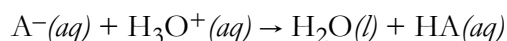


## More Buffer Problems

1. A biochemist must prepare a solution as a medium in an experiment involving acid-producing bacteria. The pH of the medium must not change by more than 0.05 pH units for every 0.0010 mol  $\text{H}_3\text{O}^+$  generated by the organisms per liter of medium. A medium consisting of 0.10 M HA and 0.10 M  $\text{A}^-$  is prepared with a total volume of 1.0 L. Will the buffer capacity be sufficient for this experiment?

With a volume of 1.0 L and concentrations of 0.10 M HA and 0.10 M  $\text{A}^-$  we know that the initial amount of HA and  $\text{A}^-$  are each 0.10 moles. Adding a strong acid to the buffer causes the following reaction



to go to completion; thus, we know that when equilibrium is reached

$$\text{moles A}^- = (\text{moles A}^-)_o - \text{moles H}_3\text{O}^+$$

$$\text{moles HA} = (\text{moles HA})_o + \text{moles H}_3\text{O}^+$$

From the Henderson-Hasselbalch equation

$$\text{pH} = \text{pK}_a + \log \frac{[\text{A}^-]}{[\text{HA}]} = \text{pK}_a + \log \frac{(\text{moles A}^-)}{(\text{moles HA})}$$

we define  $\Delta\text{pH}$  as

$$\Delta\text{pH} = \text{pH}_{\text{final}} - \text{pH}_{\text{initial}} = \left( \log \frac{(\text{moles A}^-)}{(\text{moles HA})} \right)_{\text{final}} - \left( \log \frac{(\text{moles A}^-)}{(\text{moles HA})} \right)_{\text{initial}}$$

Adding a strong acid makes the pH smaller; thus  $\Delta\text{pH}$  is  $-0.05$  and

$$-0.05 = \left( \log \frac{0.10 - \text{moles H}_3\text{O}^+}{0.10 + \text{moles H}_3\text{O}^+} \right)_{\text{final}} - \left( \log \frac{0.10}{0.10} \right)_{\text{initial}}$$

$$\frac{0.10 - \text{moles H}_3\text{O}^+}{0.10 + \text{moles H}_3\text{O}^+} = 0.891$$

$$0.10 - \text{moles H}_3\text{O}^+ = 0.0891 + 0.891 \times \text{moles H}_3\text{O}^+$$

$$\text{moles H}_3\text{O}^+ = 5.76 \times 10^{-3} \text{ mole}$$

Thus, we can add as much as  $5.76 \times 10^{-3}$  moles of strong acid and maintain a  $\Delta\text{pH}$  of 0.05; as this is better than the stated requirements, the buffer is acceptable.

2. A 1.00-L buffer is prepared that is 0.2000 M in the weak acid, HA, and 0.1500 M in the weak base NaA. The buffer has a pH of 3.35. What is the  $\text{pK}_a$  for the weak acid?

Using the Henderson-Hasselbalch equation

$$3.35 = \text{p}K_a + \log \frac{0.1500}{0.2000}$$

Gives the  $\text{p}K_a$  as 3.47.

Is this buffer better at neutralizing strong acid or strong base?

The buffer has more of its conjugate weak acid, HA, than its conjugate weak base,  $\text{A}^-$ ; therefore, the buffer is better at neutralizing strong base.

What is the buffer's capacity to neutralize strong acid?

We can add strong acid until the ratio  $[\text{A}^-]/[\text{HA}]$  is 0.1000. Adding a strong acid converts  $\text{A}^-$  to HA; thus

$$\frac{\text{moles } \text{A}^- - \text{moles } \text{H}_3\text{O}^+}{\text{moles HA} + \text{moles } \text{H}_3\text{O}^+} = 0.1000$$

Since we have 1.00-L of the buffer, we have 0.1500 moles of  $\text{A}^-$  and 0.2000 moles of HA; thus

$$\frac{0.1500 - \text{moles } \text{H}_3\text{O}^+}{0.2000 + \text{moles } \text{H}_3\text{O}^+} = 0.1000$$

$$0.1500 - \text{moles } \text{H}_3\text{O}^+ = 0.02000 + 0.1000 \times \text{moles } \text{H}_3\text{O}^+$$

$$\text{moles } \text{H}_3\text{O}^+ = 0.1182$$

A total of 0.1182 moles of strong acid can be neutralized.

What is the buffer's capacity to neutralize strong base?

We can add strong base until the ratio  $[\text{A}^-]/[\text{HA}]$  is 10.00. Adding a strong base converts HA to  $\text{A}^-$ ; thus

$$\frac{\text{moles } \text{A}^- + \text{moles } \text{OH}^-}{\text{moles HA} - \text{moles } \text{OH}^-} = 10.00$$

Since we have 1.00-L of the buffer, we have 0.1500 moles of  $\text{A}^-$  and 0.2000 moles of HA; thus

$$\frac{0.1500 + \text{moles } \text{OH}^-}{0.2000 - \text{moles } \text{OH}^-} = 10.00$$

$$0.1500 + \text{moles } \text{OH}^- = 2.000 - 10.00 \times \text{moles } \text{OH}^-$$

$$\text{moles } \text{OH}^- = 0.1682$$

A total of 0.1682 moles of strong acid can be neutralized. Note that this is consistent with the expectation that the buffer can neutralize more strong base.

What is the buffer's pH of the buffer if 0.0015 mol NaOH is added to 0.5000 L of the buffer?

Since we are working with only  $\frac{1}{2}$  of the buffer, the amount of HA and  $A^-$  are 0.1000 moles HA and 0.0750 moles  $A^-$ ; thus

$$\text{pH} = 3.47 + \log \frac{0.0750 + 0.0015}{0.1000 - 0.0015} = 3.36$$

3. An environmental chemist needs a carbonate buffer of pH 10.00 to study the effects of the acidification of limestone-rich soils. How many grams of  $\text{Na}_2\text{CO}_3$  must she add to 1.5 L of freshly prepared 0.20 M  $\text{NaHCO}_3$  to prepare this buffer?

This buffer is based on  $\text{HCO}_3^-$  and  $\text{CO}_3^{2-}$  for which the  $\text{p}K_a$  is 10.33. The total moles of  $\text{HCO}_3^-$  at the start is

$$0.20 \text{ M NaHCO}_3 \times 1.5 \text{ L} = 0.30 \text{ moles HCO}_3^-$$

Using the Henderson-Hasselbalch equation

$$10.00 = 10.33 + \log \frac{\text{moles CO}_3^{2-}}{0.30 \text{ moles HCO}_3^-}$$

Solving for the moles of  $\text{CO}_3^{2-}$  gives 0.140, which is equivalent to 14.9 g of  $\text{Na}_2\text{CO}_3$ .