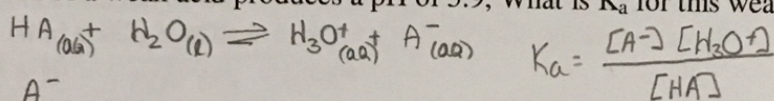


- 3) If a 0.05 M solution of a weak acid produces a pH of 5.9, What is K_a for this weak acid?

$$[H_3O^+] = 10^{-5.9} = 1.26 \times 10^{-6} = x$$



| | | | |
|---|------|---------------------------|---------------------------|
| | HA | H_3O^+ | A^- |
| I | 0.05 | 0 | 0 |
| C | -x | +x | +x |
| E | ↓ | $x = 1.26 \times 10^{-6}$ | $x = 1.26 \times 10^{-6}$ |

$0.05 - 1.26 \times 10^{-6} = 0.05$

$$K_a = \frac{[1.26 \times 10^{-6}][1.26 \times 10^{-6}]}{[0.05]} = 3.2 \times 10^{-11}$$

- 7) Calculate the molar solubility of barium fluoride (BaF_2) in water. $BaF_2(s) \xrightleftharpoons{H_2O} Ba^{2+}_{(aq)} + 2F^-_{(aq)}$

$K_{sp} = 1.6 \times 10^{-6}$

Calculate the molar solubility of this compound in an aqueous solution that is 0.2 M NaF (sodium fluoride).

Exam 3

| | | | |
|---|---------|-----------|--------|
| | BaF_2 | Ba^{2+} | $2F^-$ |
| I | solid | 0 | 0 |
| C | -x | +x | +2x |
| E | solid | x | 2x |

not on exam 2 this is ch 17 - exam 3

$$K_{sp} = 1.6 \times 10^{-6} = [Ba^{2+}][F^-]^2 = [x][2x]^2 = 4x^3$$

$$\frac{1.6 \times 10^{-6}}{4} = \frac{4x^3}{4} \quad x^3 = 4 \times 10^{-7} \quad x = 7.4 \times 10^{-3}$$

molar solubility = $7.4 \times 10^{-3} \text{ mol/L}$

How much BaF_2 dissolved

NaF complete dissociation

| | | | |
|---|---------|-----------|---------|
| | BaF_2 | Ba^{2+} | $2F^-$ |
| I | solid | 0 | 0.2M |
| C | -x | +x | +2x |
| E | solid | +x | 0.2M+2x |

going to assume can ignore

$$K_{sp} = [x][0.2+2x]^2 = x[0.2]^2$$

$$1.6 \times 10^{-6} = 0.04x$$

$$x = 4.0 \times 10^{-5} = \text{molar solubility of } BaF_2 \text{ in } 0.2M \text{ NaF}$$

- 10) What is the pH of the following solutions?

a) 0.25 M HNO_3 strong acid

$$[H_3O^+] = 0.25M$$

$$pH = -\log(0.25) = 0.602$$

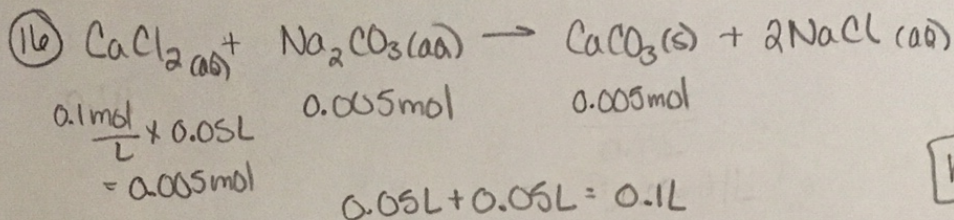
b) 0.17 M $Ba(OH)_2$ strong base

$$[OH^-] = 0.17M \times 2 = 0.34M$$

$$pOH = -\log(0.34) = 0.47$$

$$pH = 14 - 0.47 = 13.53$$

Exam 3



not on exam 2
ch 17

| | | | | | |
|---|--------------------|----------------------|----------------------|---|------------------------|
| | $\text{CaCO}_3(s)$ | \rightleftharpoons | $\text{Ca}^{2+}(aq)$ | + | $\text{CO}_3^{2-}(aq)$ |
| I | Solid | | 0 | | 0 |
| C | -x | | +x | | +x |
| E | solid | | x | | x |

$$K_{sp} = [\text{Ca}^{2+}][\text{CO}_3^{2-}]$$

$$3.8 \times 10^{-9} = (x)(x) = x^2$$

$$x = 6.2 \times 10^{-5} \text{ mol/L}$$

Dissolved $\text{Ca}^{2+} = 6.2 \times 10^{-5} \frac{\text{mol}}{L}$

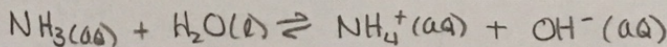
16) What is the concentration of Ca^{2+} ion remaining after CaCO_3 precipitates when 50.0 ml of 0.10 M CaCl_2 is added to 50.0 ml of 0.10 M Na_2CO_3 ? K_{sp} for CaCO_3 is 3.8×10^{-9} .

not on exam 2

What is the pH of a 0.15 M solution of NH_3 ? $K_b = 1.78 \times 10^{-5}$

weak base

| | | | |
|---|---------------|-----------------|---------------|
| | NH_3 | NH_4^+ | OH^- |
| I | 0.15 | 0 | 0 |
| C | -x | +x | +x |
| E | 0.15-x | x | x |



$$1.78 \times 10^{-5} = \frac{(x)(x)}{(0.15-x)} = \frac{x^2}{0.15}$$

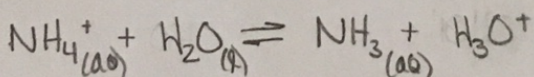
$$x^2 = 2.67 \times 10^{-6}$$

$$x = 1.634 \times 10^{-3} = [\text{OH}^-]$$

What is the pH of a 0.15 M solution of NH_4^+ ?

$$\text{pOH} = -\log(1.634 \times 10^{-3}) = 2.79$$

$$\text{pH} = 14 - 2.79 = 11.21$$



$$\text{NH}_4^+ \text{ is conjugate acid of } \text{NH}_3, \text{ so } K_a = \frac{1 \times 10^{-14}}{K_b} = \frac{1 \times 10^{-14}}{1.78 \times 10^{-5}} = 5.62 \times 10^{-10}$$

| | | | |
|---|-----------------|---------------|------------------------|
| | NH_4^+ | NH_3 | H_3O^+ |
| I | 0.15 | 0 | 0 |
| C | -x | +x | +x |
| E | 0.15-x | x | x |

$$5.62 \times 10^{-10} = \frac{[x][x]}{[0.15-x]} = \frac{x^2}{0.15}$$

$$x^2 = 8.43 \times 10^{-11}$$

$$x = 9.18 \times 10^{-6} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log[9.18 \times 10^{-6}] = 5.04$$