

18. Indicate whether the following conditions would make a solution acidic, basic or neutral. [fall 02, ex 1]

pH = 10.11	acidic / basic / neutral
$[\text{OH}^-] = 1.0 \times 10^{-11}$	acidic / basic / neutral
$[\text{H}_3\text{O}^+] = 1.0 \times 10^{-7}$	acidic / basic / neutral

The answers in order are:

- (a) basic, basic, basic (b) basic, acidic, neutral (c) acidic, basic, neutral
(d) acidic, acidic, acidic (e) basic, basic, acidic
7. Which of the following conditions would make a solution **ACIDIC**? Mark all correct answers on your scantron. [fall 01, ex1]
- (a) pH = 0.25 (b) $[\text{H}_3\text{O}^+] = 5.9 \times 10^{-5} \text{ M}$ (c) $[\text{OH}^-] = 5.9 \times 10^{-5} \text{ M}$
(d) pOH = 4.55 (e) pOH = 8.45 (ab) pH = 9.71
16. What is the $[\text{H}_3\text{O}^+]$ to the correct number of significant figures in a pH 8.7412 solution? [fall 00, ex1]
- (a) $2 \times 10^{-9} \text{ M}$ (b) $1.8 \times 10^{-9} \text{ M}$ (c) $1.81 \times 10^{-9} \text{ M}$
(d) $1.815 \times 10^{-9} \text{ M}$ (e) $1.8147 \times 10^{-9} \text{ M}$
8. What is the $[\text{OH}^-]$ to the correct number of significant figures in an aqueous solution that has a pH of 3.785? [spring 00, ex1]
- (a) $6.1 \times 10^{-11} \text{ M}$ (b) $6.14 \times 10^{-11} \text{ M}$ (c) $6.144 \times 10^{-11} \text{ M}$
(d) $1.6 \times 10^{-4} \text{ M}$ (e) $1.64 \times 10^{-4} \text{ M}$ (ab) $1.641 \times 10^{-4} \text{ M}$
15. What is the pH of a $4.5 \times 10^{-3} \text{ M}$ aqueous solution of HBr? [spring 00, ex1]
- (a) 3.26 (b) 4.5 (c) 7.95 (d) 2.35 (e) 3.95
16. What is the pH to the correct number of significant figures of a $6.75 \times 10^{-3} \text{ M}$ solution of KOH? [spring 01, ex1]
- (a) 2.171 (b) 11.83 (c) 2.17 (d) 7.246 (e) 7.24 (ab) 11.826

9. What is the pH to the correct number of significant figures in a 6.83×10^{-2} M solution of $\text{Mg}(\text{OH})_2$? [fall 01, ex1]
(a) 0.864 (b) 12.83 (c) 12.8 (d) 13.1 (e) 13.132 (ab) 0.86
12. How many grams of HBr need to be added to 100.0 mL of water to make a pH 1.558 solution? Give your answer to the correct number of significant figures. [spring 02, ex1]
(a) 0.224 g (b) 0.0277 g (c) 292.6 g (d) 0.2239 g (e) 290 g
3. 6.75 g of KOH is dissolved in 1.25 L of water. What is the pH of the solution to the correct number of significant figures? [fall 01, ex1]
(a) 13.264 (b) 13.0 (c) 13.26 (d) 12.980 (e) 1.02
17. The pH of a 0.165 M solution of a certain weak acid is 3.72. What is K_a of the acid? [fall 01, ex1]
(a) 1.8×10^{-5} (b) 1.9×10^{-4} (c) 2.2×10^{-7} (d) 5.3×10^{-10} (e) 8.1×10^{-9}
17. The pH of a 0.10 M solution of a certain weak base is 10.85. What is K_b of the base? [spring 01, ex1]
(a) 8.2×10^{-5} (b) 5.5×10^{-4} (c) 7.1×10^{-4} (d) 1.8×10^{-3} (e) 5.1×10^{-6}
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13. Phenol, C_6H_5OH , is a weak acid with $K_a = 1.3 \times 10^{-10}$. What is the pH of a 0.050 M solution of phenol? [fall 02, ex 1]
(a) 11.19 (b) 2.81 (c) 8.41 (d) 5.59 (e) 3.89
6. What is the pH of a 0.20 M solution of $CH_3CH_2CO_2H$? [spring 01, ex1]
(This is the same type of question as the previous one, but it is more difficult because you need to recognize that $CH_3CH_2CO_2H$ is a weak acid and therefore know to look at the table of K_a 's given with the exam to find K_a of $CH_3CH_2CO_2H$. K_a of $CH_3CH_2CO_2H = 1.3 \times 10^{-5}$)
(a) 3.22 (b) 11.21 (c) 5.59 (d) 8.35 (e) 2.79
7. What is the pH of a 0.20 M solution of $Na[CH_3CH_2CO_2]$? [spring 01, ex1]
(a) 3.22 (b) 9.81 (c) 9.09 (d) 7.00 (e) 4.90
10. Hydroxylamine, $HONH_2$, is a weak base with $K_b = 1.07 \times 10^{-8}$. What is the pH of a 0.15M solution of $HONH_2$? [spring 02, ex1]
(a) 9.60 (b) 4.40 (c) 13.17 (d) 7.97 (e) 6.03
17. What is the pH of a 1.0 M solution of ethylamine, $CH_3CH_2NH_2$? [fall 00, ex1]
(This is the same type of question as the previous one, but it is more difficult because you need to recognize that $CH_3CH_2NH_2$ is a weak base and therefore know to look at the table of K_b 's given with the exam to find K_b of $CH_3CH_2NH_2$. K_b of $CH_3CH_2NH_2 = 6.4 \times 10^{-4}$)
(a) 10.81 (b) 1.60 (c) 14.00 (d) 5.40 (e) 12.40
18. What is the pH of a 1.0 M solution of ethylammonium chloride, $[CH_3CH_2NH_3]Cl$? [fall 00, ex1]
(a) 5.40 (b) 1.60 (c) 12.40 (d) 7.00 (e) 3.50

17. 18. Indicate whether the following conditions would make a solution acidic, basic or neutral.

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-11}} = 1.0 \times 10^{-3}$$
 which is acidic ($> 1.0 \times 10^{-7} M$)

pH = 10.11 acidic / **basic** / neutral
 $[OH^-] = 1.0 \times 10^{-11}$ acidic / **basic** / neutral
 $[H_3O^+] = 1.0 \times 10^{-7}$ acidic / basic / **neutral**

The answers in order are:

- (a) basic, basic, basic **(b) basic, acidic, neutral** (c) acidic, basic, neutral
 (d) acidic, acidic, acidic (e) basic, basic, acidic

7. Which of the following conditions would make a solution ACIDIC? Mark all correct answers on your scantron.

- A, B, E **(a) pH = 0.25** **(b) $[H_3O^+] = 5.9 \times 10^{-5} M$** (c) $[OH^-] = 5.9 \times 10^{-5} M$
 (d) pOH = 4.55 **(e) pOH = 8.45** (ab) pH = 9.71

ACIDIC
 $pH < 7$
 $[H_3O^+] > 1.0 \times 10^{-7} M$
 $pOH > 7$
 $[OH^-] < 1.0 \times 10^{-7} M$

BASIC
 $pH > 7$
 $[H_3O^+] < 1.0 \times 10^{-7} M$
 $pOH < 7$
 $[OH^-] > 1.0 \times 10^{-7} M$

16. What is the $[H_3O^+]$ to the correct number of significant figures in a pH 8.7412 solution?

- (a) $2 \times 10^{-9} M$ (b) $1.8 \times 10^{-9} M$ (c) $1.81 \times 10^{-9} M$ $\leftarrow 4 \text{ sig figs}$
(d) $1.815 \times 10^{-9} M$ (e) $1.8147 \times 10^{-9} M$

pH is a log value - the number of sig. figs in a log is the number of places to the right of the decimal point. pH 8.7412 therefore has 4 sig. figures so the answer has to have 4 sig. figures.

8. What is the $[OH^-]$ to the correct number of significant figures in an aqueous solution that has a pH of 3.785?

- $\leftarrow 3 \text{ sig. figs!! should be 3 sig figs in answer}$
- (a) $6.1 \times 10^{-11} M$ **(b) $6.14 \times 10^{-11} M$** (c) $6.144 \times 10^{-11} M$
 (d) $1.6 \times 10^{-4} M$ (e) $1.64 \times 10^{-4} M$ (ab) $1.641 \times 10^{-4} M$

$$[H_3O^+] = 10^{-pH} = 10^{-3.785} = 1.64059 \times 10^{-4} M$$

$$[OH^-] = \frac{1.008 \times 10^{-14}}{1.64059 \times 10^{-4}} = \boxed{6.14 \times 10^{-11}}$$

15. What is the pH of a $4.5 \times 10^{-3} M$ aqueous solution of HBr?

- (a) 3.26 (b) 4.5 (c) 7.95 **(d) 2.35** (e) 3.95

HBr is a strong acid

so $[HBr] = [H_3O^+] = 4.5 \times 10^{-3}$

$pH = -\log(4.5 \times 10^{-3}) = \boxed{2.35}$

16. What is the pH to the correct number of significant figures of a 6.75×10^{-3} M solution of KOH?

strong base \uparrow 3 sig. figs
 (a) 2.171 (b) 11.83 (c) 2.17 (d) 7.246 (e) 7.24 (ab) 11.826

$$[\text{OH}^-] = [\text{KOH}] = 6.75 \times 10^{-3}$$

$$[\text{H}_3\text{O}^+] = \frac{1.008 \times 10^{-14}}{6.75 \times 10^{-3}} = 1.493 \times 10^{-12} \text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (1.493 \times 10^{-12}) = \boxed{11.826}$$

\uparrow 3 sig figs \uparrow should be 3 sig. figs in answer.
 Since pH is a log, the # of sig. figs is the # of places to the right of, decimal pt

9. What is the pH to the correct number of significant figures in a 6.83×10^{-2} M solution of $\text{Mg}(\text{OH})_2$?

STRONG BASE! \uparrow 3 sig figs
 (a) 0.864 (b) 12.83 (c) 12.8 (d) 13.1 (e) 13.132 (ab) 0.86

$$[\text{OH}^-] = 2 [\text{Mg}(\text{OH})_2] = 2 (6.83 \times 10^{-2}) = 0.1366 \text{ M}$$

\uparrow there are 2 OH⁻ in each Mg(OH)₂ unit

$$[\text{H}_3\text{O}^+] = \frac{1.008 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.008 \times 10^{-14}}{0.1366} = 7.379 \times 10^{-14}$$

$$\text{pH} = -\log (\text{H}_3\text{O}^+) = -\log (7.379 \times 10^{-14}) = \boxed{13.132}$$

want 3 places to the right of the decimal pt for 3 sig figs

12. How many grams of HBr need to be added to 100.0 mL of water to make a pH 1.558 solution? Give your answer to the correct number of significant figures.

(a) 0.224 g (b) 0.0277 g (c) 292.6 g (d) 0.2239 g (e) 290 g

HBr is a strong acid so $[\text{HBr}] = [\text{H}_3\text{O}^+]$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-1.558} = 2.7669 \times 10^{-2} \text{ M}$$

\uparrow 3 sig. figs

So - Need to make 100 mL of $2.77 \times 10^{-2} \text{ M}$ soln of HBr

$$\text{mw of HBr} = 1.008 + 79.904 = 80.912 \text{ g/mol}$$

$$100. \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{2.77 \times 10^{-2} \text{ mol HBr}}{\text{L}} \times \frac{80.912 \text{ g HBr}}{\text{mol}} = \boxed{0.224 \text{ g}}$$

of sig figs in log is # of places to right of decimal - 3 in this case.

3. 6.75 g of KOH is dissolved in 1.25 L of water. What is the pH of the solution to the correct number of significant figures?

→ 3 sig figs!
(a) 13.264 (b) 13.0 (c) 13.26 (d) 12.980 (e) 1.02

$$\text{mw of KOH} = 39.102 + 15.999 + 1.008 = 56.109 \text{ g/mol}$$

KOH is a strong base - all of it converted to OH^-
 $\text{KOH (s)} \rightarrow \text{K}^+(\text{aq}) + \text{OH}^-(\text{aq})$

$$6.75 \text{ g KOH} \rightarrow \frac{\text{mol KOH}}{56.109 \text{ g}} \times \frac{1 \text{ mol OH}^-}{1 \text{ mol KOH}} = 0.1203 \text{ mol OH}^-$$

$$[\text{OH}^-] = \frac{0.1203 \text{ mol}}{1.25 \text{ L}} = 9.624 \times 10^{-2} \text{ M}$$

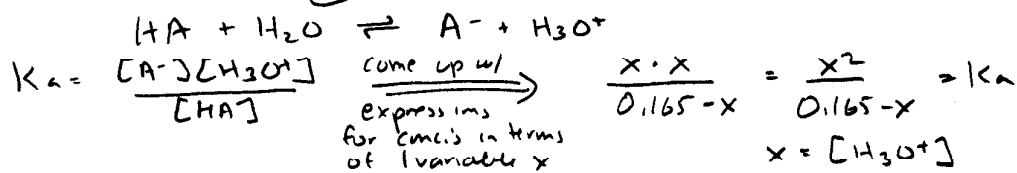
$$\text{pOH} = -\log(9.624 \times 10^{-2}) = 1.0166 \quad \text{pH} = 13.9965 - 1.0166$$

$$= 12.97986$$

$$= \boxed{12.980}$$

17. The pH of a 0.165 M solution of a certain weak acid is 3.72. What is K_a of the acid?

(a) 1.8×10^{-5} (b) 1.9×10^{-4} (c) 2.2×10^{-7} (d) 5.3×10^{-10} (e) 8.1×10^{-9}

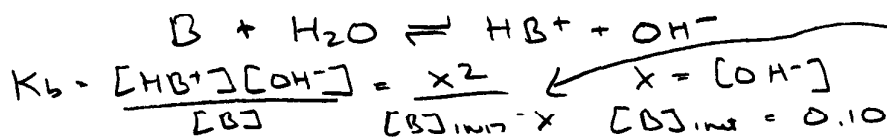


$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-3.72} = 1.905 \times 10^{-4} = x$$

$$K_a = \frac{x^2}{0.165 - x} = \frac{(1.905 \times 10^{-4})^2}{0.165 - 1.905 \times 10^{-4}} = \boxed{2.2 \times 10^{-7}}$$

17. The pH of a 0.10 M solution of a certain weak base is 10.85. What is K_b of the base?

(a) 8.2×10^{-5} (b) 5.5×10^{-4} (c) 7.1×10^{-4} (d) 1.8×10^{-3} (e) 5.1×10^{-6}



Calculate $[\text{OH}^-]$ from pH

$$[\text{H}_3\text{O}^+] = 10^{-10.85} = 1.4125 \times 10^{-11} \quad [\text{OH}^-] = \frac{1.008 \times 10^{-14}}{1.4125 \times 10^{-11}}$$

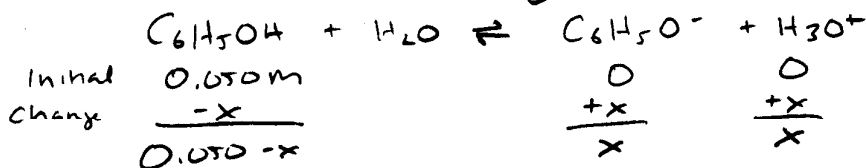
$$= 7.136 \times 10^{-4} = x$$

$$K_b = \frac{(7.136 \times 10^{-4})^2}{0.10 - 7.136 \times 10^{-4}} = \boxed{5.1 \times 10^{-6}}$$

Acid/Base Calculations

12. 15. Phenol, C_6H_5OH , is a weak acid with $K_a = 1.3 \times 10^{-10}$. What is the pH of a 0.050 M solution of phenol?

- (a) 11.19 (b) 2.81 (c) 8.41 (d) 5.59 (e) 3.89



$$K_a = \frac{[C_6H_5O^-][H_3O^+]}{[C_6H_5OH]} = \frac{x \cdot x}{0.050 - x} \approx \frac{x^2}{0.050} = 1.3 \times 10^{-10}$$

$$x = [H_3O^+]$$

$$x = \sqrt{(0.050)(1.3 \times 10^{-10})} = 2.55 \times 10^{-6} M = [H_3O^+]$$

$$pH = -\log(2.55 \times 10^{-6}) = \boxed{5.59}$$

6. What is the pH of a 0.20 M solution of $CH_3CH_2CO_2H$? WEAK ACID!

- (a) 3.22 (b) 11.21 (c) 5.59 (d) 8.35 (e) 2.79



$$K_a = \frac{[CH_3CH_2CO_2^-][H_3O^+]}{[CH_3CH_2CO_2H]} = 1.3 \times 10^{-5} = \frac{x^2}{0.20 - x} \approx \frac{x^2}{0.20}$$

$$x = [H_3O^+]$$

$$x = \sqrt{(0.20)(1.3 \times 10^{-5})} = 1.61245 \times 10^{-3} M = [H_3O^+]$$

$$pH = -\log[H_3O^+] = -\log(1.61245 \times 10^{-3}) = \boxed{2.79}$$

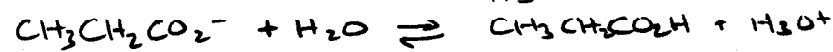
7. What is the pH of a 0.20 M solution of $Na^+CH_3CH_2CO_2^-$? Conj. base of weak acid, so its a weak base need to use K_b to calculate pH

- (a) 3.22 (b) 9.81 (c) 9.09 (d) 7.00 (e) 4.90

$$K_b K_a = K_w$$

$$K_b \text{ of } CH_3CH_2CO_2^- = \frac{K_w}{K_a \text{ of } CH_3CO_2H}$$

$$= \frac{1.008 \times 10^{-14}}{1.3 \times 10^{-5}} = 7.754 \times 10^{-10}$$



$$K_b = \frac{[CH_3CH_2CO_2H][OH^-]}{[CH_3CH_2CO_2^-]} = 7.754 \times 10^{-10} = \frac{x^2}{0.20 - x} \approx \frac{x^2}{0.2}$$

$$x = [OH^-]$$

$$[H_3O^+] = \frac{1.008 \times 10^{-14}}{1.245 \times 10^{-5}}$$

$$= 8.096 \times 10^{-10}$$

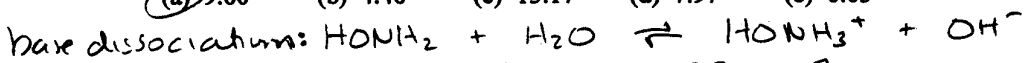
$$pH = -\log(8.096 \times 10^{-10})$$

$$= \boxed{9.09}$$

Acid/Base Calculations

10. Hydroxylamine, HONH_2 , is a weak base with $K_b = 1.07 \times 10^{-8}$. What is the pH of a 0.15M solution of HONH_2 ?

(a) 9.60 (b) 4.40 (c) 13.17 (d) 7.97 (e) 6.03



$$K_b = \frac{[\text{HONH}_3^+][\text{OH}^-]}{[\text{HONH}_2]} = 1.07 \times 10^{-8}$$

	$[\text{HONH}_2]$	$[\text{HONH}_3^+]$	$[\text{OH}^-]$
initial	0.15	0	0
change	-x	+x	+x
equil	0.15-x	x	x

$$\frac{x \cdot x}{0.15-x} = \frac{x^2}{0.15-x} = 1.07 \times 10^{-8}$$

$x = [\text{OH}^-]$
 x will be small so
 $0.15-x \approx 0.15$

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$= -\log(2.52 \times 10^{-10})$$

$$= \boxed{9.60}$$

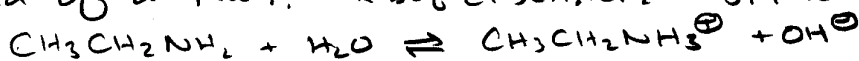
$$\frac{x^2}{0.15} = 1.07 \times 10^{-8} \Rightarrow x = \sqrt{(0.15)(1.07 \times 10^{-8})}$$

$$= 4.0 \times 10^{-5} = [\text{OH}^-]$$

$$[\text{H}_3\text{O}^+] = \frac{1.008 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.008 \times 10^{-14}}{4.0 \times 10^{-5}} = 2.52 \times 10^{-10}$$

17. What is the pH of a 1.0 M solution of ethylamine, $\text{CH}_3\text{CH}_2\text{NH}_2$? ← Organic compound with neutral N, must be a weak base!!
- (a) 10.81 (b) 1.60 (c) 14.00 (d) 5.40 (e) 12.40

Another clue that this is a weak base + not an acid - is that a K_b value is listed instead of a K_a !! K_b of $\text{CH}_3\text{CH}_2\text{NH}_2 = 6.4 \times 10^{-4}$



$$K_b = \frac{[\text{CH}_3\text{CH}_2\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{CH}_2\text{NH}_2]} \Rightarrow \frac{x^2}{1.0-x} = 6.4 \times 10^{-4}$$

$x = [\text{OH}^-]$

assume $1.0-x \approx 1.0$ then $\frac{x^2}{1.0} = 6.4 \times 10^{-4}$

$$x = \sqrt{6.4 \times 10^{-4}} = 2.53 \times 10^{-2} \text{ M} = [\text{OH}^-] \quad [\text{H}_3\text{O}^+] = \frac{1.008 \times 10^{-14}}{2.53 \times 10^{-2}}$$

18. What is the pH of a 1.0 M solution of ethylammonium chloride, $[\text{CH}_3\text{CH}_2\text{NH}_3^+]\text{Cl}^- = 3.98 \times 10^{-13}$

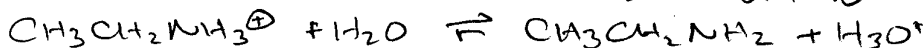
(a) 5.40 (b) 1.60 (c) 12.40 (d) 7.00 (e) 3.50

$\text{CH}_3\text{CH}_2\text{NH}_3^+$ is the conj. acid of the weak base $\text{CH}_3\text{CH}_2\text{NH}_2$ - this makes it

a weak acid. (Cl^- , on the other hand, is the conj. base of a strong acid, HCl , so it is a neutral anion.)

To solve problem, need to know K_a of $\text{CH}_3\text{CH}_2\text{NH}_3^+$
 Get this from K_b of $\text{CH}_3\text{CH}_2\text{NH}_2$

$$K_a \cdot K_b = K_w \Rightarrow K_a = \frac{K_w}{K_b} = \frac{1.008 \times 10^{-14}}{6.4 \times 10^{-4}} = 1.575 \times 10^{-11}$$



$$K_a = \frac{[\text{CH}_3\text{CH}_2\text{NH}_2][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{CH}_2\text{NH}_3^+]} \Rightarrow \frac{x^2}{1.0-x} \approx \frac{x^2}{1.0} = 1.575 \times 10^{-11}$$

$$x = \sqrt{1.575 \times 10^{-11}} = 3.97 \times 10^{-6} \text{ M} = [\text{H}_3\text{O}^+] \quad \text{pH} = -\log(3.97 \times 10^{-6})$$

$$= \boxed{5.40}$$