

What is the pH of the following solution:

0.3M boric acid (HA) and 0.9M sodium borate (A⁻)?

K_a = 6.4x10⁻¹⁰

$$pH = pK_a + \log \frac{A^-}{HA}$$

but we don't have pK_a. We have K_a. To determine pK_a

$$pK_a = -\log(K_a) = -\log(6.4 \times 10^{-10}) = 9.19$$

So, now we have everything we need:

$$pK_a = 9.19$$

$$HA = 0.3$$

$$A^- = 0.9$$

$$pH = pK_a + \log \frac{A^-}{HA} = 9.19 + \log \frac{0.9}{0.3} = \underline{\underline{9.7}}$$

What is the pH of a solution containing 0.2M weak basic drug and 0.002M of its salt. The pK_a=9.36 at 25°C.

We know that A⁻ is the base and HA is the acid. Here, A⁻ is the "weak basic drug" and HA is its (conjugate = acidic) salt.

So, we just plug and chug.

$$pK_a = 9.36$$

$$A^- = 0.2$$

$$HA = 0.002$$

$$PH = pK_a + \log \frac{A^-}{HA} = 9.36 + \log \frac{0.2}{0.002} = \underline{\underline{11.36}}$$

What is the pH of a basic drug and its conjugate salt, if their concentrations are 0.2M and 0.4M, respectively? Assume $pK_a=8.46$.

$$pK_a = 8.46$$

$$A^- = 0.2$$

$$HA = 0.4$$

$$pH = pK_a + \log \frac{A^-}{HA} = 8.46 + \log \frac{0.2}{0.4} = \underline{\underline{8.16}}$$

What is the pH of a solution containing 0.3M ammonium hydroxide and 0.015M of ammonium chloride? $K_b=1.8 \times 10^{-5}$ at 25°C. $K_w=10^{-14}$.

So who's my base and who's my acid? Let's write out each guy to find out.

Ammonium hydroxide = NH_4OH

Ammonium chloride = NH_4Cl

Any guy with OH is always a base (like $Ca(OH)_2$ or NaOH, etc). so, if NH_4OH is the base, that makes NH_4Cl its conjugate acidic salt. I'm going to stick with the A/HA notation just because I'm used to it, but feel free to use Dr. Atef's "base" and "salt" notation. It's the same thing.

So far we know:

$$A^- = 0.3$$

$$HA = 0.015$$

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

$$pH = pK_a + \log \frac{A^-}{HA}$$

but we need pK_a for H-H, not K_b , so we need to convert it.

First switch K_b to pK_b

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

To get pK_a , we use our knowledge that

$$pK_w = 14 = pK_a + pK_b$$

and with a little switcheroo

$$pKa = 14 - pKb = 14 - 4.74 = 9.26$$

Now we have everything we need:

$$A^- = 0.3$$

$$HA = 0.015$$

$$pKa = 9.26$$

$$pH = 9.26 + \log \frac{0.3}{0.015} = \underline{\underline{10.56}}$$

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- a) What is the molar ratio of salt/acid required to adjust the pH of a solution to 5.5 using Na acetate and acetic acid solution. K_a of acetic acid = 1.8×10^{-5} at 25° C.
- b) How much Na acetate (MW=82) and acetic acid (MW=60) are required to adjust the pH of a solution to 5.5 if the total buffer molar concentration is 3.35×10^{-2} mol/L. (see below)

a) We're looking for the molar ratio of salt/acid. If we have an acid, its conjugate salt will be basic, so we're solving for the ratio of A^-/HA .

So far, we have

$$pH = 5.5$$

$$K_a = 1.8 \times 10^{-5}$$

$$pH = pKa + \log \frac{A^-}{HA}$$

First we'll determine pKa from K_a :

$$pKa = -\log(K_a) = 4.75$$

Rearranging the equation for A^-/HA :

$$pH = 5.5$$

$$pKa = 4.75$$

$$pH - pKa = \log \frac{A^-}{HA}$$

$$10^{pH - pKa} = \frac{A^-}{HA} = 10^{5.5 - 4.75} = \frac{5.75}{1} = 5.75 : 1$$

If you don't round, I think you get an answer like 5.69.

If the prof asks us for the ratio strength of each, first we can determine the percentage of each guy and then go from there.

$$A^- = 5.75$$

$$HA = 1$$

$$total = 6.75$$

$$\% A^- = \frac{5.75}{6.75} \times 100\% = 85.2\%$$

$$\% HA = \frac{1}{6.75} \times 100\% = 14.8\%$$

To find the ratio strength, I just push the inverse button on my calc.

$$\% A^- = 85.2\%$$

$$ratio = 1/0.852 = 1.17$$

$$A^- \text{ ratio strength} = \frac{1}{1.17}$$

$$\% HA = 14.8\%$$

$$ratio = 1/0.148 = 6.75$$

$$HA \text{ ratio strength} = \frac{1}{6.75}$$

We already knew the second guy from above.

- b. How much Na acetate (MW=82) and acetic acid (MW=60) are required to adjust the pH of a solution to 5.5 if the total buffer molar concentration is $3.35 \times 10^{-2} \text{ mol/L}$.

So we already know the percentages of each guy, so (1) we'll just apply that to the molarity that we're given here and (2) convert to grams after that.

$$85.2\% (A^-) \text{ of } 3.35 \times 10^{-2} \frac{\text{mol}}{\text{L}} = 0.0285 \frac{\text{mol}}{\text{L}} A^-$$

$$14.8\% (HA) \text{ of } 3.35 \times 10^{-2} \frac{\text{mol}}{\text{L}} = 0.00496 \frac{\text{mol}}{\text{L}} HA$$

Remember that Na acetate is the base (A^-) and acetic acid is the acid (HA). Refer to molecular weights above.

$$\frac{0.0285 \text{ mol } A^-}{\text{L}} \left(\frac{82 \text{ g } A^-}{1 \text{ mol } A^-} \right) = \underline{\underline{2.34 \frac{\text{g}}{\text{L}} A^- = 2.34 \frac{\text{g}}{\text{L}} \text{ Na acetate}}}$$

$$\frac{0.00496 \text{ mol } HA}{\text{L}} \left(\frac{60 \text{ g } HA}{1 \text{ mol } HA} \right) = \underline{\underline{0.298 \frac{\text{g}}{\text{L}} HA = 0.298 \frac{\text{g}}{\text{L}} \text{ acetic acid}}}$$

300mL of 1% atropine sulfate is prepared. How many grams of sodium chloride should be added to make the solution isotonic? The sodium chloride equivalent of atropine is 0.13.

First determine the number of grams of atropine you have knowing the concentration is 1% (1g/100mL).

$$\frac{300 \text{ mL}}{\text{L}} \times \left(\frac{1 \text{ g atropine}}{100 \text{ mL}} \right) = 3 \text{ g atropine}$$

Next determine the number of grams NaCl that the atropine is equivalent to.

$$\frac{3 \text{ g atropine}}{\text{L}} \times \left(\frac{0.13 \text{ g NaCl}}{1 \text{ g atropine}} \right) = 0.39 \text{ g NaCl}$$

Now determine the number of grams of NaCl that would be required for the entire solution (300mL) to be isotonic (0.9% NaCl).

$$\frac{300 \text{ mL}}{\text{L}} \times \left(\frac{0.9 \text{ g NaCl}}{100 \text{ mL}} \right) = 2.7 \text{ g NaCl}$$

If we need a total of 2.7g NaCl and we only have the equivalent of 0.39g NaCl, then we need to add:

$$2.7g - 0.39g = \underline{\underline{2.31g}} \text{ additional NaCl}$$

Isotonic sodium chloride solution contains 0.9% NaCl. If a sodium chloride equivalent of boric acid is 0.52, what is the percentage strength of an isotonic solution of boric acid?

If a solution is 0.9% NaCl, then there are 0.9g NaCl per 100mL of solution.

$$\frac{0.9g \text{ NaCl}}{100mL} \times \left(\frac{1g \text{ BA}}{0.52g \text{ NaCl}} \right) = \frac{1.73g \text{ BA}}{100mL} = \underline{\underline{1.73\% \text{ BA}}}$$

If the sodium chloride equivalent of Drug A is 0.1, is an 8% solution of Drug A hypo-, hyper-, or iso-tonic?

An 8% solution of Drug A is composed of 8g/100mL.

$$\frac{0.8g \text{ DrugA}}{100mL} \times \left(\frac{0.1g \text{ NaCl}}{1g \text{ DrugA}} \right) = \frac{0.8g \text{ NaCl}}{100mL} = \underline{\underline{0.8\% \text{ NaCl}}}$$

A solution is considered isotonic at 0.9% NaCl. Because 0.8% is BELOW 0.9%, the solution is said to be HYPOTonic.

Naphazoline HCl	1%	
NaCl	qs	
Water	ad	30mL

Dispense isotonic solution.

The sodium chloride equivalent of Naphazoline is 0.27.

If a 30mL solution is 1% naphazoline, then there are 3g Naphazoline:

$$\frac{30mL}{100mL} \times \left(\frac{1g \text{ Naphazoline}}{100mL} \right) = 0.3g \text{ Naphazoline}$$

Determine how many grams NaCl this is equivalent to.

$$\frac{0.3g \text{ Naphazoline}}{1g \text{ Naphazoline}} \times \left(\frac{0.27g \text{ NaCl}}{1g \text{ Naphazoline}} \right) = 0.081g \text{ NaCl}$$

This is how many equivalent grams of NaCl we have. If we want the solution to be isotonic (0.9% NaCl) then we need 0.27g NaCl.

$$\frac{30mL}{100mL} \times \left(\frac{0.9g \text{ NaCl}}{100mL} \right) = 0.27g \text{ NaCl}$$

So how many grams of NaCl did we need to add? We have 0.081g and we need 0.27g.

$$0.27g \text{ NaCl} - 0.081g \text{ NaCl} = \underline{\underline{0.189g \text{ NaCl}}}$$

		NaCl equivalent
Phenacaine HCl	1%	0.2
Chlorbutanol	0.5%	0.14
Boric acid	qs	0.52
Purified water	qs	60mL

Prepare isotonic solution.

We're going to determine the equivalent number of grams of NaCl of each ingredient in the 60mL solution.

$$\frac{60mL}{100mL} \times \left(\frac{1g \text{ Phenacaine}}{100mL} \right) \times \left(\frac{0.2g \text{ NaCl}}{1g \text{ Phenacaine}} \right) = 0.12g \text{ NaCl}$$

$$\frac{60mL}{100mL} \times \left(\frac{0.5g \text{ Chlorbutanol}}{100mL} \right) \times \left(\frac{0.14g \text{ NaCl}}{1g \text{ Chlorbutanol}} \right) = 0.042g \text{ NaCl}$$

Determine the total number of equivalent NaCl.

$$0.12g \text{ NaCl} + 0.042g \text{ NaCl} = 0.162g \text{ NaCl}$$

Determine the total number of grams of NaCl required for an isotonic (0.9% NaCl) solution.

$$\frac{60mL}{100ml} \times \left(\frac{0.9g \text{ NaCl}}{100ml} \right) = 0.54g \text{ NaCl}$$

So we need 0.54g NaCl for an isotonic solution, but we only have 0.162g NaCl. Determine the number of grams of additional NaCl to be added.

$$0.54g \text{ NaCl} - 0.162g \text{ NaCl} = 0.378g \text{ NaCl}$$

We need an equivalent of 0.378g NaCl, but we're qs'ing with boric acid (NaCl equivalent = 0.52).

$$\frac{0.378g \text{ NaCl}}{1} \times \left(\frac{1g \text{ BA}}{0.52g \text{ NaCl}} \right) = \underline{\underline{0.727g \text{ BA}}}$$

How many mg of NaCl are required to prepare the following prescription of isotonic epinephrine solution?

The epinephrine solution is isotonic (has an 0.9% NaCl equivalent).

Atropine sulfate (E=0.14)	1%
Epinephrine solution 1:2000	8mL
Purified water	qs
	30mL

Determine the number of grams of equivalent NaCl we have in the two ingredients.

$$\frac{30mL}{100mL} \times \left(\frac{1g \text{ Atropine}}{100mL} \right) \times \left(\frac{0.14g \text{ NaCl}}{1g \text{ Atropine}} \right) = 0.042g \text{ NaCl}$$

$$\frac{30mL}{100mL} \times \left(\frac{0.9g \text{ NaCl}}{100mL} \right) = 0.072g \text{ NaCl}$$

Determine the total number of equivalent NaCl.

$$0.042g \text{ NaCl} + 0.072g \text{ NaCl} = 0.114g \text{ NaCl}$$

Determine the total number of grams of NaCl required for an isotonic (0.9% NaCl) solution.

$$\frac{30mL}{100ml} \times \left(\frac{0.9g \text{ NaCl}}{100ml} \right) = 0.27g \text{ NaCl}$$

So we need 0.27g NaCl for an isotonic solution, but we only have 0.114g NaCl. Determine the number of grams of additional NaCl to be added.

$$0.27g \text{ NaCl} - 0.114g \text{ NaCl} = \underline{\underline{0.156g \text{ NaCl}}}$$