Important Notes

- The **favored direction of the reaction** is the one in which the weaker acid/base are produced.
- The **stronger** an acid is, the weaker its conjugate base (and vice versa).
- Diluting an acid (↓ concentration) will ↑ pH and ↑ percent ionization.

Practice, practice, practice!

HCl = strong acid!

1. Calculate the pH of a 0.020 *M* solution of hydrochloric acid.

$$pH = -\log[H^+] = -\log[HCl] = -\log(\underbrace{0.020}_{2\text{ s.f.}}) = \boxed{1.70}$$

2 s.f. after decimal in pH!

$HC_2H_3O_2$ = weak acid!

2. Calculate the pH and percent ionization of a 0.020 *M* solution of acetic acid ($K_a = 1.8 \times 10^{-5}$).

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = \frac{x^2}{[HC_2H_3O_2]} = \frac{x^2}{\underbrace{0.020 - x}_{\text{Assume } x \ll 0.020}} \approx \frac{x^2}{0.020} = 1.8E-5$$

$$x = \sqrt{(1.8E-5)(0.020)} = 6.0E-4\ M = [H_3O^+]$$

$$pH = -\log(6.0E-4) = \boxed{3.22}$$

$$\%\,Ion. = \frac{[H_3O^+]}{[HC_2H_3O_2]_i} \times 100 = \frac{6.0E-4}{0.020} \times 100 = \boxed{3.0\%}$$

3. Calculate the pH for a 1.7×10^{-2} *M* solution of KOH.

KOH: strong base!

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

$$= -\log[KOH] = -\log(1.7E-2) = 1.77$$

$$pH = 14 - pOH = 14 - 1.77 = \boxed{12.23}$$

NH_3: weak base!

4. Calculate the pH and percent ionization for a 1.7×10^{-2} *M* solution of NH_3 ($K_b = 1.8 \times 10^{-5}$).

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{x^2}{[NH_3]} = \frac{x^2}{\underbrace{(1.7E-2) - x}_{\text{assume } x \ll 1.7E-2}} \approx \frac{x^2}{1.7E-2} = 1.8E-5$$

$$x = \sqrt{(1.8E-5)(1.7E-2)} = 5.5E-4\ M = [OH^-]$$

$$pOH = -\log(5.5E-4) = 3.26$$

$$pH = 14 - pOH = 14 - 3.26 = \boxed{10.74}$$

$$\%\,Ion = \frac{[OH^-]}{[NH_3]_i} \times 100 = \frac{5.5E-4}{1.7E-2} \times 100 = \boxed{3.3\%}$$