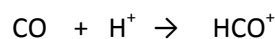
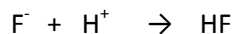


Chapter 16 exercise**Q1. Practice exercise page 671**

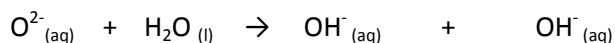
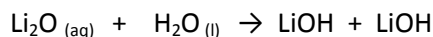
Write the formula for the conjugate acid of the following, HSO_3^- , F^- , PO_4^{3-} and CO .

Answer:

**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.

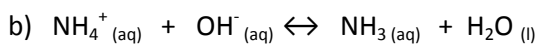
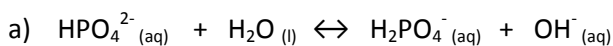
Answer:



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.



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

Answer:

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

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

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

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

Q5. Practice exercise page 675

Calculate the concentration of OH^- in solution in which

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

Answer:

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

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

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

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

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

Q6. Practice exercise page 677

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

Answer:

$\text{PH} = -\log [\text{H}^+] = 9.18$

$\text{Log} [\text{H}^+] = -9.18$

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

Q7. Practice exercise 680

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

Answer:



$$\text{pH} = -\log [\text{H}^+] = 2.34$$

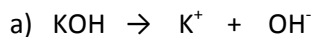
$$[\text{H}^+] 10^{-2.34} = 4.57 \times 10^{-3} \text{ M} \quad \text{concentration of } \text{HNO}_3 \text{ is } 4.57 \times 10^{-3} \text{ M}$$

Q8. Practice exercise page 680

What is the concentration of a solution of

- a) KOH for which pH is 11.89 b) $\text{Ca}(\text{OH})_2$ for which the pH is 11.68.

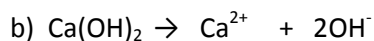
Answer:



$$\text{pH} = -\log [\text{H}^+] = 11.89$$

$$[\text{H}^+] = 10^{-11.89} = 1.29 \times 10^{-12} \text{ M}$$

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



$$\text{pH} = -\log [\text{H}^+] = 11.68 \quad [\text{H}^+] = 10^{-11.68} = 2.089 \times 10^{-12} \text{ M}$$

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

$$[\text{OH}^-] = (1.0 \times 10^{-14}) / (2.089 \times 10^{-12}) = 0.4786 \times 10^{-2} \text{ M}$$

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

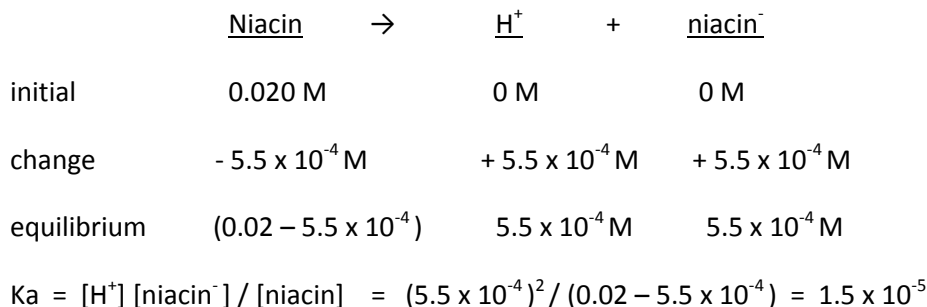
Q9. Practice exercise page 683

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 , K_a for niacin?

Answer:

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

$$[\text{H}^+] = 10^{-3.26} = 5.495 \times 10^{-4} \text{ M}$$

**Q10. Practice exercise page 684**

A 0.020 solution of niacin has pH of 3.26. Calculate the percent ionization of the niacin/

Answer:

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

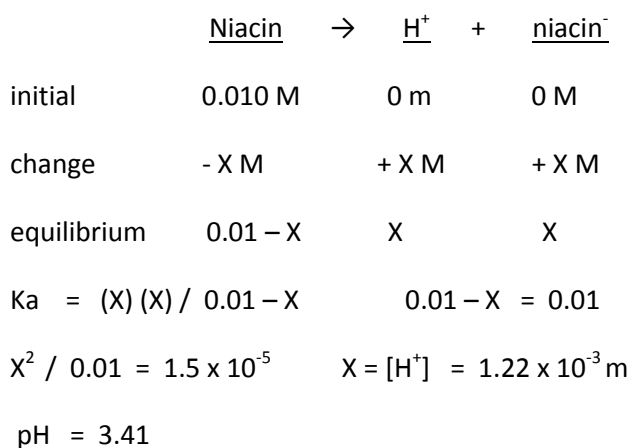
$$[\text{H}^+] = 10^{-3.26} = 5.4954 \times 10^{-4} \text{ M}$$

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

Q11. Practice exercise page 686

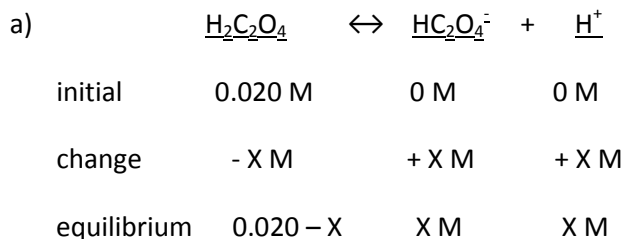
The Ka for niacin is 1.5 x 10⁻⁵. What is the pH of 0.010 M solution of niacin?

Answer:

**Q12. Practice exercise page 690**

- a) Calculate the pH of a 0.020 M solution of oxalic acid (H₂C₂O₄) Ka1 = 5.0 x 10⁻⁵, Ka2 = 6.4 x 10⁻⁵
- b) Calculate the concentration of oxalic ion, C₂O₄²⁻ in the solution.

Answer:



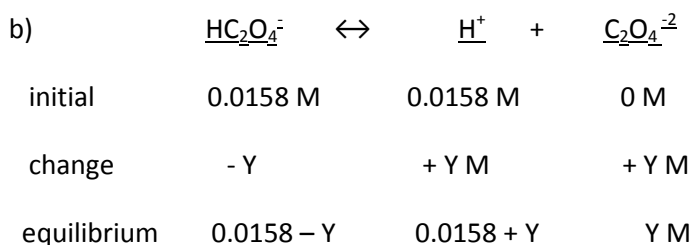
$$(X)(X) / 0.02 - X = 5.9 \times 10^{-2}$$

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

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

$$X = 0.0158 \quad \text{pH} = -\log(0.0158)$$

$$\text{pH} = 1.8$$



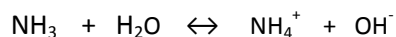
$$(Y)(0.0158 + Y) / (0.0158 - Y) = 6.4 \times 10^{-5} \quad (\text{Y is very small can be neglected})$$

$$Y * 0.0158 / 0.0158 = 6.4 \times 10^{-5} \quad Y = [\text{C}_2\text{O}_4^{2-}] = 6.4 \times 10^{-5}$$

Q13. Practice exercise page 693

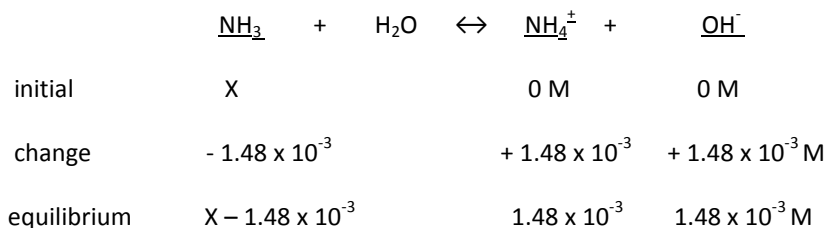
A solution of NH_3 in water has a pH = 11.17. What is the molarity of the solution?

Answer:



$$\text{POH} = 14 - \text{pH} \quad 14.00 - 11.17 = 2.83$$

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



$$K_b = \frac{[\text{NH}_4][\text{OH}^-]}{[\text{NH}_3]} = 1.8 \times 10^{-5} = \frac{(1.48 \times 10^{-3})^2}{(X - 1.48 \times 10^{-3})}$$

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

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 $\text{pK}_a = 4.9$. What is the base-dissociation constant for quinolone.

Answer:

a) NO_2^- is a conjugate base for the acid HNO_2 $K_a = 4.5 \times 10^{-4}$

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

N_3^- is a conjugate base for the acid $K_a = 1.9 \times 10^{-5}$

$$K_b = (1.0 \times 10^{-14}) (4.5 \times 10^{-4}) = 0.22 \times 10^{-10} \text{ for } \text{NO}_2^-$$

$$K_b = (1.0 \times 10^{-14}) (4.2 \times 10^{-13}) = 0.24 \times 10^{-1} \text{ for } \text{PO}_4^{3-}$$

$$K_b = (1.0 \times 10^{-14}) (1.9 \times 10^{-5}) = 0.53 \times 10^{-9} \text{ for } \text{N}_3^-$$

Largest base dissociation constant is PO_4^{3-}

b) $\text{pK}_a + \text{pK}_b = \text{pK}_w$

$$4.90 + \text{pK}_b = 14.00$$

$$\text{pK}_b = -\log K_b = 9.1$$

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

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 $\text{Fe}(\text{NO}_3)_3$ (b) KBr or KBrO (c) $\text{CH}_3\text{NH}_3\text{Cl}$ or BaCl_2 (d) NH_4NO_2 or NH_4NO_3

Answer:

a) $\text{NaNO}_3 \leftrightarrow \text{Na}^+ + \text{NO}_3^-$

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

$\text{Fe}(\text{NO}_3)_3$ 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

$\text{BrO}^- + \text{H}_2\text{O} \leftrightarrow \text{HBrO} + \text{OH}^-$ 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) $\text{NH}_3 + \text{HCl} \rightarrow \text{NH}_4\text{Cl}$

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) $\text{HIO}_3 \rightarrow$ conjugate base is IO_3^-

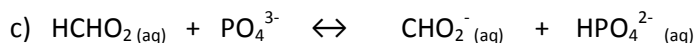
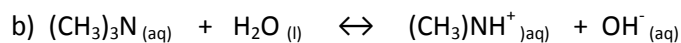
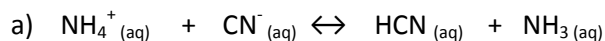
ii) $\text{NH}_4^+ \rightarrow$ conjugate base is NH_3

b) i) $O^{2-} \rightarrow$ conjugate acid is OH^-

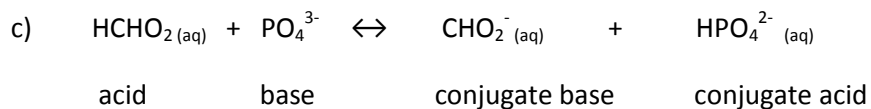
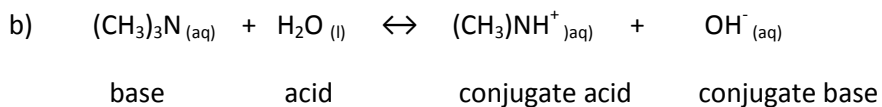
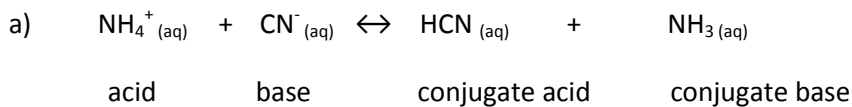
ii) $H_2PO_4^- \rightarrow$ conjugate acid is H_3PO_4

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.



Answer:

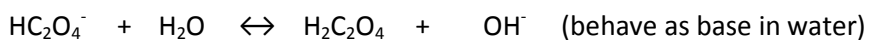
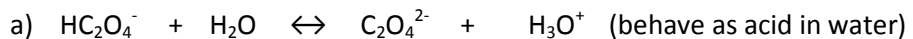


16.21

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

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

Answer:

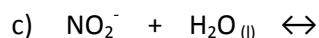
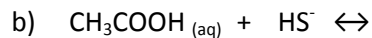
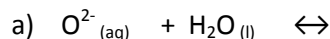


b) $H_2C_2O_4$ is a conjugate acid of $HC_2O_4^-$

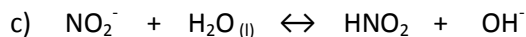
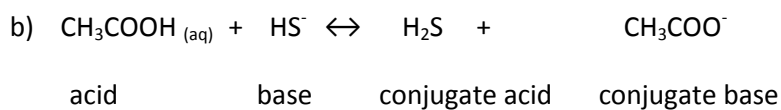
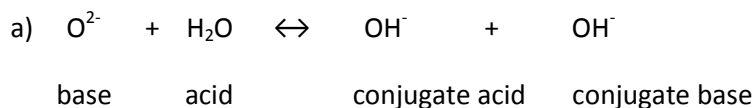
$\text{C}_2\text{O}_4^{2-}$ is a conjugate base of HC_2O_4^-

16.27

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



Answer:



The equilibrium to the left

16.31

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

A) $[\text{OH}^-] = 0.00045 \text{ M}$ b) $[\text{OH}^-] = 8.8 \times 10^{-9} \text{ M}$

c) a solution which $[\text{OH}^-]$ is 100 times greater than $[\text{H}^+]$

Answer:

a) $\text{pOH} = -\log [\text{OH}^-] = -\log (0.00045) = 3.35$

$\text{pOH} + \text{pH} = 14$

$\text{pH} = 10.65$

b) $\text{pOH} = -\log [\text{OH}^-] = -\log (8.8 \times 10^{-9}) = 7.06$

$\text{pOH} = 14.00 - 7.06 = 6.94$ the solution is acidic

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

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

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

16.33

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

Answer:

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

$$[OH^-] = [H^+] = 3.5 \times 10^{-6} \text{ M}$$

16.35

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

- a) 2.00 units b) 0.50 units

Answer:

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

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

16.39

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

<u>H⁺</u>	<u>OH⁻</u>	<u>pH</u>	<u>pOH</u>	<u>acidic or basic</u>
$7.5 \times 10^{-3} \text{ M}$				
	$3.6 \times 10^{-10} \text{ M}$			
		8.25		
			5.70	

Answer:

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

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

$$\text{pOH} = -\log(1.3 \times 10^{-12}) = 11.87$$

$$\text{pH} = 14 - 11.87 = 2.13 \quad \text{the solution is acidic}$$

<u>H⁺</u>	<u>OH⁻</u>	<u>pH</u>	<u>pOH</u>	<u>acidic or basic</u>
$7.5 \times 10^{-3} \text{ M}$	$1.3 \times 10^{-12} \text{ M}$	2.13	11.87	acidic
$2.8 \times 10^{-5} \text{ M}$	$3.6 \times 10^{-10} \text{ M}$	4.56	9.44	acidic
$5.6 \times 10^{-9} \text{ M}$	$1.8 \times 10^{-6} \text{ M}$	8.25	5.75	basic
$5.0 \times 10^{-9} \text{ M}$	$2.0 \times 10^{-6} \text{ M}$	8.30	5.70	basic

16.41

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

Answer:

$$\text{pH} = -\log[\text{H}^+] = 7.40$$

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

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

$$\text{pOH} = -\log(0.6 \times 10^{-6}) = 6.22$$

16.45

Calculate the pH of the following strong acid solutions.

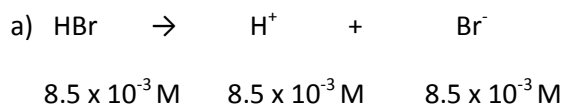
a) $8.5 \times 10^{-3} \text{ M HBr}$.

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:



$$\text{pH} = -\log(8.5 \times 10^{-3}) = 2.07$$

b) mole of $\text{HNO}_3 = (1.52 \text{ g}) / (63 \text{ g/mol}) = 0.024 \text{ mol}$

$$\text{molarity of } \text{HNO}_3 = (0.024 \text{ mol } \text{HNO}_3) / (1000 \text{ mL} / 575 \text{ mL}) = 0.042 \text{ M } [\text{H}^+]$$

$$\text{pH} = -\log [\text{H}^+] = -\log(0.042) = 1.38$$

c) $N_1 \times V_1 = N_2 \times V_2$

$$0.250 \text{ M} \times 5.00 \text{ mL} = N_2 \times 50.0 \text{ mL}$$

$$N_2 = 0.0250 \text{ M molarity of } \text{HClO}_4 = \text{molarity of } [\text{H}^+]$$

$$\text{pH} = -\log [\text{H}^+] = -\log(0.0250) = 1.60$$

d) $10 \text{ mL} + 20 \text{ mL} = 30 \text{ mL}$ volume of solution

$$N_1 \times V_1 = N_2 \times V_2$$

$$0.100 \text{ M} \times 10 \text{ mL} = N_2 \times 30 \text{ mL} \quad N_2 = 0.033 \text{ M of } [\text{HBr}] = [\text{H}^+]$$

$$N_1 \times V_1 = N_2 \times V_2$$

$$0.2 \text{ M} \times 20 \text{ mL} = N_2 \times 30 \text{ mL} \quad N_2 = 0.133 \text{ M of } \text{HCl} = [\text{H}^+]$$

$$0.033 \text{ M } [\text{H}^+] \text{ from } \text{HBr} + 0.133 \text{ M } [\text{H}^+] \text{ from } \text{HCl} = 0.166 \text{ M}$$

$$\text{pH} = -\log(0.166) = 0.778$$

16.47

Calculate $[\text{OH}^-]$ and pH for

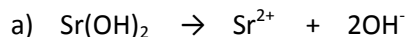
a) $1.5 \times 10^{-3} \text{ M } \text{Sr}(\text{OH})_2$

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 $9.5 \times 10^{-2} \text{ M } \text{Ca}(\text{OH})_2$

Answer:



$$[\text{OH}^-] = 2 \times 1.5 \times 10^{-3} = 3.0 \times 10^{-3} \text{ M}$$

$$\text{pOH} = -\log(3.0 \times 10^{-3}) = 2.523$$

$$\text{pH} = 14 - 2.523 = 11.48$$

b) mole of LiOH = $2.250 / 24 = 0.094$

$$[\text{LiOH}] \text{ M} = (0.094 \text{ mol LiOH}) (1000 \text{ mL} / 250 \text{ mL}) = 0.375 \text{ M of LiOH}$$



$$0.375 \qquad \qquad \qquad 0.375$$

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

$$\text{pH} = 14 - 0.426 = 13.57$$

c) $N_1 \times V_1 = N_2 \times V_2$

$$0.175 \times 1.0 \text{ mL} = N_2 \times 2000 \text{ mL} \quad N_2 = 8.75 \times 10^{-5} \text{ M NaOH}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log (8.75 \times 10^{-5}) = 4.058$$

$$\text{pH} = 14 - 4.058 = 9.942$$

d) total volume of solution $5.0 \text{ mL} + 15 \text{ mL} = 20 \text{ mL}$

for KOH $N_1 \times V_1 = N_2 \times V_2$

$$0.105 \text{ M} \times 5.0 \text{ mL} = N_2 \times 20.0 \text{ mL} \quad N_2 = 0.02625 \text{ M}$$

$$\text{KOH for Ca(OH)}_2 \quad 9.5 \times 10^{-2} \text{ M} \times 15 \text{ mL} = N_2 \times 20 \text{ mL}$$

$$N_2 = 0.07125 \text{ M of Ca(OH)}_2 = 0.07125 \text{ M } [\text{OH}^-]$$

$$\text{total concentrations of } [\text{OH}^-] = 0.02625 + 2 \times 0.07125 = 0.16875 \text{ M} = 0.17 \text{ M}$$

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

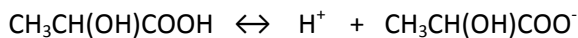
$$\text{pH} = 14 - 0.773 = 13.23$$

16.53

Lactic acid ($\text{CH}_3\text{CH}(\text{OH})\text{COOH}$) has one acidic hydrogen. A 0.10 M solution lactic acid has $\text{pH} = 2.44$.

Calculate K_a .

Answer:



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

$$[\text{H}^+] = 0.00363 \text{ M}$$

$$K_a = [\text{CH}_3\text{CH}(\text{OH})\text{COO}^-][\text{H}^+] / [\text{CH}_3\text{CH}(\text{OH})\text{COOH}] = (0.000363)(0.00363) / 0.1$$

$$K_a = 1.32 \times 10^{-4} = 1.4 \times 10^{-4}$$

16.55

A 0.10 M solution of chloroacetic acid (ClCH_2COOH) is 11.0% ionized. Using this information, calculate $[\text{ClCH}_2\text{COO}^-]$, $[\text{H}^+]$, and $[\text{ClCH}_2\text{COOH}]$ and K_a for chloroacetic acid.

Answer:

$$\text{Ionizes chloroacetic acid} = 0.10 \text{ M} \times 11/100 = 0.011 \text{ M}$$

$$\text{Remain chloroacetic acid} = 0.10 - 0.011 = 0.089 \text{ M}$$

	$\text{ClCH}_2\text{COOH}_{(\text{aq})}$	\leftrightarrow	$\text{ClCH}_2\text{COO}^-_{(\text{aq})}$	$+$	$\text{H}^+_{(\text{aq})}$
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

$$K_a = (0.011)(0.011) / (0.089) = 1.4 \times 10^{-3}$$

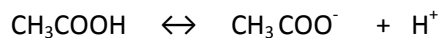
16.57

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

Answer:

$$\text{pH} - \log [\text{H}^+] = 2.90$$

$$[\text{H}^+] = 0.00126 = \text{the concentration of vinegar}$$



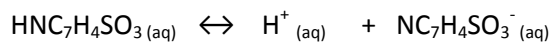
$$K_a = [\text{CH}_3\text{COO}^-][\text{H}^+] / [\text{CH}_3\text{COOH}] = 1.8 \times 10^{-5}$$

$$(0.00126)(0.00126) / [\text{CH}_3\text{COOH}] = 1.8 \times 10^{-5}$$

$$[\text{CH}_3\text{COOH}] = 0.0882 \text{ M}$$

16.63

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



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

Answer:

$$pK_a = -\log K_a = 2.32$$

$$K_a = 4.79 \times 10^{-3} = \frac{[\text{H}^+][\text{NC}_7\text{H}_4\text{SO}_3^-]}{[\text{HNC}_7\text{H}_4\text{SO}_3]}$$



initial	0.10 M	0 M	0 M
change	- X M	+ X M	+ X M
equilibrium	(0.10 - X) M	X M	X M

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

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

$$\text{pH} = -\log (2.23 \times 10^{-2}) = 1.652 = 1.7$$

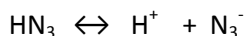
16.65

Calculate the percent ionization of hydrazoic acid (HN_3) in solutions each of the following concentrations ($K_a = 1.9 \times 10^{-5}$)

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

Answer:

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



$$K_a = \frac{[\text{H}^+][\text{N}_3^-]}{[\text{HN}_3]} = 1.9 \times 10^{-5} \quad X = [\text{H}^+] = [\text{N}_3^-]$$

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

$$\text{Percent ionization} = (2.77 \times 10^{-3}) / (0.4) \times 100 = 0.69\%$$

$$\text{b) } \frac{[\text{H}^+][\text{N}_3^-]}{[\text{HN}_3]} = 1.9 \times 10^{-5} \text{ M}$$

$$X^2 / (0.1 - X) = 1.9 \times 10^{-3} \quad X = 1.38 \times 10^{-3} \text{ M} = [\text{H}^+]$$

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

$$\text{c) } X^2 / (0.04 - X) = 1.9 \times 10^{-5} \quad X = 0.872 \times 10^{-3} \text{ M} = [\text{H}^+]$$

$$\text{percent ionization} = (0.872 \times 10^{-3}) / 0.04 \times 100 = 2.2\%$$

16.73

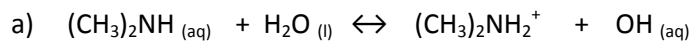
Write the chemical equations and Kb expression for the ionization of each of the following bases in aqueous solutions.

a) dimethyl amine $(\text{CH}_3)_2\text{NH}$

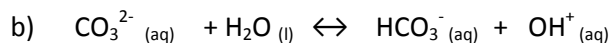
b) carbonate ion CO_3^{2-}

c) formate ion CHO_2^-

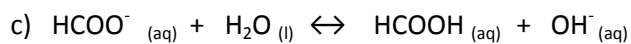
Answer:



$$K_b = [(\text{CH}_3)_2\text{NH}_2^+] [\text{OH}^-] / [(\text{CH}_3)_2\text{NH}]$$



$$K_b = [\text{HCO}_3^-] [\text{OH}^-] / [\text{CO}_3^{2-}]$$



$$K_b = [\text{HCOOH}] [\text{OH}^-] / [\text{HCOO}^-]$$

16.83

Calculate $[\text{OH}^-]$ and pH for each of the following solutions

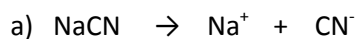
K_a for HCN = 4.9×10^{-10} , K_a for H_2CO_3 = 5.6×10^{-11} , K_a for HNO_2 = 4.5×10^{-4})

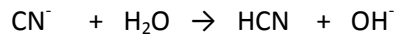
a) 0.10 M NaCN

b) 0.080 M Na_2CO_3

c) a mixture that is 0.10 M in NaNO_2 and 0.20 M in $\text{Ca}(\text{NO}_2)_2$

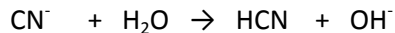
Answer:





$$K_b \times K_a = K_w \quad K_b = (1.0 \times 10^{-14}) / (4.9 \times 10^{-10}) = 0.204 \times 10^{-4}$$

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



initial	0.10 M	0 M	0 M
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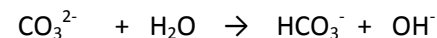
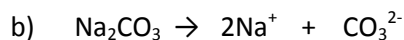
change	- X M	+ X M	+ X M
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equilibrium	0.10 - X	X M	X M
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$$(X)(X) / 0.1 - X = 0.204 \times 10^{-4} \quad X = 1.43 \times 10^{-3} \text{ M} = [\text{OH}^-]$$

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

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



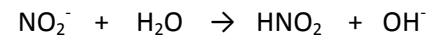
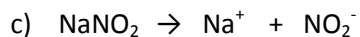
$$K_b = (1.0 \times 10^{-14}) / (5.6 \times 10^{-11}) = 0.18 \times 10^{-3}$$

$$K_b = [\text{HCO}_3^-] [\text{OH}^-] / [\text{CO}_3^{2-}] = 0.18 \times 10^{-3}$$

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

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

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

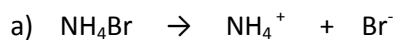


16.85

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

Answer:



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.

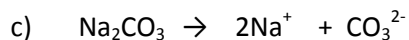
The solution of the salt is acidic



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

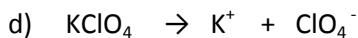
The solution of the salt is acidic



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

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

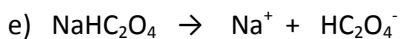
The solution of the salt is basic



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



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

HC_2O_4^- is a conjugate acid of a base $\text{C}_2\text{O}_4^{2-}$

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?

(K_b for the $\text{F}^- = 1.5 \times 10^{-11}$, K_b for $\text{ClO}^- = 0.334 \times 10^{-6}$, Cl^- is from strong acid HCl)

Answer:

$$\text{pH} + \text{pOH} = 14$$

$$8.08 + \text{pOH} = 14 \quad \text{pOH} = 5.92$$

$$\text{pOH} = -\log [\text{OH}^-] = 5.92$$

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

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

$$K_b \times K_a = 10^{-14} \quad K_a = (1.0 \times 10^{-14}) / (1.45 \times 10^{-11}) = 7.14 \times 10^{-3}$$

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

Prepared by Dr. Nabil Nassory