Let’s make a buffer

CHEM 1112

How do you make 250 mL of a 0.2M buffer with a pH of 7.6?

Volume, Molarity, pH?

1. 250 mL = volume
2. 0.2 M = molarity
3. pH = 7.6

<u>Step 1:</u> Choose your weak acid. How?

HA that has a pKa close to the desired pH

Weak acid	**K_a**
Acetic acid	1.8×10^{-5}
Phthalic acid	1.3×10^{-3}
Dihydrogen phosphate (monobasic)	6.2×10^{-8}
Monohydrogen phosphate (dibasic)	4.8×10^{-13}
Carbonic acid	4.6×10^{-7}
Citrate	8.4×10^{-4}
Dihydrogen citrate (monobasic)	1.8×10^{-5}
Monohydrogen citrate (dibasic)	4.0×10^{-6}

Step 2: Use Henderson-Hasselbalch Equation

- pH = pKa + log $[A^-]/[HA]$
- $[A^-]$ = HPO_4^{-2} (dibasic!)
- [HA] = H_2PO_4- (monobasic)

[A-] = moles A^-/total volume

[HA] = moles HA/total volume

$$pH = pKa + \log \frac{\text{moles } A^-/\text{total volume}}{\text{moles HA/total volume}}$$

pH = pKa + log moles A-/moles HA

- pH = pKa + log moles A-/moles HA
- 7.6 = 7.2 + log moles A-/moles HA
- -7.2 from both sides:

0.4 = log moles A-/moles HA

$10^{0.4}$ = moles A-/molesHA

- <u>Step 3:</u> get another equation
- Moles A^- + moles HA = total moles of buffer
- Total moles of buffer = M x vol
- Total moles of buffer = 0.2M x 250mL
- 0.05 moles of buffer = moles A^- + moles HA

- Step 4: solve two equations for two unknowns
- 0.05moles of buffer = moles A- + moles HA
- $10^{0.4}$ = moles A-/moles HA
- 2.51 = moles A-/moles HA
- 2.51 x moles HA = moles A^-
- Substitute
- 0.05 moles buffer = 2.51 moles HA + moles HA
- 0.05 moles buffer = 3.51 moles HA
- Divide both sides by 3.51
- HA moles = 1.4245×10^{-2}
- A- moles = 3.575×10^{-2}

Potassium phosphate monobasic = 136 g/mole
Potassium phosphate dibasic = 174 g/mole

- Step five: use molar mass to determine grams
- HA moles = $1.4245 \times 10^{-2} \times 136 = 1.93$ grams KH_2PO_4
- A- moles = $3.575 \times 10^{-2} \times 174 = 6.22$ grams K_2HPO_4

Weigh, put into a 250 volumetric flask, fill half way

Dissolve solids

Fill to thin line

Invert to mix.

Add acid to water

What is the pH when 10 mL of 0.15M hydrochloric acid is added to a total volume of 200 mL of water?

Moles of acid = 0.01L x 0.15M = 1.5 x 10^{-3} moles

Molarity = 1.5 x 10^{-3} moles/0.2L = 7.5 x 10^{-3} M

pH = -log (7.5 x 10^{-3}) = 2.12

The starting pH for water is about 7, so the addition of acid drops the pH a large amount.

Add the same amount of acid to your buffer

What is the pH when 10 mL of 0.15M hydrochloric acid is added to200mL of pH 7.6 0.2M phosphate buffer?

Moles of acid = 0.01L x 0.15M = 1.5×10^{-3} moles, add moles of acid to HA, subtract moles of acid from A^-

$$pH = 7.2 + \log \frac{(\text{moles } A^- - 1.5 \times 10^{-3})}{(\text{moles HA} + 1.5 \times 10^{-3})}$$

- 0.04 moles of buffer = moles A^- + moles HA
- $10^{0.4}$ = moles A^-/moles HA
- 2.51 = moles A^-/moles HA
- 2.51 x moles HA = moles A^-
- Substitute
- 0.04 moles buffer = 2.51 moles HA + moles HA
- 0.04 moles buffer = 3.51 moles HA
- Divide both sides by 3.51
- HA moles = 1.14×10^{-2}
- A- moles = 2.86×10^{-2}

$$pH = 7.2 + \log \frac{(2.86 \times 10^{-2} - 1.5 \times 10^{-3})}{(1.14 \times 10^{-2} + 1.5 \times 10^{-3})}$$

pH = 7.52

The pH drops by a tiny amount

If you add base to your buffer, use this equation:

$$pH = 7.2 + \log \frac{(2.86 \times 10^{-2} + \text{moles base})}{(1.14 \times 10^{-2} - \text{moles base})}$$