

P G6 Salt Hydrolysis

Name _____ Per _____



$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Strong Acids



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

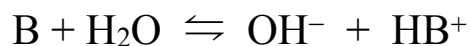
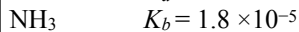
$$K_a \times K_b = K_w$$

$$M \times V = M \times V$$

**Really useful
information**

Equilibrium Konstants

K_a or K_b as appropriate for the listed substance.



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

Strong Bases



1. Which of the following salts, when dissolved in water, at 25°C, cause a change in pH? If there is a pH change, will the pH be above or below 7? Write a hydrolysis reaction that illustrates the cause of the pH change. Calculate the pH if 2.0 g of the salt is dissolved to make 100 ml of solution.

a. barium chloride, BaCl₂ (208.2 g/mol) Acidic, Basic or Neutral?

b. ammonium perchlorate, NH₄ClO₄ (117.492 g/mol) Acidic, Basic or Neutral?

c. sodium cyanide, NaCN (49.01 g/mol) Acidic, Basic or Neutral?

d. ammonium nitrate, NH₄NO₃ (80.052 g/mol) Acidic, Basic or Neutral?

e. potassium acetate, KC₂H₃O₂ (98.144 g/mol) Acidic, Basic or Neutral?

f. diethylamide bromide, (C₂H₅)₂NH₂Br (154.046 g/mol) Acidic, Basic or Neutral?

g. aluminum perchlorate, Al(ClO₄)₃ (325.33 g/mol) Acidic, Basic or Neutral?

1. Which of the following salts, when dissolved in water, at 25°C, cause a change in pH? If there is a pH change, will the pH be above or below 7? Write a hydrolysis reaction that illustrates the cause of the pH change. Calculate the pH if 2.0 g of the salt is dissolved in 100 ml of water.

a. barium chloride, BaCl₂ (208.2 g/mol)

pH will remain at 7. $\text{BaCl}_2 \rightarrow \text{Ba}^{2+} + 2 \text{Cl}^-$ (note the single arrow since BaCl₂ is a soluble ionic compound). The Ba²⁺ ions are “pathetic” meaning they do not hydrolyze in water and thus do not produce any extra H⁺ or OH⁻ ions. The Cl⁻ ions are also “pathetic” meaning they do not hydrolyze in water and thus do not produce any extra H⁺ or OH⁻ ions.

b. ammonium perchlorate, NH₄ClO₄ (117.492 g/mol)

pH will change to below 7, acidic. $\text{NH}_4\text{ClO}_3 \rightarrow \text{NH}_4^+ + \text{ClO}_3^-$ (note the single arrow since NH₄ClO₃ is a soluble ionic compound) The ClO₃⁻ ions are “pathetic” meaning they do not hydrolyze in water and thus do not produce any extra H⁺ or OH⁻ ions. But the NH₄⁺ ions are a conjugate WA and partially dissociate in water and produce H⁺ ions. $\text{NH}_4^+ \rightleftharpoons \text{H}^+ + \text{NH}_3$

$$2g \times \frac{1\text{mol}}{117.492g} = 0.017\text{mol} \quad \frac{0.017\text{mol}}{0.1L} = 0.17\text{M} \quad K_a \times K_b = K_w \quad \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10} \quad K_b = \frac{[\text{H}^+][\text{NH}_3]}{[\text{NH}_4^+]}$$

$$5.56 \times 10^{-10} = \frac{x^2}{[0.17\text{M}]} \quad x = [\text{H}^+] = 9.7 \times 10^{-6} \quad \text{pH} = 5.01$$

c. sodium cyanide, NaCN (49.01 g/mol)

pH will change to above 7, basic. $\text{NaCN} \rightarrow \text{Na}^+ + \text{CN}^-$ (note the single arrow since NaCN is a soluble ionic compound) The Na⁺ ions are “pathetic” meaning they do not hydrolyze in water and thus do not produce any extra H⁺ or OH⁻ ions. But the CN⁻ ions are a conjugate WB which hydrolyzes in water and produces extra OH⁻ ions. $\text{CN}^- + \text{H}_2\text{O} \rightleftharpoons \text{OH}^- + \text{HCN}$

$$2g \times \frac{1\text{mol}}{49.01g} = 0.0408\text{mol} \quad \frac{0.0408\text{mol}}{0.1L} = 0.408\text{M} \quad K_a \times K_b = K_w \quad \frac{1 \times 10^{-14}}{3.5 \times 10^{-4}} = 2.9 \times 10^{-11} \quad K_b = \frac{[\text{OH}^-][\text{HCN}]}{[\text{CN}^-]}$$

$$2.9 \times 10^{-11} = \frac{x^2}{[0.408]} \quad x = [\text{OH}^-] = 3.4 \times 10^{-6} \quad \text{pOH} = 5.47 \quad \text{pH} = 8.53$$

d. ammonium nitrate, NH₄NO₃ (80.052 g/mol)

pH will change to below 7, acidic. $\text{NH}_4\text{NO}_3 \rightarrow \text{NH}_4^+ + \text{NO}_3^-$ (note the single arrow since NH₄NO₃ is a soluble ionic compound) The NO₃⁻ ions are “pathetic” meaning they do not hydrolyze in water and thus do not produce any extra H⁺ or OH⁻ ions. But the NH₄⁺ ions are a conjugate WA and partially dissociate in water and produce H⁺ ions. $\text{NH}_4^+ \rightleftharpoons \text{H}^+ + \text{NH}_3$

$$2g \times \frac{1\text{mol}}{80.052g} = 0.025\text{mol} \quad \frac{0.025\text{mol}}{0.1L} = 0.250\text{M} \quad K_a \times K_b = K_w \quad \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10} \quad K_b = \frac{[\text{H}^+][\text{NH}_3]}{[\text{NH}_4^+]}$$

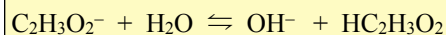
$$5.56 \times 10^{-10} = \frac{x^2}{[0.25]} \quad x = [\text{H}^+] = 1.2 \times 10^{-5} \quad \text{pH} = 4.93$$

Acid Base Equilibrium – Putting it all together

ANSWERS

e. potassium acetate, $\text{KC}_2\text{H}_3\text{O}_2$ (98.144 g/mol)

pH will change to above 7, basic. $\text{KC}_2\text{H}_3\text{O}_2 \rightarrow \text{K}^+ + \text{C}_2\text{H}_3\text{O}_2^-$ (note the single arrow since $\text{KC}_2\text{H}_3\text{O}_2$ is a soluble ionic compound) The K^+ ions are “pathetic” meaning they do not hydrolyze in water and thus do not produce any extra H^+ or OH^- ions. But the $\text{C}_2\text{H}_3\text{O}_2^-$ ions are a conjugate WB which hydrolyzes in water and produces extra OH^- ions.

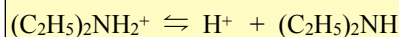


$$2g \times \frac{1\text{mol}}{98.144g} = 0.0204\text{mol} \quad \frac{0.0204\text{mol}}{0.1L} = 0.204M \quad K_a \times K_b = K_w \quad \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10} \quad K_b = \frac{[\text{OH}^-][\text{HC}_2\text{H}_3\text{O}_2]}{[\text{C}_2\text{H}_3\text{O}_2^-]}$$

$$5.56 \times 10^{-10} = \frac{x^2}{[0.204]} \quad x = [\text{OH}^-] = 1.1 \times 10^{-5} \quad pOH = 4.97 \quad pH = 9.03$$

f. diethylamide bromide, $(\text{C}_2\text{H}_5)_2\text{NH}_2\text{Br}$ (154.046 g/mol)

pH will change to below 7, acidic. $(\text{C}_2\text{H}_5)_2\text{NH}_2\text{Br} \rightarrow (\text{C}_2\text{H}_5)_2\text{NH}_2^+ + \text{Br}^-$ (note the single arrow since $(\text{C}_2\text{H}_5)_2\text{NH}_2\text{Br}$ is a soluble ionic compound, you should be able to recognize this because of the Br^- at the end of the molecule and the acid/base context of the question) The Br^- ions are “pathetic” meaning they do not hydrolyze in water and thus do not produce any extra H^+ or OH^- ions. But the $(\text{C}_2\text{H}_5)_2\text{NH}_2^+$ ions partially dissociate in water and produce H^+ ions.



$$2g \times \frac{1\text{mol}}{154.046g} = 0.013\text{mol} \quad \frac{0.013\text{mol}}{0.1L} = 0.130M \quad K_a \times K_b = K_w \quad \frac{1 \times 10^{-14}}{1.3 \times 10^{-3}} = 7.69 \times 10^{-12} \quad K_a = \frac{[\text{H}^+][(\text{C}_2\text{H}_5)_2\text{NH}]}{[(\text{C}_2\text{H}_5)_2\text{NH}_2^+]}$$

$$7.69 \times 10^{-12} = \frac{x^2}{[0.130]} \quad x = [\text{H}^+] = 9.99 \times 10^{-7} \quad pH = 6.00$$

g. aluminum perchlorate, $\text{Al}(\text{ClO}_4)_3$ (325.33 g/mol)

pH will change to below 7, acidic. $\text{Al}(\text{ClO}_4)_3 \rightarrow \text{Al}^{3+} + 3\text{ClO}_4^-$ (note the single arrow since ClO_4^- is a soluble ionic compound, you should be able to recognize this because of the ClO_4^- at the end and the acid/base context) The Al^{3+} ions interact with water and produce H^+ ions. But the ClO_4^- ions are “pathetic” meaning they do not hydrolyze in water and thus do not produce any extra H^+ or OH^- ions. ($\text{Al}(\text{H}_2\text{O})_6^{3+} \rightleftharpoons \text{Al}(\text{H}_2\text{O})_5(\text{OH})^{2+} + \text{H}^+$ (Note: You would never have to write this equation.))

$$2g \times \frac{1\text{mol}}{325.33g} = 0.00615\text{mol} \quad \frac{0.00615\text{mol}}{0.1L} = 0.0615M$$

$$K_a = \frac{[\text{H}^+][\text{Al}(\text{OH})^{2+}]}{[\text{Al}^{3+}]} 1.5 \times 10^{-5} = \frac{x^2}{[0.0615]} \quad x = [\text{H}^+] = 9.6 \times 10^{-4} \quad pH = 3.02$$