Chapter 16 exercise

Q1. Practice exercise page 671

Write the formula for the conjugate acid of the following, HSO_3^- , F^- , PO_4^{3-} and CO.

Answer:

$HSO_3^- + H^+ \rightarrow H_2SO_4$

$F^- + H^+ \rightarrow HF$

$PO_4^{3-} + H^+ \rightarrow HPO_4^{2-}$

$CO + H^+ \rightarrow HCO^+$

Q2. Practice exercise page 671

When lithium oxide (Li_2O) is dissolved in water, the solution turns basic from the reaction of oxide ion (O^{2-}) with water. Write the reaction that occurs, and identify the conjugate acid – base pair.

Answer:

$Li_2O_{(aq)} + H_2O_{(l)} \rightarrow LiOH + LiOH$

$O^{2-}_{(aq)}$	+	$H_2O_{(l)}$	$\rightarrow$	$OH^-_{(aq)}$	+	$OH^-_{(aq)}$
base		acid		conjugate acid		conjugate base

Q3. Practice exercise page 673

For the following reactions, use figure 16.4 to predict whether the equilibrium lies predominantly to the left or to the right.

a) $HPO_4^{2-}{}_{(aq)} + H_2O_{(l)} \leftrightarrow H_2PO_4^-{}_{(aq)} + OH^-_{(aq)}$

b) $NH_4^+{}_{(aq)} + OH^-_{(aq)} \leftrightarrow NH_{3(aq)} + H_2O_{(l)}$

Answer:

a) OH^- is in right column (strong base) than H_2O , H_2O is a conjugate weak acid . HPO_4^{2-}

HPO_4^{2-} is acid more than $H_2PO_4^-$, $H_2PO_4^-$ is a weak base .

OH^- more strong base than $H_2PO_4^-$, the reaction is shift to the left.

b) OH^- is a strong base than H_2O , H_2O is conjugate weak acid. OH^- strong base than NH_4^+ , the reaction shift to the right.

Q4. Practice exercise page 675

Indicate whether solutions with each of the following in concentrations are neutral, acidic or basic.

a) $[H^+] = 4 \times 10^{-9}$ M, b) $[OH^-] = 1 \times 10^{-7}$ M c) $[OH^-] = 7 \times 10^{-13}$ M

Answer:

a) $\{H^+] [OH^-] = 1.0 \times 10^{-14}$ $[OH^-] = (1.0 \times 10^{-14}) / (4 \times 10^{-9}) = 0.25 \times 10^{-5}$ M

$[OH^-]$ more concentrated than $[H^+]$, the solution is acidic.

b) $[OH^-] = 1 \times 10^{-7}$ $[H^+] = 1 \times 10^{-7}$

c) $[H^+] [OH^-] = 1.0 \times 10^{-14}$ $[H^+] = (1 \times 10^{-14}) / (7 \times 10^{-13}) = 0.143 \times 10^{-1}$ M more acidic

Q5. Practice exercise page 675

Calculate the concentration of OH^- in solution in which

a) $[H^+] = 2 \times 10^{-6}$ M b) $[H^+] = [OH^+]$ c) $[H^+] = 100 \times [OH^+]$

Answer:

a) $[OH^+] = (1.0 \times 10^{-14}) / (2 \times 10^{-6}) = 0.5 \times 10^{-8}$ M $= 5 \times 10^{-9}$ M

b) $[H^+] [OH^-] = 1.0 \times 10^{-14}$

$(x)(x) = 1.0 \times 10^{-14}$ $x^2 = 1.0 \times 10^{-14}$ $[OH^-] = 1.0 \times 10^{-7}$

c) $[OH^-] = (1.0 \times 10^{-14}) (100 \times [OH^-])$, $[OH^-]^2 = 1 \times 10^{-6}$

$[OH^-] = 1.0 \times 10^{-8}$ M

Q6. Practice exercise page 677

A solution formed by dissolving an anti-acid tablet has a pH of 9.18. Calculate $[H^+]$.

Answer:

PH = - log $[H^+]$ = 9.18

Log $[H^+]$ = - 9.18

$[H^+]$ = antilog (- 9.18) = $10^{-9.18}$ = 6.6×10^{-10} M

Q7. Practice exercise 680

An aqueous solution of HNO_3 has a pH of 2.34. What is the concentration of the acid?

Answer:

$HNO_3 \rightarrow H^+ + NO_3^-$

pH = - log [H^+] = 2.34

[H^+] $10^{-2.34}$ = 4.57×10^{-3} M concentration of HNO_3 is 4.57×10^{-3} M

Q8. Practice exercise page 680

What is the concentration of a solution of

a) KOH for which pH is 11.89 b) $Ca(OH)_2$ for which the pH is 11.68.

Answer:

a) $KOH \rightarrow K^+ + OH^-$

pH = - log [H^+] = 11.89

[H^+] = $10^{-11.89}$ = 1.29×10^{-12} M

[OH^-] = (1.0×10^{-14}) / (1.29×10^{-12}) = 0.775×10^{-2} M = 7.8×10^{-3} M

b) $Ca(OH)_2 \rightarrow Ca^{2+} + 2OH^-$

pH = - log [H^+] = 11.68 [H^+] = $10^{-11.68}$ = 2.089×10^{-12} M

[H^+] [OH^-] = 1.0×10^{-14}

[OH^-] = (1.0×10^{-14}) / (2.089×10^{-12}) = 0.4786×10^{-2} M

for $[OH^-]^2$ = $2 \times 0.4786 \times 10^{-2}$ = 2.4×10^{-3} M

Q9. Practice exercise page 683

Naicin, one of the B- vitamins, a 0.020 M solution of niacin has a pH of 3.26. What is the acid –

constant -dissociation constant , Ka for niacin?

Answer:

pH = - log [H^+] = 3.26

[H^+] = $10^{-3.26}$ = 5.495×10^{-4} M

	Niacin	→	H^+	+	$niacin^-$
initial	0.020 M		0 M		0 M
change	$- 5.5 \times 10^{-4}$ M		$+ 5.5 \times 10^{-4}$ M		$+ 5.5 \times 10^{-4}$ M
equilibrium	$(0.02 – 5.5 \times 10^{-4})$		5.5×10^{-4} M		5.5×10^{-4} M

$Ka = [H^+] [niacin^-] / [niacin] = (5.5 \times 10^{-4})^2 / (0.02 – 5.5 \times 10^{-4}) = 1.5 \times 10^{-5}$

Q10. Practice exercise page 684

A o.020 solution of niacin has pH of 3.26. Calculate the percent ionization of the niacin/

Answer:

$pH = - \log [H^+] = 3.26$

$[H^+] = 10^{- 3.26} = 5.4954 \times 10^{-4}$ M

Percent ionization $= [H^+]_{equilibrium} / [niacin] = (5.4954 \times 10^{-4})(100) / 0.02 = 2.7\%$

Q11. Practice exercise page 686

The Ka for niacin is 1.5×10^{-5}. What is the ph of 0.010 M solution of niacin?

Answer:

Niacin ↔ H^+ + $niacin^-$

	Niacin	→	H^+	+	$niacin^-$
initial	0.010 M		0 m		0 M
change	- X M		+ X M		+ X M
equilibrium	0.01 – X		X		X

$Ka = (X) (X) / 0.01 – X$ $\quad 0.01 – X = 0.01$

$X^2 / 0.01 = 1.5 \times 10^{-5}$ $\quad X = [H^+] = 1.22 \times 10^{-3}$ m

$pH = 3.41$

Q12. Practice exercise page 690

a) Calculate the pH of a 0.020 M solution of oxalic acid ($H_2C_2O_4$) Ka1 = 5.0×10^{-5}, Ka2 = 6.4×10^{-5}

b) Calculate the concentration of oxalic ion, $C_2O_4^{2-}$ in the solution.

Answer:

a)

	$H_2C_2O_4$	$\leftrightarrow$ $HC_2O_4^-$	+ H^+
initial	0.020 M	0 M	0 M
change	- X M	+ X M	+ X M
equilibrium	0.020 – X	X M	X M

$(X)(X) / 0.02 - X = 5.9 \times 10^{-2}$

$X^2 + 5.9 \times 10^{-2} X - 0.118 \times 10^{-2} = 0$

$X = (-5.9 \times 10^{-2}) \pm [\sqrt{(5.9 \times 10^{-2})^2 - 4(-0.118 \times 10^{-2})}] / 2$

$X = 0.0158$ pH = - log (0.0158)

pH = 1.8

b)

	$HC_2O_4^-$	$\leftrightarrow$ H^+	+ $C_2O_4^{-2}$
initial	0.0158 M	0.0158 M	0 M
change	- Y	+ Y M	+ Y M
equilibrium	0.0158 – Y	0.0158 + Y	Y M

$(Y)(0.0158 + Y) / (0.0158 - Y) = 6.4 \times 10^{-5}$ (Y is very small can be neglected)

$Y * 0.0158 / 0.0158 = 6.4 \times 10^{-5}$ $Y = [C_2O_4^{2-}] = 6.4$ xc 10^{-5}

Q13. Practice exercise page 693

A solution of NH_3 in water has a pH = 11.17. What is the molarity of the solution?

Answer:

$NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$

P OH = 14 - pH   14.00 – 11.17 = 2.83

$[OH^-] = 10^{-2.83} = 1.48 \times 10^{-3}$ M

	NH_3	+ H_2O	$\leftrightarrow$ $NH_4^{\pm}$	+ OH^-
initial	X		0 M	0 M
change	$- 1.48 \times 10^{-3}$		$+ 1.48 \times 10^{-3}$	$+ 1.48 \times 10^{-3}$ M
equilibrium	$X - 1.48 \times 10^{-3}$		1.48×10^{-3}	1.48×10^{-3} M

$Kb = [NH_4] [OH^-] / [NH_3] = 1.8 \times 10^{-5} = (1.48 \times 10^{-3})^2 / (X – 1.48 \times 10^{-3})$

$X = [(2.19 \times 10^{-6}) + (2.664 \times 10^{-8})] / (1.8 \times 10^{-5}) = 0.123$ M

Q14. Practice exercise page 695

a) Which of the following anions has the largest base dissociation constant NO_2^-, PO_4^{3-}, N_3^-

b) The base quinolone, its conjugate is pKa = 4.9. What is the base-dissociation constant for quinolone.

Answer:

a) NO_2^- is a conjugate base for the acid HNO_2 Ka = 4.5×10^{-4}

PO_4^{3-} is a conjugate base for H_3PO_4 has three Ka 7.5×10^{-3}, 6.2×10^{-8} and 4.2×10^{-13}

N_3^- is a conjugate base for the acid Ka = 1.9×10^{-5}

$Kb = (1.0 \times 10^{-14}) (4.5 \times 10^{-4}) = 0.22 \times 10^{-10}$ for NO_2^-

$Kb = (1.0 \times 10^{-14}) (4.2 \times 10^{-13}) = 0.24 \times 10^{-1}$ for PO_4^{3-}

$Kb = (1.0 \times 10^{-14}) (1.9 \times 10^{-5}) = 0.53 \times 10^{-9}$ for N^{3-}

Largest base dissociation constant is PO_4^{3-}

b) pKa + pKb = pKw

4.90 + pKb = 14.00

pKb = - log Kb = 9.1

$Kb = 10^{-9.1} = 7.9 \times 10^{-10}$

Q15. Practice exercise page 698

In each of the following, indicate which salt in each of the following pair will form the more acidic (or more basic).

(a) $NaNO_3$ or $Fe(NO_3)_3$ (b) KBr or KBrO (c) CH_3NH_3Cl or $BaCl_2$ (d) NH_4NO_2 or NH_4NO_3

Answer:

a) $NaNO_3 \leftrightarrow Na^+ + NO_3^-$

Na^+ ion from group 1A has no effect on pH

NO_3^- ion is the conjugate base of strong acid HNO_3 has no effect on pH

The solution is neutral

Fe^{3+} is not from group 1A or 2A , decrease the pH

NO_3^- ion is the conjugate base of strong acid HNO_3 has no effect on pH

The solution is acidic

$Fe(NO_3)_3$ more acidic than $NaNO_3$

b) K^+ ion from group 1A has no effect on pH

Br^- is a conjugate base of strong acid HBr it has no influence on pH

KBr form a neutral solution

KBrO, BrO^- is a conjugate base for a weak acid HBrO

$BrO^- + H_2O \leftrightarrow HBrO + OH^-$ the solution is basic

KBr is more acidic than KBrO

Exercises page 710

16.15

a) what is the difference between the Arrhenius and Bronsted-Lowry definition of an acid?

b) $NH_{3\,(g)}$ and $HCl_{(g)}$ react to form $NH_4Cl_{(s)}$ (figure 16.3) which substance is the Bronsted-Lowry acid in this reaction? Which is the Bronsted-Lowry base?

Answer:

a) Arrhenius base is added to water leads to an increase in the concentration of OH^-, while Arrhenius acid in water an increase in the concentration of H^+.

Bronsted – Lowry base is accept a proton from H_2O and the acid it donate a proton from H_2O.

b) $NH_3 + HCl \rightarrow NH_4Cl$

HCl is the Bronsted – Lowry acid and NH_3 is the Bronsted – Lowry base

16.17

a) Give the conjugate base of the following Bronsted – Lowry acids (i) HIO_3 (ii) NH_4^+

b) Give the conjugate acid of the following base (i) O^{2-} (ii) $H_2PO_4^-$

Answer

a) i) $HIO_3 \rightarrow$ conjugate base is IO_3^-

ii) $NH_4^+ \rightarrow$ conjugate base is NH_3

b) i) O^{2-} → conjugate acid is OH^-

ii) $H_2PO_4^-$ → conjugate acid is H_3PO_4

16.19

Designate the Bronsted – Lowry acid and Bronsted – Lowry base on the left side of each of the following equations, and also designate the conjugate acid and conjugate base on the right side.

a) $NH_4^+{}_{(aq)}$ + $CN^-{}_{(aq)}$ ↔ $HCN_{(aq)}$ + $NH_{3\,(aq)}$

b) $(CH_3)_3N_{(aq)}$ + $H_2O_{(l)}$ ↔ $(CH_3)NH^+{}_{)aq)}$ + $OH^-{}_{(aq)}$

c) $HCHO_{2\,(aq)}$ + PO_4^{3-} ↔ $CHO_2^-{}_{(aq)}$ + $HPO_4^{2-}{}_{(aq)}$

Answer:

a)	$NH_4^+{}_{(aq)}$	+	$CN^-{}_{(aq)}$	↔	$HCN_{(aq)}$	+	$NH_{3\,(aq)}$
	acid		base		conjugate acid		conjugate base
b)	$(CH_3)_3N_{(aq)}$	+	$H_2O_{(l)}$	↔	$(CH_3)NH^+{}_{)aq)}$	+	$OH^-{}_{(aq)}$
	base		acid		conjugate acid		conjugate base
c)	$HCHO_{2\,(aq)}$	+	PO_4^{3-}	↔	$CHO_2^-{}_{(aq)}$	+	$HPO_4^{2-}{}_{(aq)}$
	acid		base		conjugate base		conjugate acid

16.21

a) The hydrogen oxalate ion ($HC_2O_4^-$) is amphiprotic. Write a balance chemical equation showing how it acts as an acid toward water and another equation showing how it acts as a base towered water.

b) What is the conjugate acid of $HC_2O_4^-$? what is the conjugate base.

Answer:

a) $HC_2O_4^-$ + H_2O ↔ $C_2O_4^{2-}$ + H_3O^+ (behave as acid in water)

$HC_2O_4^-$ + H_2O ↔ $H_2C_2O_4$ + OH^- (behave as base in water)

b) $H_2C_2O_4$ is a conjugate acid of $HC_2O_4^-$

$C_2O_4^{2-}$ is a conjugate base of $HC_2O_4^-$

16.27

Predict the products of the following acid – base reactions, and predict whether the the equilibrium lies to the left or to the right of the equations.

a) $O^{2-}{}_{(aq)}$ + $H_2O_{(l)}$ ↔

b) $CH_3COOH_{(aq)}$ + HS^- ↔

c) NO_2^- + $H_2O_{(l)}$ ↔

Answer:

a) O^{2-} + H_2O ↔ OH^- + OH^-

base acid conjugate acid conjugate base

b) $CH_3COOH_{(aq)}$ + HS^- ↔ H_2S + CH_3COO^-

acid base conjugate acid conjugate base

c) NO_2^- + $H_2O_{(l)}$ ↔ HNO_2 + OH^-

The equilibrium to the left

16.31

Calculate [H^+] for such of the following solutions, and indicate whether the solution is acidic, basic or neutral.

A) [OH^-] = 0.00045 M b) [OH^-] 8.8×10^{-9} M

c) a solution which [OH^-] is 100 times greater then [H^+]

Answer:

a) pOH = - log [OH^-] = - log (0.00045) = 3.35

pOH + pH = 14

pH = 10.65

b) pOH = - log [OH^-] = - log (8.8×10^{-9}) = 7.06

pOH = 14.00 – 7.06 = 6.94 the solution is acidic

c) $[H^+]$ x 100 $[H^+]$ = 10^{-14}

$[H^+]$ = 10^{-14} / 100 = 10^{-16}

$[H^+]$ = 10^{-8} pH = 8 the solution is basic

16.33

At the freezing point of water 0˚C , K_w = 1.2 x 10^{-15} . Calculate $[H^+]$ and $[OH^-]$ for neutral solution of this temperature.

Answer:

$[OH^-][H^+]$ = 1.2 x 10^{-15} X^2 = 1.2 x 10^{-15}

$[OH^-]$ = $[H^+]$ = 3.5 x 10^{-6} M

16.35

By what factor does $[H^+]$ change for pH change of

a) 2.00 units b) 0.50 units

Answer:

pH = - log $[H^+]$ $[H^+]$ = $10^{-2.00}$ = 0.01 = 1 / 100

$[H^+]$ = $10^{-0.50}$ = 0.316 = 1 / 0.316 = 3.2

16.39

Complete the following table by calculating the missing entries and indicating whether the solution is acidic or basic

H^+	OH^-	pH	pOH	acidic or basic
7.5 x 10^{-3} M				
	3.6 x 10^{-10} M			
		8.25		
			5.70	

Answer:

$[H^+][OH^-]$ 1.0 x 10^{-14}

$[OH^-] = (1.0 \times 10^{-14}) / (7.5 \times 10^{-3}) = 1.3 \times 10^{-12}$ M

pOH = - log (1.3×10^{-12}) = 11.87

pH = 14 – 11.87 = 2.13 the solution is acidic

H^+	OH^-	pH	pOH	acidic or basic
7.5×10^{-3} M	1.3×10^{-12} M	2.13	11.87	acidic
2.8×10^{-5} M	3.6×10^{-10} M	4.56	9.44	acidic
5.6×10^{-9} M	1.8×10^{-6} M	8.25	5.75	basic
5.0×10^{-9} M	2.0×10^{-6} M	8.30	5.70	basic

16.41

The average pH of normal arterial blood is 7.40. At normal body temperature is 37˚C, K_w = 2.4×10^{-14}. Calculate $[H^+]$, and pOH for the blood at this temperature.

Answer:

PH = - log $[H^+]$ = 7.40

$[H^+] = 3.98 \times 10^{-8}$ M = 4.00×10^{-8} M

$[H^+][OH^-] = 2.4\ 10^{-14}$ $[OH^-] = 0.6 \times 10^{-6}$ M

pOH = - log (0.6×10^{-6}) = 6.22

16.45

Calculate the pH of the following strong acid solutions.

a) 8.5×10^{-3} M HBr.

b) 1.52 g of HNO_3 in 575 mL of solution.

c) 5.00 mL of 0.250 M $HClO_4$ diluted to 50.0 mL.

d) a solution formed by mixing 10.0 mL of 0.100 M HBr with 20.0 mL of 0.2 M HCl.

Answer:

a) HBr → H^+ + Br^-

8.5×10^{-3} M 8.5×10^{-3} M 8.5×10^{-3} M pH = - log (8.5×10^{-3}) = 2.07

b) mole of HNO_3 = (1.52 g) /(63 g/mol) = 0.024 mol

molarity of HNO_3 = (0.024 mol HNO_3)(1000 mL / 575 mL) = 0.042 M [H^+]

pH = – log [H^+] = - log (0.042) = 1.38

c) N1 x V1 = N2 x V2

0.250 M x 5.00 mL = N2 x 50.0 mL

N2 = 0.0250 M molarity of $HClO_4$ = molarity of [H^+]

pH = - log [H^+] = - log (0.0250) = 1.60

d) 10 mL + 20 mL = 30 mL volume of solution

N1 x V1 = N2 x V2

0.100 M x 10 mL = N2 x 30 mL N2 = 0.033 M of [HBr] = [H^+]

N1 x V1 = N2 x V2

0.2 M x 20 mL = N2 x 30 mL N2 = 0.133 M of HCl = [H^+]

0.033 M [H^+] from HBr + 0.133 M [H^+] from HCl = 0.166 M

pH = - log (0.166) = 0.778

16.47

Calculate [OH^-] and pH for

a) 1.5×10^{-3} M $Sr(OH)_2$

b) 2.250 g of LiOH in 250.0 mL of solution

c) 1.00 mL of 0.175 M NaOH diluted to 2.000 L

d) a solution formed by adding 5.00 mL of 0.105 M KOH to 15.0 mL of 9.5×10^{-2} M $Ca(OH)_2$

Answer:

a) $Sr(OH)_2 \rightarrow Sr^{2+} + 2OH^-$

[OH^-] = 2 x 1.5×10^{-3} = 3.0×10^{-3} M

pOH = - log (3.0×10^{-3}) = 2.523

pH = 14 – 2.523 = 11.48

b) mole of LiOH = 2.250 / 24 = 0.094

[LiOH] M = (0.094 mol LiOH) (1000 mL /250 mL) = 0.375 M of LiOH

$LiOH \rightarrow Li^{+} + OH^{-}$

0.375 0.375

pOH = - log [OH^{-}] = - log (0.375) = 0.426

pH = 14 – 0.426 = 13.57

c) N1 x V1 = N2 x V2

0.175 x 1.0 mL = N2 x 2000 mL N2 = 8.75×10^{-5} M NaOH

pOH = - log [OH^{-}] = - log (8.75×10^{-5}) = 4.058

pH = 14 – 4.058 = 9.942

d) total volume of solution 5.0 mL + 15 mL = 20 mL

for KOH N1 x V1 = N2 x V2

0.105 M x 5.0 mL = N2 x 20.0 mL N2 = 0.02625 M

KOH for $Ca(OH)_2$ 9.5×10^{-2} M x 15 mL = N2 x 20 mL

N2 = 0.07125 M of $Ca(OH)_2$ = 0.07125 M [OH^{-}]

total concentrations of [OH^{-}] = 0.02625 + 2x 0.07125 = 0.16875 M = 0.17 M

pOH = - log [OH^{-}] = - log (0.17) = 0.773

pH = 14 – 0.773 = 13.23

<u>16.53</u>

Lactic acid ($CH_3CH(OH)COOH$) has one acidic hydrogen. A 0.10 M solution lactic acid has pH = 2.44.

Calculate Ka.

<u>Answer</u>:

$CH_3CH(OH)COOH \leftrightarrow H^{+} + CH_3CH(OH)COO^{-}$

pH = - log [H^{+}] = 2.44

$[H^+]$ = 0.00363 M

Ka = $[CH_3CH(OH)COO^-]$ $[H^+]$ / $[CH_3CH(OH)COO]$ = (0.000363)(0.00363) / 0.1

Ka = 1.32×10^{-4} = 1.4×10^{-4}

16.55

A 0.10 M solution of chloroacetic acid ($ClCH_2COOH$) is 11.0% ionized. Using this information, calculate $[ClCH_2COO^-]$, $[H^+]$, and $[ClCH_2COOH]$ and Ka for chloroacetic acid.

Answer:

Ionizes chloroacetic acid = 0.10 M x 11/100 = 0.011 M

Remain chloroacetic acid = 0.10 – 0.011 = 0.089 M

	$ClCH_2COOH_{(aq)}$	$\leftrightarrow$ $ClCH_2COO^-_{(aq)}$	+ $H^+_{(aq)}$
initial	0.011 M	0 M	0 M
change	- 0.011 M	+ 0.011 M	+ 0.011 M
equilibrium	(0.10 – 0.011)	0.011 M	0.011 M

Ka = (0.011) (0.011) / (0.089) = 1.4×10^{-3}

16.57

A particular sample of vinegar has a pH = 2.9. If acetic acid is the only acid that vinegar contains (Ka = 1.8×10^{-5}), calculate the concentration of acetic acid in the vinegar.

Answer:

pH - log $[H^+]$ = 2.90

$[H^+]$ = 0.00126 = the concentration of vinegar

$CH_3COOH \leftrightarrow CH_3COO^- + H^+$

Ka = $[CH_3COO^{-/}]$ $[H^-]$ / $[CH_3COOH]$ = 1.8×10^{-5}

(0.00126) (0.00126) / $[CH_3COOH]$ = 1.8×10^{-5}

$[CH_3COOH]$ = 0.0882 M

16.63

Saccharin, a sugar substitute, is a weak acid with pKa = 2.32 at 25˚C. It ionizes in aqueous solution as follows:

$HNC_7H_4SO_{3\,(aq)} \leftrightarrow H^+_{(aq)} + NC_7H_4SO_3^-{}_{(aq)}$

What is the pH of a 0.10 M solution of this substance?

Answer:

pKa = - log Ka = 2.32

Ka = 4.79×10^{-3} = $[H^+][NC_7H_4SO_3^-] / [HNC_7H_4SO_3]$

	$HNC_7H_4SO_{3\,(aq)}$ $\leftrightarrow$	$H^+_{(aq)}$ +	$NC_7H_4SO_3^-{}_{(aq)}$
initial	0.10 M	0 M	0 M
change	- X M	+ X M	+ X M
equilibrium	(0.10 – X) M	X M	X M

$X^2 / (0.10 - X) = 4.79 \times 10^{-4}$

$X = [H^+] = 2.23 \times 10^{-2}$ M

pH = - log (2.23×10^{-2}) = 1.652 = 1.7

16.65

Calculate the percent ionization of hydrazoic acid (HN_3) in solutions each of the following concentrations (Ka = 1.9×10^{-5})

a) 0.400 M b) 0.100 M c) 0.0400 M

Answer:

a) percent ionization = $[H^+]_{equ} / [HA]_{initial} \times 100\%$

$HN_3 \leftrightarrow H^+ + N_3^-$

Ka = $[H^+][N_3^-] / [HN_3]$ = 1.9×10^{-5} X = $[H^+]$ = $[N_3^-]$

$X^2 / (0.4 - X) = 1.9 \times 10^{-5}$ X = 2.77×10^{-3} M = $[H^+]$

Percent ionization = (2.77×10^{-3}) / (0.4) x 100 = 0.69%

b) $[H^+][N_3^-] / [HN_3] = 1.9 \times 10^{-3}$ M

X^2 / (0.1 – X) = 1.9 x 10^{-3} X = 1.38 x 10^{-3} M = [H^+]

Percent ionization = [H^+]$_{equ}$ / [HA]$_{initial}$ x 100 = 1.4%

c) X^2 / (0.04 – X) = 1.9 x 10^{-5} X = 0.872 x 10^{-3} M = [H^+]

percent ionization = (0.872 x 10^{-3}) / 0.04 x 100 = 2.2%

16.73

Write the chemical equations and Kb expression for the ionization of each of the following bases in aqueous solutions.

a) dimethyl amine $(CH_3)_2NH$

b) carbonate ion CO_3^{2-}

c) formate ion CHO_2^-

Answer:

a) $(CH_3)_2NH_{(aq)}$ + $H_2O_{(l)}$ ↔ $(CH_3)_2NH_2^+$ + OH $_{(aq)}$

Kb = [$(CH_3)_2NH_2^+$] [OH^-] / [$(CH_3)_2NH$]

b) $CO_3^{2-}{}_{(aq)}$ + $H_2O_{(l)}$ ↔ $HCO_3^-{}_{(aq)}$ + $OH^+{}_{(aq)}$

Kb = [HCO_3^-] [OH^-] / [CO_3^{2-}]

c) $HCOO^-{}_{(aq)}$ + $H_2O_{(l)}$ ↔ $HCOOH_{(aq)}$ + $OH^-{}_{(aq)}$

Kb = [HCOOH] [OH^-] / [$HCOO^-$]

16.83

Calculate [OH^-] and pH for each of the following solutions

Ka for HCN = 4.9 x 10^{-10} , Ka for H_2CO_3 = 5.6 x 10^{-11} , Ka for HNO_2 = 4.5 x 10^{-4})

a) 0.10 M NaCN

b) 0.080 M Na_2CO_3

c) a mixture that is 0.10 M in $NaNO_2$ and 0.20 M in $Ca(NO_2)_2$

Answer:

a) NaCN → Na^+ + CN^-

$CN^- + H_2O \rightarrow HCN + OH^-$

$Kb \times Ka = Kw$ $Kb = (1.0 \times 10^{-14}) / (4.9 \times 10^{-10}) = 0.204 \times 10^{-4}$

$Kb = [HCN][OH^-] / [CN^-] = 0.204 \times 10^{-4}$

	CN^-	$+ H_2O \rightarrow$	HCN	$+ OH^-$
initial	0.10 M		0 M	0 M
change	- X M		+ X M	+ X M
equilibrium	0.10 – X		X M	X M

$(X)(X) / 0.1 - X = 0.204 \times 10^{-4}$ $X = 1.43 \times 10^{-3}$ M $= [OH^-]$

$[OH^-][H^+] = 1.0 \times 10^{-14}$ $[H^+] = (1.0 \times 10^{-14}) / (1.43 \times 10^{-3}) = 0.7 \times 10^{-11}$

$pH = -\log(0.7 \times 10^{-11}) = 11.15$

b) $Na_2CO_3 \rightarrow 2Na^+ + CO_3^{2-}$

$CO_3^{2-} + H_2O \rightarrow HCO_3^- + OH^-$

$Kb = (1.0 \times 10^{-14}) / (5.6 \times 10^{-11}) = 0.18 \times 10^{-3}$

$Kb = [HCO_3^-][OH^-] / [CO_3^{2-}] = 0.18 \times 10^{-3}$

$X^2 / (0.08 - X) = 0.18 \times 10^{-3}$ $X = 3.6 \times 10^{-3} = [OH^-]$

$[H^+][OH^-] = 10^{-14}$ $[H^+] = (1.0 \times 10^{-14}) / (3.8 \times 10^{-3}) = 0.263 \times 10^{-11}$ M

$pH = -\log[H^+] = -\log(0.263 \times 10^{-11}) = 11.58$

c) $NaNO_2 \rightarrow Na^+ + NO_2^-$

$NO_2^- + H_2O \rightarrow HNO_2 + OH^-$

<u>16.85</u>

Predict whether aqueous solutions of the following compounds are acidic, basic or neutral

a) NH_4Br b) $FeCl_3$ c) Na_2CO_3 d) $HClO_4$ e) $NaHC_2O_4$

<u>Answer</u>:

a) $NH_4Br \rightarrow NH_4^+ + Br^-$

NH_4^+ is the conjugate acid of base NH_3

Br^- is conjugate base of a strong acid HBr , it has no influence on pH.

The solution of the salt is acidic

b) $FeCl_3 \rightarrow Fe^{3+} + 3Cl^-$

Fe^{3+} is not from group 1A or group 2A, decrease the pH

Cl^- is the conjugate base of strong HCl , has no influence on pH

The solution of the salt is acidic

c) $Na_2CO_3 \rightarrow 2Na^+ + CO_3^{2-}$

Na^+ is from group 1A has no effect on pH

CO_3^{2-} is a conjugate base of weak acid H_2CO_3

The solution of the salt is basic

d) $KClO_4 \rightarrow K^+ + ClO_4^-$

K^+ is from group 1A has no effect on pH

ClO_4^- is a conjugate base of the acid $HClO_4$ has no effect on pH

The solution of the salt is neutral

e) $NaHC_2O_4 \rightarrow Na^+ + HC_2O_4^-$

Na^+ is from group 1A has no influence on pH

$HC_2O_4^-$ is a conjugate acid of a base $C_2O_4^{2-}$

The solution of the salt is acidic

16.87

An unknown salt is either NaF, NaCl , or NaOCl. When 0.05 mole of salt is dissolved in water to form 0.500 L of solution. The pH of solution is 8.08. What is the identity of the salt?

(Kb for the F^- = 1.5×10^{-11}, Kb for ClO^- = 0.334×10^{-6}, Cl^- is from strong acid HCl)

<u>Answer</u>:

$pH + pOH = 14$

$8.08 + pOH = 14$ $pOH = 5.92$

$pOH = - \log [OH^-] = 5.92$

$[OH^-] = 1.2 \times 10^{-6}$ M

$Kb = (X)(X) / 0.1 = (1.20 \times 10^{-6})(1.2 \times 10^{-6}) / 0.1 = 1.45 \times 10^{-11}$

$Kb \times Ka = 10^{-14}$ $Ka = (1.0 \times 10^{-14}) / (1.45 \times 10^{-11}) = 7.14 \times 10^{-3}$

Kb for the anion salt is 1.5×10^{-11}, Ka for the conjugate acid = 7.14×10^{-3}. The conjugate acid is F^- , the salt is NaF.

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