

Acids and Bases

Section 18.1 Introduction to Acids and Bases

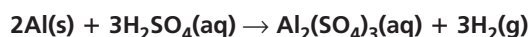
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Practice Problems

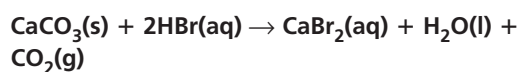
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1. Write balanced equations for reactions between the following.

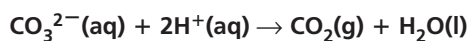
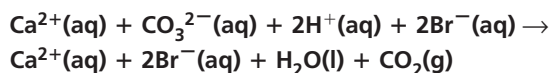
- a. aluminum and sulfuric acid



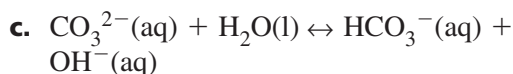
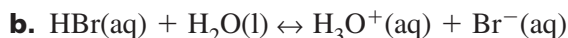
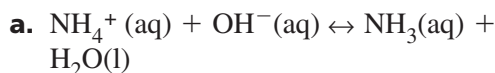
- b. calcium carbonate and hydrobromic acid



2. **Challenge** Write the net ionic equation for the reaction in question 1b.

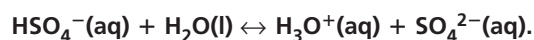


3. Identify the conjugate acid-base pairs in the following reactions.



Acid	Conjugate base	Base	Conjugate acid
a. NH_4^+	NH_3	OH^-	H_2O
b. HBr	Br^-	H_2O	H_3O^+
c. H_2O	OH^-	CO_3^{2-}	HCO_3^-

4. **Challenge** The products of an acid-base reaction are H_3O^+ and SO_4^{2-} . Write a balanced equation for the reaction and identify the conjugate acid-base pairs.



Reactant base: H_2O ; conjugate acid: H_3O^+

Reactant acid: HSO_4^- ; conjugate base: SO_4^{2-}

Section 18.1 Assessment

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5. **Explain** why many Lewis acids and bases are not classified as Arrhenius or Brønsted-Lowry acids and bases.

A Lewis acid is an electron pair acceptor and a Lewis base is an electron pair donor. A Lewis acid may not have an ionizable hydrogen ion or hydroxide ion to qualify as an Arrhenius acid or base. A Lewis acid may not have a hydrogen ion to donate, so that it could not qualify as a Brønsted-Lowry acid. However, all Lewis bases are Brønsted-Lowry bases because they can accept a hydrogen ion

6. **Compare** the physical and chemical properties of acids and bases.

Physical Properties: Acids taste sour and conduct electricity. Bases taste bitter, feel slippery, and conduct electricity.

Chemical Properties: Acids react with some metals to produce hydrogen gas and with carbonates to produce carbon dioxide. Acids turn blue litmus red. Bases react with acids and turn red litmus blue.

7. **Explain** how the concentrations of hydrogen ions and hydroxide ions determine whether a solution is acidic, basic, or neutral.

In an acidic solution, $[\text{H}^+] > [\text{OH}^-]$; in a neutral solution, $[\text{H}^+] = [\text{OH}^-]$; in a basic solution, $[\text{H}^+] < [\text{OH}^-]$.

8. Explain why many compounds that contain one or more hydrogen atoms are not classified as Arrhenius acids.

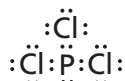
Only compounds that have one or more ionizable hydrogen atom are Arrhenius acids. An ionizable hydrogen atom is an atom that is bonded to an electronegative element such as oxygen.

9. Identify the conjugate acid-base pairs in the following equation.



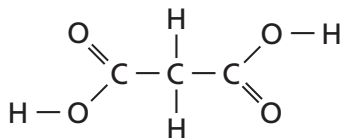
HNO_2 (acid) and NO_2^- (conjugate base); H_2O (base) and H_3O^+ (conjugate acid)

10. Write the Lewis structure for phosphorus trichloride (PCl_3). Is PCl_3 a Lewis acid, a Lewis base, or neither?



Phosphorus in PCl_3 has three electrons, which it shares with three chlorine atoms, and an unshared pair of electrons. The unshared pair can act as Lewis base.

11. Interpreting Scientific Illustrations In the accompanying structural formula, identify any hydrogen atoms that are likely to be ionizable.



The two hydrogen atoms attached to oxygen atoms are ionizable.

Section 18.2 Strengths of Acids and Bases

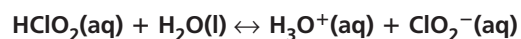
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Practice Problems

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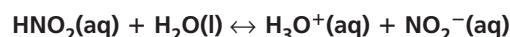
12. Write ionization equations and acid ionization constant expressions for the following acids.

a. HClO_2



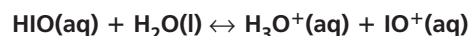
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{ClO}_2^-]}{[\text{HClO}_2]}$$

b. HNO_2



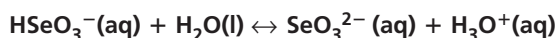
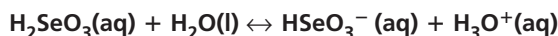
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NO}_2^-]}{[\text{HNO}_2]}$$

c. HIO



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{IO}^-]}{[\text{HIO}]}$$

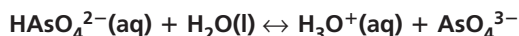
13. Write the first and second ionization equations for H_2SeO_3 .



14. Challenge Given the expression

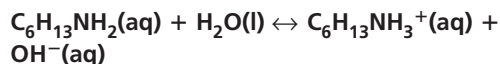
$$K_a = \frac{[\text{ASO}_4^{3-}][\text{H}_3\text{O}^+]}{[\text{HAsO}_4^{2-}]},$$

write the balanced equation for the corresponding reaction.



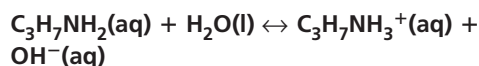
15. Write ionization equations and base ionization constant expressions for the following bases.

a. hexylamine ($C_6H_{13}NH_2$)



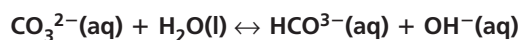
$$K_b = \frac{[C_6H_{13}NH_3^+][OH^-]}{[C_6H_{13}NH_2]}$$

b. propylamine ($C_3H_7NH_2$)



$$K_b = \frac{[C_3H_7NH_3^+][OH^-]}{[C_3H_7NH_2]}$$

c. carbonate ion (CO_3^{2-})



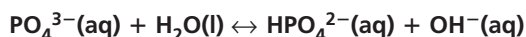
$$K_b = \frac{[HCO_3^-][OH^-]}{[CO_3^{2-}]}$$

d. hydrogen sulfite ion (HSO_3^-)



$$K_b = \frac{[H_2SO_3][OH^-]}{[HSO_3^-]}$$

16. **Challenge** Write an equation for a base equilibrium in which the base in the forward reaction is PO_4^{3-} and the base in the reverse reaction is OH^-



Section 18.2 Assessment

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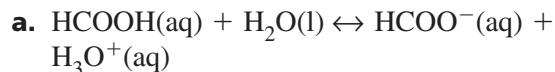
17. **Describe** the contents of dilute aqueous solutions of the strong acid HI and the weak acid HCOOH.

The solution of HI contains only H_3O^+ and I^- ions and water molecules. The solution of HCOOH contains H_3O^+ and $HCOO^-$ ions, and HCOOH and H_2O molecules.

18. **Relate** the strength of a weak acid to the strength of its conjugate base.

The stronger the acid is, the weaker its conjugate base. The weaker the acid is, the stronger its conjugate base.

19. **Identify** the conjugate acid-base pairs in each equation.



acid: HCOOH; conjugate base: $HCOO^-$; base: H_2O ; conjugate acid: H_3O^+ ;



acid: H_2O ; conjugate base: OH^- ; base: NH_3 ; conjugate acid: NH_4^+

20. **Explain** what the K_b for aniline tells you. ($K_b = 4.3 \times 10^{-10}$).

The size of aniline's K_b indicates that aniline is a weak base.

21. **Interpret Data** Use the data in Table 18.4 to put the seven acids in order according to increasing electrical conductivity.

HS^- , HCO_3^- , H_2S , H_2CO_3 , CH_3COOH , HCOOH, HF

Section 18.3 Hydrogen Ions and pH

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Practice Problems

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22. The concentration of either the H^+ ion or the OH^- ion is given for four aqueous solutions at 298 K. For each solution, calculate $[\text{H}^+]$ or $[\text{OH}^-]$. State whether the solution is acidic, basic, or neutral.

a. $[\text{H}^+] = 1.0 \times 10^{-13}M$

$$K_w = [\text{H}^+][\text{OH}^-]$$

$$1.0 \times 10^{-14} = (1.0 \times 10^{-13})[\text{OH}^-]$$

$$\frac{1.0 \times 10^{-14}}{1.0 \times 10^{-13}} = \frac{(1.0 \times 10^{-13})[\text{OH}^-]}{1.0 \times 10^{-13}}$$

$$[\text{OH}^-] = 1.0 \times 10^{-1}M$$

$[\text{OH}^-] > [\text{H}^+]$, the solution is basic.

b. $[\text{OH}^-] = 1.0 \times 10^{-7}M$

$$K_w = [\text{H}^+][\text{OH}^-]$$

$$\frac{1.0 \times 10^{-14}}{1.0 \times 10^{-7}} = \frac{[\text{H}^+](1.0 \times 10^{-7})}{1.0 \times 10^{-7}}$$

$$[\text{H}^+] = 1.0 \times 10^{-7}M$$

$[\text{OH}^-] = [\text{H}^+]$, the solution is neutral.

c. $[\text{OH}^-] = 1.0 \times 10^{-3}M$

$$K_w = [\text{H}^+][\text{OH}^-]$$

$$1.0 \times 10^{-14} = [\text{H}^+](1.0 \times 10^{-3})$$

$$\frac{1.0 \times 10^{-14}}{1.0 \times 10^{-3}} = \frac{[\text{H}^+](1.0 \times 10^{-3})}{1.0 \times 10^{-3}}$$

$$[\text{H}^+] = 1.0 \times 10^{-11}M$$

$[\text{OH}^-] > [\text{H}^+]$, the solution is basic.

d. $[\text{H}^+] = 4.0 \times 10^{-5}M$

$$K_w = [\text{H}^+][\text{OH}^-]$$

$$1.0 \times 10^{-14} = (4.0 \times 10^{-5})[\text{OH}^-]$$

$$= \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-5}} = \frac{(4.0 \times 10^{-5})[\text{OH}^-]}{(4.0 \times 10^{-5})}$$

$$[\text{OH}^-] = 2.5 \times 10^{-10}M$$

$[\text{H}^+] > [\text{OH}^-]$, the solution is acidic

23. Challenge Calculate the number of H^+ ions and the number of OH^- ions in 300 mL of pure water at 298 K.

At 298 K, $[\text{H}^+] = [\text{OH}^-] = 1.0 \times 10^{-7}M$

$$\text{Mol H}^+ = \frac{1.0 \times 10^{-7} \text{ mol}}{1L} \times \frac{1L}{1000 \text{ mL}} \times$$

$$300 \text{ mL} = 3.0 \times 10^{-8} \text{ mol}$$

$$3.0 \times 10^{-8} \text{ mol H}^+ \text{ ions} \times \frac{6.02 \times 10^{23} \text{ H}^+ \text{ ions}}{1 \text{ mol}} =$$

$$1.8 \times 10^{16} \text{ H}^+ \text{ ions}$$

Number of H^+ = number of OH^- =

$$1.8 \times 10^{16} \text{ ions}$$

24. Calculate the pH of solutions having the following ion concentrations at 298 K.

a. $[\text{H}^+] \times 1.0 \times 10^{-2}M$

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

$$\text{pH} = -\log(1.0 \times 10^{-2})$$

$$\text{pH} = 2.00$$

b. $[\text{H}^+] = 3.0 \times 10^{-6}M$

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

$$\text{pH} = -\log(3.0 \times 10^{-6})$$

$$\text{pH} = 5.52$$

25. Calculate the pH of aqueous solutions having the following $[\text{H}^+]$ at 298 K.

a. $[\text{H}^+] = 0.0055M$

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

$$\text{pH} = -\log 0.0055$$

$$\text{pH} = 2.26$$

b. $[\text{H}^+] = 0.000084M$

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

$$\text{pH} = -\log 0.000084$$

$$\text{pH} = 4.08$$

- 26. Challenge** Calculate the pH of a solution having $[\text{OH}^-] = 8.2 \times 10^{-6}M$.

$$[\text{OH}^-] = 8.2 \times 10^{-6}M$$

$$K_w = [\text{H}^+][\text{OH}^-] \times [\text{H}^+](8.2 \times 10^{-6})$$

$$[\text{H}^+] = \frac{1.0 \times 10^{-14}}{8.2 \times 10^{-6}} = 1.2 \times 10^{-9}$$

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

$$\text{pH} = -\log(1.2 \times 10^{-9})$$

$$\text{pH} = 8.92$$

- 27.** Calculate the pH and pOH of aqueous solutions with the following ion concentrations at 298 K.

a. $[\text{OH}^-] = 1.0 \times 10^{-6}M$

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

$$\text{pOH} = -\log(1.0 \times 10^{-6})$$

$$\text{pOH} = 6.00$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 6.00 = 8.00$$

b. $[\text{OH}^-] = 6.5 \times 10^{-4}M$

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

$$\text{pOH} = -\log(6.5 \times 10^{-4})$$

$$\text{pOH} = 3.19$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 3.19 = 10.81$$

c. $[\text{H}^+] = 3.6 \times 10^{-9}M$

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

$$\text{pH} = -\log(3.6 \times 10^{-9})$$

$$\text{pH} = 8.44$$

$$\text{pOH} = 14.00 - \text{pH} = 14.00 - 8.44 = 5.56$$

d. $[\text{H}^+] = 2.5 \times 10^{-2}M$

$$\text{pH} = -\log(2.5 \times 10^{-2})$$

$$\text{pH} = 1.60$$

$$\text{pOH} = 14.00 - \text{pH} = 14.00 - 1.60 = 12.40$$

- 28.** Calculate the pH and the pOH of aqueous solutions with the following concentrations at 298 K.

a. $[\text{OH}^-] = 0.000033M$

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

$$\text{pOH} = -\log(0.000033)$$

$$\text{pOH} = 4.48$$

$$\text{pH} = 14.00 - 4.48 = 9.52$$

b. $[\text{H}^+] = 0.0095M$

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

$$\text{pH} = -\log(0.0095)$$

$$\text{pH} = 2.02$$

$$\text{pOH} = 14.00 - 2.02 = 11.98$$

- 29. Challenge** Calculate pH and pOH for an aqueous solution containing 1.0×10^{-3} mol of HCl dissolved in 5.0 L of solution.

$$[\text{HCl}] = [\text{H}^+] = \frac{1.0 \times 10^{-3} \text{ mol}}{5.0 \text{ L}} = 0.00020M = 2.0 \times 10^{-4}M$$

$$\text{pH} = -\log(2.0 \times 10^{-4}) = -(-3.70) = 3.70$$

$$\text{pOH} = 14.00 - 3.70 = 10.30$$

- 30. Calculate** $[\text{H}^+]$ and $[\text{OH}^-]$ in each of the following solutions.

a. Milk, $\text{pH} = 6.50$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-6.50) = 3.2 \times 10^{-7}M$$

$$\text{pOH} = 14.00 - \text{pH} = 14.00 - 6.50 = 7.50$$

$$[\text{OH}^-] = \text{antilog}(-\text{pOH})$$

$$[\text{OH}^-] = (-7.50) = 3.2 \times 10^{-8}M$$

b. Lemon juice, $\text{pH} = 2.37$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-2.37) = 4.3 \times 10^{-3}M$$

$$\text{pOH} = 14.00 - \text{pH} = 14.00 - 2.37 = 11.63$$

$$[\text{OH}^-] = \text{antilog}(-\text{pOH})$$

$$[\text{OH}^-] = \text{antilog}(-11.63) = 2.3 \times 10^{-12}M$$

- c. Milk of magnesia, $\text{pH} = 10.50$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-10.50) = 3.2 \times 10^{-11}\text{M}$$

$$\text{pOH} = 14.00 - \text{pH} = 14.00 - 10.50 = 3.50$$

$$[\text{OH}^-] = \text{antilog}(-3.50) = 3.2 \times 10^{-4}\text{M}$$

- d. Household ammonia, $\text{pH} = 11.90$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-11.90) = 1.3 \times 10^{-12}\text{M}$$

$$\text{pOH} = 14.00 - \text{pH} = 14.00 - 11.90 = 2.10$$

$$[\text{OH}^-] = \text{antilog}(-2.10) = 7.9 \times 10^{-3}\text{M}$$

31. **Challenge** Calculate the $[\text{H}^+]$ and $[\text{OH}^-]$ in a sample of seawater with a $\text{pOH} = 5.60$.

$$[\text{OH}^-] = \text{antilog}(-\text{pOH})$$

$$[\text{OH}^-] = \text{antilog}(-5.60) = 2.5 \times 10^{-6}\text{M}$$

$$\text{pH} = 14.00 - 5.60 = 8.40$$

$$[\text{H}^+] = \text{antilog}(-8.40) = 4.0 \times 10^{-9}\text{M}$$

32. Calculate the K_a for the following acids using the given information.

- a. 0.220M solution of H_3AsO_4 , $\text{pH} = 1.50$

$$K_a = \frac{[\text{H}^+][\text{H}_2\text{AsO}_4^-]}{[\text{H}_3\text{AsO}_4]}$$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-1.50) = 3.2 \times 10^{-2}\text{M}$$

$$[\text{H}_2\text{AsO}_4^-] = [\text{H}^+] = 3.2 \times 10^{-2}\text{M}$$

$$[\text{H}_3\text{AsO}_4] = 0.220\text{M} - 3.2 \times 10^{-2}\text{M} = 0.188\text{M}$$

$$K_a = \frac{(3.2 \times 10^{-2})(3.2 \times 10^{-2})}{0.188} = 5.4 \times 10^{-3}$$

- b. 0.0400M solution of HClO_2 , $\text{pH} = 1.80$

$$K_a = \frac{[\text{H}^+][\text{ClO}_2^-]}{[\text{HClO}_2]}$$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-1.80) = 1.6 \times 10^{-2}\text{M}$$

$$[\text{ClO}_2^-] = [\text{H}^+] = 1.6 \times 10^{-2}\text{M}$$

$$[\text{HClO}_2] = 0.0400\text{M} - 1.6 \times 10^{-2}\text{M} = 0.024\text{M}$$

$$K_a = \frac{(1.6 \times 10^{-2})(1.6 \times 10^{-2})}{0.024} = 1.1 \times 10^{-2}$$

33. Calculate the K_a of the following acids using the given information.

- a. 0.00330M solution of benzoic acid ($\text{C}_6\text{H}_5\text{COOH}$), $\text{pOH} = 10.70$

$$\text{pH} = 14.00 - \text{pOH}$$

$$\text{pH} = 14.00 - 10.70 = 3.30$$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-3.30) = 5.0 \times 10^{-4}\text{M}$$

$$[\text{C}_6\text{H}_5\text{COO}^-] = [\text{H}^+] = 5.0 \times 10^{-4}\text{M}$$

$$[\text{C}_6\text{H}_5\text{COOH}] = 0.00330\text{M} - 5.0 \times 10^{-4}\text{M} = 0.0028\text{M}$$

$$K_a = \frac{[\text{H}^+][\text{C}_6\text{H}_5\text{COO}^-]}{[\text{C}_6\text{H}_5\text{COOH}]} = \frac{(5.0 \times 10^{-4})(5.0 \times 10^{-4})}{(3.5 \times 10^{-3})}$$

$$K_a = 8.9 \times 10^{-5}$$

- b. 0.100M solution of cyanic acid (HCNO), $\text{pOH} = 11.00$

$$\text{pH} = 14.00 - \text{pOH}$$

$$\text{pH} = 14.00 - 11.00 = 3.00$$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-3.00) = 1.0 \times 10^{-3}\text{M}$$

$$[\text{CNO}^-] = [\text{H}^+] = 1.0 \times 10^{-3}\text{M}$$

$$[\text{HCNO}] = 0.100 - 1.0 \times 10^{-3}\text{M} = 0.099\text{M}$$

$$K_a = \frac{[\text{H}^+][\text{CNO}^-]}{[\text{HCNO}]} = \frac{(1.0 \times 10^{-3})(1.0 \times 10^{-3})}{(0.099)}$$

$$K_a = 1.0 \times 10^{-5}$$

- c. 0.150M solution of butanoic acid ($\text{C}_3\text{H}_7\text{COOH}$), $\text{pOH} = 11.18$

$$\text{pH} = 14.00 - \text{pOH}$$

$$\text{pH} = 14.00 - 11.18 = 2.82$$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-2.82) = 1.5 \times 10^{-3}\text{M}$$

$$[\text{C}_3\text{H}_7\text{COO}^-] = [\text{H}^+] = 1.5 \times 10^{-3}\text{M}$$

$$[\text{C}_3\text{H}_7\text{COOH}] = 0.150\text{M} - 1.5 \times 10^{-3}\text{M} = 0.149\text{M}$$

$$K_a = \frac{[\text{H}^+][\text{C}_3\text{H}_7\text{COO}^-]}{[\text{C}_3\text{H}_7\text{COOH}]} = \frac{(1.5 \times 10^{-3})(1.5 \times 10^{-3})}{(0.149)}$$

$$K_a = 1.5 \times 10^{-5}$$

- 34. Challenge** Calculate the K_a of a 0.0091M solution of an unknown acid (HX) having a pOH of 11.32. Use **Table 18.4** to identify the acid.

$$\text{pH} = 14.00 - \text{pOH}$$

$$\text{pH} = 14.00 - 11.32 = 2.68$$

$$[\text{H}^+] = \text{antilog}(-\text{pH})$$

$$[\text{H}^+] = \text{antilog}(-2.68) = 2.1 \times 10^{-3}M$$

$$[\text{X}^-] = [\text{H}^+] = 2.1 \times 10^{-3}M$$

$$[\text{HX}] = 0.0091 - 0.0021 = 0.0070M$$

$$K_a = \frac{(0.0021)(0.0021)}{(0.0070)} = 6.3 \times 10^{-4}$$

The acid could be hydrofluoric acid (HF).

Section 18.3 Assessment

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- 35. Explain** why the pH of an acidic solution is always a smaller number than the pOH of the same solution.
- The sum of pH and pOH is 14.00. If a solution is acidic, its pH is less than 7.00. Therefore, pOH must be greater than 7.00.
- 36. Describe** how you can determine the pH of a solution if you know its pOH.
- Subtract the pOH from 14.00.
- 37. Explain** the significance of K_w in aqueous solutions.
- At 298 K, the product of the H^+ and OH^- concentrations in any aqueous solution equals 1.00×10^{-14} . If one ion concentration is known, the other can be calculated using the K_w expression.
- 38. Explain,** using Le Châtelier's principle, what happens to the $[\text{H}^+]$ of a 0.10M solution of acetic acid when a drop of NaOH solution is added.
- $$\text{H}_2\text{O}(\text{l}) \leftrightarrow \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$$
- The increase in OH^- ion from the drop of NaOH shifts the self-ionization of water toward the left and increases the amount of undissociated water molecules. $[\text{OH}^-]$ increases and $[\text{H}^+]$ decreases.
- 39. List** the information needed to calculate the K_a of a weak acid.
- The pH, pOH, or $[\text{H}^+]$ and the initial concentration of the acid are needed to calculate K_a . K_b can also be used.
- 40. Calculate** The pH of a tomato is approximately 4.50. What are $[\text{H}^+]$ and $[\text{OH}^-]$ in a tomato?
- $$[\text{H}^+] = \text{antilog}(-4.50) = 3.2 \times 10^{-5}M$$
- $$\text{pOH} = 14.00 - 4.50 = 9.50$$
- $$[\text{OH}^-] = \text{antilog}(-9.50) = 3.2 \times 10^{-10}M$$
- 41. Determine** the pH of a solution that contains 1.0×10^{-9} mole of OH^- ions per liter.
- $$[\text{OH}^-] = \frac{1.0 \times 10^{-9} \text{ mol}}{1 \text{ L}} = 1.0 \times 10^{-9}M$$
- $$\text{pOH} = -\log 1.0 \times 10^{-9} = 9.00$$
- $$\text{pH} = 14.00 - \text{pOH}$$
- $$\text{pH} = 14.00 - 9.00 = 5.00$$
- 42. Calculate** the pH of the following solutions.
- a. 1.0M HI
- $$[\text{H}^+] = 1.0M$$
- $$\text{pH} = -\log[\text{H}^+] = -\log 1.0$$
- $$\text{pH} = 0.00$$
- b. 0.050M HNO_3
- $$[\text{H}^+] = 0.050M$$
- $$\text{pH} = -\log[\text{H}^+] = -\log 0.050$$
- $$\text{pH} = 1.30$$
- c. 1.0M KOH
- $$[\text{OH}^-] = 1.0M$$
- $$\text{pOH} = -\log[\text{OH}^-] = -\log 1.0$$
- $$\text{pOH} = 0.00$$
- $$\text{pH} = 14.00 - 0.00 = 14.00$$
- d. $2.4 \times 10^{-5}M$ $\text{Mg}(\text{OH})_2$
- $$[\text{OH}^-] = 2 \times [\text{Mg}(\text{OH})_2] = (2)(2.4 \times 10^{-5}M) = 4.8 \times 10^{-5}M$$
- $$\text{pOH} = -\log 4.8 \times 10^{-5} = 4.32$$
- $$\text{pH} = 14.00 - 4.32 = 9.68$$

- 43. Interpret Diagrams** Refer to **Figure 18.15** to answer these questions: What happens to the $[H^+]$, $[OH^-]$, pH, and pOH as a neutral solution becomes more acidic? As a neutral solution become more basic?

As the solution becomes more acidic, $[H^+]$ increases from 10^{-7} to 1, $[OH^-]$ decreases from 10^{-7} to 10^{-14} , pH changes from 7 to 0 and pOH changes from 7 to 14. As a neutral solution becomes more basic, $[H^+]$ decreases from 10^{-7} to 10^{-14} , $[OH^-]$ increases from 10^{-7} to 1, pH changes from 7 to 14 and pOH changes from 7 to 0.

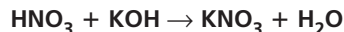
Section 18.4 Neutralization

pages 659–668

Practice Problems

pages 664–665

- 44.** What is the molarity of a nitric acid solution if 43.33 mL 0.1000M KOH solution is needed to neutralize 20.00 mL of the acid solution?



$$43.33 \text{ mL KOH} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.1000 \text{ mol KOH}}{1 \text{ L KOH}}$$

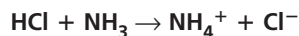
$$= 4.333 \times 10^{-3} \text{ mol KOH}$$

$$4.333 \times 10^{-3} \text{ mol KOH} \times \frac{1 \text{ mol HNO}_3}{1 \text{ mol KOH}}$$

$$= 4.333 \times 10^{-3} \text{ mol HNO}_3$$

$$M_{HNO_3} = \frac{4.333 \times 10^{-3} \text{ mol HNO}_3}{0.02000 \text{ L HNO}_3} = 0.2167M$$

- 45.** What is the concentration of a household ammonia cleaning solution if 49.90 mL of 0.5900M HCl is required to neutralize 25.00 mL of the solution?



$$49.90 \text{ mL HCl} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.5900 \text{ mol HCl}}{1 \text{ L HCl}}$$

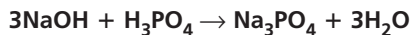
$$= 2.944 \times 10^{-2} \text{ mol HCl}$$

$$2.944 \times 10^{-2} \text{ mol HCl} \times \frac{1 \text{ mol NH}_3}{1 \text{ mol HCl}}$$

$$= 2.944 \times 10^{-2} \text{ mol NH}_3$$

$$M_{NH_3} = \frac{2.944 \times 10^{-2} \text{ mol NH}_3}{0.02500 \text{ L NH}_3} = 1.178M$$

- 46. Challenge** How many milliliters of 0.500M NaOH would neutralize 25.00 mL of 0.100M H_3PO_4 ?



$$\text{Mol } H^+ = \frac{0.100 \text{ mol } H_3PO_4}{1 \text{ L}} \times \frac{3 \text{ mol } H^+}{1 \text{ mol } H_3PO_4}$$

$$\times 25.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.00750 \text{ mol}$$

$$\text{mol } H^+ = \text{mol } OH^- = 0.00750 \text{ mol}$$

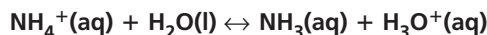
$$(L_{NaOH})(M_{NaOH}) = 0.00750 \text{ mol}$$

$$(L_{NaOH}) = \frac{(0.00750 \text{ mol})}{(0.500 \text{ mol/L})} = 0.0150 \text{ L}$$

$$0.0150 \text{ liter} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 15.0 \text{ mL NaOH}$$

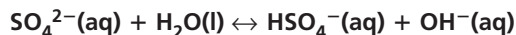
- 47.** Write equations for the salt hydrolysis reactions that occur when the following salts dissolve in water. Classify each solution as acidic, basic, or neutral.

- a. ammonium nitrate



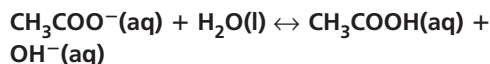
The solution is acidic.

- b. potassium sulfate



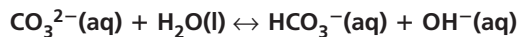
The solution is neutral.

- c. rubidium acetate



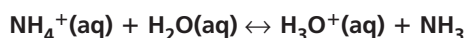
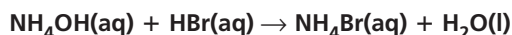
The solution is basic.

- d. calcium carbonate



The solution is basic.

48. Challenge Write the equation for the reaction that occurs in a titration of ammonium hydroxide (NH_4OH) with hydrogen bromide (HBr). Will the pH at the equivalence point be greater or less than 7?



Hydronium ions are formed so the pH will be less than 7.

Problems-Solving Lab

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1. Determine how many times greater the $[\text{H}^+]$ is if the blood's pH changes from 7.4 to 7.1.

At pH = 7.4:

$$[\text{H}^+] = \text{antilog}(-7.4) = 4.0 \times 10^{-8}\text{M}$$

At pH = 7.1

$$[\text{H}^+] = \text{antilog}(-7.1) = 7.9 \times 10^{-8}\text{M}$$

$$\frac{7.9 \times 10^{-8}}{4.0 \times 10^{-8}} = 2 \text{ times larger}$$

2. Suggest a reason why a 20:1 ratio of HCO_3^- to CO_2 in the blood is favorable for maintaining a healthy pH.

A healthy body will dump acid into the blood through increased activity. The excess hydrogen carbonate ion is available to neutralize the acid, thus driving the reaction toward the production of carbon dioxide.

3. Predict whether, for each situation, the pH of the blood will rise or fall, and which way the $\text{H}_2\text{CO}_3/\text{HCO}_3^-$ equilibrium will shift.

a. A person with severe stomach virus vomits many times during a 24-h period.

Vomit is acidic. Its removal from the body causes the pH to rise. The buffer reaction shifts to the right. The kidneys can respond by removing hydrogen carbonate ion and the person should be kept quiet to retain CO_2

b. To combat heartburn, a person foolishly takes too much (NaHCO_3).

The pH rises as the hydrogen carbonate ion levels increase. This shifts the buffer reaction to the left forming more CO_2 . The kidneys can respond by removing hydrogen carbonate ion and the person can breathe more rapidly to expel CO_2 .

Section 18.4 Assessment

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49. Explain why the net ionic equation for the neutralization reaction of any strong acid with any strong base is always the same.

After the elimination of spectator ions from the neutralization equation, each neutralization reaction is the reaction of one mole of hydrogen ion with one mole of hydroxide to form one mole of water.

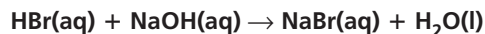
50. Explain the difference between the equivalence point and the end point of a titration.

Equivalence point is the pH at which the moles of H^+ ions from the acid equal the moles of OH^- ions from the base. The end point is the point at which the indicator used in a titration changes color.

51. Compare the results of two experiments: First, a small amount of base is added to an unbuffered solution with a pH of 7. Second, the same amount of base is added to a buffered solution with a pH of 7.

The pH of the unbuffered solution increases more than the pH of the buffered solution.

- 52. Calculate** the molarity of a solution of hydrobromic acid (HBr) if 30.35 mL of 0.1000M NaOH is required to titrate 25.00 mL of the acid to the equivalence point.



1 mol HBr reacts with 1 mol NaOH

$$\text{Mol of NaOH} = M_B \times V_B = 0.1000 \text{ mol/L} \times 0.03035 \text{ L}$$

$$\text{Mol of NaOH} = 0.003035 \text{ mol}$$

$$\text{Mol of HBr} = 0.003035 \text{ mol}$$

$$M_A \times V_A = 0.003035 \text{ mol}$$

$$(M_A)(0.02500 \text{ L}) = 0.003035$$

$$M_A = \frac{0.003035 \text{ mol}}{0.02500 \text{ L}} = 0.1214M$$

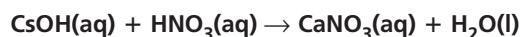
- 53. Interpret** What substances could be used to make a buffer solution with a pH of 9.4. How should the amounts of the substances be related? Use **Table 18.7**.

Use ammonia and a salt of ammonia such as ammonium nitrate or ammonium chloride. Use equal molar amounts of the acid and its salt.

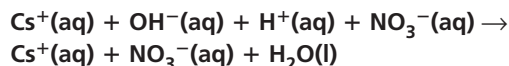
- 54. Design an Experiment** Describe how you would design and perform a titration in which you use 0.250M HNO₃ to determine the molarity of a cesium hydroxide solution. Include the formula and net ionic equations.

Place a measured volume of CsOH solution into a flask. Add an indicator such as bromthymol blue. Fill a buret with the 0.250M HNO₃ solution. Record the initial buret reading. Add HNO₃ solution slowly to the CsOH solution until the end point. Record the final buret reading. Calculate the volume of HNO₃ added. Use the volume and molarity of HNO₃ and the volume of CsOH to calculate the molarity of the CsOH solution. Refer to the solutions manual for the ionic equations.

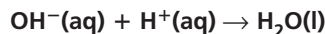
Formula equation:



Ionic equation:



Net ionic equation:



Everyday Chemistry

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Writing in Chemistry

Analyze If a recipe calls for flour, salt, sugar, bran, cereal, milk, an egg, and shortening or vegetable oil, would you use baking soda or baking powder? Explain. For more information about acids and bases in cooking, visit glencoe.com.

Baking soda should be used in the recipe. Milk is slightly acidic, with a pH of 6.5.

Chapter 18 Assessment

pages 672–675

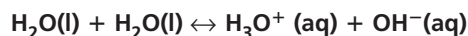
Section 18.1

Mastering Concepts

- 55.** In terms of ion concentrations, distinguish between acidic, neutral, and basic solutions.

An acidic solution has a larger concentration of H⁺ ions than OH⁻ ions. A basic solution has a larger concentration of OH⁻ ions than H⁺ ions. A neutral solution has equal concentrations of OH⁻ and H⁺ ions.

- 56.** Write a balanced chemical equation that represents the self-ionization of water.



- 57.** Classify each compound as an Arrhenius acid or an Arrhenius base.

a. H₂S
acid

b. RbOH
base

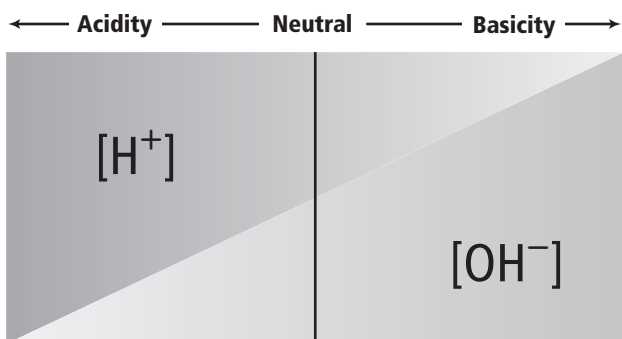
c. Mg(OH)₂
base

d. H₃PO₄
acid

- 58. Geology** When a geologist adds a few drops of HCl to a rock, gas bubbles form. What might the geologist conclude about the nature of the gas and the rock?

The gas is carbon dioxide, CO_2 ; the rock is limestone, or calcium carbonate, CaCO_3 .

- 59.** Explain the meaning of the relative sizes of the two shaded areas to the right of the dark vertical line in **Figure 18.27**.



The larger blue area means that basic solutions contain higher concentrations of hydroxide ions. The smaller red area means that basic solutions still contain hydrogen ions, but in lower concentrations than hydroxide ions.

- 60.** Explain the difference between a monoprotic acid, a diprotic acid, and a triprotic acid. Give an example of each.

A monoprotic acid can donate one H^+ ion, (HCl); a diprotic acid can donate two H^+ ions, (H_2SO_4); a triprotic acid can donate three H^+ ions, (H_3PO_4).

- 61.** Why can H^+ and H_3O^+ be used interchangeably in chemical equations?

H_3O^+ is a hydrated hydrogen ion.

- 62.** Use the symbols $<$, $>$, and $=$ to express the relationship between the concentrations of H^+ ions and OH^- ions in acidic, neutral, and basic solutions.

acidic: $[\text{H}^+] > [\text{OH}^-]$; neutral: $[\text{H}^+] = [\text{OH}^-]$; basic: $[\text{H}^+] < [\text{OH}^-]$

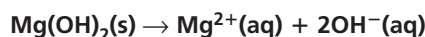
- 63.** Explain how the definition of a Lewis acid differs from the definition of a Brønsted-Lowry acid.

The Lewis model defines an acid as an electron pair acceptor, whereas in the Brønsted-Lowry model an acid is a hydrogen ion donor. A Lewis base is an electron pair donor, whereas a Brønsted-Lowry base is a hydrogen ion acceptor.

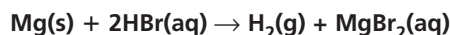
Mastering Problems

- 64.** Write a balanced chemical equation for each of the following.

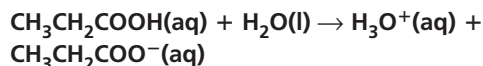
- a. the dissociation of solid magnesium hydroxide in water.



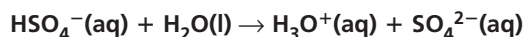
- b. the reaction of magnesium metal and hydrobromic acid.



- c. the ionization of propanoic acid ($\text{CH}_3\text{CH}_2\text{COOH}$) in water.



- d. the second ionization of sulfuric acid in water.



Section 18.2

Mastering Concepts

- 65.** Explain the difference between a strong acid and a weak acid.

In dilute aqueous solution, a strong acid ionizes completely; a weak acid ionizes slightly.

- 66.** Explain why equilibrium arrows are used in the ionization equations for some acids.

Equilibrium arrows are used for weak acids, which are partially ionized in water and create equilibria. Reaction arrows are used for strong acids, which may be considered to be 100% ionized in dilute aqueous solutions.

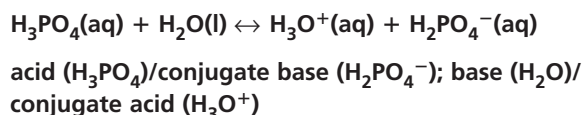
67. Which of the beakers in **Figure 18.29** might contain a solution of 0.1M hypochlorous acid? Explain your answer.

The beaker on the right might contain hypochlorous acid because hypochlorous acid is a weak acid, which means that it is only slightly ionized in aqueous solution and thus has low electrical conductivity.

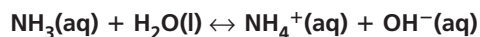
68. How would you compare the strengths of two weak acids experimentally? By looking up information in a table or a handbook?

You could compare the conductivities of equimolar solutions of the acids, and you could compare acid ionization constants of the two acids.

69. Identify the conjugate acid-base pairs in the reaction of H_3PO_4 with water.

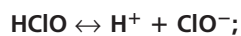


70. **Ammonia Cleaner** Write the chemical equation and K_b expression for the ionization of ammonia in water. How is it safe for a window washer to use a solution of ammonia, which is basic?



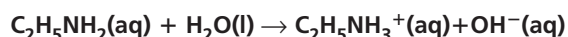
$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$; Ammonia is a weak base, therefore, its aqueous solution is not highly basic.

71. **Disinfectant** Hypochlorous acid is used as an industrial disinfectant. Write the chemical equation and the K_a expression for the ionization of hypochlorous acid in water.



$$K_a = \frac{[\text{H}^+][\text{ClO}^-]}{[\text{HClO}]}$$

72. Write the chemical equation and the K_b expression for the ionization of aniline in water. Aniline is a weak base with the formula $\text{C}_2\text{H}_5\text{NH}_2$.



$$K_b = \frac{[\text{C}_2\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{NH}_2]}$$

73. A fictional weak base, ZaH_2 , ionizes in water to yield a solution with a OH^- ion concentration of 2.68×10^{-4} mol/L. The chemical equation for the reaction is $\text{ZaH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \leftrightarrow \text{ZaH}_3^+(\text{aq}) + \text{OH}^-(\text{aq})$. If $[\text{ZaH}_2]$ at equilibrium is 0.0997 mol/L, what is the value of K_b for ZaH_2 ?

$$K_b = \frac{[\text{ZaH}_3^+][\text{OH}^-]}{[\text{ZaH}_2]} = \frac{(2.68 \times 10^{-4})(2.68 \times 10^{-4})}{0.0997 - 2.68 \times 10^{-4}} = 7.22 \times 10^{-7}$$

74. Select a strong acid, and explain how you would prepare a dilute solution of the acid. Select a weak acid, and explain how you would prepare a concentrated solution of the acid.

Students might say that the dilute solution of a strong acid is prepared by dissolving a small quantity of the strong acid in a large quantity of water, and the concentrated solution of a weak acid is prepared by dissolving a large quantity of the weak acid in a small quantity of water.

Section 18.3

Mastering Concepts

75. What is the relationship between the pOH and the OH^- ion concentration of a solution?

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

76. Solution A has a pH of 2.0. Solution B has a pH of 5.0. Which solution is more acidic? Based on the H^+ ion concentrations in the two solutions, how many times more acidic?

Solution A is more acidic than solution B. It is 10^3 , or 1000 times more acidic.

77. If the concentration of H^+ ions in an aqueous solution decreases, what must happen to the concentration of OH^- ions? Why?

$[\text{OH}^-]$ increases. $[\text{H}^+][\text{OH}^-] = K_w$

78. Use Le Chatelier's principle to explain what happens to the equilibrium $\text{H}_2\text{O}(\text{l}) \leftrightarrow \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$ when a few drops of HCl are added to pure water.

The HCl adds H^+ ions to the water, which causes the equilibrium to shift the left.

79. Common Acids and Bases Use the data in Table 18.8 to answer the following questions.

pH values	
Substance	pH
Household ammonia	11.3
Lemon juice	2.3
Antacid	9.4
Blood	7.4
Soft drinks	3.0

- a. Which substance is the most basic?
household ammonia
- b. Which substance is closest to neutral?
blood
- c. Which has a concentration of H^+
 $= 4.0 \times 10^{-10}M$?
antilog $(-9.4) = 4.0 \times 10^{-10}M$
antacid
- d. Which has a pOH of 11.0?
 $pH = 14.00 - 11.00 = 3.00$
soft drinks
- f. How many times more basic is antacid than blood?
100 times

Mastering Problems

80. What is $[OH^-]$ in an aqueous solution at 298 K in which $[H^+] = 5.40 \times 10^{-3}M$?

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

$$1.00 \times 10^{-14} = (5.40 \times 10^{-3})[OH^-]$$

$$[OH^-] = 1.85 \times 10^{-12}M$$

81. What are the pH and pOH for the solution described in Question 80?

$$pH = -\log [H^+] = -\log (5.40 \times 10^{-3}) = 2.27$$

$$pOH = -\log [OH^-] = -\log (1.85 \times 10^{-12}) = 11.7$$

82. If 5.00 mL of 6.00M HCl is added to 95.00 mL of pure water, the final volume is 100.00 mL. What is the pH of the resulting solution?

$$\text{mol HCl} = \text{mol } H^+ = 0.00500 \text{ L} \times 6.00 \text{ mol/L} = 0.0300 \text{ mol}$$

$$[H^+] = \frac{0.0300 \text{ Mol } H^+}{0.100 \text{ L}} = 0.300M$$

$$pH = -\log [H^+] = -\log (0.300) = 0.523$$

83. Given two solutions, 0.10M HCl and 0.10M HF, which solution has the greater concentration of H^+ ions? Calculate pH values for the two solutions, given that $[H^+] = 7.9 \times 10^{-3}M$ in the 0.10M HF.

$$\text{For } 0.10M \text{ HCl, a strong acid, } [H^+] = 0.10M$$

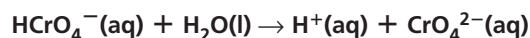
$$pH = -\log [H^+] = -\log 0.10 = 1.00$$

$$\text{For } 0.10M \text{ HF, } [H^+] = 7.9 \times 10^{-3}M$$

$$pH = -\log [H^+] = -\log 7.9 \times 10^{-3} = 2.10.$$

HCl has the greater concentration of hydrogen ions because it has the greater pH.

84. Metal Cleaner Chromic acid is used as an industrial cleaner for metals. What is K_a for the second ionization of chromic acid (H_2CrO_4) if a 0.040M solution of sodium hydrogen chromate has a pH of 3.946?



$$[H^+] = \text{antilog}(-pH)$$

$$[H^+] = \text{antilog}(-3.946) = 1.13 \times 10^{-4}M$$

$$[H^+] = [CrO_4^{2-}] = 1.13 \times 10^{-4}M$$

$$K_a = \frac{[H^+][CrO_4^{2-}]}{[HCrO_4^-]} = \frac{(1.13 \times 10^{-4})^2}{(0.040 - 1.13 \times 10^{-4})} = 3.2 \times 10^{-7}$$

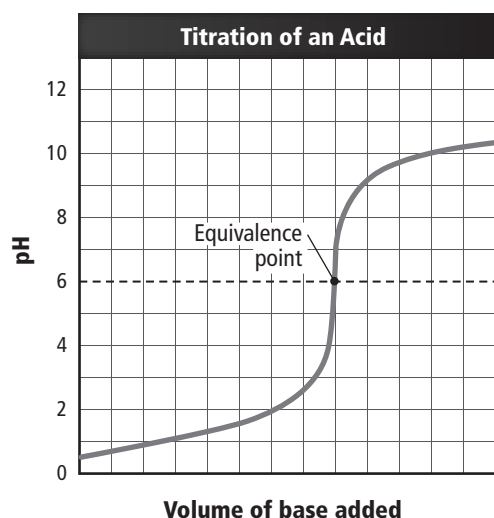
Section 18.4

Mastering Concepts

85. What acid and base must react to produce an aqueous sodium iodide solution?

Hydroiodic acid and sodium hydroxide must react.

86. What acid-base indicators, shown in Figure 18.24, would be suitable for the neutralization reaction whose titration curve is shown in Figure 18.30? Why?



Bromcresol purple or alizarin would be suitable because they would change color near a pH 6.0 equivalence point.

87. When might a pH meter be better than an indicator to determine the end point of an acid-base titration?

A pH meter could be used when there is no acid-base indicator that changes color at or near the equivalence point, or when such an indicator is not available

88. What happens when an acid is added to a solution containing the HF/F⁻ buffer system?

The acid produces hydrogen ions, which react with F⁻ ions in the solution to form HF molecules. The pH will drop slightly if the buffer is within its capacity.

89. When methyl red is added to an aqueous solution, a pink color results. When methyl orange is added to the same solution, a yellow color is produced. What is the approximate pH range of the solution? Use Figure 18.24.

The pH is between approximately 4.2 and 5.6.

90. Give the name and formula of the acid and the base from which each salt was formed.

a. NaCl

base: sodium hydroxide (NaOH); acid: hydrochloric acid (HCl)

b. KHCO₃

base: potassium hydroxide (KOH); acid: carbonic acid (H₂CO₃)

c. NH₄NO₂

base: ammonia (NH₃); acid nitrous acid (HNO₂)

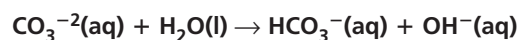
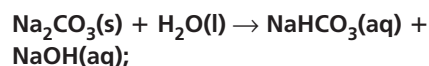
d. CaS

base: calcium hydroxide (Ca(OH)₂); acid hydrosulfuric acid (H₂S)

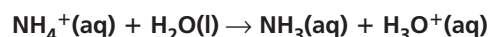
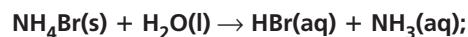
Mastering Problems

91. Write formula equations and net ionic equations for the hydrolysis of each salt in water.

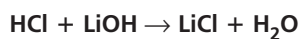
a. sodium carbonate



b. ammonium bromide



- 92. Air Purifier** Lithium hydroxide is used to purify air by removing carbon dioxide. A 25.00-mL sample of lithium hydroxide solution is titrated to an endpoint by 15.22 mL of 0.3340M hydrochloric acid solution. What is the molarity of the LiOH solution?

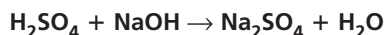


$$\text{mol HCl} = V_A \times M_A = 0.01522 \text{ L} \times 0.3340 \text{ mol/L} = 0.005083 \text{ mol}$$

$$\text{mol LiOH} = (\text{mol HCl}) \times \frac{1 \text{ mol LiOH}}{1 \text{ mol HCl}} = 0.005083 \text{ mol}$$

$$\text{LiOH } M = \frac{\text{mol LiOH}}{\text{vol LiOH}} = \frac{0.005083 \text{ mol}}{0.02500 \text{ L}} = 0.2033M$$

- 93.** In an acid-base titration, 45.78 mL of a sulfuric acid solution is titrated to the endpoint by 74.30 mL of 0.4388M sodium hydroxide solution. What is the molarity of the H₂SO₄ solution?



$$\text{mol NaOH} = V_B \times M_B = 0.07430 \text{ L} \times 0.4388M = 0.03260 \text{ mol}$$

$$\text{mol H}_2\text{SO}_4 = (\text{mol NaOH}) \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}}$$

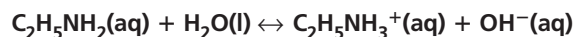
$$= \frac{0.03260}{2} = 0.01630 \text{ mol}$$

$$M_{\text{H}_2\text{SO}_4} = \frac{\text{mol H}_2\text{SO}_4}{\text{L H}_2\text{SO}_4} = \frac{0.01630 \text{ mol}}{45.78 \text{ mL}}$$

$$\times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.3561M$$

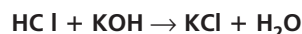
Mixed Review

- 94.** Write the equation for the ionization reaction and the base ionization constant expression for ethylamine (C₂H₅NH₂) in water.



$$K_b = \frac{[\text{C}_2\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{NH}_2]}$$

- 95.** How many milliliters of 0.225M HCl would be required to titrate 6.00 g of KOH?



$$6.00 \text{ g KOH} \times \frac{1 \text{ mol KOH}}{56.11 \text{ g KOH}} = 0.107 \text{ mol KOH}$$

$$0.107 \text{ mol KOH} \times \frac{1 \text{ mol HCl}}{1 \text{ mol KOH}} = 0.107 \text{ mol HCl}$$

$$0.107 \text{ mol HCl} \times \frac{1 \text{ L}}{0.225 \text{ mol HCl}} \times \frac{1000 \text{ mL HCl}}{1 \text{ L}} = 475 \text{ mL HCl}$$

96. What is the pH of a 0.200M solution of hypobromous acid (HBrO)?

$$K_a = 2.8 \times 10^{-9}$$

$$[\text{BrO}^-] = [\text{H}^+]; [\text{HBrO}] = 0.200M - [\text{H}^+]$$

Assume that because K_a is small, $[\text{H}^+]$ is insignificant compared to 0.200M. Thus, $[\text{HBrO}] = 0.200M$.

$$K_a = \frac{[\text{H}^+][\text{BrO}^-]}{[\text{HBrO}]}$$

$$\frac{[\text{H}^+]^2}{0.200} = 2.8 \times 10^{-9}$$

$$[\text{H}^+]^2 = 2.8 \times 10^{-9} \times 0.200$$

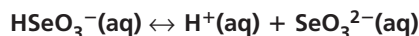
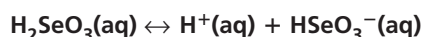
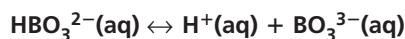
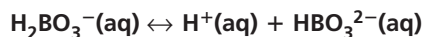
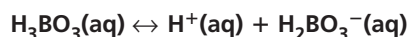
$$[\text{H}^+] = 2.4 \times 10^{-5}M$$

$$\text{pH} = -\log(2.43 \times 10^{-5}) = 4.63$$

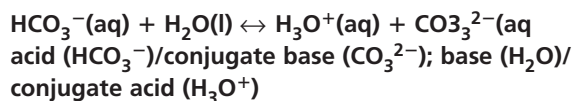
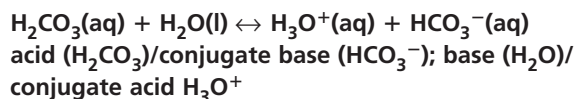
97. Which of the following are polyprotic acids? Write successive ionization equations for the polyprotic acids in water.



a and d are polyprotic acids.



98. Write balanced chemical equations for the two successive ionizations of carbonic acid in water. Identify the conjugate-base pair in each of the equations.



99. **Sugar Refining** Strontium hydroxide is used in the refining of beet sugar. Only 4.1 g of strontium hydroxide can be dissolved in 1 L of water at 273 K. Given that its solubility is so low, explain how it is possible that strontium hydroxide is considered to be a strong base.

All of the $\text{Sr}(\text{OH})_2$ that dissolves dissociates to form Sr^{2+} and OH^- ions.

100. What are the concentrations of OH^- ions in solutions having pH values of 3.00, 6.00, 9.00, and 12.00 at 298 K? What are the pOH values for the solutions?

$$\text{pH} + \text{pOH} = 14.00; \text{pOH} = 14.00 - \text{pH}; [\text{OH}^-] = \text{antilog}(-\text{pOH})$$

$$\text{pOH} = 14.00 - 3.00 = 11.00; [\text{OH}^-] = \text{antilog}(-11.00) = 1.0 \times 10^{-11}$$

$$\text{pOH} = 14.00 - 6.00 = 8.00; [\text{OH}^-] = \text{antilog}(-8.00) = 1.0 \times 10^{-8}$$

$$\text{pOH} = 14.00 - 9.00 = 5.00; [\text{OH}^-] = \text{antilog}(-5.00) = 1.0 \times 10^{-5}$$

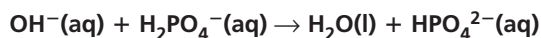
$$\text{pOH} = 14.00 - 12.00 = 2.00; [\text{OH}^-] = \text{antilog}(-2.00) = 1.0 \times 10^{-2}$$

101. The pH probe in Figure 18.31 is immersed in a 0.200M solution of a monoprotic acid, HA, at 303 K. What is the value of K_a for the acid at 303 K?

$$[\text{H}^+] = \text{antilog}(-\text{pH}) = \text{antilog}(-3.10) = 7.9 \times 10^{-4}M$$

$$K_a = \frac{(7.9 \times 10^{-4})(7.9 \times 10^{-4})}{(0.200 - 7.9 \times 10^{-4})} = 3.1 \times 10^{-6}$$

102. Write the chemical equation for the reaction that would occur when a base is added to a solution containing the $\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$ buffer system.



- 103.** An aqueous solution buffered by benzoic acid (C_6H_5COOH) and sodium benzoate (C_6H_5COONa) is $0.0500M$ in both compounds. Given that benzoic acid's $K_a = 6.4 \times 10^{-5}$, what is the pH of the solution?

$$K_a = 6.4 \times 10^{-5} = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

$$[H^+] = (6.4 \times 10^{-5}) \times \frac{[CH_3COOH]}{[CH_3COO^-]} = (6.4 \times 10^{-5}) \times \frac{0.0500}{0.0500} = 6.4 \times 10^{-5}M$$

$$pH = -\log [H^+] = -\log 6.4 \times 10^{-5}M = 4.19$$

Think Critically

- 104. Critique** the following statement: "A substance whose chemical formula contains a hydroxyl group must be considered to be a base."

The statement is misleading. If the substance dissociates or reacts with water to produce hydroxide ions in solution, it is considered a base. However, substances such as organic acids contain hydroxyl groups bonded in such a way that they donate hydrogen ions in water and produce acidic solutions.

- 105. Analyze and Conclude** Is it possible that an Arrhenius acid is not a Brønsted-Lowry acid? Is it possible that an acid according to the Brønsted-Lowry model is not an Arrhenius acid? Is it possible that a Lewis acid could not be classified as either an Arrhenius or a Brønsted-Lowry acid? Explain and give examples.

All Arrhenius acids are Brønsted-Lowry acids. Most Brønsted-Lowry acids are Arrhenius acids when in aqueous solution. Examples, HCl , H_2SO_4 , H_3PO_4 . Lewis acids are electron pair acceptors. Because the hydrogen ion is an electron pair acceptor, all Arrhenius and Brønsted-Lowry acids are Lewis acids. But some Lewis acids are not Arrhenius or Brønsted-Lowry acids, for example, BF_3

- 106. Apply Concepts** Use the ion product constant of water at 298 K to explain why a solution with a pH of 3.0 must have a pOH of 11.0.

A solution having a pH of 3.0 has a hydrogen ion concentration of $1.00 \times 10^{-3}M$. Substitute $1.00 \times 10^{-3}M$ for $[H^+]$ in the expression $K_w = [H^+][OH^-]$. $[OH^-] = 1.00 \times 10^{-11}M$; $pOH = 11.0$

- 107. Identify** the Lewis acids and bases in the following reactions.



a. Lewis acids: H^+ and H_2O ; Lewis base: OH^-

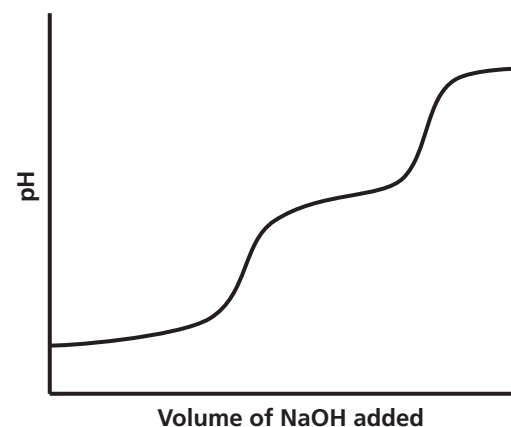


b. Lewis acid: BCl_3 ; Lewis bases: Cl^- , BCl_4^-



c. Lewis acid: SO_3 ; Lewis base: H_2O

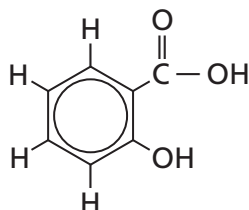
- 108. Interpret Scientific Illustrations** Sketch the shape of the approximate pH v. volume curve that would result from titrating a diprotic acid with a $0.10M$ NaOH solution.



- 109. Recognize Cause and Effect** Illustrate how a buffer works using the $C_2H_5NH_3^+ / C_2H_5NH_2$ buffer system. Show with equations how the weak base/conjugate acid system is affected when small amounts of acid and base are added to a solution containing this buffer system.

$C_2H_5NH_3^+(aq) \leftrightarrow H^+(aq) + C_2H_5NH_2(aq)$ When an acid is added, the buffer equilibrium shifts to the left. When a base is added, the added OH^- ions react with H^+ ions and the reaction shifts to the right.

- 110. Predict** Salicylic acid, shown in **Figure 18.32**, is used to manufacture acetylsalicylic acid, commonly known as aspirin. Evaluate the hydrogen atoms in the salicylic acid molecule based on your knowledge about the ionizable hydrogen in the ethanoic acid molecule, CH_3COOH . Predict which of salicylic acid's hydrogen atoms is likely to be ionizable.



Only the hydrogen atom in the COOH group is likely to be ionizable.

- 111. Apply Concepts** Like all equilibrium constants, K_w varies with temperature. K_w equals 2.92×10^{-15} at 10°C , 1.00×10^{-14} at 25°C , and 2.92×10^{-14} at 40°C . In light of this information, calculate and compare the pH values for pure water at these three temperatures. Based on your calculations, is it correct to say that the pH of pure water is always 7.0? Explain your answer.

$K_w = [\text{H}^+][\text{OH}^-]$; In pure water, $[\text{H}^+] = [\text{OH}^-]$ because each water molecule ionizes to produce one hydrogen ion and one hydroxide ion. For pure water, therefore, $K_w = [\text{H}^+]^2$

At 10°C :

$$2.92 \times 10^{-15} = [\text{H}^+]^2; [\text{H}^+] = 5.40 \times 10^{-8}; \text{pH} = -\log [\text{H}^+] = -\log (5.40 \times 10^{-8}) = 7.268$$

At 25°C :

$$1.00 \times 10^{-14} = [\text{H}^+]^2; [\text{H}^+] = 1.00 \times 10^{-7}; \text{pH} = -\log [\text{H}^+] = -\log (1.00 \times 10^{-7}) = 7.000$$

At 40°C :

$$2.92 \times 10^{-14} = [\text{H}^+]^2; [\text{H}^+] = 1.71 \times 10^{-7}; \text{pH} = -\log [\text{H}^+] = -\log (1.71 \times 10^{-7}) = 6.767$$

At 10°C , the pH of pure water is 7.268. At 25°C , pH = 6.998. At 40°C , pH = 6.767. It is incorrect to say that the pH of pure water is always 7.0; the pH of pure water is 7.0 only at 25°C , or 298 K.

Challenge Problem

- 112.** You have 20.0 mL of a solution of a weak acid, HX, whose $K_a = 2.14 \times 10^{-6}$. The pH of the solution is found to be 3.800. How much distilled water would you have to add to the solution to increase the pH to 4.000?

Original solution:

$$[\text{H}^+] = \text{antilog}(-\text{pH}) = \text{antilog}(-3.800) = 1.58 \times 10^{-4}$$

$$\frac{[\text{H}^+][\text{X}^-]}{[\text{HX}]} = \frac{(1.58 \times 10^{-4})^2}{M_{\text{initial}}} = 2.14 \times 10^{-6}$$

$$M_{\text{initial}} = \frac{(1.58 \times 10^{-4})^2}{(2.14 \times 10^{-6})} = 0.0117M$$

Diluted solution:

$$[\text{H}^+] = \text{antilog}(-\text{pH}) = \text{antilog}(-4.000) = 1.00 \times 10^{-4}$$

$$2.14 \times 10^{-6} \cong \frac{(1.00 \times 10^{-4})^2}{M_{\text{final}}}$$

$$M_{\text{final}} = \frac{(1.00 \times 10^{-4})^2}{2.14 \times 10^{-6}} = 0.00467M$$

Moles of HX in the initial and final solutions are equal.

$$(M_{\text{initial}})(V_{\text{initial}}) = (M_{\text{final}})(V_{\text{final}})$$

$$(0.0117M \times 20.00 \text{ mL}) = (0.00467M \times V_{\text{final}})$$

$$V_{\text{final}} = \frac{(0.0117 \text{ mol/L})(20.00 \text{ mL})}{0.00467 \text{ mol/L}}$$

Add 30.1 mL of distilled water to the original 20.0 mL.

Cumulative Review

- 113.** What factors determine whether a molecule is polar or nonpolar? (*Chapter 8*)

A molecule is polar if it contains bonds in which the electrons are not shared equally because of differences in the electronegativity of the bonding atom. Also the geometry of the molecule is important because the effects of polar bonds can be cancelled or enhanced depending on the arrangement of the bonds in the molecule.

- 114.** What property of some liquids accounts for the meniscus that forms at the surface of a solution in a buret. (*Chapter 12*)

Adhesion and cohesion forces are at work. Adhesion describes the force between molecules that are different. Cohesion describes the force between molecules that are the same. If the force of adhesion between a liquid and the glass of the buret are greater than the force between the molecules of the liquid, a meniscus forms.

- 115.** Which of the following physical processes are exothermic for water—freezing, boiling, condensing, subliming, evaporating? (*Chapter 12*)

freezing and condensing

- 116.** Explain why an air pump gets hot when you pump air into a bicycle's tires. (*Chapter 13*)

Gay-Lussac's law says that when volume is a constant, temperature increases with pressure.

- 117.** When 5.00 g of a compound was burned in a calorimeter, the temperature of 2.00 kg of water increased from 24.5°C to 40.5°C. How much heat would be released by the combustion of 1.00 mol of the compound (molar mass = 46.1 g/mol)? (*Chapter 15*)

$$q = c \times m \times \Delta T$$

$$\Delta T = 40.5^\circ\text{C} - 24.5^\circ\text{C} = 16.0^\circ\text{C}$$

$$2.00 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} = 2000 \text{ g H}_2\text{O}$$

$$5.00 \text{ g compound} \times \frac{1 \text{ ml compound}}{46.1 \text{ g compound}} = 0.108 \text{ mol compound}$$

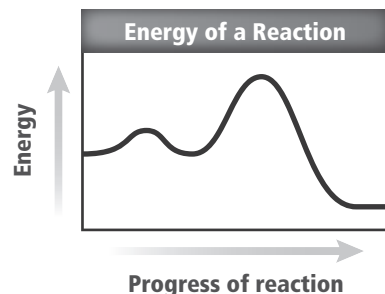
$$q = (4.184 \text{ J}/(\text{g} \cdot ^\circ\text{C}))(2.00 \times 10^3 \text{ g})(16.0^\circ\text{C}) = 1.34 \times 10^5 \text{ J}$$

$$\frac{1.34 \times 10^5 \text{ J}}{0.108 \text{ mol}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 1240 \text{ kJ/mol released}$$

- 118.** What is the difference between an exothermic and an endothermic reaction? (*Chapter 15*)

Energy is released in an exothermic reaction. Energy is absorbed in an endothermic reaction.

- 119.** Figure 18.33 shows how energy changes during the progress of a reaction.



- a.** Is the reaction exothermic or endothermic? (*Chapter 15*)

The reaction is exothermic because the reactants are at a higher energy than the products.

- b.** How many steps are in the reaction mechanism for the reaction? (*Chapter 16*)

two steps because the graph shows two activation energy barriers.

- c.** Explain how you could use the graph to identify the rate-determining step. (*Chapter 16*)

the rate-determining step is the one with the higher activation energy and therefore, the second step.

- 120.** Hydrogen and fluorine react to form HF according to the following equilibrium equation.



Will raising the temperature cause the amount of product to increase? Explain. (*Chapter 17*)

The reaction is exothermic because the sign of ΔH is negative. Therefore, heat is a product of the reaction. According to Le châtlier's principle, the addition of a product by, in this case, raising the temperature, will cause the equilibrium to shift to the left toward the reactants. Less product will be formed.

Additional Assessment

Writing in Chemistry

121. Acid/Base Theories Suppose that you are the Danish chemist Johannes Brønsted. The year is 1923, and you've formulated a new theory of acids and bases. Write a letter to the Swedish chemist Svante Arrhenius in which you discuss the differences between your theory and his and point out the advantages of yours.

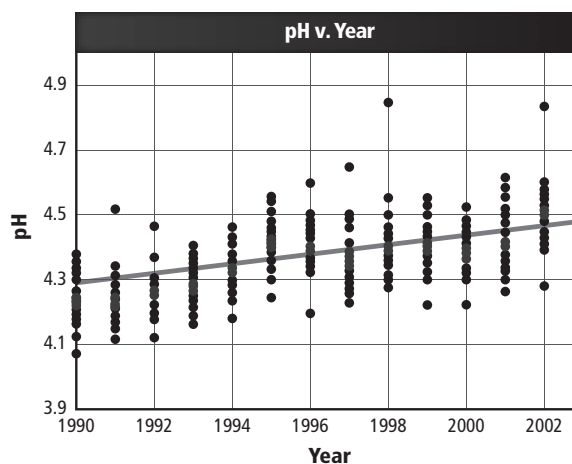
Students' letters should explain that the Brønsted-Lowry theory included all the acids and bases that were defined by the Arrhenius theory. The Brønsted theory went further by explaining why some substances such as ammonia produce basic solutions but do not contain a hydroxide ion in their structures. Brønsted's theory also explains the role of water and the hydronium ion in acidic and basic solutions.

122. Ammino Acids Twenty amino acids combine to form proteins in living systems. Research the structures and K_a values for five amino acids. Compare the strengths of these acids with the weak acids in Table 18-4.

Student answers will vary. For example, K_a for valine, whose structure is found on page 776, is 2.51×10^{-4} at 298 K.

Document-Based Question

Rainwater Figure 18.34 shows pH measurements made from a number of the monitoring sites in New York State. The pink dot represents the average of the measurement taken at all of the sites at a particular time.



123. In general, what is the trend in the average pH for the years 1990 to 2003?

The pH has increased gradually, from approximately 4.25 in 1990 to approximately 4.55 in 2003.

124. Calculate the $[H^+]$ for the lowest and the highest pH measurements recorded on the graph. How many times more acidic is the rainwater having the highest reading than the rainwater with the lowest?

Lowest pH (in 1990) = 4.08

$$[H^+] = \text{antilog}(-4.08) = 8.3 \times 10^{-5}M$$

Highest pH (in 1998) = 4.85

$$[H^+] = \text{antilog}(-4.85) = 1.4 \times 10^{-5}M$$

$$\frac{8.3 \times 10^{-5}}{1.4 \times 10^{-5}} = 5.9 \text{ times more acidic}$$

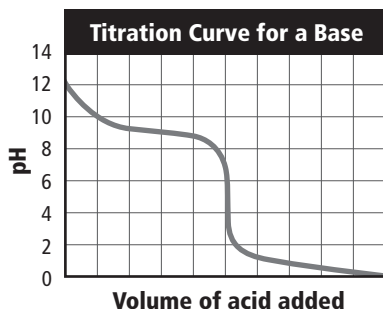
125. What is the pH of the trend line in 2003? How much has the average pH changed between 1990 and 2003?

The trend line passes through 4.48 in 2003. The average pH changed from 4.39 in 1990 to 4.48 in 2003, a change of 0.09.

Standardized Test Practice

pages 676–677

Use the graph below to answer Questions 1 and 2.



1. What is the pH at the equivalent point of this titration?

c

2. Which indicator would be effective for detecting the end point of this titration?

c

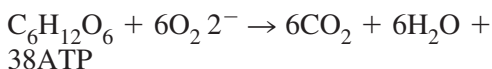
3. Hydrogen bromide (HBr) is a strong, highly corrosive acid. What is the pOH of a 0.0375M HBr?

a

$$\text{pH} = -\log(0.0375) = 1.430$$

$$\text{pOH} = 14 - 1.43 = 12.574$$

4. Cellular respiration produces about 38 mol of ATP for every mole of glucose consumed:



If each mole of ATP can release 30.5 kJ of energy, how much energy can be obtained from a candy bar containing 130.0 g of glucose?

b

$$130.0 \text{ g glu} \times (1 \text{ mol}/180.18 \text{ g}) \times (38 \text{ mol ATP}/1 \text{ mol glu}) \times (30.5 \text{ kJ}/1 \text{ mol ATP}) = 836 \text{ kJ}$$

Use the table below to answer Questions 5–7.

Ionization Constants and pH Data for Several Weak Organic Acids

Acid	pH of 1.000 M Solution	K_a
Formic	1.87	1.78×10^{-4}
Cyanoacetic	?	3.55×10^{-3}
Propanoic	2.43	?
Lutidinic	1.09	7.08×10^{-3}
Barbituric	2.01	9.77×10^{-5}

5. Which acid is the strongest?

- a. formic acid
b. cyanoacetic acid
c. lutidinic acid
d. barbituric acid

c

6. What is the acid dissociation constant of propanoic acid?

- a. 1.4×10^{-5}
b. 2.43×10^9
c. 3.72×10^{-3}
d. 7.3×10^4

c

7. What is the pH of a 0.40M solution of cyanoacetic acid?

- a. 2.06
b. 1.22
c. 2.45
d. 1.42

d

$$3.55 \times 10^{-3} \times 0.40 \text{ M} = [\text{H}^+]^2$$

$$[\text{H}^+] = 0.038 \text{ M}$$

$$\text{pH} = -\log(0.038) = 1.42$$

8. What does a value of K_{eq} greater than 1 mean?
- More reactants than products exist at equilibrium.
 - More products than reactants exist at equilibrium.
 - The rate of the forward reaction is high at equilibrium.
 - The rate of the reverse reaction is high at equilibrium.
- b
9. Magnesium sulfate ($MgSO_4$) is often added to water-insoluble liquid product of chemical reactions to remove unwanted water. $MgSO_4$ readily absorbs water to form two different hydrates. One of them is found to contain 13.0% H_2O and 87.0% $MgSO_4$. What is the name of this hydrate?
- magnesium sulfate monohydrate
 - magnesium sulfate dihydrate
 - magnesium sulfate hexahydrate
 - magnesium sulfate heptahydrate

a

Use the description of an experiment below to answer Questions 10-12.

Two 0.050-mol samples of gas at 20 degrees Celcius are released from the end of a long tube at the same time. One gas is xenon (Xe), and the other is sulfur dioxide (SO_2).

10. Explain which gas will have traveled farther after 5 seconds. How can you tell?

Sulfur dioxide will have traveled farther because SO_2 has the smaller mass. At 20°C both gases have the same kinetic energy which equals $\frac{1}{2}mv^2$. Thus, the gas with the smaller mass must have the greater velocity and travel farther in 5 seconds.

11. How will increasing the temperature of this experiment affect the rate of effusion of each gas?

Increasing the temperature increases the kinetic energy of the gases. The masses will not change but the velocities will and the rate of effusion will increase.

12. If the pressure on the xenon at the end of the experiment is 0.092 atm, what volume will it occupy?

$$PV = nRT$$

$$P = 0.092 \text{ atm}$$

$$T_K = 20^\circ\text{C} + 273 \text{ K} = 293 \text{ K}$$

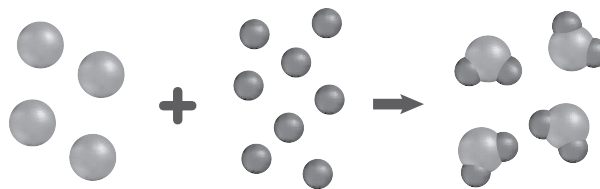
$$R = 0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K}$$

$$n = 0.050 \text{ mol}$$

$$V = \frac{nRT}{P} = \frac{(0.050 \text{ mol})(0.0821 \text{ L} \cdot \text{atm/mol} \cdot \text{K})(293 \text{ K})}{0.092 \text{ atm}}$$

$$V = 13 \text{ L}$$

Use the figure below to answer Question 13.



Atoms of Element A

Atoms of Element B

13. Explain how the chemical reaction shown in this figure demonstrates the law of conservation of mass.

All the atoms of element A are found in the molecules of product. Similarly, all the atoms of element B are found in the molecules of the product. Thus, no atoms have been lost and no atoms been gained.

- 14.** Describe lab procedures for preparing a 0.50M aqueous solution of NaOH and a 0.50 m aqueous solution of NaOH.

To prepare a 0.50M solution of NaOH, calculate the mass of NaOH corresponding to one mole. Add the mass to a 1-L volumetric flask partially filled with water. Swirl to dissolve the solid. Then, fill the flask to the zero mark with water.

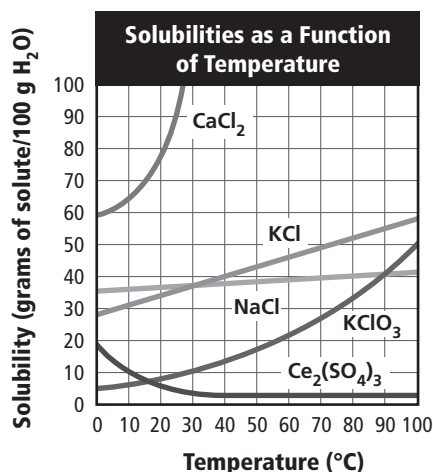
To prepare a 0.5 m solution of NaOH, measure 1000 mL of water, which corresponds to 1 kg. Add the mass of 0.50 mol of NaOH and swirl to dissolve.

- 15.** Explain how you could express the concentration of the 0.50 m solution in Question 14 as a mole fraction.

To convert the unit 0.050 m to mole fraction, calculate the number of moles of water contained in 1.00 kg of water. The mole fraction of NaOH is the moles of NaOH (0.050 mol) divided by the sum of the moles of water and the moles of NaOH. The mole fraction of water is the moles of water, divided by the sum of the moles of water and the moles of NaOH.

- 16.** Water has an unusually high boiling point compared to other compounds of similar molar mass because of
- hydrogen bonding.
 - adhesive forces.
 - covalent bonding.
 - dispersion forces.
 - pi bonds.
- a

Use the graph below to answer Questions 17 and 18.



- 17.** Which compound has a solubility of 38g/100 g H₂O at 50 degrees Celcius?
- CaCl₂
 - KCl
 - NaCl
 - KClO₃
 - Ce₂(SO₄)₃
- c
- 18.** Which has the greatest increase in solubility as temperature increases?
- Ce₂(SO₄)₃
 - CaCl₂
 - KClO₃
 - NaCl
 - KCl
- b