, \text{HC}_2\text{H}_3\text{O}_2} + \log \left(\frac{[\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} \right)$$

ex: basic buffer



$$\text{pH} = \text{p}K_{a, \text{NH}_4^+} + \log \left(\frac{[\text{NH}_3]}{[\text{NH}_4^+]} \right)$$

$$\text{p}K_a + \text{p}K_b = 14,00 \quad \dots \text{is the } -\log \text{ of } K_a \times K_b = K_w$$