

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 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 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 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.)

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1. Understanding the eBook Guide 16 Solution Chemistry Answer Key
 - The Rise of Digital Reading Guide 16 Solution Chemistry Answer Key
 - Advantages of eBooks Over Traditional Books
2. Identifying Guide 16 Solution Chemistry Answer Key
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in an Guide 16 Solution Chemistry Answer Key
 - User-Friendly Interface
4. Exploring eBook Recommendations from Guide 16 Solution Chemistry Answer Key
 - Personalized Recommendations
 - Guide 16 Solution Chemistry Answer Key User Reviews and Ratings
 - Guide 16 Solution Chemistry Answer Key and Bestseller Lists
5. Accessing Guide 16 Solution Chemistry Answer Key Free and Paid eBooks
 - Guide 16 Solution Chemistry Answer Key Public Domain eBooks
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 - Guide 16 Solution Chemistry Answer Key Enhanced eBook Features
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- Adjustable Fonts and Text Sizes of Guide 16 Solution Chemistry Answer Key
 - Highlighting and Note-Taking Guide 16 Solution Chemistry Answer Key
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- Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Guide 16 Solution Chemistry Answer Key
9. Balancing eBooks and Physical Books Guide 16 Solution Chemistry Answer Key
- Benefits of a Digital Library
 - Creating a Diverse Reading Collection Guide 16 Solution Chemistry Answer Key
10. Overcoming Reading Challenges
- Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
11. Cultivating a Reading Routine Guide 16 Solution Chemistry Answer Key
- Setting Reading Goals Guide 16 Solution Chemistry Answer Key
 - Carving Out Dedicated Reading Time
12. Sourcing Reliable Information of Guide 16 Solution Chemistry Answer Key
- Fact-Checking eBook Content of Guide 16 Solution Chemistry Answer Key
 - Distinguishing Credible Sources
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- Utilizing eBooks for Skill Development
 - Exploring Educational eBooks
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 - Interactive and Gamified eBooks

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