Key Worksheet 11 Acids & Bases: Defined, Conjugates, Strength, Kw, pH, and pOH

Objectives

Definitions

Arrhenius: An acid produces hydrogen ions in water. A base produces hydroxide ions in water.
Bronsted – Lowry: An acid is a proton door. A base is a proton acceptor.

Lewis: An acid is an electron pair acceptor. A base is an electron pair donor.

Conjugates

The conjugate base of an acid has a proton removed from the acid.

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

Acid **Conjugate Base**

The conjugate acid of a base has a proton added to the base.

$$CH_3NH_2(aq) + H^+(aq) \rightleftharpoons CH_3NH_3^+(aq)$$

Base **Conjugate Acid**

Strength

The strength of an acid depends on two (related) factors. How easy it is to break the bond to the acidic proton and how stable the conjugate base is. The longer a bond is the easier it is to break and the stronger the acid. The more polar a bond is the easier it is to break and the stronger the acid. The more stable the conjugate base is the easier it is to form and the stronger the acid. The more delocalized the charge is on the conjugate base the more stable it is and the stronger the acid. The more electronegative an element is the more stable is is with a negative charge and the stronger it's conjugate acid.

K_w

$[H^+][OH^-] = 1.00 \times 10^{-14}$ at 25 °C (Different at different temperatures!)

In a neutral solution at 25 °C $[H^+] = [OH^-] = 1.00 \times 10^{-7}$ *M* in pure water (different with acidic or basic solutes added or at a different temperature).

pH & pOH

$$\text{pH} = -\log\left[H^+\right] \qquad \text{pOH} = -\log\left[OH^-\right]$$

$$[H^+] = 10^{-\text{pH}} \qquad [OH^-] = 10^{-\text{pOH}} \qquad \text{pH} + \text{pOH} = 14.00\ (\text{at } 25\ ^\circ\text{C})$$

1.) Label the acid and the base in the following reactions. Label the conjugate base (C.B.) and the conjugate acid (C.A.) in the first three equations.

$$HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$$

Acid **Base** **C.B.** **C.A.**

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^- aq + H_3O^+(aq)$$

Acid **Base** **C.B.** **C.A.**

$$H_2O(l) + KF(aq) \rightleftharpoons KOH(aq) + HF(aq)$$

Acid **Base** **C.B.** **C.A.**

$$BF_3(aq) + NaF(aq) \rightarrow NaBF_4(aq)$$

Acid **Base**

$$CH_3CH_2CH_2CH_2CH_2CH_2Cl + AlCl_3 \rightarrow CH_3CH_2CH_2CH_2CH_2CH_2^+ + AlCl_4^-$$

Base **Acid**

2.) Label the stronger acid of each pair. Explain why it is the stronger acid.

For all of these you should draw out the Lewis structures of both acids.

HNO_3 or $HClO_4$

$HClO_4$ *is the stronger of the two because the conjugate base has more resonance forms, the negative charge is more delocalized.*

$CH_2ClCOOH$ or CH_3COOH

$CH_2ClCOOH$ *is the stronger of the two acids because the* Cl *is more electron withdrawing than* H*, making the* O–H *bond weaker and stabilizing the conjugate base.*

HNO_2 or HCN

HNO_2 *is the stronger acid because there are more resonance forms for the conjugate base. The* O's *are also more electron withdrawing than the* C.

HF or HCl

HCl *is the stronger acid because the Cl atom is larger, making the* H–Cl *bond longer, weaker, and easier to break.*

$HSbF_6$ or HI

$HSbF_6$ *is much stronger (it is called a super acid) because the 6 highly electronegative F atoms are very electron withdrawing and delocalize the negative charge on the conjugate base very effectively, making the conjugate base very stable and thus easy to form.*

3.) At 100 °C the ion–product constant for water, K_w, is 5.13 x 10^{-13}. What is the pH of pure water when it boils at 100 °C?

$$[H^+][OH^-] = x^2 = 5.13 \times 10^{-13} \Rightarrow x = 7.16_2 \times 10^{-7}$$

$$x = [H^+] = 7.16_2 \times 10^{-7}$$

$$pH = -\log(7.16_2 \times 10^{-7}) = 6.144_9$$

pH = <u>6.145</u>

4.) What is the pH of a solution that is made by dissolving 5.199 g of NaOH (a strong electrolyte) in enough water to make 475 mL of solution?

$$\frac{5.199 \text{ g NaOH}}{39.997 \text{ g/mol}} = 0.1299_8 \text{ mol NaOH} \Rightarrow 0.1299_8 \text{ mol OH}^-$$

$$\frac{0.1299_8 \text{ mol OH}^-}{0.475 \text{ L}} = 0.273_6 \; M \text{ OH}^-$$

$$\text{pOH} = -\log(0.273_6) = 0.562_8$$

$$\text{pH} = 14.00 - 0.562_8 = 13.43_7$$

pH = <u>13.44</u>

5.) What is the $[OH^-]$ in a solution that has a pH of 4.922?

$$\text{pOH} = 14.00 - \text{pH} = 14.00 - 4.922 = 9.07_8$$

$$[OH^-] = 10^{-\text{pOH}} = 10^{-9.07_8} = 8.3_5 \times 10^{-10} \; M$$

$[OH^-]$ = <u>8.4 x 10^{-10} *M*</u>

6.) What is the pOH of a solution that is made by adding 729 mL of D.I. water to 25.0 mL of 15.8 *M* nitric acid?

First find the concentration of nitric acid using the dilution equation. Since nitric acid is a strong acid we know that the $[H^+] = [HNO_3]$. *We can then get the* pH, *and then the* pOH.

$$C_1V_1 = C_2V_2 \Rightarrow C_2 = \frac{C_1V_1}{V_2} = \frac{(15.8 \; M)(25.0 \text{ mL})}{754 \text{ mL}} = 0.523_8 \; M = [H^+]$$

$$\text{pH} = -\log(0.523_8) = 0.280_7$$

$$\text{pOH} = 14.00 - \text{pH} = 14.00 - 0.280_7 = 13.71_9$$

pOH = 13.72

7.) Consider the following reaction:

$$CH_3COO^- + HF \rightleftharpoons CH_3COOH + F^-$$

Given that K_a for HF is 6.6×10^{-4} and the K_a for CH_3COOH is 1.8×10^{-5} (both at 25°C), is K for this reaction at 25°C less than, greater than, or equal to 1? Explain.

Because the K_a for hydrofluoric acid is greater than that of acetic acid we can see that the equilibrium written above will favor the right side of the arrows. This is because a larger Ka means hydrofluoric acid is a stronger acid than acetic acid. This means that hydrofluoric acid releases it's proton more readily than acetic acid does. It alsdo means that the conjugate base of acetic acid, the acetate ion, is a stronger base than the conjugate base of hydrofluoric acid, the fluoride ion. The stronger conjugate base (acetate) pulls the proton off of the acid (hydrofluoric acid) in the left side of the arrows more strongly than the fluoride ion pulls them off of acetic acid on the right side of the arrows. This means that at equilibrium there will be more products than reactants, making $K > 1$.