

1. e 2. a 3. c 4. e (b, d) 5. b b. c 7. d 9. c 10. d 11. d 13. e

14. a

16. amphiprotic means that it can act as an acid or a base



17. 1. Write the reaction of HF and water

2. Determine the equilibrium concentrations

3. Substitute into the K_a equation



$$K_a = \frac{[\text{F}^-][\text{H}_3\text{O}^+]}{[\text{HF}]}$$

$$= \frac{(0.0059)(0.0059)}{0.0941}$$

$$= 3.7 \times 10^{-4}$$

\therefore the K_a is 3.7×10^{-4}



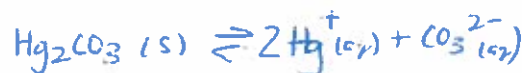
I			
C	-x	+x	+x
E		+x	+x

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

$$1.42 \times 10^{-7} = (x)(x)$$

$$x = 3.77 \times 10^{-4} \text{ mol/L}$$

NiCO_3 is more soluble



I			
C	-x	+2x	x
E		2x	x

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

$$3.6 \times 10^{-17} = (2x)^2 (x)$$

$$3.6 \times 10^{-17} = 4x^3$$

$$x = 2.08 \times 10^{-6} \text{ mol/L}$$

24. The Arrhenius definition did not take into account molecules and ions that acted as bases but didn't contain an OH^- .
 In the Bronsted-Lowry theory, the base didn't need an OH^- .

25. The acid donates an H^+ to water

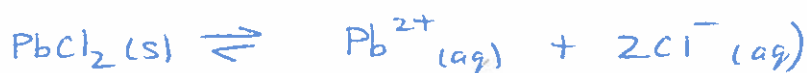


27. Common-ion effect \rightarrow the equilibrium will shift according to Le Chatelier principle when an ion is added that is part of the equilibrium system (away)

28. endpoint - when the indicator changes colour

equivalence point - when the moles acid = moles base stoichiometrically

32. a) pure water



I

C $-x$ $+x$ $+2x$

E $-x$ $+x$ $+2x$

$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{Cl}^-]^2$$

$$1.7 \times 10^{-5} = (x)(2x)^2$$

$$1.7 \times 10^{-5} = 4x^3$$

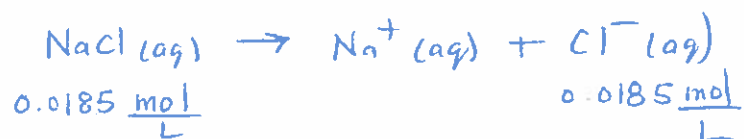
$$x = 1.6 \times 10^{-2} \text{ mol/L}$$

Check

$$\frac{0.0185}{1.7 \times 10^{-5}} = 1088 > 1000$$

$\therefore x$ is small w.r.t to 0.0185

b) in 0.0185 mol/L solution of NaCl



$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{Cl}^-]^2$$

$$1.7 \times 10^{-5} = (x)(0.0185 + 2x)^2$$

$$1.7 \times 10^{-5} = (x)(0.0185)^2$$

$$x = 0.04967$$

$$= 5.0 \times 10^{-2} \text{ mol/L}$$



I

0.0185

C $-x$ $+x$ $+2x$

0.0185 + 2x

E $-x$ $+x$

c) In b the equilibrium will shift to the left to offset the additional Cl^-
 \therefore the molar solubility decreases

$$38. [B] = 7.25 \text{ g} \times \frac{1 \text{ mol}}{146.21 \text{ g}} \times \frac{1}{1.00 \text{ L}}$$

$$= 0.04959 \frac{\text{mol}}{\text{L}}$$

$$\text{pH} = 9.12$$

$$\text{pOH} = 14 - 9.12$$

$$= 4.88$$

$$[\text{OH}^-] = 10^{-\text{pOH}}$$

$$= 10^{-4.88}$$

$$= 1.318 \times 10^{-5} \frac{\text{mol}}{\text{L}}$$



$$\text{I} \quad 0.04959$$

$$\text{C} \quad -1.318 \times 10^{-5}$$

$$1.318 \times 10^{-5} \quad + 1.318 \times 10^{-5}$$

$$\text{E} \quad 0.04959 - 1.318 \times 10^{-5}$$

$$1.318 \times 10^{-5} \quad 1.318 \times 10^{-5}$$

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

$$= \frac{(1.318 \times 10^{-5})^2}{(0.04959 - 1.318 \times 10^{-5})}$$

$$= 3.51 \times 10^{-9}$$

$$\therefore K_b \text{ is } 3.51 \times 10^{-9}$$

$$40. M_{\text{Ba}(\text{IO}_3)_2} = 487.129 \frac{\text{g}}{\text{mol}}$$

$$n_{\text{Ba}(\text{IO}_3)_2} = \frac{0.36 \text{ g}}{\text{L}} \times \frac{\text{mol}}{487.129 \text{ g}}$$

$$= 7.390 \times 10^{-4} \frac{\text{mol}}{\text{L}}$$

$$K_{sp} = [\text{Ba}^{2+}][\text{IO}_3^-]^2$$

$$= (7.390 \times 10^{-4})(1.478 \times 10^{-3})^2$$

$$= 1.61 \times 10^{-9}$$

$$\therefore K_{sp} \text{ is } 1.61 \times 10^{-9}$$



$$7.390 \times 10^{-4} \frac{\text{mol}}{\text{L}} \quad 7.390 \times 10^{-4} \frac{\text{mol}}{\text{L}} \quad 2(7.390 \times 10^{-4} \frac{\text{mol}}{\text{L}})$$

$$= 1.478 \times 10^{-3} \frac{\text{mol}}{\text{L}}$$

42. Look at table 8.1 to see who is a stronger acid HF or CH_3COOH . \therefore HF is stronger the equilibrium will lie on the side that has F^- (reactant side).

64. $K_b = 3.17 \times 10^{-5}$



I	$0.1 \frac{\text{mol}}{\text{L}}$	0	0
C	$-x$	$+x$	$+x$
E	$0.1 - x$	x	x

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

$$3.17 \times 10^{-5} = \frac{(x)(x)}{(0.1-x)}$$

$$\frac{0.1}{3.17 \times 10^{-5}} = 3155 > 1000$$

$\therefore x$ is small wrt to 0.1

$$3.17 \times 10^{-5} = \frac{x^2}{0.1}$$

$$x = 1.78 \times 10^{-3} \quad x \text{ can't be -ve}$$

$$[\text{OH}^-] = 1.78 \times 10^{-3}$$

$$\begin{aligned} \text{pOH} &= -\log [\text{OH}^-] \\ &= -\log (1.78 \times 10^{-3}) \\ &= 2.75 \end{aligned}$$

$$\begin{aligned} \text{pH} &= 14 - 2.75 \\ &= 11.25 \end{aligned}$$

\therefore pH is 11.25



$$\begin{array}{ccc} 2.8 \times 10^{-5} \frac{\text{mol}}{\text{L}} & 2(2.8 \times 10^{-5} \frac{\text{mol}}{\text{L}}) & 2.8 \times 10^{-5} \frac{\text{mol}}{\text{L}} \\ & = 5.6 \times 10^{-5} \frac{\text{mol}}{\text{L}} & \end{array}$$

$$\begin{aligned} K_{sp} &= [\text{X}^+]^2 [\text{Y}^{2-}] \\ &= (5.6 \times 10^{-5})^2 (2.8 \times 10^{-5}) \\ &= 8.8 \times 10^{-14} \end{aligned}$$

$\therefore K_{sp}$ is 8.8×10^{-14}



NH_4^+ came from a weak base NH_3

$\text{HNO}_3 \rightarrow$ strong acid



\therefore acidic pH

weak base + strong acid



$$n = 0.125 \text{ L} \times 3.5 \times 10^{-4} \frac{\text{mol}}{\text{L}}$$

$$= 4.375 \times 10^{-5} \text{ mol}$$

$$[\text{Cu}^+] = \frac{4.375 \times 10^{-5} \text{ mol}}{0.2 \text{ L}}$$

$$= 2.188 \times 10^{-4} \frac{\text{mol}}{\text{L}}$$

$$\text{total } V = 125 + 75$$

$$= 200 \text{ mL}$$

$$= 0.2 \text{ L}$$



$$n = 0.075 \text{ L} \times 8.1 \times 10^{-4} \frac{\text{mol}}{\text{L}}$$

$$= 6.075 \times 10^{-5} \text{ mol}$$

$$[\text{Cl}^-] = \frac{6.075 \times 10^{-5} \text{ mol}}{0.2 \text{ L}}$$

$$= 3.038 \times 10^{-4} \frac{\text{mol}}{\text{L}}$$



$$Q_{\text{sp}} = [\text{Cu}^+][\text{Cl}^-]$$

$$= (2.188 \times 10^{-4})(3.038 \times 10^{-4})$$

$$= 6.6 \times 10^{-8}$$

$$Q_{\text{sp}} < K_{\text{sp}}$$

$$6.6 \times 10^{-8} < 1.7 \times 10^{-7}$$

\therefore a precipitate will not form

