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Skills Worksheet

Problem Solving

pH

In 1909, Danish biochemist S. P. L Sørensen introduced a system in which acidity was expressed as the negative logarithm of the H^+ concentration. In this way, the acidity of a solution having H^+ concentration of 10^{-5} M would have a value of 5. Because the *power* of 10 was now a part of the number, the system was called *pH*, meaning *power of hydrogen*.

Taking the negative logarithm of the hydronium ion concentration will give you a solution's pH as given by the following equation.

$$\text{pH} = -\log[H_3O^+]$$

Likewise, the value pOH equals the negative logarithm of the hydroxide ion concentration.

$$\text{pOH} = -\log[OH^-]$$

Water molecules interact with each other and ionize. At the same time, the ions in solution reform molecules of water. This process is represented by the following reversible equation.

$$H_2O(l) + H_2O(l) \rightleftarrows H_3O(aq) + OH^-(aq)$$

In pure water the concentrations of hydroxide ions and hydronium ions will always be equal. These two quantities are related by a term called the *ion product constant* for water, K_w.

$$K_w = [H_3O^+][OH^-]$$

The ion product constant for water can be used to convert from pOH to pH. The following equations derive a simple formula for this conversion.

Pure water has a pH of 7. Rearranging the equation for pH, you can solve for the hydronium ion concentration of pure water, which is equal to the hydroxide ion concentration. These values can be used to obtain a numerical value for K_w.

$$K_w = [H_3O^+][OH^-] = [1 \times 10^{-7}][1 \times 10^{-7}] = 1 \times 10^{-14}$$

Rearrange the K_w expression to solve for $[OH^-]$.

$$[OH^-] = \frac{K_w}{[H_3O^+]}$$

Substitute this value into the equation for pOH.

$$\text{pOH} = -\log\frac{K_w}{[H_3O^+]}$$

The logarithm of a quotient is the difference of the logarithms of the numerator and the denominator.

$$\text{pOH} = -\log K_w + \log[H_3O^+]$$

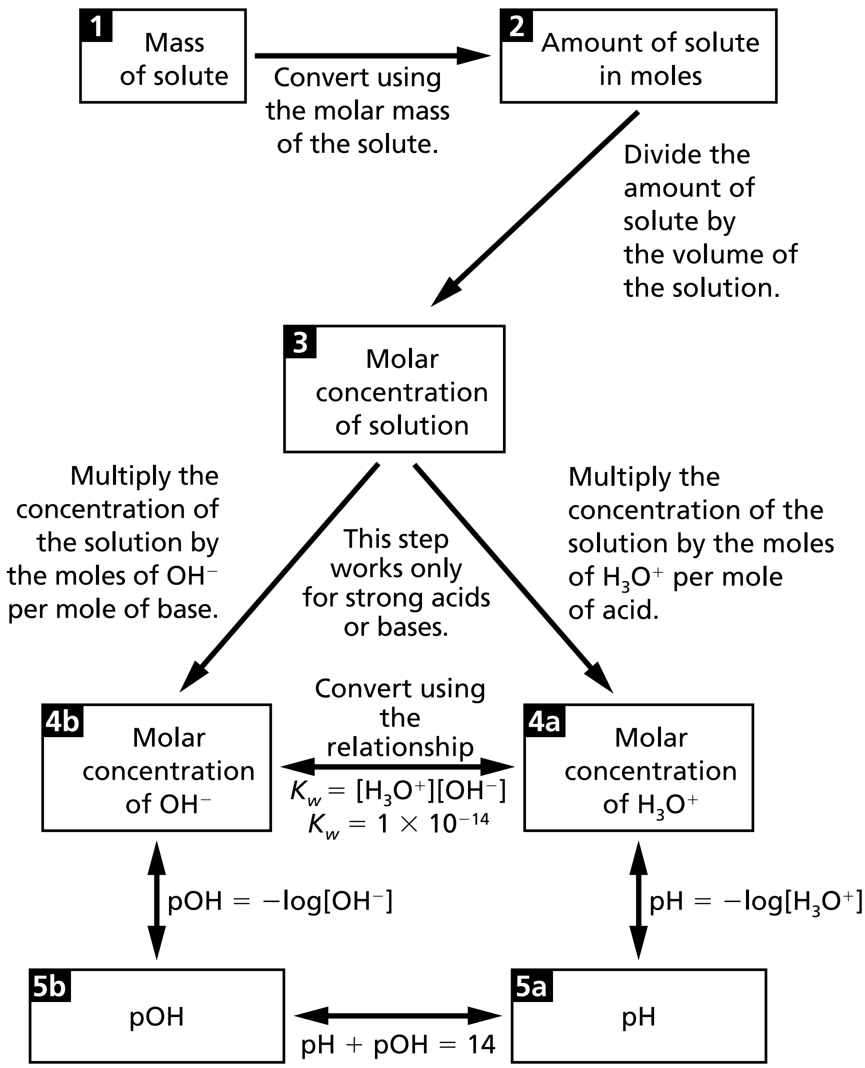
Substitute the value for K_w and pH for $+\log[H_3O^+]$.

$$\text{pOH} = -\log(1.0 \times 10^{-14}) - \text{pH}$$

The negative logarithm of 10^{-14} is 14. Substitute this value, and rearrange.

$$pOH + pH = 14$$

General Plan for Solving pH Problems



Sample Problem 1

A HCl solution has a concentration of 0.0050 M. Calculate [OH^-] and [H_3O^+] for this solution. HCl is a strong acid, so assume it is 100% ionized.

Solution

ANALYZE

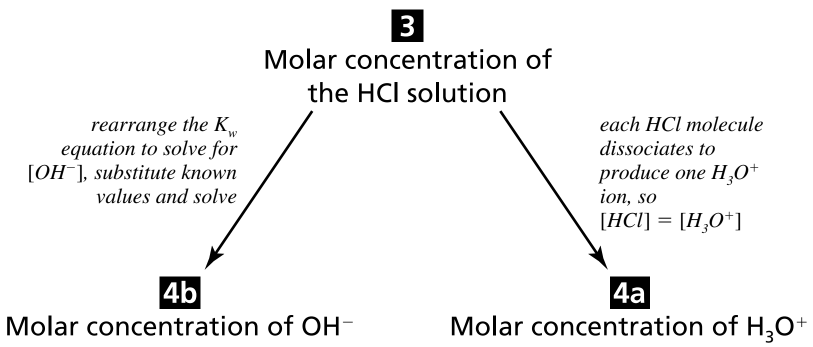
What is given in the problem? **the molarity of the HCl solution, and the fact that HCl is a strong acid**

What are you asked to find? **[H_3O^+] and [OH^-]**

Items	**Data**
Identity of solute	HCl
Concentration of solute	0.0050 M
Acid or base	acid
K_w	1.0×10^{-14}
[H_3O^+]	? M
[OH^-]	? M

PLAN

What steps are needed to calculate the concentration of H_3O^+ and OH^-?
Determine [H_3O^+] from molarity and the fact that the acid is strong. Use K_w to calculate [OH^-].



$$\overset{given}{[\mathrm{HCl}]} = [\mathrm{H_3O^+}]$$

$$\overset{K_w}{1 \times 10^{-14}} = [\mathrm{OH^-}] \times \overset{\substack{calculated\\above}}{[\mathrm{H_3O^+}]}$$

$$[\mathrm{OH^-}] = \frac{\overset{K_w}{1 \times 10^{-14}}}{\underset{calculated\ above}{[\mathrm{H_3O^+}]}}$$

COMPUTE

$$[H_3O^+] = 0.0050 \text{ M}$$

$$[OH^-] = \frac{1.0 \times 10^{-14}}{0.0050} = 2.0 \times 10^{-12} \text{ M}$$

EVALUATE

Are the units correct?

Yes; molarity, or mol/L, was required.

Is the number of significant figures correct?

Yes; the number of significant figures is correct because molarity of HCl was given to two significant figures.

Is the answer reasonable?

Yes; the two concentrations multiply to give 10×10^{-15}, which is equal to 1.0×10^{-14}.

Practice

1. The hydroxide ion concentration of an aqueous solution is 6.4×10^{-5} M. What is the hydronium ion concentration? **ans: 1.6×10^{-10} M**

2. Calculate the H_3O^+ and OH^- concentrations in a 7.50×10^{-4} M solution of HNO_3, a strong acid. **ans: $[H_3O^+] = 7.50 \times 10^{-4}$ M; $[OH^-] = 1.33 \times 10^{-11}$ M**

Sample Problem 2

Calculate the pH of a 0.000 287 M solution of H_2SO_4. Assume 100% ionization.

Solution

ANALYZE

What is given in the problem? **the molarity of the H_2SO_4 solution, and the fact that H_2SO_4 is completely ionized**

What are you asked to find? **pH**

Items	Data
Identity of solute	H_2SO_4
Concentration of solute	0.000 287 M
Acid or base	acid
K_w	1.0×10^{-14}
$[H_3O^+]$	? M
pH	?

PLAN

What steps are needed to calculate the concentration of H_3O^+ and the pH?

Determine $[H_3O^+]$ from molarity and the fact that the acid is 100% ionized. Determine the pH, negative logarithm of the concentration.

3
Molar concentration of the H_2SO_4 solution

each H_2SO_4 molecule dissociates to produce two H_3O^+ ions, so $2 \times [H_2SO_4] = [H_3O^+]$

↓

4a
Molar concentration of H_3O^+

$pH = -\log[H_3O^+]$

↓

5a
pH

$$2 \times \overset{\text{given}}{[H_2SO_4]} = [H_3O^+]$$

$$-\log\overset{\text{calculated above}}{[H_3O^+]} = pH$$

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COMPUTE

$$[H_3O^+] = 2 \times 0.000\ 287\ \text{M} = 5.74 \times 10^{-4}\ \text{M}$$
$$-\log[5.74 \times 10^{-4}] = 3.24$$

EVALUATE

Are the units correct?

Yes; there are no units on a pH value.

Is the number of significant figures correct?

Yes; the number of significant figures is correct because molarity of H_2SO_4 was given to three significant figures.

Is the answer reasonable?

Yes. You would expect the pH of a dilute H_2SO_4 solution to be below 7.

Practice

1. Determine the pH of a 0.001 18 M solution of HBr. **ans: 2.93**

2. What is the pH of a solution that has a hydronium ion concentration of 1.0 M? **ans: 0.00**

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3. What is the pH of a 2.0 M solution of HCl, assuming the acid remains 100% ionized? **ans: −0.30**

4. What is the theoretical pH of a 10.0 M solution of HCl? **ans: −1.00**

Sample Problem 3

A solution of acetic acid has a pH of 5.86. What are the pOH and $[OH^-]$ of the solution?

Solution

ANALYZE

What is given in the problem? **the pH of the acetic acid solution**

What are you asked to find? **pOH and $[OH^-]$**

Items	Data
Identity of solute	acetic acid
Acid or base	acid
pH	5.86
pOH	?
$[OH^-]$	? M

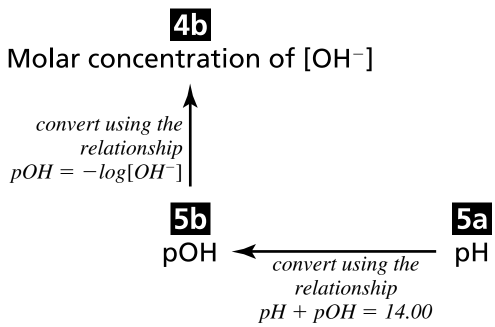
PLAN

What steps are needed to calculate the pOH?

The sum of the pH and pOH of any solution is 14.00. Use this relationship to find the pOH of the acetic acid solution.

What steps are needed to calculate $[OH^-]$?

The pOH of a solution is the negative logarithm of the hydroxide ion concentration. Therefore, calculate $[OH^-]$ using the inverse logarithm of the negative pOH.



$$\overset{given}{pH} + pOH = 14.00$$

$$14.00 - \overset{given}{pH} = pOH$$

$$\overset{\text{calculated above}}{pOH} = -\log[OH^-]$$

$$10^{-\overset{\text{calculated above}}{pOH}} = [OH^-]$$

COMPUTE

$$14.00 - 5.86 = 8.14$$
$$10^{-8.14} = 7.2 \times 10^{-9}\ \mathrm{M}$$

EVALUATE

Are the units correct?

Yes; there are no units on a pOH value, and $[OH^-]$ has the correct units of molarity.

Is the number of significant figures correct?

Yes; the number of significant figures is correct because the data were given to three significant figures.

Is the answer reasonable?

Yes; you would expect the pOH of an acid to be above 7, and the hydroxide ion concentration to be small.

Practice

1. What is the pH of a solution with the following hydroxide ion concentrations?

a. 1×10^{-5} M **ans: 9.0**

b. 5×10^{-8} M **ans: 6.7**

c. 2.90×10^{-11} M **ans: 3.46**

2. What are the pOH and hydroxide ion concentration of a solution with a pH of 8.92? **ans: pOH = 5.08, $[OH^-] = 8.3 \times 10^{-6}$ M**

3. What are the pOH values of solutions with the following hydronium ion concentrations?

a. 2.51×10^{-13} M **ans: 1.40**

b. 4.3×10^{-3} M **ans: 11.6**

c. 9.1×10^{-6} M **ans: 8.96**

d. 0.070 M **ans: 12.8**

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Sample Problem 4

Determine the pH of a solution made by dissolving 4.50 g NaOH in a 0.400 L aqueous solution. NaOH is a strong base.

Solution

ANALYZE

What is given in the problem? **the mass of NaOH, and the solution volume**

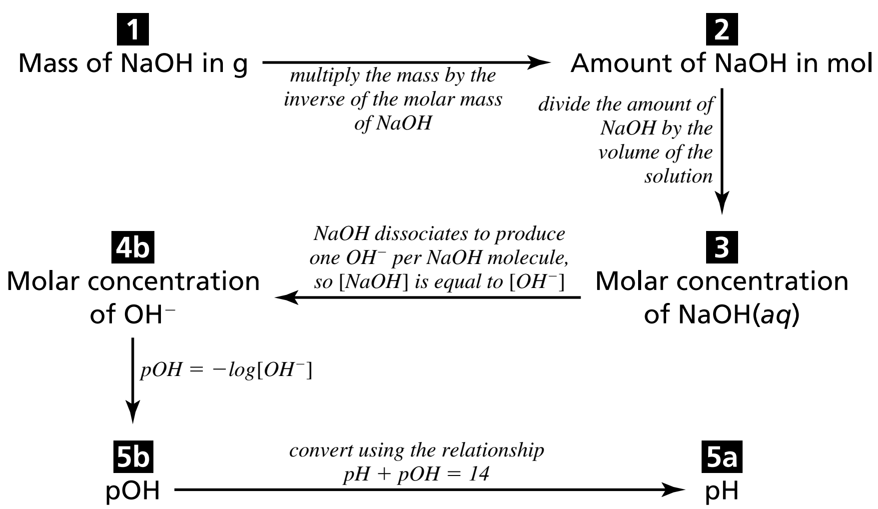
What are you asked to find? **pH**

Items	Data
Identity of solute	NaOH
Mass of solute	4.50 g
Molar mass of solute	40.00 g/mol
Volume of solution	0.400 L
Concentration of solute	? M
Acid or base	base
$[OH^-]$	? M
pOH	?
pH	?

PLAN

What steps are needed to calculate the pH?

First determine the concentration of the solution. Then find the concentration of hydroxide ions. Calculate the pOH of the solution, and use this to find the pH.



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$$\overset{given}{\text{g NaOH}} \times \overset{\frac{1}{\textit{molar mass of NaOH}}}{\frac{\text{1 mol NaOH}}{\text{40.00 g NaOH}}} \times \frac{1}{\underset{given}{\text{L solution}}} \times \frac{\text{1 mol OH}^-}{\text{1 mol NaOH}} = [\text{OH}^-]$$

$$\text{pOH} = -\log\overset{\substack{calculated\\above}}{[\text{OH}^-]}$$

$$\text{pH} + \overset{\substack{calculated\\above}}{\text{pOH}} = 14.00$$

$$14.00 - \overset{\substack{calculated\\above}}{\text{pOH}} = \text{pH}$$

COMPUTE

$$4.50\ \cancel{\text{g NaOH}} \times \frac{\text{1 mol}\ \cancel{\text{NaOH}}}{40.00\ \cancel{\text{g NaOH}}} \times \frac{1}{\text{0.400 L solution}} \times \frac{\text{1 mol OH}^-}{1\ \cancel{\text{mol NaOH}}} = 0.281\ \text{M}$$

$$-\log[0.281] = 0.551$$

$$14.00 - 0.551 = 13.45$$

EVALUATE

Are the units correct?

Yes; pH has no units.

Is the number of significant figures correct?

Yes; the number of significant figures is correct because the value 14.00 has two decimal places.

Is the answer reasonable?

Yes; NaOH is a strong base, so you would expect it to have a pH around 14.

Practice

1. A solution is prepared by dissolving 3.50 g of sodium hydroxide in water and adding water until the total volume of the solution is 2.50 L. What are the OH^- and H_3O^+ concentrations? **ans: $[OH^-] = 0.0350$ M, $[H_3O^+] = 2.86 \times 10^{-13}$ M**

2. If 1.00 L of a potassium hydroxide solution with a pH of 12.90 is diluted to 2.00 L, what is the pH of the resulting solution? **ans: 13.20**

Additional Problems—pH

1. Calculate the H_3O^+ and OH^- concentrations in the following solutions. Each is either a strong acid or a stong base.

a. 0.05 M sodium hydroxide

b. 0.0025 M sulfuric acid

c. 0.013 M lithium hydroxide

d. 0.150 M nitric acid

e. 0.0200 M calcium hydroxide

f. 0.390 M perchloric acid

g. What is the pH of each solution in items 1a. to 1f.?

2. Calculate $[H_3O^+]$ and $[OH^-]$ in a 0.160 M solution of potassium hydroxide. Assume that the solute is 100% dissociated at this concentration.

3. The pH of an aqueous solution of sodium hydroxide is 12.9. What is the molarity of the solution?

4. What is the pH of a 0.001 25 M HBr solution? If 175 mL of this solution is diluted to a total volume of 3.00 L, what is the pH of the diluted solution?

5. What is the pH of a 0.0001 M solution of NaOH? What is the pH of a 0.0005 M solution of NaOH?

6. A solution is prepared using 15.0 mL of 1.0 M HCl and 20.0 mL of 0.50 M HNO_3. The final volume of the solution is 1.25 L. Answer the following questions.

a. What are the $[H_3O^+]$ and $[OH^-]$ in the final solution?

b. What is the pH of the final solution?

7. A container is labeled 500.0 mL of 0.001 57 M nitric acid solution. A chemist finds that the container was not sealed and that some evaporation has taken place. The volume of solution is now 447.0 mL.

a. What was the original pH of the solution?

b. What is the pH of the solution now?

8. Calculate the hydroxide ion concentration in an aqueous solution that has a 0.000 35 M hydronium ion concentration.

9. A solution of sodium hydroxide has a pH of 12.14. If 50.00 mL of the solution is diluted to 2.000 L with water, what is the pH of the diluted solution?

10. An acetic acid solution has a pH of 4.0. What are the $[H_3O^+]$ and $[OH^-]$ in this solution?

11. What is the pH of a 0.000 460 M solution of $Ca(OH)_2$?

12. A solution of strontium hydroxide with a pH of 11.4 is to be prepared. What mass of $Sr(OH)_2$ would be required to make 1.00 L of this solution?

13. A solution of NH_3 has a pH of 11.00. What are the concentrations of hydronium and hydroxide ions in this solution?

14. Acetic acid does not completely ionize in solution. Percent ionization of a substance dissolved in water is equal to the moles of ions produced as a percentage of the moles of ions that would be produced if the substance were completely ionized. Calculate the percent ionization of acetic acid the following solutions.

a. 1.0 M acetic acid solution with a pH of 2.40

b. 0.10 M acetic acid solution with a pH of 2.90

c. 0.010 M acetic acid solution, with a pH of 3.40

15. Calculate the pH of an aqueous solution that contains 5.00 g of HNO_3 in 2.00 L of solution.

16. A solution of HCl has a pH of 1.50. Determine the pH of the solutions made in each of the following ways.

a. 1.00 mL of the solution is diluted to 1000. mL with water.

b. 25.00 mL is diluted to 200 mL with distilled water.

c. 18.83 mL of the solution is diluted to 4.000 L with distilled water.

d. 1.50 L is diluted to 20.0 kL with distilled water.

17. An aqueous solution contains 10 000 times more hydronium ions than hydroxide ions. What is the concentration of each ion?

18. A potassium hydroxide solution has a pH of 12.90. Enough acid is added to react with half of the OH^- ions present. What is the pH of the resulting solution? Assume that the products of the neutralization have no effect on pH and that the amount of additional water produced is negligible.

19. A hydrochloric acid solution has a pH of 1.70. What is the $[H_3O^+]$ in this solution? Considering that HCl is a strong acid, what is the HCl concentration of the solution?

20. What is the molarity of a solution of the strong base $Ca(OH)_2$ in a solution that has a pH of 10.80?

21. You have a 1.00 M solution of the strong acid, HCl. What is the pH of this solution? You need a solution of pH 4.00. To what volume would you dilute 1.00 L of the HCl solution to get this pH? To what volume would you dilute 1.00 L of the pH 4.00 solution to get a solution of pH 6.00? To what volume would you dilute 1.00 L of the pH 4.00 solution to get a solution of pH 8.00?

22. A solution of perchloric acid, $HClO_3$, a strong acid, has a pH of 1.28. How many moles of NaOH would be required to react completely with the $HClO_3$ in 1.00 L of the solution? What mass of NaOH is required?

23. A solution of the weak base NH_3 has a pH of 11.90. How many moles of HCl would have to be added to 1.00 L of the ammonia to react with all of the OH^- ions present at pH 11.90?

24. The pH of a citric acid solution is 3.15. What are the $[H_3O^+]$ and $[OH^-]$ in this solution?

10. a. $-15.0°C$
  b. $104.1°C$
11. 292 g/mol
12. $-47.0°C$
13. 190 g/mol
14. $107.2°C$
15. $-27.9°C$
16. 204.4 g/mol; $C_{16}H_{10}$
17. 9.9 g $CaCl_2$; 49 g glucose
18. $C_3H_6O_3$
19. $-29.2°C$
20. 2.71 kg; $104.9°C$
21. a. 2.2 m
  b. 1.7 m
22. 92.0 g H_2O
23. 150 g/mol; $C_5H_{10}O_5$
24. 124 g/mol; $C_6H_6O_3$

Chemical Equilibrium

EQUILIBRIUM

1. 1.98×10^{-7}
2. 2.446×10^{-12}
3. 3.97×10^{-5}
4. a. 8.13×10^{-4} M
  b. 0.0126 M
5. a. The concentrations are equal.
  b. *K* will increase.
6. 0.09198 M
7. a. 0.7304 M
  b. 6.479×10^{-4} M
8. a. $[A] = [B] = [C] = 1/2[A]_{initial}$
  b. [A], [B], and [C] will increase equally. K remains the same.
9. a. $K_{eq} = [HBr]^2/[H_2][Br_2]$
  b. 2.11×10^{-10} M
  c. Br_2 and H_2 will still have the same concentration. HBr will have a much higher concentration than the two reactants; at equilibrium, essentially only HBr will be present.
10. 1.281×10^{-6}
11. 4.61×10^{-3}
12. a. $K_{eq} = [HCN]/[HCl]$
  b. 3.725×10^{-7} M
13. a. The reaction yields essentially no products at 25°C; as a result, the equilibrium constant is very small. At 110 K, the reaction proceeds to some extent.
  b. 2.51 M
14. 0.0424
15. 0.0390

EQUILIBRIUM OF SALTS, K_{sp}

1. 1.0×10^{-4}
2. 1.51×10^{-7}
3. a. 1.6×10^{-11}
  b. 0.49 g
4. 2.000×10^{-7}
5. a. 8.9×10^{-4} M
  b. 0.097 g
6. 0.036 M
7. a. 1.1×10^{-4} M
  b. 4.6 L
8. a. 5.7×10^{-4} M
  b. 2.1 g
9. 0.83 g remains
10. 2.8×10^{-3} M
11. a. 1.2×10^{-5}
  b. Yes
12. a. 1.1×10^{-8}
  b. No
13. a. $Mg(OH)_2 \rightarrow Mg^{2+} + 2OH^-$
  b. 11 L
  c. A suspension contains undissolved $Mg(OH)_2$ suspended in a saturated solution of $Mg(OH)_2$. As hydroxide ions are depleted by titration, the dissociation equilibrium continues to replenish them until all of the $Mg(OH)_2$ is used up.
14. a. 0.184 M
  b. 46.8 g
15. 5.040×10^{-3}
16. a. 1.0
  b. 0.25
  c. 0.01
  d. 1×10^{-6}
17. 4×10^{-4}; 3×10^{-5}; 3×10^{-5}; 1×10^{-6}; 3×10^{-7}
18. 7.1×10^{-7} M; 1.3×10^{-3} g
19. a. 0.011 M
  b. 0.022 M
  c. 12.35
20. 0.063 g

Acids and Bases

pH

1. a. $[OH^-] = 0.05$ M, $[H_3O^+] = 2 \times 10^{-13}$ M
  b. $[OH^-] = 2.0 \times 10^{-12}$ M, $[H_3O^+] = 5.0 \times 10^{-3}$ M
  c. $[OH^-] = 0.013$ M, $[H_3O^+] = 7.7 \times 10^{-13}$ M

d. $[OH^-] = 6.67 \times 10^{-14}$ M,
$[H_3O^+] = 0.150$ M
e. $[OH^-] = 0.0400$ M,
$[H_3O^+] = 2.50 \times 10^{-13}$ M
f. $[OH^-] = 2.56 \times 10^{-14}$ M,
$[H_3O^+] = 0.390$ M
g. 10, 2.3, 12.11, 0.824, 12.602, 0.409

2. $[OH^-] = 0.160$ M
$[H_3O^+] = 6.25 \times 10^{-14}$ M
3. 0.08 M
4. 2.903; 4.137
5. 10.0; 10.7
6. a. $[H_3O^+] = 0.020$ M,
$[OH^-] = 5.0 \times 10^{-13}$ M
b. 1.7
7. a. 2.804
b. 2.755
8. 2.9×10^{-11} M
9. 10.54
10. $[H_3O^+] = 1 \times 10^{-4}$ M
$[OH^-] = 1 \times 10^{-10}$ M
11. 10.96
12. 0.2 g
13. $[H_3O^+] = 1.0 \times 10^{-3}$ M
$[OH^-] = 1.0 \times 10^{-4}$ M
14. a. 0.40%
b. 1.3%
c. 4.0%
15. 1.402
16. a. 4.50
b. 2.40
c. 3.83
d. 5.63
17. $[H_3O^+] = 1 \times 10^{-9}$ M
$[OH^-] = 1 \times 10^{-5}$ M
18. 11.70
19. $[H_3O^+] = 0.020$ M
$[HCl] = 0.020$ M
20. 3.2×10^{-4} M
21. pH = 0.000
10. kL
1.0×10^2 L
10. kL
22. 0.052 mol NaOH
2.1 g NaOH
23. 7.1×10^{-3}
24. $[H_3O^+] = 7.1 \times 10^{-4}$ M
$[OH^-] = 1.4 \times 10^{-11}$ M

TITRATIONS

1. 0.4563 M KOH
2. 2.262 M CH_3COOH
3. 2.433 M NH_3
4. a. 20.00 mL base
b. 10.00 mL acid
c. 80.00 mL acid
5. 5.066 M HF
6. 0.2216 M oxalic acid
7. 1.022 M H_2SO_4
8. 0.1705 M KOH
9. 0.5748 M citric acid
10. 0.3437 M KOH
11. 43.2 mL NaOH
12. 23.56 mL H_2SO_4
13. 0.06996 mol KOH; 3.926 g KOH;
98.02% KOH
14. 51.7 g $Mg(OH)_2$
15. 0.514 M NH_3
16. 20.85 mL oxalic acid
17. a. 0.5056 M HCl
b. 0.9118 M RbOH
18. 570 kg $Ca(OH)_2$
19. 16.9 M HNO_3
20. 33.58 mL

EQUILIBRIUM OF ACIDS AND BASES, K_a AND K_b

1. pH = 2.857
$K_a = 1.92 \times 10^{-5}$
2. pH = 1.717
$K_a = 2.04 \times 10^{-3}$
3. a. $[H_3O^+] = 3.70 \times 10^{-11}$ M
pH = 10.431
$K_b = 1.82 \times 10^{-7}$
b. $[B] = 4.66 \times 10^{-3}$ M
$K_b = 1.53 \times 10^{-4}$
pH = 10.93
c. $[OH^-] = 1.9 \times 10^{-3}$ M
$[H_3O^+] = 5.3 \times 10^{-12}$ M
$[B] = 0.0331$ M
$K_b = 1.1 \times 10^{-4}$
d. $[B]_{initial} = 7.20 \times 10^{-3}$ M
$K_b = 1.35 \times 10^{-4}$
pH = 10.96
4. 6.4×10^{-5}
5. $K_b = 3.1 \times 10^{-5}$
$[H_2NCH_2CH_2OH] = 6.06 \times 10^{-3}$ M
6. 1.3×10^{-4}
7. $[OH^-] = 1.58 \times 10^{-5}$ M
pH = 9.20
8. $K_a = 2.2 \times 10^{-3}$
$[HB]_{initial} = 0.0276$ M
9. 4.63×10^{-3}
10. 2.62×10^{-4}
11. $[H_3O^+] = 0.0124$ M
pH = 1.907