③ (a) HY is the strong acid. The diagram showed
(H+1) and (Y-1) ions, but did not show any (H-Y)
molecules. So, all the H-Y molecules had split up
into ions, which a strong acid would do.

(b) of the two remaining acids, HX and HZ,

HX shows   6 HX molecules, & 2 H+1, and 2 X-1

HZ shows   5 HZ molecules, 3 H+1, and 3 Z-1

so HX splits up less than HZ, and is therefore a
weaker acid. HX, the weaker acid, has the lower Ka value.

(c) HX will have the highest pH.
all 3 solutions are acidic, so all 3 will have pH <7.
but HX is the least acidic. It has the fewest H+1 ions,
so will have the highest pH (the most neutral pH / the pH
closest to 7!)

⑧ Which diagram shows a solution of NaF?

Since HF is a weak acid, F-1 will be appreciably basic.

$$F^{-1} + H_2O \rightleftharpoons HF + OH^{-1}$$

fluoride is a weak base; the above rxn does not go to completion

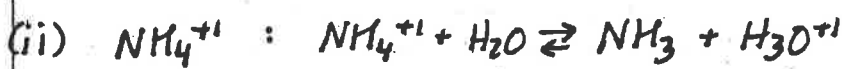so the solution would contain Na+1 ions, F-1 ion, HF molecules,
and OH-1 ions (and H2O).

Solution A doesn't work; it only shows Na+1, F-1, but no HF, OH-

Solution B doesn't work; it shows Na+1 F-1, OH-, but no HF.
acc to stoichiometry, the solution should have one HF for every one OH-

Solution C works.   It does have Na+1, F-1, OH-, and HF.

(15) (a) Find the conjugate base of each
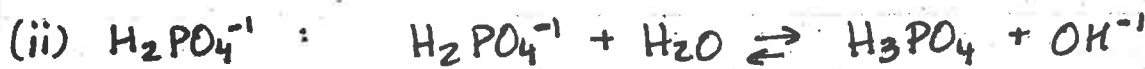(the conjugate base is what forms after the acid donates an H^{+1}/proton)

(i) HIO_3 : $\qquad$ $HIO_3 + H_2O \rightleftarrows H_3O^{+1} + \underline{\underline{IO_3^{-1}}}$

(ii) NH_4^{+1} : $NH_4^{+1} + H_2O \rightleftarrows NH_3 + H_3O^{+1}$

$\boxed{IO_3^{-1} \text{ is the conjugate base of } HIO_3}$

$\boxed{NH_3 \text{ is the conj. base of } NH_4^{+1}}$

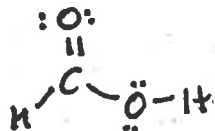(b) Find conj. acid of each (the conj. acid is what forms after the base accepts the H^{+1}/proton)

(i) O^{-2} : $\qquad$ $O^{-2} + H_2O \rightleftarrows OH^{-1} + OH^{-1}$

$\boxed{OH^{-1} \text{ is the conjugate acid of } O^{-2}}$

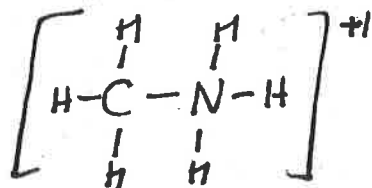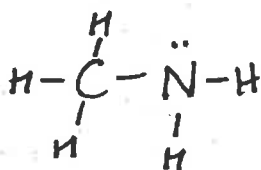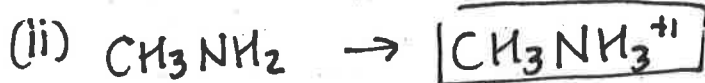(ii) $H_2PO_4^{-1}$: $H_2PO_4^{-1} + H_2O \rightleftarrows H_3PO_4 + OH^{-1}$

$\boxed{H_3PO_4 \text{ is the conjugate acid of } H_2PO_4^{-1}}$

(16.) (a) give the conjugate base of each

(i) $HCOOH$ $\rightarrow$ $\boxed{HCOO^{-1}}$

(ii) HPO_4^{-2} $\rightarrow$ $\boxed{PO_4^{-3}}$

(b) give the conjugate acid of each

(i) SO_4^{-2} $\rightarrow$ $\boxed{HSO_4^{-1}}$

(ii) CH_3NH_2 $\rightarrow$ $\boxed{CH_3NH_3^{+1}}$

(17) ID Bronsted acid, base on LHS, and c.Acid and c.base on RHS

(a) $NH_{4(aq)}^{+1}$ + $CN_{(aq)}^{-1}$ ⇌ $HCN_{(aq)}$ + $NH_{3(aq)}$

    acid   base   conj. acid   conj.base

(b) $(CH_3)_3 N$ + H_2O ⇌ $(CH_3)_3NH^{+1}$ + OH^{-1}

    base   acid   c. acid   c.base

(c) $HCOOH$ + PO_4^{-3} ⇌ $HCOO^{-1}$ + HPO_4^{-2}

    acid   base   conj.base   conj. acid
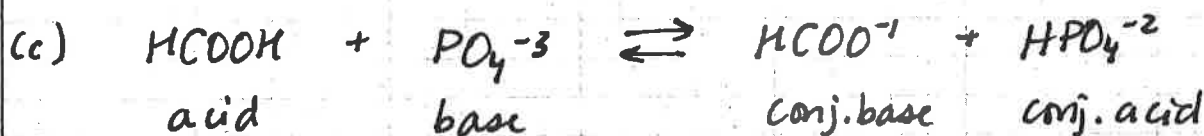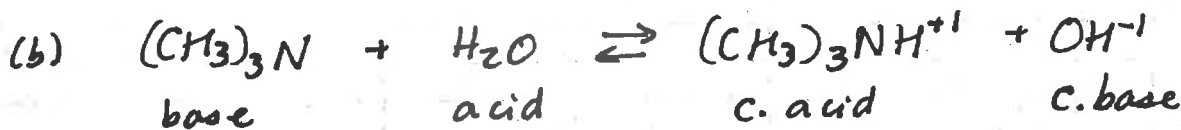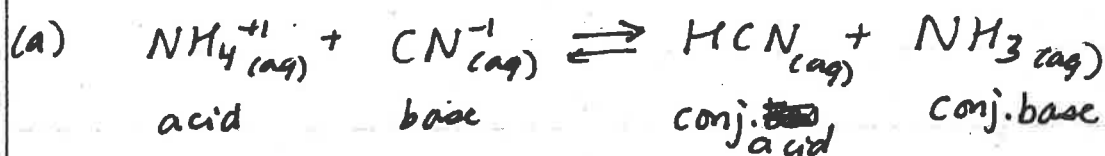
(19) (a) $HC_2O_4^{-1}$ ("hydrogen oxalate ion") is amphiprotic

Rxn where it acts as an _acid_ toward H_2O

$$HC_2O_4^{-1} + H_2O ⇌ C_2O_4^{-2} + H_3O^{+1}$$

⌐ acting as _acid_ since it donates a proton / H^{+1}
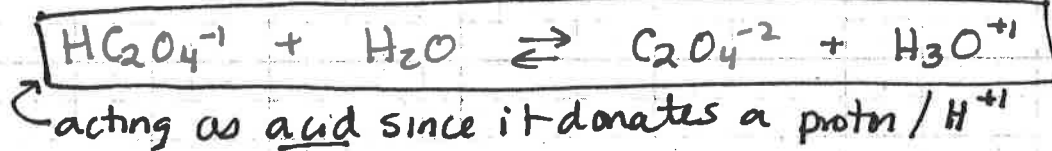
Rxn where it acts as a _base_ toward H_2O :

$$HC_2O_4^{-1} + H_2O ⇌ H_2C_2O_4 + OH^{-1}$$

⌐ acting as a _base_ since it accepts an H^{+1}/proton.

(b) The conjugate acid of $HC_2O_4^{-1}$ is $H_2C_2O_4$

the conj. base of $HC_2O_4^{-1}$ is $C_2O_4^{-2}$

(27) A neutral solution (of H_2O) has pH = 7.00.
Then is heated to 50°C - as it is heated, pH drops to 6.63
Does this mean $[H^+]$ is now > $[OH^-]$? Explain.

No.   Neutral means that $[H^+] = [OH^-]$.

At/near 25°C, $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$

so for neutral solutions at 25°C, $x^2 = 1.0 \times 10^{-14}$

$$x = [H^+] = [OH^-] = 1.0 \times 10^{-7} M$$
$$\text{So pH} = -\log (1.0 \times 10^{-7}) = 7.0 \quad \text{at } 25°C$$

Assuming it is still pure H_2O at 50°C, it will still be neutral; $[H^+]$ and $[OH^-]$ will be equal.

but, since K_w has a larger value at 50°C, $[H^+]$ and $[OH^-]$ will both be higher at 50°C than at 25°C.
They will both be equal to $10^{-6.63}$ or $2.3(44) \times 10^{-7} M$
instead of $1.0 \times 10^{-7} M$.

So $K_w = [H^+][OH^-] = (2.3(44) \times 10^{-7})^2 = \underline{5.5 \times 10^{-14}} (M^2)$ at 50°C.

(29) (C) Calculate $[H^+]$ and classify as acidic/basic/neutral, for a solution in which $[OH^-]$ is 100× greater than $[H^+]$.

$[OH^-] = \underline{[H^+](100)}$ (given)

$[H^+][OH^-] = 1.0 \times 10^{-14}$ (assuming 25°C)

$[H^+][H^+](100) = 1.0 \times 10^{-14}$

$[H^+]^2 = 1.0 \times 10^{-16}$ $[H^+] = (1.0 \times 10^{-16})^{1/2} = \boxed{1.0 \times 10^{-8} M}$

The solution is $\boxed{basic}$ since $[OH^-] > [H^+]$

(30) c) Calculate $[OH^-]$ and classify as acidic/basic/neutral if $[H^+]$ is 1000× greater than $[OH^-]$ in a solution.

$[H^+] = 1000 \cdot [OH^-]$

$[H^+][OH^-] = 1.0 \times 10^{-14}$
$1000 \cdot [OH^-][OH^-] = 1.0 \times 10^{-14}$
$[OH^-]^2 = 1.0 \times 10^{-17}$

$[OH^-] = (1.0 \times 10^{-17})^{1/2} = \boxed{3.2 \times 10^{-9} M}$

This solution is $\boxed{acidic}$
since $[H^+] > [OH^-]$

(33) By what factor does $[H^+]$ change, if the pH changes:
(a) 2.00 units   (b) 0.50 units.

$$[H^+] = 10^{-pH} \quad \text{so} \quad \frac{[H^+]_{final}}{[H^+]_{initial}} = \frac{10^{-pH\,final}}{10^{-pH\,initial}}$$

$$\frac{[H^+]_f}{[H^+]_i} = 10^{-(pH_f - pH_i)}$$

so (a) if the pH changes 2.00 pts, $[H^+]$ will change by a factor of $10^{2.00}$ or 100 (1.0×10^2).

if pH increases 2 points, $[H^+]$ will decrease by a factor of 100, and if pH decreases 2 pts, $[H^+]$ will increase by a factor of 100.

(b) if pH changes 0.50 units, $[H^+]$ will change by a factor of $10^{0.50}$ or 3.2.

(#35 is on next page)

(37)

$[H^+]$	$[OH^-]$	pH	pOH	Acidic or Basic ?
7.5×10^{-3} M	1.3×10^{-12} M	2.12	11.88	Acidic
2.8×10^{-5} M	3.6×10^{-10} M	4.56	9.44	Acidic
5.6×10^{-9} M	1.8×10^{-6} M	8.25	5.75	Basic
5.0×10^{-9} M	2.0×10^{-6} M	8.30	5.70	Basic

given values not underlined. work below

$[H^+] = 7.5 \times 10^{-3}$ M → $pH = -\log(7.5 \times 10^{-3}$ M$) = 2.1249$
$pOH = 14.00 - pH = 14.00 - 2.1249 = 11.8751$
$[OH^-] = 10^{-pOH} = 10^{-11.8751} = 1.333 \times 10^{-12}$ M

acidic since $[H^+] > [OH^-]$

check: $[H^+][OH^-] = (7.5 \times 10^{-3})(1.333 \times 10^{-12}) = 1.0 \times 10^{-14}$ ✓

$[OH^-] = 3.6 \times 10^{-10}$ M
$pOH = -\log(3.6 \times 10^{-10}) = 9.4437$
$pH = 14 - pOH = 14 - 9.4437 = 4.5563$
$[H^+] = 10^{-pH} = 10^{-4.5563} = 2.777 \times 10^{-5}$ M

$pH = 8.25$ $[H^+] = 10^{-8.25} = 5.623 \times 10^{-9}$ M
$pOH = 14 - pH = 14 - 8.25 = 5.75$
$[OH^-] = 10^{-pOH} = 10^{-5.75} = 1.778 \times 10^{-6}$ M

$[H^+] < [OH^-]$ so it is basic

$pOH = 5.70$ $[OH^-] = 10^{-5.70} = 1.995 \times 10^{-6}$ M
$pH = 14 - 5.70 = 8.30$
$[H^+] = 10^{-8.30} = 5.012 \times 10^{-9}$ M

(35) (a) if NaOH is added to H_2O, how will $[OH^-]$ and pH change?

If NaOH is added, $[Na^+]$ and $[OH^-]$ will <u>increase</u>.

$[H^+] = K_w/[OH^-]$ so   as $[OH^-]$ increases, $[H^+]$ will decrease

$pH = -\log[H^+]$ so as $[H^+]$ decreases, <u>pH will increase.</u>

(b) estimate pH of a solution if $[H^+] = .0006\ M$

.0006 M is between 10^{-4} and 10^{-3}, so the pH
will be somewhere between 4 and 3. maybe about
3.3 since .0006 is closer to 10^{-3} than 10^{-4}?

(calculated pH $= -\log(.0006) = 3.22$)

(c) If pH $= 5.2$, estimate, then calculate, $[H^+]$ and $[OH^-]$

$[H^+] = 10^{-pH} = 10^{-5.2}$

if pH $= 5$, $[H^+] = 10^{-5} = 1.0 \times 10^{-5}$
but pH is higher than 5, so $[H^+]$ will be a bit
lower than 1.0×10^{-5}.. maybe 6 or 7×10^{-6}?

(Actual: $[H^+] = 10^{-5.2} = \underline{6.(3) \times 10^{-6}\ M}$)

$pOH = 14 - pH = 8.8$

So $[OH^-] = 10^{-8.8}$ so a bit large than 1×10^{-9}
   maybe 2×10^{-9}?

(Actual: $[OH^-] = 10^{-8.8} = 1.58 \times 10^{-9} \rightarrow \underline{2 \times 10^{-9}\ M}$)

(#37) is on the previous page!

(#39) pH of arterial blood is 7.40
$K_w = 2.4 \times 10^{-14}$ at 37°C (body temp)] calculate $[H^+]$, $[OH^-]$
and pOH

$[H^+] = 10^{-pH} = 10^{-7.40} = 3.98107 \times 10^{-8}\ M \rightarrow \boxed{[H^+] = 4.0 \times 10^{-8}\ M}$

$[OH^-] = K_w/[H^+] = 2.4 \times 10^{-14}/3.98^{07} \times 10^{-8}\ M = 6.0285 \times 10^{-7}\ M$

$\boxed{[OH^-] = 6.0 \times 10^{-7}\ M}$

$pOH = -\log[OH^-] = -\log(6.0285 \times 10^{-7})$
$= 6.2198$ $\rightarrow \boxed{pOH = 6.22}$

41 (a) What is a strong acid?

A strong acid is a compound that splits up ("dissociates") 100% into H^{+1} ions and conjugate base anions when dissolved in H_2O.

(b) What is $[H^{+1}]$ in a 0.500 M solution of HCl?

HCl is a strong acid, so ~~0.500 M~~ HCl will split up into H^{+1} and Cl^{-1} ions 100%, so $[H^+]$ and $[Cl^-]$ will each be 0.500 M.

(c) Which of these (HF, HCl, HBr, HI) is/are strong acids.
HCl, HBr, and HI are strong acids! HF is not. (HF is a weak acid)
(other strong acids include HNO_3, $HClO_3$, $HClO_4$, H_2SO_4 .. the 1st H only)

43 Calculate pH of each strong acid soln.
(a) 8.5×10^{-3} M HBr
$[H^+] = [Br^-] = 8.5 \times 10^{-3}$ M. $pH = -\log(8.5 \times 10^{-3}) =$ $\boxed{2.07}$

(b) 1.52g HNO_3 in 575 mL soln.

$$\frac{(1.52g\ HNO_3)\left(\frac{1 mole}{63.0128\ g}\right)}{0.575 L} = 0.041951\ M\ HNO_3 \nearrow .041951\ M\ H^{+1} \searrow .041951 M\ NO_3^{-1}$$

$pH = -\log(.041951) = \boxed{1.377}$

(c) 5.00 mL of 0.250 M $HClO_4$ diluted to 50.0 mL

$M_1 V_1 = M_2 V_2$ (0.250 M)(5.00 mL) = M_2(50.0mL)

$M_2 = 0.0250$M $HClO_4$ $\nearrow [H^+] = .0250$M $\searrow [ClO_4^-] = .0250$M

$pH = -\log[H^+] = -\log(.0250) = \boxed{1.602}$
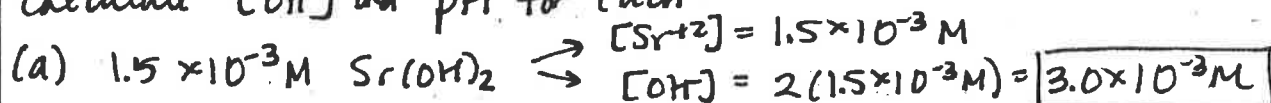
(d) A soln made by mixing   10.0 mL of .100 M HBr w/ 20.0 mL of .200 M HCl

Assume volumes are additive, so total V = 30.0 mL

$$\frac{\overbrace{(.0100\ L)(.100\ \tfrac{mol}{L})}^{moles\ HBr} + \overbrace{(.0200 L)(.200\ \tfrac{mol}{L})}^{moles\ HCl}}{.0300 L} = \frac{(.00100 + .00400)moles\ acid}{.0300\ L}$$

$[acid] = \dfrac{.00500\ mole}{.0300\ L} = 0.167\ M$ $pH = -\log(.167) = \boxed{0.778}$

(45) Calculate [OH] and pH for each:

(a) 1.5×10^{-3} M $Sr(OH)_2$ → $[Sr^{+2}] = 1.5 \times 10^{-3}$ M
$[OH^-] = 2(1.5 \times 10^{-3} M) = \boxed{3.0 \times 10^{-3} M}$

$pOH = -\log(3.0 \times 10^{-3}) = 2.52.$ $pH = 14 - pOH = 14 - 2.52 = \boxed{11.48}$

(b) 2.250 g LiOH in 250.0 mL of soln.

$$\frac{(2.250 g\ LiOH)\left(\frac{1 mole}{23.9483 g}\right)}{(0.2500 L)} = 0.37581\ M\ LiOH$$

→ $[Li^+] = .37581$ M
→ $[OH^-] = .37581$ M

$\boxed{[OH^-] = 0.3758\ M}$

$pOH = -\log(.37581) = 0.425032$

$\left(K_w = 1.01 \times 10^{-14} \text{ at } 25°C \right)$
$\left(pK_w = 13.995(7) \text{ at } 25°C \right)$

$pH = 13.995(7) - pOH$
$= 13.995(7) - .425032$
$\boxed{pH = 13.571}$

(c) 1.00 mL of 0.175 M NaOH diluted to 2.00 L

$M_1 V_1 = M_2 V_2$ $M_2 = \frac{M_1 V_1}{V_2} = \frac{(.175 M)(1.00 mL)\left(\frac{1 L}{1000. mL}\right)}{(2.00 L)}$

$= \boxed{8.75 \times 10^{-5}\ M\ NaOH}$

$pOH = -\log(8.75 \times 10^{-5}) = 4.05799$ $pH = 13.995(7) - 4.05799$
$\boxed{pH = 9.938}$

(d) A soln formed by adding 5.00 mL of 0.105 M KOH
to 15.0 mL of 9.5×10^{-2} M $Ca(OH)_2$

assume volumes are additive so total V = 20.0 mL.

$(.00500 L)(.105 \frac{mol}{L}) = 0.000525$ moles KOH → .000525 mole K^{+1}
.000525 mole OH^-

$(.0150 L)(9.5 \times 10^{-2} \frac{mol}{L}) = .001425$ moles $Ca(OH)_2$ → .001425 mole Ca^{+2}
.002850 mole OH^-

$[OH^-] = \frac{(.000525 + .002850)\ moles\ OH^-}{(.0200 L)} = 0.16875\ M → \boxed{.169\ M}$

$pOH = -\log(.16875) = 0.772756$

$pH = 13.9957 - .772756 = \boxed{13.223}$

(48) Calculate conc'n of $Ca(OH)_2$ if pH = 10.05
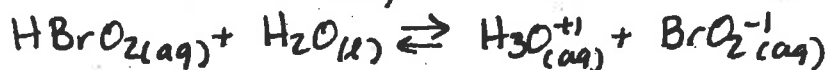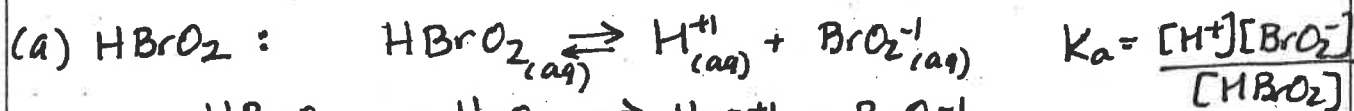
$$[H^+] = 10^{-pH} = 10^{-10.05} = 8.9125 \times 10^{-11} M$$

$$[H^+][OH^-] = K_w = 1.01 \times 10^{-14}$$

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.01 \times 10^{-14}}{(8.9125 \times 10^{-11})} = 0.000113324 \ M$$

$$\left(0.000113324 \ \frac{mole \ OH^-}{L}\right)\left(\frac{1 \ mole \ Ca(OH)_2}{2 \ mole \ OH^-}\right) = 0.00005662 M$$

$$\boxed{0.000057 \ M \ Ca(OH)_2}$$

(49) Write chemical eqn and K_a expression..
Do each twice: once w/ H^{+1} as product,
once w/ H_3O^{+1} as product.

(a) $HBrO_2$:

$$HBrO_{2(aq)} \rightleftharpoons H^{+1}_{(aq)} + BrO_2^{-1}_{(aq)} \qquad K_a = \frac{[H^+][BrO_2^-]}{[HBrO_2]}$$

$$HBrO_{2(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^{+1}_{(aq)} + BrO_2^{-1}_{(aq)}$$

$$K_a = \frac{[H_3O^{+1}][BrO_2^{-1}]}{[HBrO_2]}$$

(b) C_2H_5COOH

$$C_2H_5COOH_{(aq)} \rightleftharpoons H^{+1}_{(aq)} + C_2H_5COO^{-1}_{(aq)}$$

$$C_2H_5COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^{+1}_{(aq)} + C_2H_5COO^{-1}_{(aq)}$$
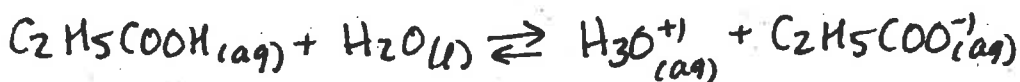
$$K_a = \frac{[H^+][C_2H_5COO^{-1}]}{[C_2H_5COOH]} \qquad or \qquad K_a = \frac{[H_3O^{+1}][C_2H_5COO^{-1}]}{[C_2H_5COOH]}$$

(51) Lactic acid $CH_3CH(OH)COOH$ has one acidic hydrogen.
a 0.10 M soln has pH = 2.44. calculate K_a.

$$CH_3CH(OH)COOH_{(aq)} \rightleftharpoons H^{+1}_{(aq)} + CH_3CH(OH)COO^{-1}_{(aq)} \qquad \leftarrow \text{"lactate ion"}$$

$$(.10 - .003631)M \qquad\qquad .003631 \ M \qquad .003631 \ M$$
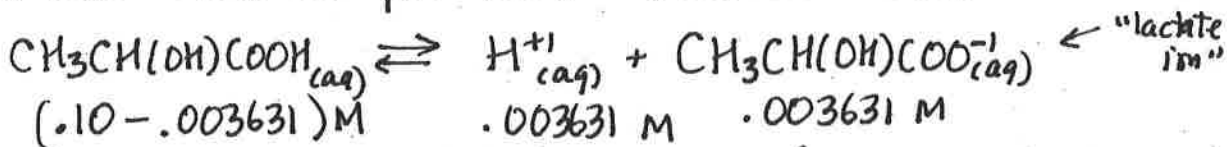
$$[H^+] = 10^{-2.44} = .003631 \ M \ \nearrow$$

By stoichiometry, $[H^+] = [\text{lactate ion}]$

$$K_a = \frac{[H^+][CH_3CHOHCOO^-]}{[CH_3CHOHCOOH]} = \frac{(.003631)(.003631)}{(.10 - .003631)} = .00013 6793$$

$$\boxed{K_a = 1.4 \times 10^{-4}}$$
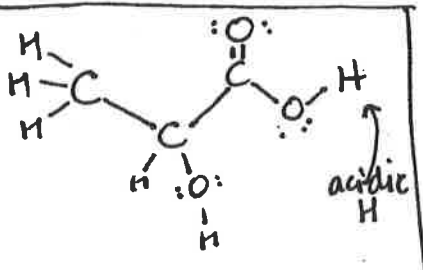
(53) 0.100 M chloroacetic acid ($ClCH_2COOH$) is __11.0%__ ionized. find $[ClCH_2COO^-]$, $[H^+]$, $[ClCH_2COOH]$ in this soln, and K_a.

$(.100 M)(.110) = 0.011 M$

$$ClCH_2COOH_{(aq)} \rightleftharpoons H^{+1}_{(aq)} + ClCH_2COO^{-1}_{(aq)}$$

	$ClCH_2COOH$	H^+	$ClCH_2COO^-$
I	.100 M	$\approx .0$ M	0 M
A	$-.011$ M	$+.011$ M	$+.011$ M
E	.089 M	0.011 M	0.011 M

$\boxed{[ClCH_2COOH] = .089\ M}$ $\boxed{[H^+] = [ClCH_2COO^-] = .011\ M}$

$$K_a = \frac{[H^+][ClCH_2COO^-]}{[ClCH_2COOH]} = \frac{(.011)^2}{(.089)} = \boxed{1.4 \times 10^{-3}}$$

(57) K_a of benzoic acid, C_6H_5COOH, is 6.3×10^{-5}
Find $[H_3O^+]$, $[C_6H_5COO^-]$, $[C_6H_5COOH]$
at eqm if initial acid concn is .050 M

$$C_6H_5COOH_{(aq)} \rightleftharpoons H^{+1}_{(aq)} + C_6H_5COO^{-1}_{(aq)}$$
$$.050 - x \qquad\qquad x \qquad\qquad x$$
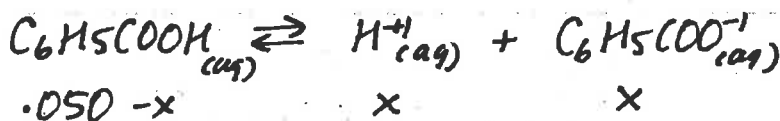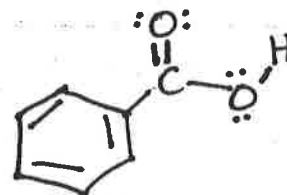
$$K_a = \frac{[H^+][C_6H_5COO^-]}{[C_6H_5COOH]}$$

$6.3 \times 10^{-5} = \dfrac{x^2}{.050 - x}$ ← assume $x \ll .050$

$\dfrac{x^2}{.050} = 6.3 \times 10^{-5}$ $\longrightarrow x = .0017748\ M$

plug in

$\dfrac{x^2}{.050 - .0017748} = 6.3 \times 10^{-5}$ $\longrightarrow x = .001743\ M$

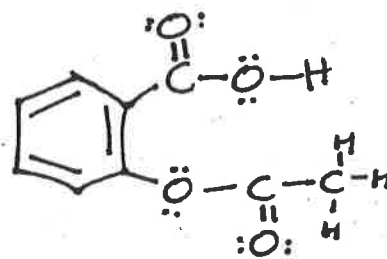$\dfrac{x^2}{.050 - .0017748} = 6.3 \times 10^{-5} \longrightarrow x = .001744\ M$

or $[H_3O^+]$

$\boxed{[H^+] = [C_6H_5COO^-] = x = .0017\ M}$

$\boxed{[C_6H_5COOH] = .050 - x = .050 - .001744 = 0.048\ M}$

(62) Aspirin = acetyl salicylic acid
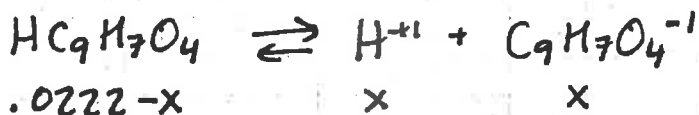
$HC_9H_7O_4$ / $C_9H_8O_4$

$Ka = 3.3 \times 10^{-4}$ @ 25°C

Find pH of a soln:

2 × 500 mg aspirin tablets into 250 mL H_2O

So we have 1000 mg aspirin in 250 mL

$$[Aspirin] = \frac{(1\ g)(\frac{1\ mole}{180.15989})}{0.25\ L} = 0.02(2203)\ M$$

$$HC_9H_7O_4 \rightleftharpoons H^{+1} + C_9H_7O_4^{-1}$$
$$.0222-x \qquad\qquad x \qquad x$$

$$Ka = \frac{[H^+][C_9H_7O_4^{-1}]}{[HC_9H_7O_4]} = \frac{x^2}{.0222-x} = 3.3 \times 10^{-4}$$

try assuming $x \ll .0222$

$$\frac{x^2}{.0222} = 3.3 \times 10^{-4} \longrightarrow x = \underline{.0027068}$$

a bit large to be "neglected"
so we could just do
quadratic formula..
or successive
approximations

$$\frac{x^2}{.0222-.0027068} = 3.3 \times 10^{-4}$$

plug in!
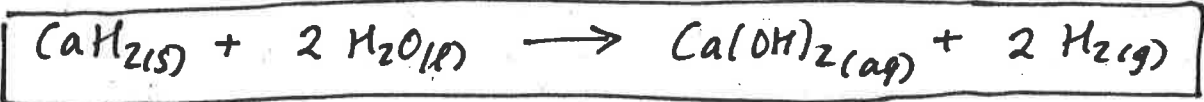
$x = .002536$
again: $x = .002547$
again: $x = .002547$ ..

$$[H^+] = x = .002547\ M$$

$$pH = -\log(.002547) = 2.5940$$

pH 2.59

(or if they really meant 1 SF on the "500mg"
then pH = 2.6)

(66) Calcium hydride: $Ca^{+2}/H^{-1} \rightarrow CaH_2$

(a) Rxn w/ H_2O

$$CaH_{2(s)} + 2 H_2O_{(l)} \longrightarrow Ca(OH)_{2(aq)} + 2 H_{2(g)}$$

(b) $(4.500 \text{ g } H_2)\left(\dfrac{1 \text{ mole}}{2.01589}\right)\left(\dfrac{1 \text{ mole } CaH_2}{2 \text{ mole } H_2}\right)\left(\dfrac{42.09589}{\text{mole}}\right) = \boxed{46.99 \text{ g } CaH_2 \text{ needed}}$

(c) find pH if 0.40 g CaH_2 dissolve to make 1.0 L soln

$(0.40 \text{ g } CaH_2)\left(\dfrac{1 \text{ mole}}{42.09589}\right)\left(\dfrac{1 \text{ mole } Ca(OH)_2}{1 \text{ mole } CaH_2}\right) = 0.0095021 \text{ moles } Ca(OH)_2$

$(.0095021 \text{ mol } Ca(OH)_2)\left(\dfrac{2 \text{ mole } OH^-}{1 \text{ mole } Ca(OH)_2}\right) = .019004 \text{ moles } OH^-$

$pOH = -\log[OH^-] = -\log(.019004) = 1.72115$

moles = molarity since volume was 1.0 L

$pH = pK_W - pOH = 13.9957 - 1.72115 = 12.2745$

$\longrightarrow \boxed{pH = 12.27}$

(d) Verify the $Ca(OH)_2$ can actually be (aq) (dissolved)

K_{sp} of $Ca(OH)_2 = 5.5 \times 10^{-6} = [Ca^{+2}][OH^-]^2$ @ eqm

$Q_{sp} = [Ca^{+2}][OH^-]^2 = \left(\dfrac{.0095021 \text{ mol}}{1.0 \text{ L}}\right)\left(\dfrac{.019004 \text{ mole}}{1.0 \text{ L}}\right)^2$

$Q_{sp} = \underline{3.4 \times 10^{-6}}$

$Q_{sp} < K_{sp}$ so the $Ca(OH)_2$ will be able to dissolve

(e) Volume occupied by 4.50 g H_2, collected by H_2O displacemt at 25°C, $P_{TOTAL} = 745 \text{ mmHg}$

$P_{TOTAL} = P_{H_2} + VP_{H_2O}$

looked up in apphdix B page 1058

$P_{H_2} = P_{TOTAL} - VP_{H_2O} = 745 \text{ mmHg} - 23.76 \text{ mmHg} = 721.24 \text{ mmHg}$

$(4.50 \text{ g } H_2)\left(\dfrac{1 \text{ mole}}{2.01589}\right) = 2.23236 \text{ moles } H_2$

$V = \dfrac{n_{H_2} RT}{P_{H_2}} = \dfrac{(2.23236 \text{ mole})(.0821 \frac{l \cdot atm}{mol \cdot k})(298 K)}{\left(\frac{721.24}{760.} atm\right)} = 57.552$

$$\boxed{V_{H_2} = 57.6 \text{ Liters}}$$

(69) hydroxylamine (a base) is NH_2OH
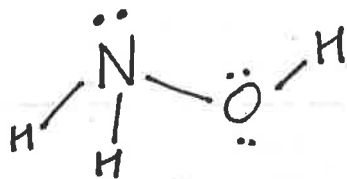
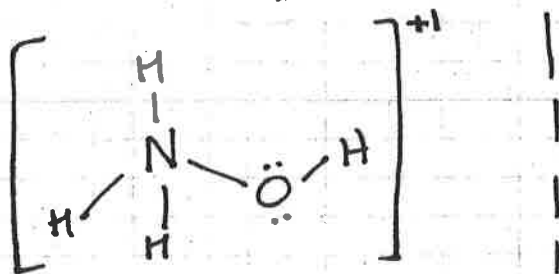(a) The conjugate acid is $\boxed{HNH_2OH^{+1}\ /\ NH_3OH^{+1}\ /\ HONH_3^{+1}}$
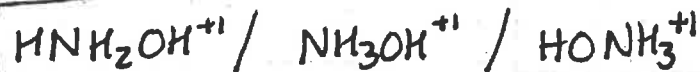
(b) Which atom in NH_2OH accepts the proton when NH_2OH acts as a base?
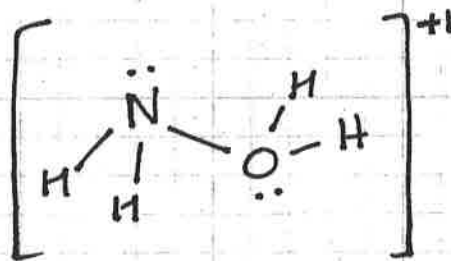
Since it is an _amine_, it would be the Nitrogen that accepts the proton / H^{+1}, since the lone pair of e- on the N attracts the hydrogen ion.

(c) Explain why the N makes a better proton acceptor than the O (since they both have at least one lone pair). Use _Lewis Dots_ and _formal charge_ to answer.

If the N accepts the H^{+1}:   If the O accepts the H^{+1}:

$$Formal\ charge = \left(\begin{array}{c}the\ \#\ of\ valence\ e- \\ in\ the\ unbonded\ atom\end{array}\right) - \left(\begin{array}{c}\#\ of\ valence\ e-\ the \\ atom\ "owns"\ in\ the\ molecule\end{array}\right)$$

Each H : FC = 1-1 = 0
N : FC = 5-4 = 1
O : FC = 6-6 = 0

assume each atom owns all the e- in lone pairs, and _half_ the e- in bonds to that atom

each H : FC = 1-1 = 0
N : FC = 5-5 = 0
O : FC = 6-5 = 1

Both Dot structures have one atom with a Formal charge of +1.
Oxygen is more electronegative than nitrogen (3.5 vs 3.0)
So oxygen should _not_ be the element with the positive formal charge
So the _left_ structure is better; which shows that the N's lone pair / the N accepted the H^{+1}

(72) K_b expression, equation, dotstructures for each base and conjugate acid.

(a) propylamine : $C_3H_7NH_2 + H_2O \rightleftharpoons C_3H_7NH_3^{+1} + OH^{-1}$
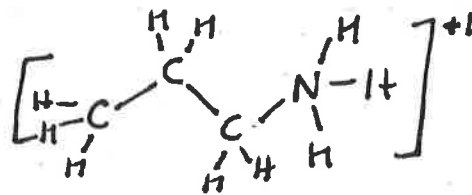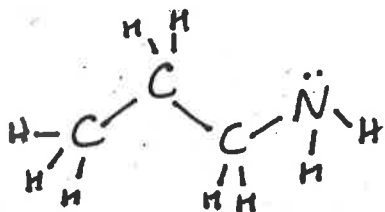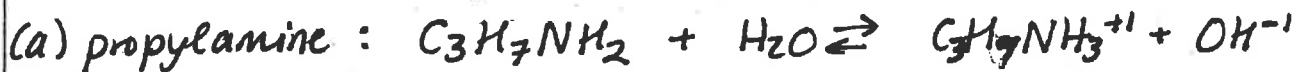
$$K_b = \frac{[C_3H_7NH_3^{+1}][OH^-]}{[C_3H_7NH_2]}$$

(b) monohydrogen phosphate ion

$$HPO_4^{-2} + H_2O \rightleftharpoons H_2PO_4^{+1} + OH^{-1}$$
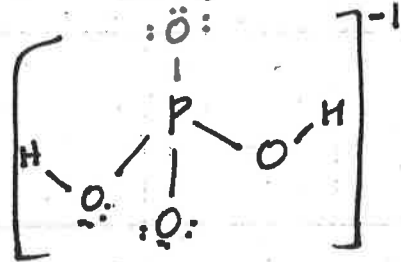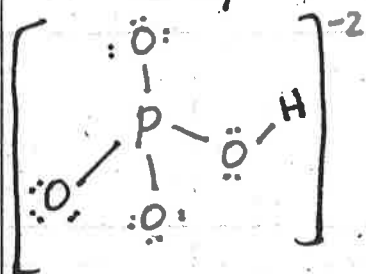
$$K_b = \frac{[H_2PO_4^{-1}][OH^-]}{[HPO_4^{-2}]}$$

(c) benzoate ion :

$$C_6H_5CO_2^{-1} + H_2O \rightleftharpoons C_6H_5COOH + OH^{-1}$$

← benzoic acid
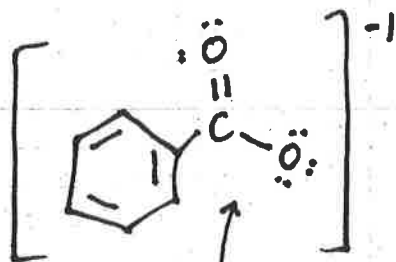
resonance ..
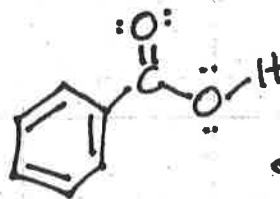
$$K_b = \frac{[C_6H_5COOH][OH^-]}{[C_6H_5COO^{-1}]}$$

and chemical equations

(71) Write the K_b expression for (a) $(CH_3)_2NH$ b) CO_3^{-2} c) CHO_2^{-1}

(a) dimethylamine, $(CH_3)_2NH$

$$(CH_3)_2NH_{(aq)} + H_2O_{(l)} \rightleftharpoons OH^{-1}_{(aq)} + (CH_3)_2NH_2^{+1}_{(aq)} \qquad K_b = \frac{[OH^-][(CH_3)_2NH_2^{+1}]}{[(CH_3)_2NH]}$$

(b) carbonate ion, CO_3^{-2}
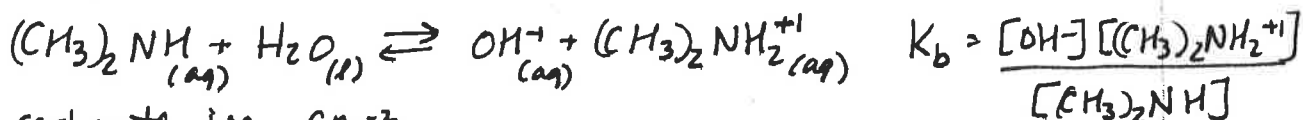
$$CO_3^{-2} + H_2O_{(l)} \rightleftharpoons OH^- + HCO_3^{-1} \qquad K_b = \frac{[OH^-][HCO_3^-]}{[CO_3^{-2}]}$$

(c) formate ion, CHO_2^{-1}

(72 is on previous page)

$$CHO_2^{-1} + H_2O \rightleftharpoons OH^- + HCHO_2 \qquad K_b = \frac{[OH^-][HCHO_2]}{[CHO_2^{-1}]}$$
$$(CHO_2H)$$
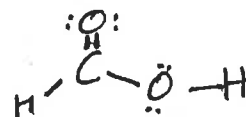$$(HCOOH) \leftarrow \text{formic acid / methanoic acid}$$

$(HC_2H_3O_2)$
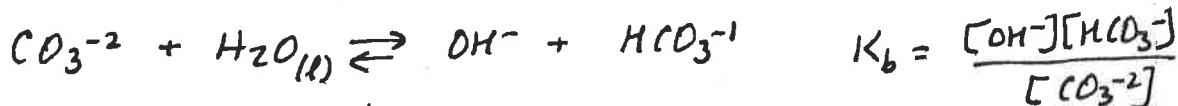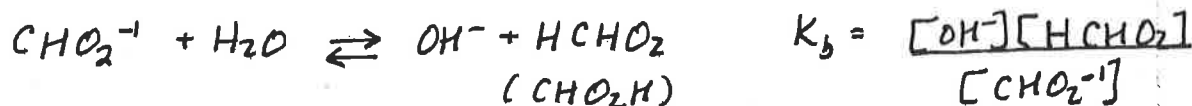
(79) Acetic acid: $K_a = 1.8 \times 10^{-5}$

Hypochlorous acid $(HOCl)$ $K_a = 3.0 \times 10^{-8}$

(a) | Acetic acid | is a stronger acid than hypochlorous acid. Acetic acid has a higher K_a value.

(b) hypochlorite ion is a stronger base than acetate ion.

   Since acetic acid is the stronger acid, its conjugate base will be weaker than that of the other acid.

(c) $K_b(C_2H_3O_2^{-1}) = \dfrac{K_w}{K_a(HC_2H_3O_2)} = \dfrac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = \boxed{5.6 \times 10^{-10} \ K_b \text{ of acetate ion}}$

   $K_b(ClO^{-1}) = \dfrac{K_w}{K_a(HOCl)} = \dfrac{1.0 \times 10^{-14}}{3.0 \times 10^{-8}} = \boxed{3.3 \times 10^{-7} \ K_b \text{ of hypochlorite ion}}$

(80) K_b of $NH_3 = 1.8 \times 10^{-5}$
K_b of $NH_2OH = 1.1 \times 10^{-8}$
(hydroxylamine)

(a) | Ammonia (NH_3) | is the stronger base, since it has the higher K_b value.

(b) | H_3NOH^{+1} | is a stronger acid than NH_4^{+1}. Since NH_3 is a stronger base than NH_2OH, ammonia's conjugate acid will be weaker than that of the other base.

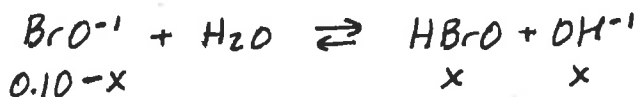(c) K_a of $NH_4^{+1} = \dfrac{K_w}{K_b(NH_3)} = \dfrac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = \boxed{5.6 \times 10^{-10} \ K_a \text{ of } NH_4^{+1}}$

   K_a of $H_3NOH^{+1} = \dfrac{K_w}{K_b(H_2NOH)} = \dfrac{1.0 \times 10^{-14}}{1.1 \times 10^{-8}} = \boxed{9.1 \times 10^{-7} \ K_a \text{ of } H_3NOH^{+1}}$

(76) Codeine $C_{18}H_{21}NO_3$ (weak base)

pH of a 5.0×10^{-3} M solution is 9.95.
Calculate K_b and pK_b of codeine.

$$C_{18}H_{21}NO_{3\,(aq)} + H_2O_{(l)} \rightleftharpoons C_{18}H_{21}NO_3H^{+1}_{(aq)} + OH^{-1}_{(aq)}$$

inital   .0050 M   0   "0"

-8.9125×10^{-5} M $+8.9125 \times 10^{-5}$ M $+8.9125 \times 10^{-5}$ M

.0049109 8.9125×10^{-5} M 8.9125×10^{-5} M

$$K_b = \frac{[C_{18}H_{21}NO_3H^{+1}][OH^-]}{[C_{18}H_{21}NO_3]}$$

$$K_b = \frac{(8.9125 \times 10^{-5})^2}{(.0049109)}$$

pH = 9.95
pOH = 14 - pH = 14 - 9.95 = 4.05
$[OH^-] = 10^{-4.05} = 8.9125 \times 10^{-5}$ M

$$\boxed{K_b = 1.6 \times 10^{-6}}$$

(1.6175×10^{-6})

$$pK_b = -\log(K_b) = -\log(1.6175 \times 10^{-6})$$

$$\boxed{pK_b = 5.79}$$

(81) (a) Calculate pH of 0.10 M NaBrO (!)

from appendix D, K_a of HBrO (hypobromous acid) is 2.5×10^{-9}

$$K_b (BrO^{-1}) = \frac{K_w}{K_a(HBrO)} = \frac{1.0 \times 10^{-14}}{2.5 \times 10^{-9}} = 4.0 \times 10^{-6}$$

$$BrO^{-1} + H_2O \rightleftharpoons HBrO + OH^{-1}$$
$$0.10 - x \qquad\qquad x \qquad x$$

$$K_b = \frac{[HBrO][OH^-]}{[BrO^{-1}]} \qquad\qquad \frac{x^2}{.10 - x} = 4.0 \times 10^{-6}$$

← assume $x \ll 0.10$

$$\frac{x^2}{.10} = 4.0 \times 10^{-6}$$

$$x = 6.3(2) \times 10^{-4} M = [OH^-]$$

$$pOH = -\log(6.32 \times 10^{-4}) = 3.19(897)$$

$$\boxed{pH = 14 - pOH = 10.80}$$

(c) a mixture is

0.10 M $NaNO_2$ + 0.20 M $Ca(NO_2)_2$

$[Na^+] = 0.10 M$ $[Ca^{+2}] = 0.20 M$
$[NO_2^-] = 0.10 M$ $[NO_2^-] = 0.40 M$

Total $[NO_2^-] = 0.10 M + 0.40 M = 0.50 M$

Look up K_a of HNO_2: 4.5×10^{-4}

so K_b of $NO_2^{-1} = \dfrac{K_w}{4.5 \times 10^{-4}} = 2.2(22) \times 10^{-11}$
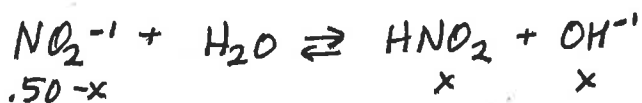
$$NO_2^{-1} + H_2O \rightleftharpoons HNO_2 + OH^{-1} \qquad K_b = \frac{x^2}{.50 - x} = 2.2(22) \times 10^{-11}$$
$$.50 - x \qquad\qquad x \qquad x \qquad\qquad\qquad\qquad\qquad \leftarrow ass.\ x \ll 0.50$$

$$\frac{x^2}{0.50} = 2.2(22) \times 10^{-11}$$

(Note: I am ignoring the
Na^{+1} and Ca^{+2} ions, since

IA and IIA cations are
essentially neutral.)

$$x = [OH^-] = 3.3(33) \times 10^{-6} M$$
$$pOH = 5.48$$
$$\boxed{pH = 14 - pOH = 8.52}$$

(82) b). Calculate pH of a solution of 0.035 M Na_2S

K_{a_1} of H_2S = 9.5×10^{-8}
K_{a_2} of H_2S = 1×10^{-19}

K_b of S^{-2} = $\dfrac{K_w}{K_a(HS^{-1})}$ = $\dfrac{K_w}{K_{a_2} \text{of} H_2S}$ = $\dfrac{1.00 \times 10^{-14}}{1 \times 10^{-19}}$ = 1×10^5 !

since K_b is so large, S^{-2} is essentially a strong base; we'll assume

$S^{-2} + H_2O \rightleftarrows HS^{-1} + OH^-$ ← that this rxn goes essentially to completion.
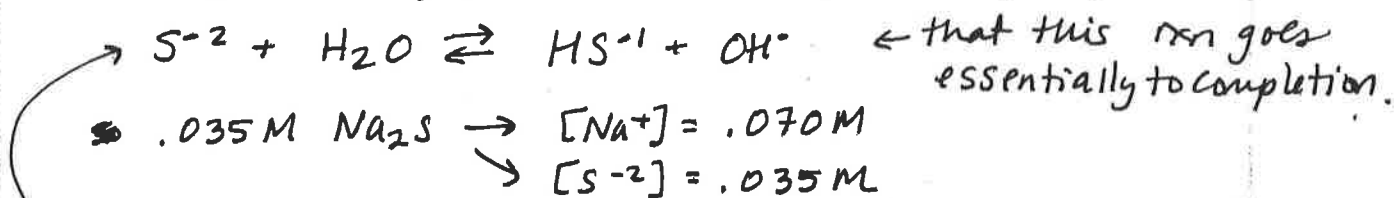
so .035 M Na_2S → $[Na^+]$ = .070 M
$[S^{-2}]$ = .035 M

If this rxn goes to completion, .035 M OH^- will be produced.

$pOH = -\log (.035) = 1.46$ | pH = 14 - pOH = 12.54 |

(HS^{-1} could actually accept another H^{+1}, increasing $[OH^-]$, but it has a K_b of $\dfrac{K_w}{9.5 \times 10^{-8}}$ = 1.1×10^{-7}, so will not increase OH^- enough to worry about it, since we only have 2 SF)

(c) Calculate pH of a solution : $\begin{cases} .045 M & CH_3COONa \\ .055 M & (H_3COO)_2Ba \end{cases}$

.045 M CH_3COONa / $NaC_2H_3O_2$:

$[Na^+]$ = .045 M, $[C_2H_3O_2^-]$ = 0.045 M

0.055 M $(CH_3COO)_2Ba$ / $Ba(C_2H_3O_2)_2$:

$[Ba^{+2}]$ = 0.055 M, $[C_2H_3O_2^-]$ = 0.11 M

total $[C_2H_3O_2^-]$ = .045 + 0.011 = 0.15(5) M

$K_b (C_2H_3O_2^-)$ = $\dfrac{K_w}{K_a (HC_2H_3O_2)}$ = $\dfrac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}}$ = $5.5(56) \times 10^{-10}$

$C_2H_3O_2^- + H_2O \rightleftarrows OH^- + HC_2H_3O_2$
.155~x   x   x
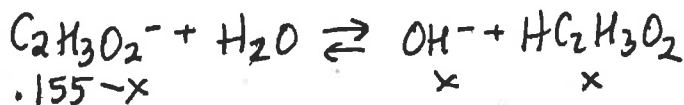
$\dfrac{x^2}{.155-x}$ = 5.556×10^{-10}

ass. x ≪ 0.155

$\dfrac{x^2}{.155}$ = 5.556×10^{-10}

$x = [OH^-] = 9.2(796) \times 10^{-6}$ M
$pOH = -\log (9.2796 \times 10^{-6}) = 5.03$
pH = 14 - pOH = | 8.97 |

N	A	B	B

16.86 An unknown salt is either KBr, NH_4Cl, KCN, or K_2CO_3.
a. If a 0.100 M solution of the salt is neutral, what is the identify of the salt? **KBr**
Explain. Classify each salt as acidic, basic, or neutral as part of your explanation.

(this is more than you need to do) ☺

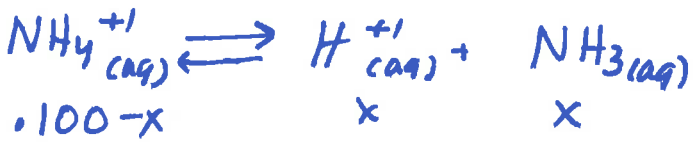(K^{+1} ion is not acidic or basic ; IA ions are neutral)

Br^{-1} is neutral since its the conjugate base of a strong acid (HBr)
So KBr is neutral.

CN^{-1} and CO_3^{-2} are basic since they are the conjugate bases
of weak acids. So KCN and K_2CO_3 are basic.

NH_4^{+1} is acidic since it is a positive ion containing H^{+1}
(and is the conjugate acid of a weak base, NH_3)
Cl^{-1} is neutral since its the conj base of a strong acid. So NH_4Cl is acidic.

b. One of the salts should have been acidic. Calculate the pH of the one that is acidic, still using a concentration of 0.100 M.

$\underline{NH_4Cl.}$ K_a of $NH_4^{+1} = \dfrac{K_w}{K_b(NH_3)} = \dfrac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.5 \times 10^{-10}$

$$NH_4^{+1}{}_{(aq)} \rightleftharpoons H^{+1}{}_{(aq)} + NH_3{}_{(aq)}$$
$$.100-x \qquad\qquad x \qquad\quad x$$

$K_a = \dfrac{[H^+][NH_3]}{[NH_4^+]} = \dfrac{x^2}{.100-x} = 5.5 \times 10^{-10}$

(actually 2 SF)

assume $x \ll 0.100$

$\dfrac{x^2}{.100} = 5.5 \times 10^{-10}$
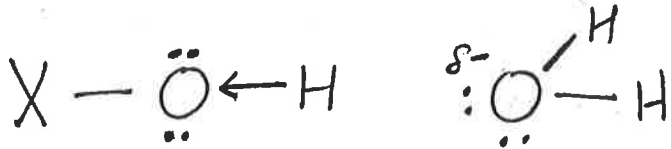
$x = 7.45 \times 10^{-6} M = [H^+]$

$pH = -\log(7.45 \times 10^{-6}) = \boxed{5.13}$

(2 SF)

(87) how does the acidi strength of an oxyacid depend on
(a) the electronegativity (EN) of the central atom?
(b) the number of non   protonated oxygens on the molecule?

as the EN increases, acid strength increases.
as the number of oxygens increases, acid strength increases.

$$X - \ddot{O} \leftarrow H \qquad \overset{\delta-}{:}\ddot{O} \overset{\nearrow H}{-} H$$

the more EN this is, and the
more oxygens attached, the more the
H-O bond in the acid becomes polarized away from
the Hydrogen. as that happens, the H becomes less attracted
to the rest of the acid, and more likely to stick to the
$\delta-$ lone pair on a nearby water molecule, leaving
the acid to form a H_3O^{+1} ion!

(91) Select the stronger base out of each pair.

(The stronger an acid, the weaker the conjugate base!)

(a) BrO^{-1} or ClO^{-1}

These are conj. bases of $HBrO$ and $HClO$. Since Cl is more EN than Br,
$HClO$ is expected to be stronger than $HBrO$. So
$\boxed{BrO^{-1}}$ would be a stronger base than ClO^-, since BrO^{-1}
is the conj. base of a weaker acid.

(b) BrO^{-1} or BrO_2^{-1}

$HBrO_2$ should be more acidic than $HBrO$, since $HBrO_2$ has 1 more oxygen.
So $\boxed{BrO^{-1}}$ should be a stronger base than BrO_2^{-1}

(c) HPO_4^{-2} or $H_2PO_4^{-1}$
                    $\nwarrow$ conj. base of H_3PO_4, which has a Ka of .0075 (Ka_1)
        $\uparrow$
conj. base of $H_2PO_4^{-1}$
which has a Ka of 6.2×10^{-8}
(Ka_2 of H_3PO_4)
This $\boxed{HPO_4^{-2}}$ will be the
stronger base. it is the conj base of
a weaker acid, so will be more basic.
(this makes sense since it is more negatively charged,
so more "attractive" to H^{+1} ions / more able to "accept" a proton)

(106) Calculate pH if 2.50 g Li_2O used to make 1.500 L soln (w/ H_2O)

$$\frac{(2.50g\ Li_2O)\left(\frac{1\ mole}{29.8814g}\right)}{(1.500\ L)} = .055776\ M \leftarrow \text{mole } Li_2O \text{ per liter of solution}$$

Li_2O is "dibasic" since 1 mole Li_2O produces 2 moles $LiOH$ acc to

$$Li_2O + H_2O \rightarrow 2\ LiOH$$
$$(or\ O^{-2} + H_2O \rightarrow 2\ OH^{-})$$
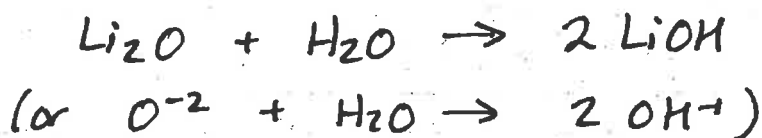
$$(.055776\ M\ Li_2O)\left(\frac{2\ mole\ LiOH}{1\ mole\ Li_2O}\right) = .11155\ M\ LiOH$$

$$[Li^+] = .11155M, \quad [OH^-] = .11155M$$

$$pOH = -\log(.11155M) = 0.95252$$
$$pH = pK_w - pOH = 13.9957 - .95252 = 13.0432$$

$$\boxed{pH = 13.043}$$