5.E: Acid-Base Equilibria (Exercises)

5.E.1: Brønsted-Lowry Acids and Bases

1. Write equations that show NH_3 as both a conjugate acid and a conjugate base.

Answer

One example for NH_3 as a conjugate acid: $NH_2^- + H^+ \longrightarrow NH_3$; as a conjugate base: $NH_4^+(aq) + OH^-(aq) \longrightarrow NH_3(aq) + H_2O(l)$

2. Write equations that show $H_2PO_4^-$ acting both as an acid and as a base.

3. Show by suitable net ionic equations that each of the following species can act as a Brønsted-Lowry acid:

a. H_3O^+
b. HCl
c. NH_3
d. CH_3CO_2H
e. NH_4^+
f. HSO_4^-

Answer

a. $H_3O^+(aq) \longrightarrow H^+(aq) + H_2O_{(l)}$;
b. $HCl(l) \longrightarrow H^+(aq) + Cl^-(aq)$;
c. $NH_3(aq) \longrightarrow H^+(aq) + NH_2^-(aq)$;
d. $CH_3CO_2H(aq) \longrightarrow H^+(aq) + CH_3CO_2^-(aq)$;
e. $NH_4^+(aq) \longrightarrow H^+(aq) + NH_3(aq)$;
f. $HSO_4^-(aq) \longrightarrow H^+(aq) + SO_4^{2-}(aq)$

4. Show by suitable net ionic equations that each of the following species can act as a Brønsted-Lowry acid:

a. HNO_3
b. PH_4^+
c. H_2S
d. CH_3CH_2COOH
e. $H_2PO_4^-$
f. HS^-

5. Show by suitable net ionic equations that each of the following species can act as a Brønsted-Lowry base:

a. H_2O
b. OH^-
c. NH_3
d. CN^-
e. S^{2-}
f. $H_2PO_4^-$

Answer

a. $H_2O_{(l)} + H^+(aq) \longrightarrow H_3O^+(aq)$
b. $OH^-(aq) + H^+(aq) \longrightarrow H_2O_{(l)}$
c. $NH_3(aq) + H^+(aq) \longrightarrow NH_4^+(aq)$;
d. $CN^-(aq) + H^+(aq) \longrightarrow HCN(aq)$
e. $S^{2-}(aq) + H^+(aq) \longrightarrow HS^-(aq)$
f. $H_2PO_4^-(aq) + H^+(aq) \longrightarrow H_3PO_4(aq)$

6. Show by suitable net ionic equations that each of the following species can act as a Brønsted-Lowry base:

a. HS^-
b. PO_4^{3-}
c. NH_2^-
d. C_2H_5OH
e. O^{2-}
f. $H_2PO_4^-$

7. What is the conjugate acid of each of the following? What is the conjugate base of each?

a. OH^-
b. H_2O
c. HCO_3^-
d. NH_3
e. HSO_4^-
f. H_2O_2
g. HS^-
h. $H_5N_2^+$

Answer

H_2O, O^{2-}; H_3O^+, OH^- ; H_2CO_3, CO_3^{2-}; NH_4^+, NH_2^-; H_2SO_4, SO_4^{2-}; $H_3O_2^+$, HO_2^-; H_2S; S^{2-}; $H_6N_2^{2+}$, H_4N_2

8. What is the conjugate acid of each of the following? What is the conjugate base of each?

a. H_2S
b. $H_2PO_4^-$
c. PH_3
d. HS^-
e. HSO_3^-
f. $H_3O_2^+$
g. H_4N_2
h. CH_3OH

9. Identify and label the Brønsted-Lowry acid, its conjugate base, the Brønsted-Lowry base, and its conjugate acid in each of the following equations:

a. $HNO_3 + H_2O \longrightarrow H_3O^+ + NO_3^-$
b. $CN^- + H_2O \longrightarrow HCN + OH^-$
c. $H_2SO_4 + Cl^- \longrightarrow HCl + HSO_4^-$
d. $HSO_4^- + OH^- \longrightarrow SO_4^{2-} + H_2O$
e. $O^{2-} + H_2O \longrightarrow 2\,OH^-$
f. $[Cu(H_2O)_3(OH)]^+ + [Al(H_2O)_6]^{3+} \longrightarrow [Cu(H_2O)_4]^{2+} + [Al(H_2O)_5(OH)]^{2+}$
g. $H_2S + NH_2^- \longrightarrow HS^- + NH_3$

Answer

The labels are Brønsted-Lowry acid = BA; its conjugate base = CB; Brønsted-Lowry base = BB; its conjugate acid = CA. HNO_3(BA), H_2O(BB), H_3O^+(CA), NO_3^-(CB); CN^-(BB), H_2O(BA), HCN(CA), OH^- (CB); H_2SO_4(BA), Cl^-(BB), HCl(CA), HSO_4^-(CB); HSO_4^-(BA), OH^-(BB), SO_4^{2-}(CB), H_2O(CA); O^{2-}(BB), H2O(BA) OH^- (CB and CA); $[Cu(H_2O)_3(OH)]^+$ (BB), $[Al(H_2O)_6]^{3+}$(BA), $[Cu(H_2O)_4]^{2+}$(CA), $[Al(H_2O)_5(OH)]^{2+}$(CB); H_2S(BA), NH_2^-(BB), HS^-(CB), NH_3(CA)

10. Identify and label the Brønsted-Lowry acid, its conjugate base, the Brønsted-Lowry base, and its conjugate acid in each of the following equations:

a. $NO_2^- + H_2O \longrightarrow HNO_2 + OH^-$
b. $HBr + H_2O \longrightarrow H_3O^+ + Br^-$
c. $HS^- + H_2O \longrightarrow H_2S + OH^-$

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d. $H_2PO_4^- + OH^- \longrightarrow HPO_4^{2-} + H_2O$
e. $H_2PO_4^- + HCl \longrightarrow H_3PO_4 + Cl^-$
f. $[Fe(H_2O)_5(OH)]^{2+} + [Al(H_2O)_6]^{3+} \longrightarrow [Fe(H_2O)_6]^{3+} + [Al(H_2O)_5(OH)]^{2+}$
g. $CH_3OH + H^- \longrightarrow CH_3O^- + H_2$

11. What are amphiprotic species? Illustrate with suitable equations.

Answer

Amphiprotic species may either gain or lose a proton in a chemical reaction, thus acting as a base or an acid. An example is H_2O.

As an acid: $H_2O(aq) + NH_3(aq) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$.

As a base: $H_2O(aq) + HCl(aq) \rightleftharpoons H_3O^+(aq) + Cl^-(aq)$

12. State which of the following species are amphiprotic and write chemical equations illustrating the amphiprotic character of these species:

a. H_2O
b. $H_2PO_4^-$
c. S^{2-}
d. CO_3^{2-}
e. HSO_4^-

13. State which of the following species are amphiprotic and write chemical equations illustrating the amphiprotic character of these species.

a. NH_3
b. HPO_4^-
c. Br^-
d. NH_4^+
e. ASO_4^{3-}

Answer

amphiprotic: $NH_3 + H_3O^+ \longrightarrow NH_4OH + H_2O$, $NH_3 + OCH_3^- \longrightarrow NH_2^- + CH_3OH$;
$HPO_4^{2-} + OH^- \longrightarrow PO_4^{3-} + H_2O$, $HPO_4^{2-} + HClO_4 \longrightarrow H_2PO_4^- + ClO_4^-$; not amphiprotic: Br^-; NH_4^+; AsO_4^{3-}

14. Is the self ionization of water endothermic or exothermic? The ionization constant for water (K_w) is 2.9×10^{-14} at 40 °C and 9.6×10^{-14} at 60 °C.

5.E.2: pH and pOH

15. Explain why a sample of pure water at 40 °C is neutral even though $[H_3O^+] = 1.7 \times 10^{-7}$ *M*. K_w is 2.9×10^{-14} at 40 °C.

Answer

In a neutral solution $[H_3O^+] = [OH^-]$. At 40 °C,

$[H_3O^+] = [OH^-] = (2.910^{-14})^{1/2} = 1.7 \times 10^{-7}$.

16. The ionization constant for water (K_w) is 2.9×10^{-14} at 40 °C. Calculate $[H_3O^+]$, $[OH^-]$, pH, and pOH for pure water at 40 °C.

17. The ionization constant for water (K_w) is 9.614×10^{-14} at 60 °C. Calculate $[H_3O^+]$, $[OH^-]$, pH, and pOH for pure water at 60 °C.

Answer

x = 3.101×10^{-7} M = $[H_3O^+] = [OH^-]$

pH = -log $3.101 \times 10^{-7} = -(-6.5085) = 6.5085$

pOH = pH = 6.5085

18. Calculate the pH and the pOH of each of the following solutions at 25 °C for which the substances ionize completely:

a. 0.200 M HCl
b. 0.0143 M NaOH
c. 3.0 M HNO_3
d. 0.0031 M $Ca(OH)_2$

19. Calculate the pH and the pOH of each of the following solutions at 25 °C for which the substances ionize completely:

a. 0.000259 M $HClO_4$
b. 0.21 M NaOH
c. 0.000071 M $Ba(OH)_2$
d. 2.5 M KOH

Answer

pH = 3.587; pOH = 10.413; pH = 0.68; pOH = 13.32; pOH = 3.85; pH = 10.15; pH = −0.40; pOH = 14.4

20. What are the pH and pOH of a solution of 2.0 M HCl, which ionizes completely?

21. What are the hydronium and hydroxide ion concentrations in a solution whose pH is 6.52?

Answer

$[H_3O^+] = 3.0 \times 10^{-7}$ M; $[OH^-] = 3.3 \times 10^{-8}$ M

22. Calculate the hydrogen ion concentration and the hydroxide ion concentration in wine from its pH. See below Figure for useful information.

$[H_3O^+]$ (M)	$[OH^-]$ (M)	pH	pOH	Sample Solution	
10^{1}	10^{-15}	–1	15		
10^{0} or 1	10^{-14}	0	14	1 M HCl	acidic
10^{-1}	10^{-13}	1	13		
10^{-2}	10^{-12}	2	12	gastric juice; lime juice; 1 M CH_3CO_2H (vinegar)	
10^{-3}	10^{-11}	3	11	stomach acid	
10^{-4}	10^{-10}	4	10	wine; orange juice	
10^{-5}	10^{-9}	5	9	coffee	
10^{-6}	10^{-8}	6	8	rain water	
10^{-7}	10^{-7}	7	7	pure water	neutral
10^{-8}	10^{-6}	8	6	blood; ocean water	
10^{-9}	10^{-5}	9	5	baking soda	
10^{-10}	10^{-4}	10	4		
10^{-11}	10^{-3}	11	3	Milk of Magnesia	
10^{-12}	10^{-2}	12	2	household ammonia, NH_3	
10^{-13}	10^{-1}	13	1	bleach	
10^{-14}	10^{0} or 1	14	0	1 M NaOH	basic
10^{-15}	10^{1}	15	–1		

23. Calculate the hydronium ion concentration and the hydroxide ion concentration in lime juice from its pH. See Figure for useful information.

Answer

$[H_3O^+] = 1 \times 10^{-2}$ M; $[OH^-] = 1 \times 10^{-12}$ M

24. The hydronium ion concentration in a sample of rainwater is found to be 1.7×10^{-6} M at 25 °C. What is the concentration of hydroxide ions in the rainwater?

25. The hydroxide ion concentration in household ammonia is 3.2×10^{-3} M at 25 °C. What is the concentration of hydronium ions in the solution?

Answer

$[OH^-] = 3.1 \times 10^{-12}\ M$

5.E.3: Relative Strengths of Acids and Bases

5.E.3.1: Q14.3.1

Explain why the neutralization reaction of a strong acid and a weak base gives a weakly acidic solution.

5.E.3.2: Q14.3.2

Explain why the neutralization reaction of a weak acid and a strong base gives a weakly basic solution.

The salt ionizes in solution, but the anion slightly reacts with water to form the weak acid. This reaction also forms OH^-, which causes the solution to be basic.

5.E.3.3: Q14.3.3

Use this list of important industrial compounds (and Figure) to answer the following questions regarding: CaO, $Ca(OH)_2$, CH_3CO_2H, CO_2, HCl, H_2CO_3, HF, HNO_2, HNO_3, H_3PO_4, H_2SO_4, NH_3, NaOH, Na_2CO_3.

a. Identify the strong Brønsted-Lowry acids and strong Brønsted-Lowry bases.
b. List those compounds in that can behave as Brønsted-Lowry acids with strengths lying between those of H_3O^+ and H_2O.
c. List those compounds in that can behave as Brønsted-Lowry bases with strengths lying between those of H_2O and OH^-.

5.E.3.4: Q14.3.4

The odor of vinegar is due to the presence of acetic acid, CH_3CO_2H, a weak acid. List, in order of descending concentration, all of the ionic and molecular species present in a 1-*M* aqueous solution of this acid.

5.E.3.5: S14.3.4

$[H_2O] > [CH_3CO_2H] > [H_3O^+] \approx [CH_3CO_2^-] > [OH^-]$

5.E.3.6: Q14.3.5

Household ammonia is a solution of the weak base NH_3 in water. List, in order of descending concentration, all of the ionic and molecular species present in a 1-*M* aqueous solution of this base.

5.E.3.7: Q14.3.4

Explain why the ionization constant, K_a, for H_2SO_4 is larger than the ionization constant for H_2SO_3.

5.E.3.8: S14.3.4

The oxidation state of the sulfur in H_2SO_4 is greater than the oxidation state of the sulfur in H_2SO_3.

5.E.3.9: Q14.3.7

Explain why the ionization constant, K_a, for HI is larger than the ionization constant for HF.

5.E.3.10: Q14.3.8

Gastric juice, the digestive fluid produced in the stomach, contains hydrochloric acid, HCl. Milk of Magnesia, a suspension of solid $Mg(OH)_2$ in an aqueous medium, is sometimes used to neutralize excess stomach acid. Write a complete balanced equation for the neutralization reaction, and identify the conjugate acid-base pairs.

5.E.3.11: S14.3.8

$$\underset{\text{BB}}{\mathrm{Mg(OH)_2}(s)} + \underset{\text{BA}}{\mathrm{HCl}(aq)} \longrightarrow \underset{\text{CB}}{\mathrm{Mg^{2+}}(aq)} + 2\,\underset{\text{CA}}{\mathrm{Cl^-}(aq)} + 2\,\mathrm{H_2O}(l)$$

5.E.3.12: Q14.3.9

Nitric acid reacts with insoluble copper(II) oxide to form soluble copper(II) nitrate, $Cu(NO_3)_2$, a compound that has been used to prevent the growth of algae in swimming pools. Write the balanced chemical equation for the reaction of an aqueous solution of HNO_3 with CuO.

5.E.3.13: Q14.3.10

What is the ionization constant at 25 °C for the weak acid $CH_3NH_3^+$, the conjugate acid of the weak base CH_3NH_2, K_b = 4.4 × 10^{-4}.

5.E.3.14: S14.3.10

$K_a = 2.3 \times 10^{-11}$

5.E.3.15: Q14.3.11

What is the ionization constant at 25 °C for the weak acid $(CH_3)_2NH_2^+$, the conjugate acid of the weak base $(CH_3)_2NH$, K_b = 7.4 × 10^{-4}?

5.E.3.16: Q14.3.12

Which base, CH_3NH_2 or $(CH_3)_2NH$, is the strongest base? Which conjugate acid, $(CH_3)_2NH_2^+$ or $(CH_3)_2NH$, is the strongest acid?

5.E.3.17: S14.3.12

The strongest base or strongest acid is the one with the larger K_b or K_a, respectively. In these two examples, they are $(CH_3)_2NH$ and $CH_3NH_3^+$.

5.E.3.18: Q14.3.3

Which is the stronger acid, NH_4^+ or HBrO?

5.E.3.19: Q14.3.14

Which is the stronger base, $(CH_3)_3N$ or $H_2BO_3^-$?

5.E.3.20: S14.3.14

triethylamine.

5.E.3.21: Q14.3.15

Predict which acid in each of the following pairs is the stronger and explain your reasoning for each.

a. H_2O or HF
b. $B(OH)_3$ or $Al(OH)_3$
c. HSO_3^- or HSO_4^-
d. NH_3 or H_2S
e. H_2O or H_2Te

5.E.3.22: Q14.3.16

Predict which compound in each of the following pairs of compounds is more acidic and explain your reasoning for each.

a. HSO_4^- or $HSeO_4^-$
b. NH_3 or H_2O
c. PH_3 or HI
d. NH_3 or PH_3
e. H_2S or HBr

5.E.3.23: S14.3.16

a. HSO_4^-; higher electronegativity of the central ion. H_2O;
b. NH_3 is a base and water is neutral, or decide on the basis of K_a values. HI;
c. PH_3 is weaker than HCl; HCl is weaker than HI. Thus, PH_3 is weaker than HI.
d. PH_3; in binary compounds of hydrogen with nonmetals, the acidity increases for the element lower in a group.
e. HBr; in a period, the acidity increases from left to right; in a group, it increases from top to bottom. Br is to the left and below S, so HBr is the stronger acid.

5.E.3.24: Q14.3.17

Rank the compounds in each of the following groups in order of increasing acidity or basicity, as indicated, and explain the order you assign.

a. acidity: HCl, HBr, HI
b. basicity: H_2O, OH^-, H^-, Cl^-
c. basicity: $Mg(OH)_2$, $Si(OH)_4$, $ClO_3(OH)$ (Hint: Formula could also be written as $HClO_4$).
d. acidity: HF, H_2O, NH_3, CH_4

5.E.3.25: Q14.3.18

Rank the compounds in each of the following groups in order of increasing acidity or basicity, as indicated, and explain the order you assign.

a. acidity: $NaHSO_3$, $NaHSeO_3$, $NaHSO_4$
b. basicity: BrO_2^-, ClO_2^-, IO_2^-
c. acidity: HOCl, HOBr, HOI
d. acidity: HOCl, HOClO, $HOClO_2$, $HOClO_3$
e. basicity: NH_2^-, HS^-, HTe^-, PH_2^-
f. basicity: BrO^-, BrO_2^-, BrO_3^-, BrO_4^-

5.E.3.26: S14.3.18

a. $NaHSeO_3 < NaHSO_3 < NaHSO_4$; in polyoxy acids, the more electronegative central element—S, in this case—forms the stronger acid. The larger number of oxygen atoms on the central atom (giving it a higher oxidation state) also creates a greater release of hydrogen atoms, resulting in a stronger acid. As a salt, the acidity increases in the same manner.
b. $ClO_2^- < BrO_2^- < IO_2^-$; the basicity of the anions in a series of acids will be the opposite of the acidity in their oxyacids. The acidity increases as the electronegativity of the central atom increases. Cl is more electronegative than Br, and I is the least electronegative of the three.
c. HOI < HOBr < HOCl; in a series of the same form of oxyacids, the acidity increases as the electronegativity of the central atom increases. Cl is more electronegative than Br, and I is the least electronegative of the three.
d. HOCl < HOClO < $HOClO_2$ < $HOClO_3$; in a series of oxyacids of the same central element, the acidity increases as the number of oxygen atoms increases (or as the oxidation state of the central atom increases).
e. $HTe^- < HS^- << PH_2^- < NH_2^-$; PH_2^- and NH_2^- are anions of weak bases, so they act as strong bases toward H^+. HTe^- and HS^- are anions of weak acids, so they have less basic character. In a periodic group, the more electronegative element has the more basic anion.
f. $BrO_4^- < BrO_3^- < BrO_2^- < BrO^-$; with a larger number of oxygen atoms (that is, as the oxidation state of the central ion increases), the corresponding acid becomes more acidic and the anion consequently less basic.

5.E.3.27: Q14.3.19

Both HF and HCN ionize in water to a limited extent. Which of the conjugate bases, F^- or CN^-, is the stronger base? See Table.

5.E.3.28: Q14.3.20

The active ingredient formed by aspirin in the body is salicylic acid, $C_6H_4OH(CO_2H)$. The carboxyl group ($-CO_2H$) acts as a weak acid. The phenol group (an OH group bonded to an aromatic ring) also acts as an acid but a much weaker acid. List, in order of descending concentration, all of the ionic and molecular species present in a 0.001-*M* aqueous solution of $C_6H_4OH(CO_2H)$.

$$[H_2O] > [C_6H_4OH(CO_2H)] > [H^+]_0 > [C_6H_4OH(CO_2)^-] \gg [C_6H_4O(CO_2H)^-] > [OH^-]$$

What do we represent when we write:

$$CH_3CO_2H(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3CO_2^-(aq)? \tag{5.E.1}$$

5.E.3.29: Q14.3.21

Explain why equilibrium calculations are not necessary to determine ionic concentrations in solutions of certain strong electrolytes such as NaOH and HCl. Under what conditions are equilibrium calculations necessary as part of the determination of the concentrations of all ions of some other strong electrolytes in solution?

5.E.3.30: S14.3.21

Strong electrolytes are 100% ionized, and, as long as the component ions are neither weak acids nor weak bases, the ionic species present result from the dissociation of the strong electrolyte. Equilibrium calculations are necessary when one (or more) of the ions is a weak acid or a weak base.

5.E.3.31: Q14.3.22

Are the concentrations of hydronium ion and hydroxide ion in a solution of an acid or a base in water directly proportional or inversely proportional? Explain your answer.

5.E.3.32: Q14.3.23

What two common assumptions can simplify calculation of equilibrium concentrations in a solution of a weak acid?

5.E.3.33: S14.3.23

1. Assume that the change in initial concentration of the acid as the equilibrium is established can be neglected, so this concentration can be assumed constant and equal to the initial value of the total acid concentration.
2. Assume we can neglect the contribution of water to the equilibrium concentration of H_3O^+.

5.E.3.34: Q14.3.24

What two common assumptions can simplify calculation of equilibrium concentrations in a solution of a weak base?

5.E.3.35: Q14.3.25

Which of the following will increase the percent of NH_3 that is converted to the ammonium ion in water (Hint: Use LeChâtelier's principle.)?

a. addition of NaOH
b. addition of HCl
c. addition of NH_4Cl

5.E.3.36: S14.3.25

The addition of HCl

5.E.3.37: Q14.3.26

Which of the following will increase the percent of HF that is converted to the fluoride ion in water?

a. addition of NaOH
b. addition of HCl
c. addition of NaF

5.E.3.38: Q14.3.27

What is the effect on the concentrations of NO_2^-, HNO_2, and OH^- when the following are added to a solution of KNO_2 in water:

a. HCl
b. HNO_2
c. NaOH
d. NaCl
e. KNO

The equation for the equilibrium is:

$$\mathrm{NO_2^-}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{HNO_2}(aq) + \mathrm{OH^-}(aq) \tag{5.E.2}$$

5.E.3.39: S14.3.27

a. Adding HCl will add H_3O^+ ions, which will then react with the OH^- ions, lowering their concentration. The equilibrium will shift to the right, increasing the concentration of HNO_2, and decreasing the concentration of NO_2^- ions.
b. Adding HNO_2 increases the concentration of HNO_2 and shifts the equilibrium to the left, increasing the concentration of NO_2^- ions and decreasing the concentration of OH^- ions.

c. Adding NaOH adds OH^- ions, which shifts the equilibrium to the left, increasing the concentration of NO_2^- ions and decreasing the concentrations of HNO_2.
d. Adding NaCl has no effect on the concentrations of the ions.
e. Adding KNO_2 adds NO_2^- ions and shifts the equilibrium to the right, increasing the HNO_2 and OH^- ion concentrations.

5.E.3.40: Q14.3.28

What is the effect on the concentration of hydrofluoric acid, hydronium ion, and fluoride ion when the following are added to separate solutions of hydrofluoric acid?

a. HCl
b. KF
c. NaCl
d. KOH
e. HF

The equation for the equilibrium is:

$$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq) \tag{5.E.3}$$

5.E.3.41: Q14.3.29

Why is the hydronium ion concentration in a solution that is 0.10 *M* in HCl and 0.10 *M* in HCOOH determined by the concentration of HCl?

5.E.3.42: S14.3.29

This is a case in which the solution contains a mixture of acids of different ionization strengths. In solution, the HCO_2H exists primarily as HCO_2H molecules because the ionization of the weak acid is suppressed by the strong acid. Therefore, the HCO_2H contributes a negligible amount of hydronium ions to the solution. The stronger acid, HCl, is the dominant producer of hydronium ions because it is completely ionized. In such a solution, the stronger acid determines the concentration of hydronium ions, and the ionization of the weaker acid is fixed by the $[H_3O^+]$ produced by the stronger acid.

5.E.3.43: Q14.3.30

From the equilibrium concentrations given, calculate K_a for each of the weak acids and K_b for each of the weak bases.

CH_3CO_2H: $[H_3O^+]$ = 1.34 × 10^{-3} *M*;

$[CH_3CO_2^-]$ = 1.34 × 10^{-3} *M*;

$[CH_3CO_2H]$ = 9.866 × 10^{-2} *M*;

ClO^-: $[OH^-]$ = 4.0 × 10^{-4} *M*;

[HClO] = 2.38 × 10^{-5} *M*;

$[ClO^-]$ = 0.273 *M*;

HCO_2H: $[HCO_2H]$ = 0.524 *M*;

$[H_3O^+]$ = 9.8 × 10^{-3} *M*; $[HCO_2^-]$ = 9.8 × 10^{-3} *M*;

$C_6H_5NH_3^+$: $[C_6H_5NH_3^+]$ = 0.233 *M*;

$[C_6H_5NH_2]$ = 2.3 × 10^{-3} *M*;

$[H_3O^+]$ = 2.3 × 10^{-3} *M*

From the equilibrium concentrations given, calculate K_a for each of the weak acids and K_b for each of the weak bases.

NH_3: $[OH^-]$ = 3.1 × 10^{-3} *M*;

$[NH_4^+]$ = 3.1 × 10^{-3} *M*;

$[NH_3]$ = 0.533 *M*;

HNO_2: $[H_3O^+]$ = 0.011 *M*;

$[NO_2^-]$ = 0.0438 *M*;

$[HNO_2]$ = 1.07 *M*;

$(CH_3)_3N$: $[(CH_3)_3N]$ = 0.25 *M*;

$[(CH_3)_3NH^+]$ = 4.3 × 10^{-3} *M*;

$[OH^-]$ = 4.3 × 10^{-3} *M*;

NH_4^+ : $[NH_4^+]$ = 0.100 *M*;

$[NH_3]$ = 7.5 × 10^{-6} *M*;

$[H_3O^+]$ = 7.5 × 10^{-6} *M*

a. $K_b = 1.8 \times 10^{-5}$;
b. $K_a = 4.5 \times 10^{-4}$;
c. $K_b = 7.4 \times 10^{-5}$;
d. $K_a = 5.6 \times 10^{-10}$

5.E.3.44: Q14.3.31

Determine K_b for the nitrite ion, NO_2^-. In a 0.10-*M* solution this base is 0.0015% ionized.

5.E.3.45: Q14.3.32

Determine K_a for hydrogen sulfate ion, HSO_4^-. In a 0.10-*M* solution the acid is 29% ionized.

5.E.3.46: S14.3.32

$K_a = 1.2 \times 10^{-2}$

5.E.3.47: Q14.3.33

Calculate the ionization constant for each of the following acids or bases from the ionization constant of its conjugate base or conjugate acid:

a. F^-
b. NH_4^+
c. AsO_4^{3-}
d. $(CH_3)_2NH_2^+$
e. NO_2^-
f. $HC_2O_4^-$ (as a base)

5.E.3.48: Q14.3.52

Calculate the ionization constant for each of the following acids or bases from the ionization constant of its conjugate base or conjugate acid:

1. HTe^- (as a base)
2. $(CH_3)_3NH^+$
3. $HAsO_4^{3-}$ (as a base)
4. HO_2^- (as a base)
5. $C_6H_5NH_3^+$
6. HSO_3^- (as a base)

5.E.3.49: S14.3.52

a. $K_b = 4.3 \times 10^{-12}$;
b. $K_a = 1.4 \times 10^{-10}$;
c. $K_b = 1 \times 10^{-7}$;
d. $K_b = 4.2 \times 10^{-3}$;
e. $K_b = 4.2 \times 10^{-3}$;
f. $K_b = 8.3 \times 10^{-13}$

5.E.3.50: Q14.3.53

For which of the following solutions must we consider the ionization of water when calculating the pH or pOH?

a. 3×10^{-8} *M* HNO_3
b. 0.10 g HCl in 1.0 L of solution
c. 0.00080 g NaOH in 0.50 L of solution
d. 1×10^{-7} *M* $Ca(OH)_2$
e. 0.0245 *M* KNO_3

5.E.3.51: Q14.3.54

Even though both NH_3 and $C_6H_5NH_2$ are weak bases, NH_3 is a much stronger acid than $C_6H_5NH_2$. Which of the following is correct at equilibrium for a solution that is initially 0.10 *M* in NH_3 and 0.10 *M* in $C_6H_5NH_2$?

a. $[OH^-] = [NH_4^+]$
b. $[NH_4^+] = [C_6H_5NH_3^+]$
c. $[OH^-] = [C_6H_5NH_3^+]$
d. $[NH_3] = [C_6H_5NH_2]$
e. both a and b are correct

is the correct statement.

5.E.3.52: Q14.3.55

Calculate the equilibrium concentration of the nonionized acids and all ions in a solution that is 0.25 *M* in HCO_2H and 0.10 *M* in HClO.

5.E.3.53: Q14.3.56

Calculate the equilibrium concentration of the nonionized acids and all ions in a solution that is 0.134 *M* in HNO_2 and 0.120 *M* in HBrO.

5.E.3.54: S14.3.56

$[H_3O^+] = 7.5 \times 10^{-3}$ *M*

$[HNO_2] = 0.126$ $[OH^-] = 1.3 \times 10^{-12}$ *M* $[BrO^-] = 3.2 \times 10^{-8}$ *M* $[HBrO] = 0.120$ *M*

5.E.3.55: Q14.3.57

Calculate the equilibrium concentration of the nonionized bases and all ions in a solution that is 0.25 *M* in CH_3NH_2 and 0.10 *M* in C_5H_5N ($K_b = 1.7 \times 10^{-9}$).

5.E.3.56: Q14.3.58

Calculate the equilibrium concentration of the nonionized bases and all ions in a solution that is 0.115 *M* in NH_3 and 0.100 *M* in $C_6H_5NH_2$.

5.E.3.57: S14.3.58

$[OH^-] = [NO_4^+] = 0.0014$ *M*

$[NH_3] = 0.144$ *M* $[H_3O^+] = 6.9 \times 10^{-12}$ *M* $[C_6H_5NH_3^+] = 3.9 \times 10^{-8}$ *M* $[C_6H_5NH_2] = 0.100$ *M*

5.E.3.58: Q14.3.59

Using the K_a values in Appendix H, place $Al(H_2O)_6^{3+}$ in the correct location in Figure.

5.E.3.59: Q14.3.60

Calculate the concentration of all solute species in each of the following solutions of acids or bases. Assume that the ionization of water can be neglected, and show that the change in the initial concentrations can be neglected. Ionization constants can be found in Appendix H and Appendix I.

a. 0.0092 *M* HClO, a weak acid
b. 0.0784 *M* $C_6H_5NH_2$, a weak base

c. 0.0810 M HCN, a weak acid
d. 0.11 M $(CH_3)_3N$, a weak base
e. 0.120 M $Fe(H_2O)_6^{2+}$ a weak acid, $K_a = 1.6 \times 10^{-7}$

5.E.3.60: S14.3.60

$$\frac{[H_3O^+][ClO^-]}{[HClO]} = \frac{(x)(x)}{(0.0092 - x)} \approx \frac{(x)(x)}{0.0092} = 3.5 \times 10^{-8}$$

Solving for x gives 1.79×10^{-5} M. This value is less than 5% of 0.0092, so the assumption that it can be neglected is valid. Thus, the concentrations of solute species at equilibrium are:

$[H_3O^+] = [ClO] = 1.8 \times 10^{-5}$ M[HClO] = 0.00092 M $[OH^-] = 5.6 \times 10^{-10}$ M;

$$\frac{[C_6H_5NH_3^+][OH^-]}{[C_6H_5NH_2]} = \frac{(x)(x)}{(0.0784 - x)} \approx \frac{(x)(x)}{0.0784} = 4.6 \times 10^{-10}$$

Solving for x gives 6.01×10^{-6} M.

This value is less than 5% of 0.0784, so the assumption that it can be neglected is valid. Thus, the concentrations of solute species at equilibrium are: $[CH_3CO_2^-]$ = $[OH^-]$ = 6.0×10^{-6} M $[C_6H_5NH_2]$ = 0.00784 $[H_3O^+]$ = 1.7×10^{-9} M; $\frac{[H_3O^+][CN^-]}{[HCN]} = \frac{(x)(x)}{(0.0810 - x)} \approx \frac{(x)(x)}{0.0810} = 4 \times 10^{-10}$ Solving for x gives 5.69×10^{-6} M. This value is less than 5% of 0.0810, so the assumption that it can be neglected is valid. Thus, the concentrations of solute species at equilibrium are: $[H_3O^+]$ = $[CN^-] = 5.7 \times 10^{-6}$ M [HCN] = 0.0810 M $[OH^-] = 1.8 \times 10^{-9}$ M; $\frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]} = \frac{(x)(x)}{(0.11 - x)} \approx \frac{(x)(x)}{0.11} = 7.4 \times 10^{-5}$ Solving for x gives 2.85×10^{-3} M. This value is less than 5% of 0.11, so the assumption that it can be neglected is valid. Thus, the concentrations of solute species at equilibrium are: $[(CH_3)_3NH^+] = [OH^-] = 2.9 \times 10^{-3}$ M $[(CH_3)_3N] = 0.11$ M $[H_3O^+] = 3.5 \times 10^{-12}$ M; $\frac{[Fe(H_2O)_5(OH)^+][H_3O^+]}{[Fe(H_2O)_6^{2+}]} = \frac{(x)(x)}{(0.120 - x)} \approx \frac{(x)(x)}{0.120} = 1.6 \times 10^{-7}$ Solving for x gives 1.39×10^{-4} M. This value is less than 5% of 0.120, so the assumption that it can be neglected is valid. Thus, the concentrations of solute species at equilibrium are: $[Fe(H_2O)_5(OH)^+] = [H_3O^+] = 1.4 \times 10^{-4}$ M $[Fe(H_2O)_6^{2+}] = 0.120$ M $[OH^-] = 7.2 \times 10^{-11}$ M

5.E.3.61: Q14.3.61

Propionic acid, $C_2H_5CO_2H$ ($K_a = 1.34 \times 10^{-5}$), is used in the manufacture of calcium propionate, a food preservative. What is the hydronium ion concentration in a 0.698-M solution of $C_2H_5CO_2H$?

5.E.3.62: Q14.3.62

White vinegar is a 5.0% by mass solution of acetic acid in water. If the density of white vinegar is 1.007 g/cm^3, what is the pH?

5.E.3.63: S14.3.62

pH = 2.41

5.E.3.64: Q14.3.63

The ionization constant of lactic acid, $CH_3CH(OH)CO_2H$, an acid found in the blood after strenuous exercise, is 1.36×10^{-4}. If 20.0 g of lactic acid is used to make a solution with a volume of 1.00 L, what is the concentration of hydronium ion in the solution?

5.E.3.65: Q14.3.64

Nicotine, $C_{10}H_{14}N_2$, is a base that will accept two protons ($K_1 = 7 \times 10^{-7}$, $K_2 = 1.4 \times 10^{-11}$). What is the concentration of each species present in a 0.050-M solution of nicotine?

5.E.3.66: S14.3.64

$[C_{10}H_{14}N_2] = 0.049$ M

$[C_{10}H_{14}N_2H^+] = 1.9 \times 10^{-4}$ M $[C_{10}H_{14}N_2H_2^{2+}] = 1.4 \times 10^{-11}$ M $[OH^-] = 1.9 \times 10^{-4}$ M $[H_3O^+] = 5.3 \times 10^{-11}$ M

5.E.3.67: Q14.3.65

The pH of a 0.20-*M* solution of HF is 1.92. Determine K_a for HF from these data.

5.E.3.68: Q14.3.66

The pH of a 0.15-*M* solution of HSO_4^- is 1.43. Determine K_a for HSO_4^- from these data.

5.E.3.69: S14.3.66

$K_a = 1.2 \times 10^{-2}$

5.E.3.70: Q14.3.67

The pH of a 0.10-*M* solution of caffeine is 11.16. Determine K_b for caffeine from these data:

$$C_8H_{10}N_4O_2(aq) + H_2O(l) \rightleftharpoons C_8H_{10}N_4O_2H^+(aq) + OH^-(aq)$$

5.E.3.71: Q14.3.68

The pH of a solution of household ammonia, a 0.950 M solution of NH_3, is 11.612. Determine K_b for NH_3 from these data.

5.E.3.72: S14.3.68

$K_b = 1.77 \times 10^{-5}$

5.E.4: Hydrolysis of Salt Solutions

5.E.4.1: Q14.4.1

Determine whether aqueous solutions of the following salts are acidic, basic, or neutral:

a. $Al(NO_3)_3$
b. RbI
c. $KHCO_2$
d. CH_3NH_3Br

5.E.4.2: Q14.4.2

Determine whether aqueous solutions of the following salts are acidic, basic, or neutral:

a. $FeCl_3$
b. K_2CO_3
c. NH_4Br
d. $KClO_4$

5.E.4.3: S14.4.2

acidic; basic; acidic; neutral

5.E.4.4: Q14.4.3

Novocaine, $C_{13}H_{21}O_2N_2Cl$, is the salt of the base procaine and hydrochloric acid. The ionization constant for procaine is 7×10^{-6}. Is a solution of novocaine acidic or basic? What are $[H_3O^+]$, $[OH^-]$, and pH of a 2.0% solution by mass of novocaine, assuming that the density of the solution is 1.0 g/mL.

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