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Chapter 13
Acids and Bases

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Outline

1. Brønsted-Lowry acid-base model
2. The ion product of water
3. pH and pOH
4. Weak acids and their equilibrium constants
5. Weak bases and their equilibrium constants
6. Acid-base properties of salt solutions
7. Extending the concept of acids and bases: the Lewis model

Review from Chapter 4

- The Arrhenius definition of acid and base
  - Acids produce H^+ in water
  - Bases produce OH^- in water
  - H^+ from acids combines with OH^- from bases to produce water in a reaction called a neutralization

Brønsted-Lowry Acid-Base Model

- Brønsted-Lowry
  - Johannes Brønsted (1879-1947)
  - Thomas Lowry (1874-1936)
- Brønsted-Lowry model focuses on the ***reaction that takes place between acid and base***, rather than on the independent nature of the acid or base, as the Arrhenius model does
  - Acids donate H^+ to bases
  - Bases accept H^+ from acids

The Nature of H^+

- The H^+ ion is the medium of exchange in a Brønsted-Lowry reaction
  - H^+ can also be called a proton
  - ***Acid-base reactions involve proton exchange***

Conjugate Pairs

- The species that forms when a proton is removed from an acid is called the ***conjugate base*** of the acid
  - If the acid is HB, the conjugate base is B^-
- The only difference between the members of a conjugate acid-base pair is the position of the proton
- A species that can either accept or donate a proton is called amphiprotic
  - Consider water:
    - $OH^- \leftarrow H_2O \rightarrow H_3O^+$

      Remove H^+ Add H^+

Examples of Conjugate Acid-Base Pairs

Conjugate Acid	Conjugate Base
HF	F^-
HSO_4^-	SO_4^{2-}
NH_4^+	NH_3

Example 13.1, (Cont'd)

SOLUTION

(a) HNO_2 conjugate base	$HNO_2 \longrightarrow NO_2^{0-1} \longrightarrow NO_2^-$
F^- conjugate acid	$F^- \longrightarrow HF^{-1+1} \longrightarrow HF$
(b) HCO_3^- conjugate base	$HCO_3^- \longrightarrow CO_3^{-1-1} \longrightarrow CO_3^{2-}$
HCO_3^- conjugate acid	$HCO_3^- \longrightarrow H_2CO_3^{-1+1} \longrightarrow H_2CO_3$

The Hydronium Ion

- Another way to write the H^+ ion is as H_3O^+
  - H_3O^+ is the hydronium ion
    - H^+ exists in water as hydronium ion, since H^+ itself would not be stable in water
  - Depending on the reason for writing the reaction, either H^+ or H_3O^+ can be used, and interchangeably
  - The only difference is the inclusion or exclusion of the H_2O molecule

The Ion Product of Water

- Water can react with itself in a reaction called ***autoionization***
- Water can react with itself in an acid-base reaction:
  - $H_2O + H_2O \rightleftharpoons H_3O^+ (aq) + OH^- (aq)$
- An alternate way to write the reaction is:
  - $H_2O \rightleftharpoons H^+ (aq) + OH^- (aq)$

Example 13.1

EXAMPLE 13.1

(a) What is the conjugate base of HNO_2? The conjugate acid of F^-?

(b) The HCO_3^- ion, like the H_2O molecule, is amphiprotic. What is its conjugate base? Its conjugate acid?

STRATEGY

1. Form the conjugate base by removing one H atom. Decrease the charge by one unit (e.g., −1 to −2).
2. Form the conjugate acid by adding one H atom. Increase the charge by one unit (e.g., −1 to 0).

continued

Equilibrium and the Auto-Ionization of Water

- $H_2O \rightleftharpoons H^+ (aq) + OH^- (aq)$
  - Recall that concentrations can be used to write equilibrium constant expressions
  - K for this reaction is $[H^+][OH^-]$
  - This K is called the ***ion product constant of water, K_w***
    - $K_w = [H^+][OH^-]$
    - At 25 ºC, $K_w = 1.0 \times 10^{-14}$

Concentrations of H^+ and OH^- in pure water

- For water, $[H^+][OH^-] = 1.0 \times 10^{-14}$
- In pure water, the two concentrations are equal:
  - $[H^+] = 1.0 \times 10^{-7}$ M
  - $[OH^-] = 1.0 \times 10^{-7}$ M
- Since one concentration must rise as the other falls,
  - If $[H^+] > 1.0 \times 10^{-7}$ M, then $[OH^-] < 1.0 \times 10^{-7}$ M and the solution is acidic
  - If $[OH^-] > 1.0 \times 10^{-7}$ M, then $[H^+] < 1.0 \times 10^{-7}$ M and the solution is basic

Figure 13.2

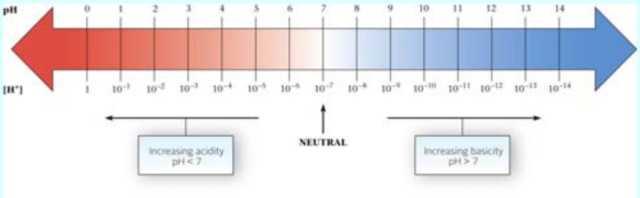
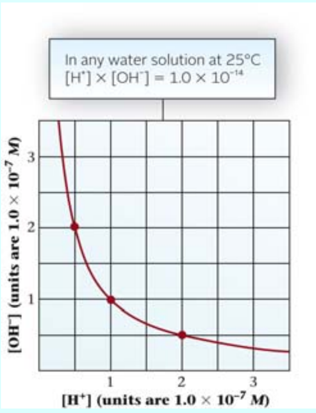


Figure 13.1



pOH

- Defining equation for pH

$$pOH = -\log[OH^-]$$
$$[OH^-] = \text{antilog}[-pOH] = 10^{-pOH}$$

- The higher the pOH, the more basic the solution
- The lower the pOH, the more acidic the solution

pH

- Defining equation for pH

$$pH = -\log[H^+]$$
$$[H^+] = anti\log(-pH) = 10^{-pH}$$

- The higher the pH, the less acidic the solution
- The lower the pH, the more acidic the solution

Relationship between pH and pOH

$$[H^+][OH^-] = 1.0 \times 10^{-14}$$
$$pH + pOH = 14.00$$

Table 13.1: pH of Some Common Materials

TABLE 13.1 pH of Some Common Materials

Lemon juice	2.2–2.4	Urine, human	4.8–8.4
Wine	2.8–3.8	Cow's milk	6.3–6.6
Vinegar	3.0	Saliva, human	6.5–7.5
Tomato juice	4.0	Drinking water	5.5–8.0
Beer	4–5	Blood, human	7.3–7.5
Cheese	4.8–6.4	Seawater	8.3

Example 13.2

EXAMPLE 13.2 GRADED

Calculate, at 25°C

(a) the $[H^+]$ and pH of a tapwater sample in which $[OH^-] = 2.0 \times 10^{-7}$.

(b) the $[H^+]$ and $[OH^-]$ of human blood at pH 7.40.

(c) the pOH of a solution in which $[H^+] = (5.0)[OH^-]$.

(a)

ANALYSIS

Information given:	$[OH^-]$ (2.0×10^{-7} *M*)
Information implied:	K_w value
Asked for:	$[H^+]$ and pH

STRATEGY

1. Substitute into Equation 13.1 to obtain $[H^+]$.
2. Substitute into Equation 13.3 to convert $[H^+]$ to pH.

SOLUTION

1. $[H^+]$	$K_w = 1.0 \times 10^{-14} = [H^+][OH^-] = [H^+](2.0 \times 10^{-7})$ $[H^+] = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-7}} = 5.0 \times 10^{-8}\,M$
2. pH	$pH = -\log_{10}[H^+] = -\log_{10}(5.0 \times 10^{-8}) = 7.30$

Example 13.2, (Cont'd)

(b)

ANALYSIS

Information given:	pH (7.40)
Information implied:	K_w value
Asked for:	$[H^+]$ and $[OH^-]$

STRATEGY

1. Substitute into Equation 13.3 to convert pH to $[H^+]$.
2. Substitute into Equation 13.1 to obtain $[OH^-]$.

SOLUTION

1. $[H^+]$	$pH = -\log_{10}[H^+] \longrightarrow 7.40 = -\log_{10}[H^+] \longrightarrow [H^+] = 10^{-7.40} = 4.0 \times 10^{-8}\,M$
2. $[OH^-]$	$1.0 \times 10^{-14} = [OH^-](4.0 \times 10^{-8}) \longrightarrow [OH^-] = 2.5 \times 10^{-7}\,M$

continued

Example 13.2, (Cont'd)

(c)

ANALYSIS

Information given:	$[H^+]$ and $[OH^-]$ relation ($[H^+] = 5.0\,[OH^-]$)
Information implied:	K_w value
Asked for:	pOH

STRATEGY

1. Substitute into Equation 13.1 to obtain $[OH^-]$.
2. Substitute into Equation 13.4 to convert $[OH^-]$ to pOH.

SOLUTION

1. $[OH^-]$	$1.0 \times 10^{-14} = [OH^-][H^+] = [OH^-](5.0[OH^-]) = 5.0[OH^-]^2$ $[OH^-]^2 = \frac{1.0 \times 10^{-14}}{5.0} \longrightarrow [OH^-] = 4.5 \times 10^{-8}\,M$
2. pOH	$pOH = -\log_{10}[OH^-] = -\log_{10}(4.5 \times 10^{-8}) = 7.35$

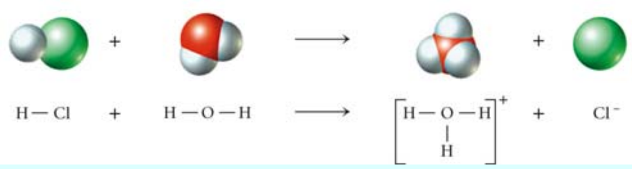
pH and Blood

- From the previous example, it is seen that the $[H^+]$ in blood is very small, about 4.0 X 10^{-8} M
  - Small changes in $[H^+]$ can have dramatic physiological effects
    - Many biological reactions depend on $[H^+]$
    - An increase in $[H^+]$ from 4.0 X 10^{-8} to 4.0 X 10^{-7} can increase the reaction rate by a power of 10
  - Small increases in $[H^+]$ can lead to ***acidosis***; small decreases in $[H^+]$ can lead to ***alkalosis***
  - ***Effective control of many physiological reactions depends on pH control***

pH of Strong Acids

- Recall from Chapter 4 that some acids are strong
  - HCl, HBr, HI, $HClO_4$, HNO_3, H_2SO_4
    - These completely ionize in water
    - $[H^+]$ is equal to the $[H^+]$ of the acid
  - A 0.10 M solution of HCl has $[H^+]$ = 0.10 M, so the pH of the solution is 1.00

Figure 13.3



Example 13.3, (Cont'd)

b STUDENT B

ANALYSIS

Information given:	pH (13.51) mass $Ba(OH)_2$ added by Student A (4.23 g); volume of solution (455 mL)
Information implied:	molar mass of $Ba(OH)_2$; K_w
Asked for:	mass $Ba(OH)_2$ added compared with Student A

continued

STRATEGY

1. The pathway to follow is the reverse of that in part (a):

$$pH \xrightarrow{Eq.\ 13.3} [H^+] \xrightarrow{K_w} [OH^-] \xrightarrow{2[OH^-]/[Ba(OH)_2]} [Ba(OH)_2] \xrightarrow{V} mol\ Ba(OH)_2 \xrightarrow{MM} mass$$

2. Compare masses used by Students A and B.

SOLUTION

$[OH^-]$	$[H^+] = 10^{-13.51} = 3.1 \times 10^{-14}\ M$; $[OH^-] = \frac{1.0 \times 10^{-14}}{3.1 \times 10^{-14}} = 0.32\ M$
Mass $Ba(OH)_2$ (Student B)	$\frac{0.32\ mol\ OH^-}{1\ L} \times \frac{1\ mol\ Ba(OH)_2}{2\ mol\ OH^-} \times 0.455\ L \times \frac{171.3\ g\ Ba(OH)_2}{1\ mol} = 12\ g$
Comparison	Student A: 4.32 g; Student B: 12 g 12 − 4.32 = 8 g more $Ba(OH)_2$ were added by Student B.

pH of Strong Bases

- Recall as well that some bases are strong:
  - LiOH, NaOH, KOH, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$
    - These bases ionize completely to OH^-
    - pOH is dependent on the concentration of the strong base
  - For an 0.10 M solution of NaOH,
    - $[Na^+] = [OH^-] = 0.10$ M
    - pOH = 1.00
    - pH = 13.00

Example 13.3, (Cont'd)

c STUDENT C

ANALYSIS

Information given:	from part (a): $[OH^-]$ due to $Ba(OH)_2$ (0.109 *M*); volume of solution (455 mL) mass of NaOH added (0.60 g)
Information implied:	molar mass of NaOH; K_w
Asked for:	pH of the solution

STRATEGY

1. Find moles OH^- contributed by $Ba(OH)_2$ from part (a).
2. Find moles OH^- contributed by NaOH.
3. Find $[OH^-]$ after NaOH addition.

$$\frac{(mol\ OH^-)_{NaOH} + (mol\ OH^-)_{Ba(OH)_2}}{V}$$

4. Find $[H^+]$ and pH.

SOLUTION

1. mol OH^- from $Ba(OH)_2$	$\frac{0.109\ mol\ OH^-}{1\ L} \times 0.455\ L = 0.0496$
2. mol OH^- from NaOH	$0.60\ g\ NaOH \times \frac{1\ mol}{40.0\ g} \times \frac{1\ mol\ OH^-}{1\ mol\ NaOH} = 0.015$
3. $[OH^-]$	$\frac{0.0496\ mol + 0.015\ mol}{0.455\ L} = 0.14\ M$
4. $[H^+]$ and pH	$[H^+] = \frac{1.0 \times 10^{-14}}{0.14} = 7.0 \times 10^{-14}\ M$ $pH = -\log_{10}(7.0 \times 10^{-14}) = 13.15$

Example 13.3

EXAMPLE 13.3 GRADED

Consider barium hydroxide, $Ba(OH)_2$, a white, powdery substance. Student A prepares a solution of $Ba(OH)_2$ by dissolving 4.23 g of $Ba(OH)_2$ in enough water to make 455 mL of solution.

a What is the pH of Student A's solution?

b Student B was asked to prepare the same solution as Student A. Student B's solution had a pH of 13.51. Did Student B add more or less $Ba(OH)_2$ to his solution? How much more or less $Ba(OH)_2$ was added?

c Student C was asked to add 0.60 g of NaOH to Student A's solution. What is the pH of Student C's solution? (Assume no volume change.)

a STUDENT A

ANALYSIS

Information given:	mass $Ba(OH)_2$ (4.23 g); volume of solution (455 mL)
Information implied:	molar mass of $Ba(OH)_2$; K_w
Asked for:	pH of the solution

STRATEGY

1. Start by expressing the concentration in g $Ba(OH)_2$/L of solution.
2. Follow the following pathway:

$$mass\ Ba(OH)_2/L \xrightarrow{MM} [Ba(OH)_2] \xrightarrow{2[OH^-]/[Ba(OH)_2]} [OH^-] \xrightarrow{K_w} [H^+] \xrightarrow{Eq.\ 13.3} pH$$

SOLUTION

$[OH^-]$	$\frac{4.23\ g\ Ba(OH)_2}{0.455\ L} \times \frac{1\ mol\ Ba(OH)_2}{171.3\ g} \times \frac{2\ mol\ OH^-}{1\ mol\ Ba(OH)_2} = 0.109\ mol/L = 0.109\ M$
$[H^+]$	$[H^+] = \frac{1.0 \times 10^{-14}}{0.109} = 9.2 \times 10^{-14}\ M$
pH	$pH = -\log_{10}(9.2 \times 10^{-14}) = 13.04$

Measuring pH

- pH can be measured with a pH meter
  - Translates $[H^+]$ into an electrical signal
  - Signal is shown on an analog or digital meter calibrated in pH units

Figure 13.4: pH of Carbonated Soda

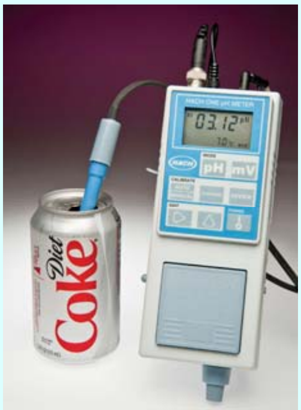


Figure 13.6

pH Indicators

- Universal indicator
  - Mixture of substances that change color depending on the concentration of H^+
  - Less accurate than pH meter
  - Depending on the indicator used, can display pH over a narrow or wide range of $[H^+]$
- Some plants can act as pH indicators
  - Color of some flowers in plants is dependent on the pH of the soil in which the plant is grown

Weak Acids and their Equilibrium Constants

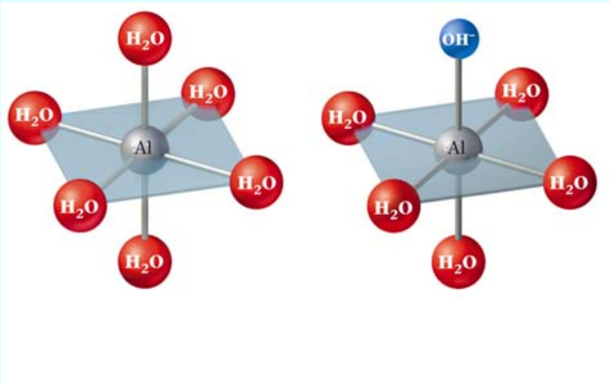
- Weak acids ionize only partially
- Prototype reaction
  - $HB\ (aq) + H_2O \rightleftharpoons H_3O^+\ (aq) + B^-\ (aq)$
- Two types of species that behave as weak acids
  1. Molecules with an ionizable hydrogen atom
     - $HNO_2\ (aq) + H_2O \rightleftharpoons H_3O^+\ (aq) + NO_2^-\ (aq)$
  2. Cations
     - $NH_4^+\ (aq) + H_2O \rightleftharpoons H_3O^+\ (aq) + NH_3\ (aq)$

Figure 13.5: Universal Indicator

Metal Cations as Acids

- Many metal cations act as weak acids in water solution as well
  - **$Zn(H_2O)_4^{2+}\ (aq) + H_2O \rightleftharpoons H_3O^+\ (aq) + Zn(H_2O)_3(OH)^+\ (aq)$**
  - The bond that forms between the oxygen and the metal ion weakens the O-H bond
  - H^+ is more easily ionized as a result of the weakened bond

Figure 13.7



pK_a

- $pK_a = -\log K_a$
- The smaller pK_a is, the stronger the acid
  - pK_a follows the trend for pH

Equilibrium Constants for Weak Acids

- HB (aq) + $H_2O \rightleftharpoons H_3O^+$ (aq) + B^- (aq)
  - K_a is the acid equilibrium constant
  - Simplifying the above to HB (aq) ⇔ H^+ (aq) + B^- (aq)

$$K_a = \frac{[H^+][B^-]}{[HB]}$$

- K_a values are related to the weak acid strength
  - The smaller K_a is, the weaker the acid is

Example 13.4

EXAMPLE 13.4

Consider acetic acid, $HC_2H_3O_2$, and the hydrated zinc cation, $Zn(H_2O)_4^{2+}$.

(a) Write equations to show why these species are acidic.

(b) Which is the stronger acid?

(c) What is the pK_a of $Zn(H_2O)_4^{2+}$?

(a)

STRATEGY AND SOLUTION

1. To prove that a species is acidic, you must produce a hydronium ion (H_3O^+) obtained by transferring an H atom to water.

transfers to H_2O

$HC_2H_3O_2(aq) + H_2O \rightleftharpoons C_2H_3O_2^-(aq) + HH_2O$ (H_3O^+)

$HC_2H_3O_2(aq) + H_2O \rightleftharpoons C_2H_3O_2^-(aq) + H_3O^+$

2. For the hydrated cation, one of the water molecules in the ion donates an H atom to an unattached water molecule. Think of $Zn(H_2O)_4^{2+}$ as $Zn(H_2O)(H_2O)_3^{2+}$.

transfers to H_2O

$Zn(H_2O)(H_2O)_3^{2+}(aq) + H_2O \rightleftharpoons Zn(OH^-)(H_2O)_3^{2+}(aq) + HH_2O$ ($Zn(OH)(H_2O)_3^+$) (H_3O^+)

$Zn(H_2O)(H_2O)_3^{2+}(aq) + H_2O \rightleftharpoons Zn(OH)(H_2O)_3^+(aq) + H_3O^+$

Table 13.2

TABLE 13.2 Equilibrium Constants for Weak Acids and Their Conjugate Bases

	Acid	K_a	Base	K_b
Sulfurous acid	H_2SO_3	1.7×10^{-2}	HSO_3^-	5.9×10^{-13}
Hydrogen sulfate ion	HSO_4^-	1.0×10^{-2}	SO_4^{2-}	1.0×10^{-12}
Phosphoric acid	H_3PO_4	7.1×10^{-3}	$H_2PO_4^-$	1.4×10^{-12}
Hexaaquairon(III) ion	$Fe(H_2O)_6^{3+}$	6.7×10^{-3}	$Fe(H_2O)_5OH^{2+}$	1.5×10^{-12}
Hydrofluoric acid	HF	6.9×10^{-4}	F^-	1.4×10^{-11}
Nitrous acid	HNO_2	6.0×10^{-4}	NO_2^-	1.7×10^{-11}
Formic acid	$HCHO_2$	1.9×10^{-4}	CHO_2^-	5.3×10^{-11}
Lactic acid	$HC_3H_5O_3$	1.4×10^{-4}	$C_3H_5O_3^-$	7.1×10^{-11}
Benzoic acid	$HC_7H_5O_2$	6.6×10^{-5}	$C_7H_5O_2^-$	1.5×10^{-10}
Acetic acid	$HC_2H_3O_2$	1.8×10^{-5}	$C_2H_3O_2^-$	5.6×10^{-10}
Hexaaquaaluminum(III) ion	$Al(H_2O)_6^{3+}$	1.2×10^{-5}	$Al(H_2O)_5OH^{2+}$	8.3×10^{-10}
Carbonic acid	H_2CO_3	4.4×10^{-7}	HCO_3^-	2.3×10^{-8}
Dihydrogen phosphate ion	$H_2PO_4^-$	6.2×10^{-8}	HPO_4^{2-}	1.6×10^{-7}
Hydrogen sulfite ion	HSO_3^-	6.0×10^{-8}	SO_3^{2-}	1.7×10^{-7}
Hypochlorous acid	HClO	2.8×10^{-8}	ClO^-	3.6×10^{-7}
Hydrocyanic acid	HCN	5.8×10^{-10}	CN^-	1.7×10^{-5}
Ammonium ion	NH_4^+	5.6×10^{-10}	NH_3	1.8×10^{-5}
Tetraaquazinc(II) ion	$Zn(H_2O)_4^{2+}$	3.3×10^{-10}	$Zn(H_2O)_3OH^+$	3.0×10^{-5}
Hydrogen carbonate ion	HCO_3^-	4.7×10^{-11}	CO_3^{2-}	2.1×10^{-4}
Hydrogen phosphate ion	HPO_4^{2-}	4.5×10^{-13}	PO_4^{3-}	2.2×10^{-2}

$HB(aq) \rightleftharpoons H^+(aq) + B^-(aq) \quad K_a = \frac{[H^+] \times [B^-]}{[HB]}$

$B^-(aq) + H_2O \rightleftharpoons HB(aq) + OH^-(aq) \quad K_b = \frac{[HB] \times [OH^-]}{[B^-]}$

Example 13.4, (Cont'd)

(b)

SOLUTION

K_a values from Table 13.2	$HC_2H_3O_2(K_a = 1.8 \times 10^{-5})$ vs $Zn(H_2O)_4^{2+}(K_a = 3.3 \times 10^{-10})$ $1.8 \times 10^{-5} > 3.3 \times 10^{-10}$ $HC_2H_3O_2$ is the stronger acid.

(c)

SOLUTION

pK_a	$pK_a = -\log_{10} K_a = -\log_{10}(3.3 \times 10^{-10}) = 9.48$

Example 13.5

EXAMPLE 13.5

Aspirin, a commonly used pain reliever, is a weak organic acid whose molecular formula may be written as $HC_9H_7O_4$. An aqueous solution of aspirin has total volume 350.0 mL and contains 1.26 g of aspirin. The pH of the solution is found to be 2.60. Calculate K_a for aspirin.

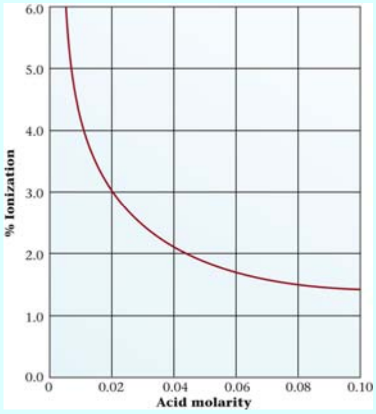
ANALYSIS

Information given:	molecular formula for aspirin ($HC_9H_7O_4$); mass of aspirin (1.26 g); volume of solution (350.0 mL); pH of solution (2.60)
Information implied:	molar mass of aspirin; $[H^+]$
Asked for:	K_a

STRATEGY

1. Determine the original concentration, $[\ ]_o$, of aspirin.
2. $pH = [H^+]_{eq}$
3. Draw a table as illustrated in Example 12.4. Substitute $[\ ]_o$ for P_o, $\Delta[\ ]$ for ΔP, and $[\ ]_{eq}$ for P_{eq}. Since only one H atom ionizes at a time, $\Delta[\ ]$ for all species is the same. Recall that $[\ ]$ stands for the concentration in molarity.
4. Write the K expression for the ionization and calculate K_a.

Figure 13.8 - % Ionization and Concentration



Example 13.5, (Cont'd)

SOLUTION

1. $[\ ]_o$ for aspirin: $\frac{1.26\text{ g}}{0.3500\text{ L}} \times \frac{1\text{ mol}}{180.15\text{ g}} = 0.0200\ M$
2. $[H^+]_{eq}$: $2.60 = -\log_{10}[H^+]$; $[H^+] = 10^{-2.60} = 2.5 \times 10^{-3}\ M$
3. Table

	$HC_9H_7O_4(aq)$ $\rightleftharpoons$	$H^+(aq)$ +	$C_9H_7O_4^-(aq)$
$[\ ]_o$	0.0200	0.0000	0.0000
$\Delta[\ ]$	−0.0025	+0.0025	+0.0025
$[\ ]_{eq}$	0.0175	0.0025	0.0025

4. K expression: $HC_9H_7O_4(aq) \rightleftharpoons H^+(aq) + C_9H_7O_4^-(aq)$

K_a: $K_a = \frac{[H^+][C_9H_7O_4^-]}{[HC_9H_7O_4]} = \frac{(0.0025)(0.0025)}{0.0175} = 3.6 \times 10^{-4}$

END POINT

Aspirin is a relatively *strong* weak acid. It would be located near the top of Table 13.2.

Example 13.6

EXAMPLE 13.6 CONCEPTUAL

The box below shows a system at equilibrium. It has a volume of 0.50 L and the symbol (red circle) represents 0.10 mol of a weak acid, HB. The symbol (blue circle) represents 0.10 mol of the conjugate base, B^-. Hydronium ions and water molecules are not shown. What is the percent ionization of the acid?

STRATEGY

1. Since there are 4 red circles (HB) and 1 blue circle (B^-) at equilibrium, there must have been 5 red circles to start with.
2. Substitute into Equation 13.7.

$$\%\text{ ionization} = \frac{[\ ]_{eq}}{[\ ]_o} \times 100$$

SOLUTION

% ionization: 1 blue circle $= [B^-]_{eq}$; 5 red circles $= [HB]_o$

$$\%\text{ ionization} = \frac{1}{5} \times 100\% = 20\%$$

Percent Ionization

- The percent ionization of a weak acid is defined as

$$\%\text{ionization} = \frac{[H^+]_{equilibrium}}{[HB]_{initial}} \times 100\%$$

- For the calculation in example 13.5, the percent ionization is about 12 %
  - Note that the percent ionization depends on the molarity of the weak acid

Calculating [H+] in a Water Solution of a Weak Acid

- We can use the process for calculating equilibrium pressure for gaseous reactions that we looked at in Chapter 12 to calculate the equilibrium concentration of $[H^+]$ for a weak acid
- The relationship between [HB], $[H^+]$ and $[B^-]$ is given in the equilibrium expression itself

Example 13.7

EXAMPLE 13.7

Nicotinic acid, $HC_6H_4O_2N$ ($K_a = 1.4 \times 10^{-5}$), is another name for niacin, an important member of the vitamin B group. Determine $[H^+]$ in a solution prepared by dissolving 3.0 g of nicotinic acid (MM = 123.11 g/mol), HNic, in enough water to form 245 mL of solution.

ANALYSIS

Information given:	molar mass for nicotinic acid, HNic (123.11 g/mol); mass of HNic (3.0 g) volume of solution (245 mL) K_a (1.4×10^{-5})
Information implied:	$[HNic]_o$
Asked for:	$[H^+]_{eq}$

STRATEGY

1. Determine the original concentration, $[\ ]_o$, of HNic.
2. Let $x = \Delta[H^+]$.
   Since all the coefficients in the reaction are one, $\Delta[HNic]$ and $\Delta[Nic^-]$ also equal x.
3. Draw a table as illustrated in Example 13.5.
4. Write the K expression for the ionization and subsitute the equilibrium concentrations for HNic, Nic^-, and H^+ obtained from the table.
5. Solve for x. (Assume $x << [Nic]_o$ to avoid the quadratic equation.)
6. Substitute the value for x in $[H^+]_{eq}$.

Algebra Review – Quadratic Equations

- Recall that for a quadratic equation in the form

$$ax^2 + bx + c = 0$$

- The roots are

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

Example 13.7, (Cont'd)

SOLUTION

1. $[\ ]_o$ for HNic: $\frac{3.0\ g}{0.245\ L} \times \frac{1\ mol}{123.11\ g} = 0.10\ M$
2. $\Delta[\ ]$: $\Delta[HNic] = \Delta[Nic^-] = \Delta[H^+] = x$
3. Table

	$HNic(aq)$	$\rightleftharpoons$	$H^+(aq)$	+	$Nic^-(aq)$
$[\ ]_o$	0.10		0.00		0.00
$\Delta[\ ]$	$-x$		$+x$		$+x$
$[\ ]_{eq}$	$0.10 - x$		x		x

4. K_a expression: $1.4 \times 10^{-5} = \frac{[H^+][Nic^-]}{[HNic]} = \frac{(x)(x)}{0.10 - x}$
5. Assume $x << 0.10$: $1.4 \times 10^{-5} = \frac{(x)(x)}{0.10} \longrightarrow x = 0.0012\ M$

continued

Approximations Used in Calculations

- The value of K_a is usually known no more accurately than about ±5%
- When solving for the unknowns used to work the equilibrium problem, for the expression

$$K_a = \frac{x^2}{a - x}$$

- Where a is the initial concentration of weak acid, you can neglect x in the denominator if doing so does not introduce an error of more than 5%, i.e.,

$$if\ \frac{x}{a} \le 0.05, then$$

$$a - x \approx a$$

Example 13.7, (Cont'd)

END POINT

Note that the concentration of H^+, 0.0012 M, is

- much smaller than the original concentration of the weak acid, 0.10 M. In this case, then, the approximation $0.10 - x \approx 0.10$ is justified. This will usually, but not always, be the case (see Example 13.8).
- much larger than $[H^+]$ in pure water, 1×10^{-7} M, justifying the assumption that the ionization of water can be neglected. *This will always be the case, provided $[H^+]$ from the weak acid is $\ge 10^{-6}$ M.*

Approximations and Percent Ionization

- When

$$\frac{x}{a} = \frac{[H^+]_{eq}}{[HB]_o}$$

- Multiplying by 100% will give the percent ionization:

$$\frac{x}{a}\% = \frac{[H^+]_{eq}}{[HB]_o} \times 100\%$$

- If the percent ionization is 5% or less, you may make the approximation.
- If the percent ionization is greater than 5%, the quadratic formula or the successive approximation method is required

Example 13.8

EXAMPLE 13.8

Calculate $[H^+]$ in a 0.100 M solution of nitrous acid, HNO_2, for which $K_a = 6.0 \times 10^{-4}$.

ANALYSIS	
Information given:	$[HNO_2]_o$ (0.100 M); K_a (6.0×10^{-4})
Asked for:	$[H^+] = [H^+]_{eq}$

continued

Polyprotic Weak Acids

- Acids containing more than one ionizable hydrogen are called polyprotic
  - The anion formed in one step produces another H^+ in a successive ionization step
  - The equilibrium constant becomes smaller with each successive step

Example 13.8, (Cont'd)

STRATEGY

1. The table setup is identical to that of Example 13.7 giving the equilibrium expression

$$K_a = \frac{[H^+][NO_2^-]}{[HNO_2]} = \frac{(x)(x)}{0.100 - x}$$

2. Assume $x << 0.100$ and solve for x.
3. Check your assumption by calculating % ionization.

   If the % ionization is less than 5%, your assumption is valid. If not, solve for x using the quadratic equation.

Triprotic Acid

- Phosphoric acid
  - $H_3PO_4\ (aq) \rightleftharpoons H^+\ (aq) + H_2PO_4^-\ (aq)$ K_{a1}
  - $H_2PO_4^-\ (aq) \rightleftharpoons H^+\ (aq) + HPO_4^{2-}\ (aq)$ K_{a2}
  - $HPO_4^{2-}\ (aq) \rightleftharpoons H^+\ (aq) + PO_4^{3-}\ (aq)$ K_{a3}

  - $K_{a1} > K_{a2} > K_{a3}$
  - With each successive step, ***the acid becomes progressively weaker***

Example 13.8, (Cont'd)

SOLUTION	
Equilibrium expression	$6.0 \times 10^{-4} = \frac{[H^+][NO_2^-]}{[HNO_2]} = \frac{(x)(x)}{0.100 - x}$
Assume $x << 0.100$.	$6.0 \times 10^{-4} = \frac{x^2}{0.100} \longrightarrow x = 0.0077\ M = [H^+]$
Check the assumption.	$\%\ \text{ionization} = \frac{[H^+]_{eq}}{[HB]_o} \times 100\% = \frac{0.0077}{0.100} \times 100\% = 7.7\%$
	$7.7\% > 5.0\% \longrightarrow$ The assumption is not valid.
Use the quadratic equation.	$x^2 + (6.0 \times 10^{-4})x - (6.0 \times 10^{-5}) = 0$
	$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-(6.0 \times 10^{-4}) \pm \sqrt{(6.0 \times 10^{-4})^2 - 4(6.0 \times 10^{-5})}}{2}$
	$x = 0.0074\ M$ or $-0.0080\ M$
	$-0.0080\ M$ is physically impossible, so $x = [H^+] = 0.0074\ M$

Table 13.3

TABLE 13.3 Equilibrium Constants for Some Weak Polyprotic Acids at 25°C

Acid	Formula	K_{a1}	K_{a2}	K_{a3}
Carbonic acid*	H_2CO_3	4.4×10^{-7}	4.7×10^{-11}	
Oxalic acid	$H_2C_2O_4$	5.9×10^{-2}	5.2×10^{-5}	
Phosphoric acid	H_3PO_4	7.1×10^{-3}	6.2×10^{-8}	4.5×10^{-13}
Sulfurous acid	H_2SO_3	1.7×10^{-2}	6.0×10^{-8}	

*Carbonic acid is a water solution of carbon dioxide:

$$CO_2(g) + H_2O \rightleftharpoons H_2CO_3(aq)$$

Example 13.9

EXAMPLE 13.9

The distilled water you use in the laboratory is slightly acidic because of dissolved CO_2, which reacts to form carbonic acid, H_2CO_3. Calculate the pH of a 0.0010 *M* solution of H_2CO_3 and $[CO_3^{2-}]$ at equilibrium.

ANALYSIS

Information given:	$[H_2CO_3]_o$ (0.0010 *M*)
Information implied:	two-step ionization, K_a value for each ionization (Table 13.3)
Asked for:	pH, $[CO_3^{2-}]_{eq}$

STRATEGY

1. Write the ionization reactions and the K_a expression for each ionization. Recall that ionization of polyprotic acids takes place one H^+ at a time and that most of the H^+ is obtained from the first ionization.
2. Find $[H^+]_{eq}$ for the first ionization and convert to pH.
3. Note $[H^+] = [HCO_3^-]$. Substitute into the expression for the second ionization to obtain $[CO_3^{2-}]_{eq}$.

SOLUTION

First ionization	$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq) \quad K_{a1} = 4.4 \times 10^{-7}$ $4.4 \times 10^{-7} = \frac{[H^+][HCO_3^-]}{[H_2CO_3]} = \frac{x^2}{0.0010 - x}$
Assume $x << 0.0010$.	$4.4 \times 10^{-7} = \frac{x^2}{0.0010} \longrightarrow x^2 = (0.0010)(4.4 \times 10^{-7}) \longrightarrow x = 2.1 \times 10^{-5}\ M$
Check assumption.	% ionization = 2.1%; assumption is valid.
pH	$[H^+] = [HCO_3^-] = 2.1 \times 10^{-5}\ M$; pH = 4.68

continued

Example 13.10

EXAMPLE 13.10

Write an equation to explain why each of the following produces a basic water solution.

(a) NO_2^- (b) Na_2CO_3 (c) $KHCO_3$

STRATEGY

1. React each basic anion with a water molecule.
2. The weak base picks up the proton (H^+) and increases its charge by one unit to create its conjugate acid.
3. OH^- is the other product of the reaction.

SOLUTION

(a) $NO_2^-(aq) + H_2O \rightarrow HNO_2^{-1+1}(aq) + OH^-(aq) \rightarrow HNO_2(aq) + OH^-(aq)$
HOH

(b) $CO_3^{2-}(aq) + H_2O \rightarrow HCO_3^{-2+1}(aq) + OH^-(aq) \rightarrow HCO_3^-(aq) + OH^-(aq)$
HOH

(c) $HCO_3^-(aq) + H_2O \rightarrow HHCO_3^{-1+1}(aq) + OH^-(aq) \rightarrow H_2CO_3(aq) + OH^-(aq)$
HOH

END POINT

The presence of OH^- as a product is the reason these anions in water are considered to be basic.

Example 13.9, (Cont'd)

Second ionization	$HCO_3^-(aq) \rightleftharpoons H^+(aq) + CO_3^{2-}(aq) \quad K_{a2} = 4.7 \times 10^{-11}$ $4.7 \times 10^{-11} = \frac{[H^+][CO_3^{2-}]}{[HCO_3^-]}$
$[CO_3^{2-}]$	From first ionization: $[H^+] = [HCO_3^-]$, so $K_{a2} = [CO_3^{2-}] = 4.7 \times 10^{-11}\ M$

Weak Base Equilibrium Constant

- NH_3 (aq) + $H_2O \rightleftharpoons NH_4^+$ (aq) + OH^- (aq)
  - The ***base equilibrium constant, K_b*** is

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

- For a generic weak base where
  - B (aq) + $H_2O \Leftrightarrow HB^+$ (aq) + OH^- (aq)

$$K_b = \frac{[HB^+][OH^-]}{[B]}$$

Weak Bases and their Equilibrium Expressions

- Types of weak bases
  - Molecules
    - Ammonia, NH_3, and amines
    - NH_3 (aq) + $H_2O \rightleftharpoons NH_4^+$ (aq) + OH^- (aq)
  - Anions
    - Anions derived from weak acids are weak bases
    - I^- (aq) + $H_2O \rightleftharpoons$ HI (aq) + OH^- (aq)

K_b by the Numbers

- As K_b becomes larger, base strength increases
- As with acids and K_a, we can define a pK_b:
  - $pK_b = -\log K_b$
  - As pK_b becomes smaller, base strength increases

Calculation of $[OH^-]$ in a Weak Base Solution

- The process of calculating the $[OH^-]$ in a weak base solution is the same as the process for calculating $[H^+]$ in a weak acid solution

Example 13.11

EXAMPLE 13.11 GRADED

Consider sodium hypochlorite, NaOCl, the main component in household bleach. The hypochlorite ion, OCl^-, has $K_b = 3.6 \times 10^{-7}$. A solution is prepared by dissolving 12.0 g of NaOCl (MM = 74.45 g/mol) in enough water to make 835 mL of solution.

(a) What is the pH of the solution?

(b) Household bleach is 5.25% NaOCl by mass. Assuming that its density is 1.00 g/mL, is household bleach more alkaline than the prepared solution?

(a)

ANALYSIS

Information given:	K_b for OCl^- (3.6×10^{-7}) mass of NaOCl (12.0 g); molar mass of NaOCl (74.45 g/mol) volume of solution (0.835 L)
Information implied:	K_w
Asked for:	pH of the solution

Example 13.11, (Cont'd)

STRATEGY

1. Determine the original concentration, $[\ ]_o$, of NaOCl.
2. Draw a table as illustrated in Example 13.5.
3. Write the K expression and find $[OH^-]$ assuming $[OH^-]_{eq} << [OCl^-]_o$. Check the validity of the assumption by finding % ionization.
4. Find $[H^+]$ using Equation 13.1 and pH using Equation 13.3.

SOLUTION

1. $[\ ]_o$ for NaOCl: $\frac{12.0\ g}{0.835\ L} \times \frac{1\ mol}{74.45\ g} = 0.193\ M$

2. Table

	$OCl^-(aq)$	$+ H_2O \rightleftharpoons$	$HOCl(aq)$	+	$OH^-(aq)$
$[\ ]_o$	0.193		0.00		0.00
$\Delta[\ ]$	$-x$		$+x$		$+x$
$[\ ]_{eq}$	$0.193 - x$		x		x

continued

Example 13.11, (Cont'd)

3. K expression	$K_b = \frac{[HOCl][OH^-]}{[OCl^-]} = 3.6 \times 10^{-7} = \frac{(x)(x)}{0.193 - x}$
Assume $x << 0.193$	$x^2 = 0.193(3.6 \times 10^{-7}) \longrightarrow x = 2.6 \times 10^{-4}$
Check assumption	% ionization = 0.14%; the assumption is justified. $[OH^-] = 2.6 \times 10^{-4}\ M$
4. $[H^+]$; pH	$1.0 \times 10^{-14} = [H^+](2.6 \times 10^{-4}) \longrightarrow [H^+] = 3.9 \times 10^{-11} \longrightarrow pH = 10.41$

(b)

ANALYSIS

Information given:	NaOCl content of household bleach (5.25% by mass) density of bleach (1.00 g/mL) pH of solution in part (a) (10.41)
Asked for:	Compare pH of solution (a) and pH of bleach.

Example 13.11, (Cont'd)

STRATEGY

1. Assume 100.0 g (= 100.0 mL) of bleach. Thus, there are 5.25 g of NaOCl in 100.0 mL of solution.
2. Find $[OH^-]$, $[H^+]$, and pH of bleach as in part (a).
3. Compare the pH of both solutions. The solution with a higher pH is more alkaline.

SOLUTION

$[NaOCl]_o = [OCl^-]_o$	$\frac{5.25\ g}{0.100\ L} \times \frac{1\ mol}{74.45\ g} = 0.705\ M$
K expression	(as in part (a)): $K_b = \frac{[HOCl][OH^-]}{[OCl^-]} = 3.6 \times 10^{-7} = \frac{(x)(x)}{0.705 - x}$
Assume $x << 0.705$	$x^2 = 0.705(3.6 \times 10^{-7}) \longrightarrow x = 5.0 \times 10^{-4}$
Check assumption.	% ionization = 0.071%; the assumption is justified. $[OH^-] = 5.0 \times 10^{-4}\ M$
$[H^+]$; pH	$1.0 \times 10^{-14} = [H^+](5.0 \times 10^{-4}) \longrightarrow [H^+] = 2.0 \times 10^{-11} \longrightarrow pH = 10.70$
Comparison	pH of the solution in part (a) = 10.41; pH of bleach = 10.70 10.70 > 10.41; bleach is more alkaline than the solution prepared in part (a).

Relation between K_a and K_b

- Consider the relation between a conjugate acid-base pair
  - $HB\ (aq) \rightleftharpoons H^+\ (aq) + B^-\ (aq)$ $K_I = K_a$ of HB
  - $B^-\ (aq) + H_2O \rightleftharpoons HB\ (aq) + OH^-\ (aq)$ $K_{II} = K_b$ of B^-
- These add to
  - $H_2O \rightleftharpoons H^+\ (aq) + OH^-\ (aq)$ $K_{III} = K_w$
- Since $K_I K_{II} = K_{III}$, $K_a K_b = K_w = 1.0 \times 10^{-14}$
  - for a conjugate acid base pair only
  - In log form, $pK_a + pK_b = pK_w = 14.00$

Notes on Acid-Base Strength

- K_a and K_b are inversely related
  - The larger K_a is, the smaller K_b is
- Features
  - Brønsted-Lowry acids
    - Strong acids
    - Weak acids
    - Acids weaker than water (conjugates of strong bases)
  - Brønsted-Lowry bases
    - Strong bases
    - Weak bases
    - Bases weaker than water (conjugates of strong acids)

Acid-Base Properties of Solutions of Salts

- A salt is an ionic solid containing a cation other than H^+ and an anion other than OH^- or O^{2-}
- We can predict whether a salt will be acidic, basic or neutral by
  1. Deciding what effect the cation has on water
     - Is it acidic or is it neutral?
  2. Deciding what effect the anion has on water
     - Is it basic or is it neutral?
  3. Combining the two effects to decide the behavior of the salt in water

Table 13.4

TABLE 13.4 Relative Strengths of Brønsted-Lowry Acids and Bases

K_a	Conjugate Acid	Conjugate Base	K_b
Very large	$HClO_4$	ClO_4^-	Very small
Very large	HCl	Cl^-	Very small
Very large	HNO_3	NO_3^-	Very small
	H_3O^+	**H_2O**	
6.9×10^{-4}	HF	F^-	1.4×10^{-11}
1.8×10^{-5}	$HC_2H_3O_2$	$C_2H_3O_2^-$	5.6×10^{-10}
1.2×10^{-5}	$Al(H_2O)_6^{3+}$	$Al(H_2O)_5(OH)^{2+}$	8.3×10^{-10}
4.4×10^{-7}	H_2CO_3	HCO_3^-	2.3×10^{-8}
2.8×10^{-8}	$HClO$	ClO^-	3.6×10^{-7}
5.6×10^{-10}	NH_4^+	NH_3	1.8×10^{-5}
4.7×10^{-11}	HCO_3^-	CO_3^{2-}	2.1×10^{-4}
	H_2O	**OH^-**	
Very small	C_2H_5OH	$C_2H_5O^-$	Very large
Very small	OH^-	O^{2-}	Very large
Very small	H_2	H^-	Very large

Cations

- Weak acid or spectator ion?
  - Most cations are acidic
    - These will change the pH by more than 0.5 pH units in a 0.1 M solution
  - Exceptions – these are spectators
    - Alkali metal cations
    - Heavier alkaline earth cations (Ca^{2+}, Sr^{2+}, Ba^{2+})

Hydride ion

- Reaction of water with CaH_2
  - H^- is the conjugate base of H_2, a very weak acid
  - As a result, H^- is an extremely strong base

Anions

- Weak base or spectator ion?
  - Many anions are weak bases
    - These will change the pH by more than 0.5 pH units at 0.1 M
  - Exceptions – these are spectators
    - Anions of very strong acids: Cl^-, Br^-, I^-, NO_3^-, ClO_4^-

Sodium Chloride Solution - Neutral

Amphiprotic Anions

- HCO_3^-
  - $K_a = 4.7 \times 10^{-11}$
  - $K_b = 2.3 \times 10^{-8}$
- Because $K_b > K_a$, a solution of $NaHCO_3$ will be basic

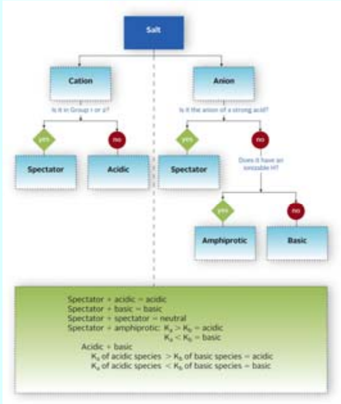
Table 13.5

TABLE 13.5 Acid-Base Properties of Ions* in Water Solution

	Spectator		Basic		Acidic	
Anion	Cl^- Br^- I^-	NO_3^- ClO_4^-	$C_2H_3O_2^-$ F^-	CO_3^{2-} PO_4^{3-} Many others		
Cation	Li^+ Na^+ K^+	Ca^{2+} Sr^{2+} Ba^{2+}			NH_4^+ Mg^{2+} Transition metal ions	Al^{3+}

*For the acid-base properties of amphiprotic anions such as HCO_3^- or $H_2PO_4^-$, see the discussion at the end of this section.

Figure 13.9: Flowchart for Acid-Base Properties of Salts



Salts: Acidic, Basic or Neutral

- By comparing the K_a of an acidic cation with the K_b of a basic anion, the salt of both can be classified as acidic, basic or neutral
  - If $K_a > K_b$, the salt is acidic
    - NH_4F, $K_a = 5.6 \times 10^{-10}$; $K_b = 1.4 \times 10^{-11}$
  - If $K_b > K_a$, the salt is basic
    - NH_4ClO, $K_a = 5.6 \times 10^{-10}$; $K_b = 3.6 \times 10^{-7}$

Example 13.12

EXAMPLE 13.12

Consider aqueous solutions of the following salts:

(a) $Zn(NO_3)_2$ (b) $KClO_4$ (c) Na_3PO_4 (d) NH_4F (e) $NaHCO_3$

Which of theses solutions are acidic? basic? neutral?

STRATEGY

Follow the flowchart in Figure 13.9.

Example 13.12, (Cont'd)

SOLUTION

(a) cation: Zn^{2+}	not in group 1 or 2 ⟶ acidic
anion: NO_3^-	anion of a strong acid ⟶ spectator
$Zn(NO_3)_2$	acidic + spectator ⟶ acidic
(b) cation: K^+	group 1 ⟶ spectator
anion: ClO_4^-	anion of a strong acid ⟶ spectator
$KClO_4$	spectator + spectator ⟶ neutral
(c) cation: Na^+	group 1 ⟶ spectator
anion: PO_4^{3-}	not an anion of a strong acid; no ionizable H ⟶ basic
Na_3PO_4	spectator + basic ⟶ basic
(d) cation: NH_4^+	not in group 1 or 2 ⟶ acidic
anion: F^-	not an anion of a strong acid; no ionizable H ⟶ basic
NH_4F	acidic + basic ⟶ Check K_a for NH_4^+ and K_b for F^-
	K_a for $NH_4^+ = 5.6 \times 10^{-10}$; K_b for $F^- = 1.4 \times 10^{-11}$
	$K_a > K_b$ ⟶ NH_4F is acidic.
(e) cation: Na^+	group 1 ⟶ spectator
anion: HCO_3^-	not an anion of a strong acid; has an ionizable H ⟶ amphiprotic
$NaHCO_3$	spectator + amphiprotic ⟶ Check K_a and K_b for HCO_3^-
	K_a for $HCO_3^- = 4.7 \times 10^{-11}$; K_b for $HCO_3^- = 2.3 \times 10^{-8}$
	$K_a < K_b$ ⟶ $NaHCO_3$ is basic.

Lewis Acids

- The concept of a Lewis acid extends the acid-base model
  - The Lewis model greatly expands the number of species considered to be acids
  - For example, metal cations are *not* Brønsted-Lowry acids but are Lewis acids

Extending the Concept of Acids and Bases

- The Lewis Model
  - A Lewis base donates a pair of electrons
  - A Lewis acid accepts a pair of electrons

Table 13.6

TABLE 13.6 Alternative Definitions of Acids and Bases

Model	Acid	Base
Arrhenius	Supplies H^+ to water	Supplies OH^- to water
Brønsted-Lowry	H^+ donor	H^+ acceptor
Lewis	Electron pair acceptor	Electron pair donor

Lewis Bases

- The concept of a Lewis base does not structurally differ from that of a Brønsted-Lowry base
  - For a species to accept a proton, it must contain an atom that possesses a lone pair
  - Lewis bases are also Brønsted-Lowry bases

Key Concepts

1. Classify a substance as a Brønsted-Lowry acid or base and write the net ionic equation to support the classification
2. Given $[H^+]$, $[OH^-]$, pH or pOH, calculate the three other quantities
3. Given the pH and original concentration of a weak acid, calculate K_a
4. Given the K_a and original concentration of a weak acid, calculate $[H^+]$
5. Given the K_b and original concentration of a weak base, calculate $[OH^-]$

Key Concepts, (Cont'd)

6. Given K_a for a weak acid, calculate K_b for its conjugate base (or vice-versa).
7. Predict whether a salt will be acidic, basic or neutral.
8. Understand the similarities and differences between Lewis and Brønsted-Lowry acids and bases