**ANSWER KEY.**

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# pH and Buffer Virtual Lab

Chemistry and AP Chemistry | ANSWER KEY

Open activities.zornscience.org/ph-buffer-lab/ and follow each step. Record pH to two decimal places, as the lab shows it.

## Objective

Use the pH and Buffer Lab to compare strong and weak acids and bases, build buffers with the Henderson-Hasselbalch equation, and test how water and a buffer respond to added acid. Find what limits a buffer's capacity.

## NGSS alignment

HS-PS1-6 (changes to a system at equilibrium). Practices: planning and carrying out investigations, analyzing and interpreting data, using mathematics. Crosscutting concept: stability and change.

## AP alignment

Topics 8.1, 8.2, 8.3, 8.8, 8.9, 8.10. Skills 2.D, 5.F, 6.G.

## Meet the simulation

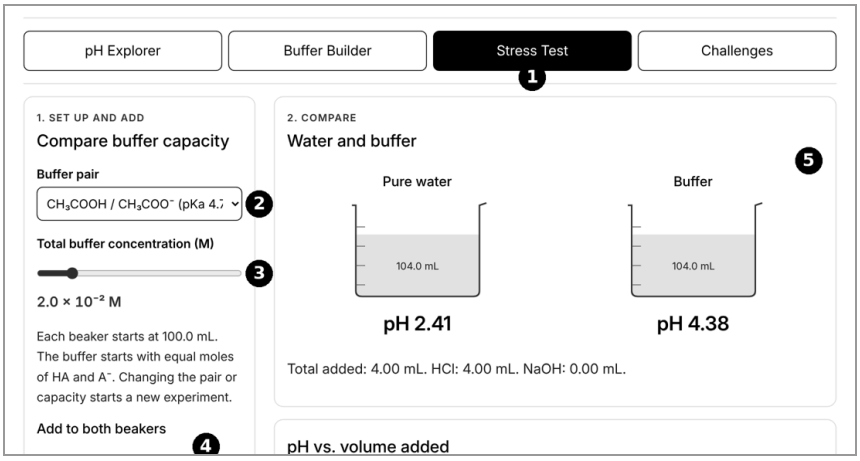



**Figure 1.** The Stress Test. (1) Mode tabs: pH Explorer, Buffer Builder, Stress Test, Challenges. (2) Buffer pair. (3) Total buffer concentration, which sets the capacity. (4) Buttons that add 0.1, 1.0, or 5.0 mL of 0.10 M HCl or NaOH to both beakers. (5) The pH of each beaker. Below these are the pH graph, the moles of HA and $A^-$, and a data table you can copy.

## Pre-lab

**1.** Define a strong acid, a weak acid, and a buffer.

**A strong acid ionizes 100% in water. A weak acid ionizes only partly (mostly molecules). A buffer is a weak acid and its conjugate base (or a weak base and its conjugate acid) that resists changes in pH.**

**2.** Predict the pH of 0.010 M HCl. Will 0.010 M $CH_3COOH$ have a higher or lower pH? Why?

**pH = −log(0.010) = 2.00. Acetic acid will have a higher pH (3.38) because it is a weak acid: only a small fraction ionizes, so $[H_3O^+]$ is much lower.**


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**3.** A buffer has equal concentrations of $CH_3COOH$ and $CH_3COO^-$ ($pK_a = 4.74$). Use Henderson-Hasselbalch to predict its pH.

**pH = 4.74 + log(1) = 4.74**

## Part A. pH Explorer

**4.** Choose the pH Explorer tab. Set each solute and concentration, then record the pH and % ionized. Then answer: what happens to the % ionized of $CH_3COOH$ when it is diluted from 0.10 M to 0.0010 M? How do strong acids compare?

| Solute | Concentration (M) | pH | % ionized |
|---|---|---|---|
| HCl | 0.10 | **1.00** | **100** |
| HCl | 0.0010 | **3.00** | **100** |
| $CH_3COOH$ | 0.10 | **2.88** | **1.33** |
| $CH_3COOH$ | 0.0010 | **3.90** | **12.5** |
| HClO | 0.10 | **4.26** | **0.0548** |
| NaOH | 0.10 | **13.00** | **100** |
| $NH_3$ | 0.10 | **11.12** | **1.33** |

**Diluting acetic acid raises its % ionized (1.33% to 12.5%). Strong acids stay at 100% at any concentration. Diluting HCl 100 times raises the pH by exactly 2.00.**

**5.** Select $CH_3COOH$ at 0.10 M and check Show student method. Record the shortcut pH, the exact pH, and x/C. Repeat for HF at 0.20 M. When is the shortcut $x = \sqrt{(K_a \times C)}$ good enough?

**Acetic acid: shortcut 2.87, exact 2.88, x/C = 1.34%, close enough. HF: shortcut 1.93, exact 1.95, x/C = 5.83%, not reliable. The shortcut works when x is a small percent of C (weak acid, not too dilute).**


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## Part B. Buffer Builder

**6.** Choose Buffer Builder and the $CH_3COOH$ / $CH_3COO^-$ pair. Set each pair of concentrations and record the pH. Then answer: what happens to the pH when both concentrations are divided by 10? What happens when there is more base than acid?

| **[$CH_3COOH$] (M)** | **[$CH_3COO^-$] (M)** | **pH** |
|---|---|---|
| 0.10 | 0.10 | **4.74** |
| 0.10 | 0.30 | **5.22** |
| 0.30 | 0.10 | **4.26** |
| 0.010 | 0.010 | **4.74** |

**The pH stays 4.74: only the ratio [$A^-$]/[HA] matters. More base than acid makes the pH higher than the $pK_a$ (5.22); more acid makes it lower (4.26).**

**7.** You need a buffer at pH 9.00. Which pair should you choose: acetate (4.74), phosphate (7.21), or ammonium (9.25)? What ratio [$NH_3$]/[$NH_4^+$] do you need? Check it in the lab.

**Ammonium, because its $pK_a$ is within 1 of 9.00. 9.00 = 9.25 + log(ratio), so ratio = $10^{-0.25}$ = 0.56. For example, 0.10 M $NH_4^+$ and 0.056 M $NH_3$.**

## Part C. Stress Test

**8.** Choose Stress Test, the acetate pair, and 0.020 M total buffer. Add HCl and record both pH values at each volume. Then press Reset, set the total buffer to 0.20 M, and record the buffer pH again.

| **HCl added (mL)** | **0** | **1.0** | **2.0** | **5.0** | **8.0** | **10.0** | **12.0** |
|---|---|---|---|---|---|---|---|
| pH water | **7.00** | **3.00** | **2.71** | **2.32** | **2.13** | **2.04** | **1.97** |
| pH buffer, 0.020 M | **4.74** | **4.65** | **4.57** | **4.27** | **3.82** | **3.25** | **2.71** |
| pH buffer, 0.20 M | **4.74** | **4.73** | **4.72** | **4.70** | **4.67** | **4.65** | **4.64** |

**The 1.0 mL of HCl drops water's pH by 4.00 units but the buffer's by only 0.09.**


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**9.** Look at the 0.020 M buffer row. Near what volume does it stop working? Use the moles of $CH_3COO^-$ in the buffer (shown in the lab) to explain that volume.

**Around 8 to 10 mL, where the pH starts to drop fast. The buffer has 0.020 M × 0.1000 L ÷ 2 = $1.0 \times 10^{-3}$ mol $CH_3COO^-$. That reacts with $1.0 \times 10^{-3}$ mol HCl, which is 10.0 mL of 0.10 M HCl. After that, no base is left to absorb the acid.**

**10.** Write the net ionic equation for the reaction that happens in the buffer when HCl is added, and when NaOH is added.

**HCl: $CH_3COO^- + H_3O^+ \rightarrow CH_3COOH + H_2O$**

**NaOH: $CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$**

## Analysis

**11.** Compare the 0.020 M and 0.20 M buffers. Did they start at the same pH? Which has more capacity, and about how many mL of HCl would it take to use it up?

**Both start at 4.74 (same ratio). The 0.20 M buffer has 10 times more HA and $A^-$, so it has 10 times the capacity: about 100 mL of 0.10 M HCl.**

**12.** A buffer has more $CH_3COO^-$ than $CH_3COOH$. Is it better at absorbing added acid or added base? Explain.

**Added acid. The conjugate base $CH_3COO^-$ is what reacts with $H_3O^+$, so more base means more capacity for acid.**

**13.** Blood is buffered by $H_2CO_3/HCO_3^-$ ($pK_a$ = 6.35 in this lab). What ratio $[HCO_3^-]/[H_2CO_3]$ gives pH 7.40? Try it in Challenges.

**7.40 = 6.35 + log(ratio), so ratio = $10^{1.05}$ = 11.2. About 11 times more $HCO_3^-$ than $H_2CO_3$.**

**14.** Name two sources of error if you made and tested this buffer in a real school lab.

**Any two: the pH meter is not calibrated or the probe is not rinsed; volumes or masses are measured inaccurately; the temperature is not 25 °C, so $K_w$ and $K_a$ change; $CO_2$ from the air dissolves in the water.**

## Conclusion

**15.** Write a claim, evidence, and reasoning: why does a buffer resist changes in pH, and what limits how much acid it can absorb?

**Claim: a buffer resists pH change until one of its parts runs out. Evidence: 1.0 mL of HCl dropped water from 7.00 to 3.00 but the buffer only from 4.74 to 4.65; the 0.020 M buffer crashed near 10 mL. Reasoning: $A^-$ reacts with added $H_3O^+$ (and HA with added $OH^-$), so the ratio $[A^-]/[HA]$ changes only a little. Once the $A^-$ is used up, the added acid stays in solution and the pH drops fast. More concentrated buffers have more capacity.**


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