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

pH = 10.11	acidic / basic / neutral
$[OH^-] = 1.0\times10^{-11}$	acidic / basic / neutral
$[H_3O^+] = 1.0\times10^{-7}$	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) $[H_3O^+] = 5.9\times10^{-5}$ M (c) $[OH^-] = 5.9\times10^{-5}$ M
(d) pOH = 4.55 (e) pOH = 8.45 (ab) pH = 9.71

16. What is the $[H_3O^+]$ to the correct number of significant figures in a pH 8.7412 solution? [fall 00, ex1]

(a) 2×10^{-9} M (b) 1.8×10^{-9} M (c) 1.81×10^{-9} M
(d) 1.815×10^{-9} M (e) 1.8147×10^{-9} M

8. What is the $[OH^-]$ to the correct number of significant figures in an aqueous solution that has a pH of 3.785? [spring 00, ex1]

(a) 6.1×10^{-11} M (b) 6.14×10^{-11} M (c) 6.144×10^{-11} M
(d) 1.6×10^{-4} M (e) 1.64×10^{-4} M (ab) 1.641×10^{-4} M

15. What is the pH of a 4.5×10^{-3} 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 signifcant figures of a 6.75×10^{-3} 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 $Mg(OH)_2$? [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 Ka 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 Kb 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}

13. Phenol, C_6H_5OH, is a weak acid with Ka = 1.3×10^{-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 Ka's given with the exam to find K_a of $CH_3CH_2CO_2H$. K_a of $CH_3CH_2CO_2H$ = 1.3×10^{-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 $K_b = 1.07\times10^{-8}$. 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 $CH_3CH_2NH_2 = 6.4\times10^{-4}$)
(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.**

$[H_3O^+] = \frac{K_w}{[OH^-]}$

	pH = 10.11	acidic / (basic) / neutral
←	$[OH^-] = 1.0\times10^{-11}$	(acidic) / basic / neutral
	$[H_3O^+] = 1.0\times10^{-7}$	acidic / basic / (neutral)

$= \frac{1.008\times10^{-14}}{1.8\times10^{-11}} = 1.0\times10^{-3}$ which is acidic ($> 1.0\times10^{-7}$ M)

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) $[H_3O^+] = 5.9\times10^{-5}$ M** **(c) $[OH^-] = 5.9\times10^{-5}$ M**
(d) pOH = 4.55 **(e) pOH = 8.45** **(ab) pH = 9.71**

ACIDIC
pH < 7
$[H_3O^+] > 1.0\times10^{-7}$ M
pOH > 7
$[OH^-] < 1.0\times10^{-7}$ M

BASIC
pH > 7
$[H_3O^+] < 1.0\times10^{-7}$ M
pOH < 7
$[OH^-] > 1.0\times10^{-7}$ M

16. **What is the $[H_3O^+]$ to the correct number of significant figures in a pH 8.7412 solution?** ← 4 sig figs

(a) 2×10^{-9} M **(b) 1.8×10^{-9} M** **(c) 1.81×10^{-9} M**
(d) 1.815×10^{-9} M 4 sig figs! **(e) 1.8147×10^{-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?** ← 3 sig. figs !! should be 3 sig figs in answer

(a) 6.1×10^{-11} M **(b) 6.14×10^{-11} M** **(c) 6.144×10^{-11} M**
(d) 1.6×10^{-4} M **(e) 1.64×10^{-4} M** **(ab) 1.641×10^{-4} M**

$[H_3O^+] = 10^{-pH} = 10^{-3.785} = 1.64059\times10^{-4}$ M

$[OH^-] = \frac{1.008\times10^{-14}}{1.64059\times10^{-4}} = \boxed{6.14\times10^{-11}}$

15. **What is the pH of a 4.5×10^{-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 $[HBr] = [H_3O^+] = 4.5\times10^{-3}$
$pH = -\log(4.5\times10^{-3}) = \boxed{2.35}$

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

(a) 2.171 (b) 11.83 (c) 2.17 (d) 7.246 (e) 7.24 (ab) 11.826

strong base ↑ 3 sig. figs

$[OH^-] = [KOH] = 6.75\times10^{-3}$

$[H_3O^+] = \frac{1.008\times10^{-14}}{6.75\times10^{-3}} = 1.493\times10^{-12}$ M

$pH = -\log[H_3O^+] = -\log(1.493\times10^{-12}) = \boxed{11.826}$

↑ 3 sig figs ↑ 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 $Mg(OH)_2$?**

STRONG BASE! 3 sig Fig's

(a) 0.864 (b) 12.83 (c) 12.8 (d) 13.1 (e) 13.132 (ab) 0.86

$[OH^-] = 2[Mg(OH)_2] = 2(6.83\times10^{-2}) = 0.1366$ M

↑ there are 2 OH^- in each $Mg(OH)_2$ unit

$[H_3O^+] = \frac{1.008\times10^{-14}}{[OH^-]} = \frac{1.008\times10^{-14}}{0.1366} = 7.379\times10^{-14}$

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

$pH = -\log(H_3O^+) = -\log(7.379\times10^{-14}) = \boxed{13.132}$

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.**

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

(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 $[HBr] = [H_3O^+]$

$[H_3O^+] = 10^{-pH} = 10^{-1.558} = 2.769\times10^{-2}$ M

↑ 3 sig. fig's

So - Need to make 100 mL of a 2.77×10^{-2} M soln of HBr

mw of HBr = 1.008 + 79.904 = 80.912 g/mole

$100.\text{mL}\times\frac{1\text{L}}{1000\text{mL}}\times\frac{2.77\times10^{-2}\text{ mole HBr}}{\text{L}}\times\frac{80.912\text{ g HBr}}{\text{mole}} = \boxed{0.224\text{g}}$

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?

→ 3 sig figs!

(a) 13.264 (b) 13.0 (c) 13.26 (d) 12.980 (e) 1.02

mw of KOH = 39.102 + 15.999 + 1.008 = 56.109 g/mol

KOH is a strong base - all of it converted to OH^-

$KOH(s) \rightarrow K^+(aq) + OH^-(aq)$

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

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

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

$= 12.97986$

$= \boxed{12.980}$

17. The pH of a 0.165 M solution of a certain weak acid is 3.72. What is Ka 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}

$HA + H_2O \rightleftharpoons A^- + H_3O^+$

$K_a = \frac{[A^-][H_3O^+]}{[HA]}$ → come up w/ expressions for concs in terms of 1 variable x → $\frac{x \cdot x}{0.165 - x} = \frac{x^2}{0.165 - x} = K_a$

$x = [H_3O^+]$

$[H_3O^+] = 10^{-pH} = 10^{-3.72} = 1.905 \times 10^{-4} = x$

$K_a = \frac{x^2}{0.165 - x} = \frac{(1.905\times10^{-4})^2}{0.165 - 1.905\times10^{-4}} = \boxed{2.2\times10^{-7}}$

17. The pH of a 0.10 M solution of a certain weak base is 10.85. What is Kb 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}

$B + H_2O \rightleftharpoons HB^+ + OH^-$

$K_b = \frac{[HB^+][OH^-]}{[B]} = \frac{x^2}{[B]_{init} - x}$ $\quad x = [OH^-]$ $\quad [B]_{init} = 0.10$

Calculate $[OH^-]$ from pH

$[H_3O^+] = 10^{-10.85} = 1.4125 \times 10^{-11}$ $\quad [OH^-] = \frac{1.008\times10^{-14}}{1.4125\times10^{-11}}$

$= 7.136\times10^{-4} = x$

$K_b = \frac{(7.136\times10^{-4})^2}{0.10 - 7.136\times10^{-4}} = \boxed{5.1\times10^{-6}}$

Acid/Base Calculations

12 ~~13~~. Phenol, C_6H_5OH, is a weak acid with Ka = 1.3×10^{-10}. What is the pH of a 0.050 M solution of phenol? ← since its a weak acid you need to use Ka to calculate pH

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

$$C_6H_5OH + H_2O \rightleftharpoons C_6H_5O^- + H_3O^+$$

	C_6H_5OH		$C_6H_5O^-$		H_3O^+
Initial	0.050 M		0		0
Change	$-x$		$+x$		$+x$
	$0.050-x$		x		x

$$K = \frac{[C_6H_5O^-][H_3O^+]}{[C_6H_5OH]} = \frac{x\cdot x}{0.05-x} \sim \frac{x^2}{0.050} = 1.3\times10^{-10}$$

$x = [H_3O^+]$

$$x = \sqrt{(0.050)(1.3\times10^{-10})} = 2.\bar{5}5\times10^{-6}\ M = [H_3O^+]$$

$$pH = -\log(2.\bar{5}5\times10^{-6}) = \boxed{5.59}$$

6. What is the pH of a 0.20 M solution of $CH_3CH_2CO_2H$? ← WEAK ACID! need to use Ka to calculate pH

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

$$CH_3CH_2CO_2H + H_2O \rightleftharpoons CH_3CH_2CO_2^- + H_3O^+$$

$$K_a = \frac{[CH_3CH_2CO_2^-][H_3O^+]}{[CH_3CH_2CO_2H]} = 1.3\times10^{-5} = \frac{x^2}{0.20-x} \approx \frac{x^2}{0.20}$$

$x = [H_3O^+]$

$$x = \sqrt{(0.20)(1.3\times10^{-5})} = 1.\bar{6}1245\times10^{-3}\ M = [H_3O^+]$$

$$pH = -\log[H_3O^+] = -\log(1.\bar{6}1245\times10^{-3}) = \boxed{2.79}$$

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 Kb to calculate pH; neutral

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

$K_b K_a = K_w$

$$K_b \text{ of } CH_3CH_2CO_2^- = \frac{K_w}{K_a \text{ of } CH_3CH_2CO_2H} = \frac{1.008\times10^{-14}}{1.3\times10^{-5}} = 7.\bar{7}54\times10^{-10}$$

$$CH_3CH_2CO_2^- + H_2O \rightleftharpoons CH_3CH_2CO_2H + H_3O^+$$

$$K_b = \frac{[CH_3CH_2CO_2H][H_3O^+]}{[CH_3CH_2CO_2^-]} = 7.\bar{7}54\times10^{-10} = \frac{x^2}{0.20-x} \approx \frac{x^2}{0.2}$$

$x = [OH^-]$

$$x = \sqrt{(0.20)(7.\bar{7}54\times10^{-10})} = 1.\bar{2}45\times10^{-5} = [OH^-]$$

$$[H_3O^+] = \frac{1.008\times10^{-14}}{1.\bar{2}45\times10^{-5}} = 8.\bar{0}96\times10^{-10}$$

$$pH = -\log(8.\bar{0}96\times10^{-10}) = \boxed{9.09}$$

Acid/Base Calculations

10. **Hydroxylamine, $HONH_2$, is a weak base with $K_b = 1.07\times10^{-8}$. 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

base dissociation: $HONH_2 + H_2O \rightleftharpoons HONH_3^+ + OH^-$

$$K_b = \frac{[HONH_3^+][OH^-]}{[HONH_2]} = 1.07\times10^{-8}$$

	$[HONH_2]$	$[HONH_3^+]$	$[OH^-]$
initial	0.15	0	0
change	$-x$	$+x$	$+x$
equil	$0.15-x$	x	x

$$\frac{x\cdot x}{0.15-x} = \frac{x^2}{0.15-x} = 1.07\times10^{-8} \qquad x = [OH^-]$$

x will be small so $0.15 - x \approx 0.15$

$$\frac{x^2}{0.15} = 1.07\times10^{-8} \Rightarrow x = \sqrt{(0.15)(1.07\times10^{-8})} = 4.0\times10^{-5} = [OH^-]$$

$$[H_3O^+] = \frac{1.008\times10^{-14}}{[OH^-]} = \frac{1.008\times10^{-14}}{4.0\times10^{-5}} = 2.52\times10^{-10}$$

$pH = -\log[H_3O^+]$

$= -\log(2.52\times10^{-10})$

$= \boxed{9.60}$

17. **What is the pH of a 1.0 M solution of ethylamine, $CH_3CH_2NH_2$?**

(a) 10.81 (b) 1.60 (c) 14.00 (d) 5.40 (e) 12.40

← organic compound with neutral N, must be a weak base!!

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 $CH_3CH_2NH_2 = 6.4\times10^{-4}$

$$CH_3CH_2NH_2 + H_2O \rightleftharpoons CH_3CH_2NH_3^+ + OH^-$$

$$K_b = \frac{[CH_3CH_2NH_3^+][OH^-]}{[CH_3CH_2NH_2]} \text{ or } \frac{x^2}{1.0-x} = 6.4\times10^{-4} \qquad x = [OH^-]$$

assume $1.0 - x \cong 1.0$ then $\frac{x^2}{1.0} = 6.4\times10^{-4}$

$x = \sqrt{6.4\times10^{-4}} = 2.53\times10^{-2}\,M = [OH^-]$ $[H_3O^+] = \frac{1.008\times10^{-14}}{2.53\times10^{-2}} = 3.98\times10^{-13}$

18. **What is the pH of a 1.0 M solution of ethylammonium chloride, $[CH_3CH_2NH_3]Cl$?**

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

$pH = -\log(3.98\times10^{-13})$

$= \boxed{12.40}$

$CH_3CH_2NH_3^+$ is the conj. acid of the weak base $CH_3CH_2NH_2$ - 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 $CH_3CH_2NH_3^+$

Get this from K_b of $CH_3CH_2NH_2$

$$K_a\cdot K_b = K_w \Rightarrow K_a = \frac{K_w}{K_b} = \frac{1.008\times10^{-14}}{6.4\times10^{-4}} = 1.575\times10^{-11}$$

$$CH_3CH_2NH_3^+ + H_2O \rightleftharpoons CH_3CH_2NH_2 + H_3O^+$$

$$K_a = \frac{[CH_3CH_2NH_2][H_3O^+]}{[CH_3CH_2NH_3^+]} \text{ or } \frac{x^2}{1.0-x} \approx \frac{x^2}{1.0} = 1.575\times10^{-11}$$

$x = \sqrt{1.575\times10^{-11}} = 3.97\times10^{-2}\,M = [H_3O^+]$ $pH = -\log(3.97\times10^{-2})$

$= \boxed{5.40}$