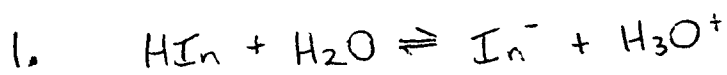


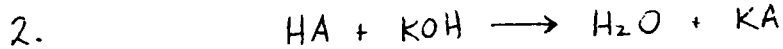
Acid Base III Worksheet 2



(a) yellow in acid solution - eq shifts left - HIn is predominant form

(b) blue in base solution - eq shifts right - In^- is predominant form

(c) in solutions with high $[\text{H}_3\text{O}^+]$ yellow HIn is in greater concentration. As $[\text{H}_3\text{O}^+]$ decreased solution reached a point where $[\text{HIn}] = [\text{In}^-]$ at this point the amount of yellow and blue is equal and green colour is observed. As $[\text{H}_3\text{O}^+]$ decreases further the blue In^- is the predominant form and the solution is blue.



(a) mol KOH = $18.4 \text{ mL} \times 0.100 \text{ M} = 1.84 \text{ mmol}$

ratio is 1:1 so mol HA = 1.84 mmol

$$[\text{HA}] = \frac{1.84 \text{ mmol}}{10.0 \text{ mL}} = 0.184 \text{ M}$$

(b) $0.220 \text{ g} \div 1.84 \text{ mmol} = 120 \text{ g/mole}$

(c) pH at stoichiometric endpoint is 9.600 \therefore acid is weak.

3. mol $\text{Ba}(\text{OH})_2 = 50.0 \text{ mL} \times 0.0200 \text{ M} = 1.00 \text{ mmol Ba}(\text{OH})_2$

mol $\text{OH}^- = 2 \times 1.00 \text{ mmol} = 2.00 \text{ mmol OH}^-$

mol $\text{H}^+ = 30.0 \text{ mL} \times 0.0100 \text{ M} = 0.300 \text{ mmol H}^+$

mol OH^- in excess = $2.00 - 0.300 = 1.70 \text{ mmol}$

$$[\text{OH}^-] = \frac{1.70 \text{ mmol}}{(50.0 + 30.0) \text{ mL}} = 0.0213 \text{ M} \quad \text{pOH} = 1.673 \quad \text{pH} = 12.327$$

4. $[\text{HCl}] = 10^{-2.63} = 2.34 \times 10^{-3} \text{ M}$

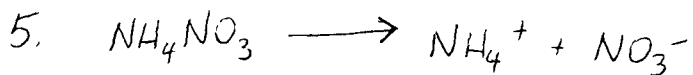
$$[\text{NaOH}] = 10^{-(14 - 10.82)} = 6.61 \times 10^{-4} \text{ M}$$

mol $\text{H}^+ = 2.34 \times 10^{-3} \text{ M} \times 30.0 \text{ mL} = 7.02 \times 10^{-2} \text{ mmol}$

mol $\text{OH}^- = 6.61 \times 10^{-4} \text{ M} \times 70.0 \text{ mL} = 4.63 \times 10^{-2} \text{ mmol}$

mol H^+ in excess = $7.02 \times 10^{-2} - 4.63 \times 10^{-2} = 2.39 \times 10^{-2} \text{ mmol}$

$$[\text{H}^+] = \frac{2.39 \times 10^{-2} \text{ mmol}}{(30.0 + 70.0) \text{ mL}} = 2.39 \times 10^{-4} \text{ M} \quad \text{pH} = 3.62$$



NO_3^- does not hydrolyze.

R	$\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$
I	1.00 0 0
C	-x +x +x
E	$1.00 - x \approx 1.00$ if x is small

$$[\text{NH}_4^+]_{\text{initial}} = \frac{2.0\text{g} \div 80.0\text{g/mol}}{0.0250\text{L}} = 1.00\text{M}$$

$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = 5.6 \times 10^{-10} = \frac{x^2}{1.00}$$

$$x = [\text{H}_3\text{O}^+] = 2.37 \times 10^{-5}\text{M} \quad \text{pH} = -\log(2.37 \times 10^{-5}\text{M}) = 4.62$$

6.

R	$\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$
I	2.0M 0 0
C	-x +x +x
E	$2.0\text{M} - x \approx 2.0\text{M}$ if x is small x x

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-10}} = 1.79 \times 10^{-5}$$

$$K_b = \frac{x^2}{2.0} = 1.79 \times 10^{-5} \quad [\text{OH}^-] = x = 5.98 \times 10^{-3}\text{M}$$

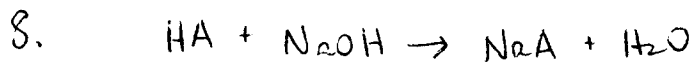
$$\text{pOH} = -\log(5.98 \times 10^{-3}) = 2.224$$

$$\text{pH} = \text{pK}_w - \text{pOH} = 14.00 - 2.224 = 11.78$$

7. Several answers apply:

eg. titration

pH at endpoint of titration



(a) $\text{mol OH}^- = 0.10\text{M} \times 20.0\text{mL} = 2.00\text{mmol}$

$\text{mol HA} = \text{mol OH}^- = 2.00\text{mmol}$

$$[\text{HA}] = \frac{2.00\text{mmol}}{10.0\text{mL}} = 0.20\text{M}$$

(b) $\text{pH at midpoint} = 5.67 = \text{pK}_a$

$$K_a = 10^{-5.67} = 2.1 \times 10^{-6}$$