Chemistry 112

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?

7) Calculate the molar solubility of barium fluoride (BaF_2) in water.
$K_{sp} = 1.6 \times 10^{-6}$.
Calculate the molar solubility of this compound in an aqueous solution that is 0.2 M NaF (sodium fluoride).

10) What is the pH of the following solutions?
a) 0.25 M HNO_3
b) 0.17 M $Ba(OH)_2$

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}.

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

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

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

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?

$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$ $\quad K_a = \frac{[A^-][H_3O^+]}{[HA]}$

	HA	H_3O^+	A^-
I	0.05	0	0
C	−x	+x	+x
E	↓	x = 1.26×10^{-6}	x = 1.26×10^{-6}

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

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

7) Calculate the molar solubility of barium fluoride (BaF_2) in water. $K_{sp} = 1.6 \times 10^{-6}$.
Calculate the molar solubility of this compound in an aqueous solution that is 0.2 M NaF (sodium fluoride). not on exam 2 — this is ch 17 - exam 3

$BaF_{2(s)} \xrightleftharpoons{H_2O} Ba^{2+}_{(aq)} + 2F^-_{(aq)}$

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

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

$\frac{1.6 \times 10^{-6}}{4} = \frac{4x^3}{4}$ $\quad x^3 = 4 \times 10^{-7}$ $\quad x = 7.4 \times 10^{-3}$ = How much BaF_2 dissolved

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

NaF complete dissociation

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

$K_{sp} = [x][0.2 + 2x]^2$ going to assume can ignore

$= x[0.2]$

$1.6 \times 10^{-6} = 0.2x$ $\quad \boxed{x = 8.0 \times 10^{-6} = \text{molar solubility of } BaF_2 \text{ in } 0.2M \text{ NaF}}$

10) What is the pH of the following solutions?

a) 0.25 M HNO_3 strong acid

$[H_3O^+] = 0.25M$

$pH = -\log(0.25)$

$= \boxed{0.602}$

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

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

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

$pH = 14 - 0.47 = \boxed{13.53}$

(16) $CaCl_{2(aq)} + Na_2CO_{3(aq)} \rightarrow CaCO_{3(s)} + 2NaCl_{(aq)}$

$\frac{0.1\,mol}{L} \times 0.05L = 0.005\,mol$ 0.005mol 0.005mol

0.05L + 0.05L = 0.1L

not on exam

ch17

$CaCO_{3(s)} \rightleftharpoons Ca^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)}$

	$CaCO_3$	Ca^{2+}	CO_3^{2-}
I	Solid	0	0
C	−x	+x	+x
E	solid	x	x

$K_{sp} = [Ca^{2+}][CO_3^{2-}]$

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

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

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

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

weak base

What is the pH of a 0.15 M solution of NH_3? Kb = 1.78×10^{-5}

$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$

	NH_3	NH_4^+	OH
I	0.15	0	0
C	−x	+x	+x
E	0.15−x	x	x

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

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

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

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

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

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

$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons NH_{3(aq)} + H_3O^+$

NH_4 is conjugate acid of NH_3, 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

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

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

$x = 9.18 \times 10^{-6} = [H_3O^+]$

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