

Buffers

How can a solution neutralize both acids and bases?

Why?

Buffer solutions are a mixture of substances that have a fairly constant pH regardless of addition of acid or base. They are used in medicine, industry and manufacturing to keep the pH of a solution stable regardless of other reactions that might be occurring. You have buffers in your blood, for example, because the proper function of enzymes is dependent on the pH of your blood staying in a narrow range. Eating, breathing and exercising can all alter the acidity or basicity of your blood, but the buffers in your blood help control the pH and keep all of your enzymes working properly.

Model 1 – Comparing Solutions:

Beaker A: 100 mL Pure Water



| | WB | WA | ↔ | CA | CB |
|---|----|----|---|------------------------|-----------------------|
| | | | | H_3O^+ | OH^- |
| I | — | — | | 0 | 0 |
| C | — | — | | $+1.0 \times 10^{-7}$ | $+1.0 \times 10^{-7}$ |
| E | — | — | | 1.0×10^{-7} | 1.0×10^{-7} |

Beaker B: 100 mL of 1.00 M HCl



| | SA | WB | → | CA | CB |
|---|-------|----|---|------------------------|---------------|
| | HCl | | | H_3O^+ | Cl^- |
| I | 1.00 | — | | 0 | 0 |
| C | -1.00 | — | | +1.00 | +1.00 |
| F | 0 | — | | 1.00 | 1.00 |

Beaker C: 100 mL of 1.00 M HNO_2

$$K_a \text{ of } \text{HNO}_2 = 4.6 \times 10^{-4}$$



| | WA | WB | ↔ | CA | CB |
|---|----------------|----|---|------------------------|----------------------|
| | HNO_2 | | | H_3O^+ | NO_2^- |
| I | 1.00 | — | | 0 | 0 |
| C | -x | — | | +x | +x |
| E | $1.00 - x$ | — | | x | x |
| | 0.98 | — | | 2.1×10^{-2} | 2.1×10^{-2} |

$$K_a = 4.6 \times 10^{-4} = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{x \cdot x}{1.00 - x}$$

$$4.6 \times 10^{-4} = \frac{x^2}{1.00} \text{ and } x = 2.1 \times 10^{-2}$$

Beaker D: 100 mL of 1.00 M HNO_2 mixed with 100 mL of 1.00 M NaNO_2 (Be careful-dilution!!!)



| | WA | WB | ↔ | CA | CB |
|---|----------------|----|---|------------------------|-----------------|
| | HNO_2 | | | H_3O^+ | NO_2^- |
| I | 0.500 | — | | 0 | 0.500 |
| C | -x | — | | +x | +x |
| E | $0.500 - x$ | — | | x | $0.500 + x$ |
| | 0.500 | — | | 4.6×10^{-4} | 0.500 |

$$K_a = 4.6 \times 10^{-4} = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{x(0.500 + x)}{(0.500 - x)}$$

$$4.6 \times 10^{-4} = \frac{x(0.500)}{0.500} \text{ and } x = 4.6 \times 10^{-4}$$

- For each beaker in Model 1, *underneath each reaction equation*, write SA (strong acid), SB (strong base), WA (weak acid), or WB (weak base) for the reactants, and write CA (conjugate acid) or CB (conjugate base) for the products.
- For Beaker A and B, fill in the ICE/ICF charts to determine all equilibrium/final concentrations of the molecules or ions. ICF ("final") is used for reactions that go to completion, such as reactions with strong acids or bases. For these the change will stoichiometrically consume the limiting reactant.
- For Beaker C and D, fill in the ICE charts, using "x" values as needed. Then, underneath each ICE chart, write the K_a expression and show calculations to determine the value of "x." Then, go back and modify values in the Equilibrium row so that each concentration is a numerical value. (Substitute in all "x" values.)

Guide 16 Solution Chemistry Answer Key

Moffatt and Paige



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