

Class Announcements



8.10 What Are Buffers? – Identifying Buffer Solutions

How many of the following aqueous solutions are buffers? **3!**



a) 0.24 M HI + 0.18 M NaI **X ... HI is a strong acid**

d) 0.10 M CH₃COOH + 0.18 M CH₃COOK **✓ ... Weak acid and its conjugate base**

c) 0.27 M NH₄Br + 0.31 M NH₃ **✓ ... Weak acid and its conjugate base**

b) 0.34 M NH₄NO₃ + 0.39 M NaNO₃ **X ... NO₃⁻ is not the conjugate base of NH₄⁺**

d) 0.10 M HCl + 0.21 M NaF **✓!** SA + NB = 100% $\text{H}_3\text{O}^+ + \text{F}^- = \text{H}_2\text{O}(\text{l}) + \text{HF}(\text{aq})$
$$\underbrace{0.1\text{M} \quad 0.21\text{M}}_{0.11\text{M}} = \quad 0.1\text{M}$$

After reaction 0.1M HF and 0.11M F⁻ remain!

8.10 What Are Buffers? – How Do They Resist Drastic pH Changes

(d)

A 1L solution contains 0.25 mol of NaCN and 0.15 mol of HCN.

1. Increase significantly
3. Decrease significantly
5. Increase



2. Increase slightly
4. Decrease slightly
6. Decrease

a) Addition of 0.1 mol of HCl will case the [HCN] to –



b) Addition of 0.1 mol of HCl will case the pOH to –

2. $\text{pH} \downarrow$ slightly with the addition of H_3O^+ ... $\text{pH} + \text{pOH} = 14$... $\text{pOH} \uparrow$ slightly

c) Addition of 0.1 mol of NaOH will case the [HCN] to –



d) Addition of 0.2 mol of NaOH will case the pH to –

1. Buffer capacity exceeded ... pH will increase significantly.



8.11 How do We Calculate the pH of a Buffer?



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

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+]$$

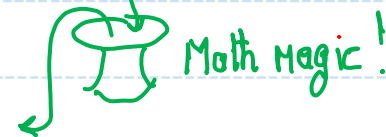
$$[\text{H}_3\text{O}^+] = K_a \left(\frac{[\text{HA}]}{[\text{A}^-]} \right)$$

$$\log_{10} [\text{H}_3\text{O}^+] = \log_{10} K_a + \log_{10} \frac{[\text{HA}]}{[\text{A}^-]}$$

$$-\log_{10} [\text{H}_3\text{O}^+] = -\log_{10} K_a - \log_{10} \frac{[\text{HA}]}{[\text{A}^-]}$$

$$\text{pH} = \text{p}K_a - \log_{10} \frac{[\text{HA}]}{[\text{A}^-]}$$

$$\text{pH} = \text{p}K_a + \log_{10} \frac{[\text{A}^-]}{[\text{HA}]}$$



HA = weak acid, A⁻ is HA's conjugate base

Buffer pH: $\text{pH} = \text{p}K_a + \log_{10} \frac{[\text{Buffer base}]}{[\text{Buffer acid}]}$

↳ HENDERSON-HASSELBACK EQUATION

8.11 How do We Calculate the pH of a Buffer?

MAIN QUESTION

Question

A solution contains the following components

0.208 M HCO_2H ... Acid
0.376 M NaHCO_2 ... Base

What is the pH of the solution?

$K_a \text{HCO}_2\text{H} = 1.8 \times 10^{-4}$

Answer

Enter a response, then Submit.

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$$\text{pH} = \text{p}K_a + \log_{10} \frac{[\text{Buffer base}]}{[\text{Buffer acid}]}$$

$$\text{pH} = -\log_{10}(1.8 \times 10^{-4}) + \log_{10} \frac{[\text{HCO}_2^-]}{[\text{HCO}_2\text{H}]}$$

$$\begin{aligned} \text{pH} &= 3.745 + \log_{10} \left(\frac{0.376}{0.208} \right) \\ &= 3.745 + \log_{10}(1.807) \\ &= 3.745 + 0.257 \\ &= 4.002 \end{aligned}$$



8.11 How do We Calculate the pH of a Buffer?

Acid Base

A buffer solution made from HF and KF has a pH of 2.84. If the pKa for HF is 3.14, what is the $[F^-]/[HF]$ in the buffer?

$$\frac{[F^-]}{[HF]} = 0.5$$



$$pH = pK_a + \log_{10} \frac{[\text{Buffer base}]}{[\text{Buffer acid}]}$$

$$2.84 = 3.14 + \log_{10} \frac{[F^-]}{[HF]}$$

$$2.84 - 3.14 = \log_{10} \frac{[F^-]}{[HF]}$$

$$-0.3 = \log_{10} \frac{[F^-]}{[HF]}$$

$$\frac{[F^-]}{[HF]} = 10^{-0.3} = 0.501$$